CHEMISTRY 143 - INTRODUCTORY CHEMISTRY



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Chemistry 143 - Introductory Chemistry (Bunag)

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1: INTRODUCTION TO CHEMISTRY

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1.1: Defining Chemistry



Figure 1.1.1 (Public Domain; User:Phanton/Wikipedia via https://commons.wikimedia.org/wiki/F...tive_Soaps.jpg)

How does soap relate to chemistry?

Chemistry affects every area of our lives. Here is just one example of chemistry in action – the making of soaps. Soap was once made by boiling animal fat in ashes – the product was hard on the skin and not very pleasant to use. Today, soap manufacture involves complicated chemical processes to provide a wide variety of soaps for different skin types. Colors and odors can be custom-made for that individual experience.

"What is this made of? How can we produce this material quickly and at a low cost? Will this product harm us or help us?"—these are all questions that can be answered using the science of chemistry.

What is Chemistry?

Science is a general term used to describe the principled, rigorous study of the natural world. Many interconnected disciplines fall under this broader concept. For example, physics is the study of motion and forces. Biology is the study of living things. Geology is the study of the Earth and the rocks and minerals of which it is comprised. Chemistry is the study of the composition of matter and the changes that matter undergoes. Matter is anything that has mass and takes up space. Virtually everything around us is matter, including both living and nonliving things. Chemistry affects nearly everything we see and every action we take. Chemistry explains why milk that is left in the refrigerator for too long turns sour. Chemistry explains why certain pollutants called chlorofluorocarbons have done lasting damage to the ozone layer of our planet. Chemistry explains why the leaves of deciduous trees turn from green in the summer to various shades of red and yellow in the autumn (Figure below).

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Figure 1.1.2: Chemical reactions in the leaves of deciduous trees cause them to change color from green to red, orange, or yellow before dropping to the ground. (Public Domain; User:Jongleur100/Wikimedia Commons via https://commons.wikimedia.org/wiki/F...untry_lane.jpg)

Chemistry touches every area of our lives. The medicines we take, the food we eat, the clothes we wear—all of these materials and more are, in some way or another, products of chemistry.

YouTube Video

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Chemists look at the world in two ways, often simultaneously. The two worlds of the chemist are the **macroscopic** world and the **microscopic** world. Macroscopic refers to substances and objects that can be seen, touched, and measured directly. Microscopic refers to the small particles that make up all matter. Chemists must observe matter and do experiments macroscopically; then make generalizations and propose explanations that are microscopic in nature. For example, anyone can observe the physical change in appearance that occurs as an iron object, such as a tractor, is left out in the elements and gradually turns to rust. However, a chemist looking at the rusting tractor considers the individual atoms that make up the iron, and how they are changing as a result of exposure to oxygen in the air, and water from rain. Throughout the study of chemistry, there is often a switch back and forth between the macroscopic and microscopic worlds.



Figure 1.1.3: Rusting artillery shells. (the US Marine Corps, Source: http://commons.wikimedia.org/wiki/File:USMC-060507-M-7799R-003.jpg; License: Public Domain)

Summary

- Chemistry is the study of matter and the changes it undergoes.
- Chemistry considers both macroscopic and microscopic information.

Review

1. Give two examples of chemistry in your everyday life.
2. What is the macroscopic world?
3. What is the microscopic world?

Explore More

1. Read the label on a prepared food product (for example: bread, cereal, dessert). List all of the ingredients in the product. Look up each ingredient on the Internet and write down what that material is doing in the food product.
2. Select your favorite hobby or activity. List all of the items you use in that activity or hobby. For each item, find out how chemistry has contributed to the creation or better operation of that item.

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1.2: History of Chemistry

Where did chemistry come from? Early "chemists" focused on practical problems—how to make dyes and perfumes, soap manufacture, uses of metals, and glass production, among others. The goal was not to understand the physical world—that came later. People just wanted to make things that would improve their lives in some way.

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 and soaps did not require much **theory**, just a good recipe and careful attention to detail. There was no standard way of naming materials (and no periodic table that everyone could agree on). However, the science developed over the centuries.

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



Figure 1.2.1: Robert Boyle. (Public Domain; Johann Kerseboom via Wikipedia)

The field of chemistry began to develop rapidly in the 1700s. 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".

Chemists continued to discover new compounds in the 1800s. 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 theory and application) developed through major contributions by Humphry Davy (1778-1829) and Michael Faraday (1791-1867). Other areas of the discipline also progressed rapidly.

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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.

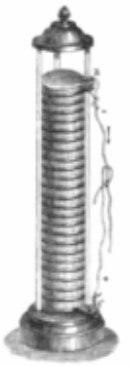


Figure 1.2.2: Battery developed by Volta. (Credit: Adolphe Ganot; Source: http://commons.wikimedia.org/wiki/File:Voltaic_pile_battery.png)

Charles Goodyear (1800-1860) discovered the process of vulcanization, allowing a stable rubber product to be produced for the tires of all the vehicles that 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. 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, which is widely used for inexpensive and sturdy dinnerware.

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Figure 1.2.3: Dynamite explosion in Panama, Central America (1908). (The Field Museum Library; Source: http://commons.wikimedia.org/wiki/File:ExplosionInProgress.jpg; License: Public Domain)

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Summary

- 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 products of modern chemistry.

Review

1. Who invented the first battery?
2. What contribution to chemistry did Robert Boyle make?
3. Who invented dynamite?
4. What was the first synthetic resin used for?

Explore More

Use this resource to answer the following questions: http://www.columbia.edu/itc/chemistr.../chemhist.html

1. Who published the atomic theory?
2. What elements are proposed by Aristotle?
3. Who discovered the proton?
4. What new elements did Marie Curie discover?
5. Who is called the "Father of Modern Chemistry"?

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1.3: 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
- Organic chemistry
- Inorganic chemistry
- Analytical chemistry
- Biochemistry

Physical Chemistry

Physical chemistry is the study of macroscopic properties, atomic properties, and phenomena in chemical systems. As examples, a physical chemist may study 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 chemicals found in every living organism are based on carbon.

Inorganic Chemistry

Inorganic chemistry is the study of chemicals that do not, in general, contain 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.

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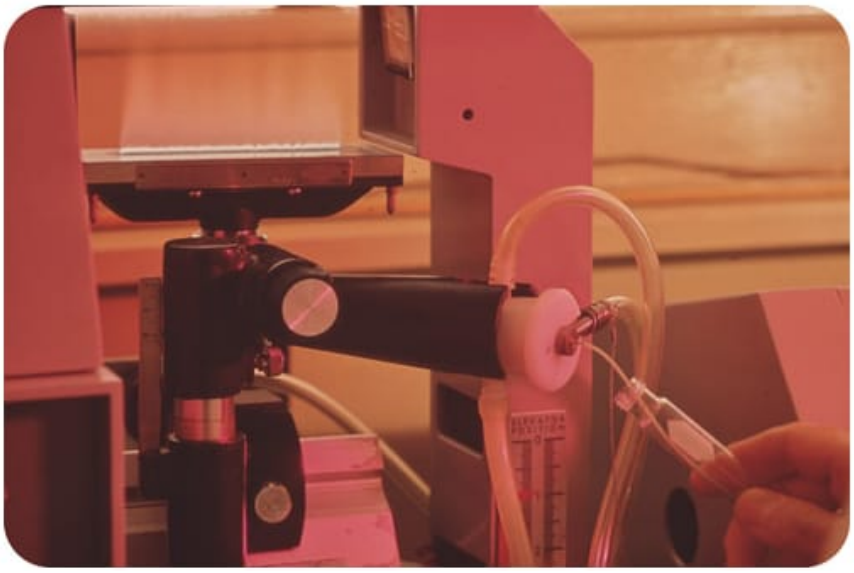


Figure 1.3.1: Measurement of trace metals using atomic spectroscopy. (Bill Shrout, Environmental Protection Agency;n Source: http://commons.wikimedia.org/wiki/File:EPA_GULF_BREEZE_LABORATORY,_THE_MICROBIOLOGY_LAB._AN_ATOMIC_ABSORPTION_SYSTEM,_WHICH_MEASURES_MINUTE_QUAN_NARA_-_546269.tif; License: Public Domain)

Biochemistry

Biochemistry is the study of chemical processes that occur in living things. Research may cover basic cellular processes up to understanding disease states, so that better treatments can be developed.



Figure 1.3.2: Measuring hormone concentrations. (Fred W. Baker III, U.S. Department of Defense; Source: http://commons.wikimedia.org/wiki/File:Defense.gov_photo_essay_090506-D-1852B-019.jpg; License: Public Domain)

In practice, chemical research is often not limited to just one of the five major disciplines. A 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.

Summary

- Five areas of chemistry are described:
  - Physical chemistry
  - Organic chemistry
  - Inorganic chemistry
  - Analytical chemistry
  - Biochemistry

Review

Match the project with the proper chemistry discipline.

Match the project with the proper chemistry discipline.

a. measuring mercury in seawater	1. biochemistry
b. studying enzymes in cells	2. organic chemistry
c. measuring the electrical properties of solutions	3. inorganic chemistry
d. synthesizing new carbon compounds	4. physical chemistry
e. making new compounds for energy processes	5. analytical chemistry

Explore More

1. Do an internet search using one of the five chemistry areas as a search term. List two significant contributions made to chemistry by that area.

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1.4: Pure and Applied Chemistry

"How did chemistry develop? What is happening in the field of chemistry today? What can I do with a chemistry degree?"—these are all good questions, and should be asked by students interested in chemistry. Research in chemistry (or any other field, for that matter) is interesting and challenging. But there are different directions a student can take as they explore research opportunities.

Pure and Applied Research

The study of modern chemistry can be split into two types of research: pure and applied. Chemists who study **pure chemistry** do research primarily to advance mankind's understanding of chemistry. Pure chemistry is concerned with a greater understanding of the theories behind how matter is changing in chemical reactions. Pure chemists tend to be less concerned with direct applications of the research that they are doing. That is not to say that pure chemistry can never lead to a real-world application but, rather, that a potential application is not the primary motivation for doing the research in the first place. **Applied chemistry** is chemistry that is directed toward a specific practical goal or application. The video below further describes applied research.

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Pure Research Examples

The early history of chemistry contains many examples of pure research. The ancient Greek philosophers debated the composition of matter (earth? air? fire? water? all of the above?). They weren't going to do anything with their knowledge – they just wanted to know. Studies on the elements (especially after Mendeleev's periodic table was published) were primarily "pure" research types of experiments. Does this element exist? What are its properties? The scientists did not have any practical application in mind, but were curious about the world around them. Below are some examples of questions that where pure research would be used:

- How was the universe formed?
- Is there life on Mars?
- What are protons, neutrons, and electrons composed of?
- What are the properties of boron?

Applied Research Examples

There is a great deal of "applied" research taking place today. In general, no new science principles are discovered, but existing knowledge is used to develop a new product. Research on laundry detergents will probably not give us any new concepts about soap, but will help us to develop materials that get our clothes cleaner, use less water, and create lower amounts of pollution. Petroleum companies use applied research to find better ways to power vehicles, better lubricants to cut down on engine wear, and better ways to lower air pollution. These companies will use information that is readily available to come up with new products.

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Figure 1.4.1: U.S. Air Force Technical Sergeant Shanice Spearman at Aerospace Fuels Laboratory NCO prepares fuel samples for flashpoint testing. (Public Domain; Airman 1st Class Greg Erwin via https://www.kadena.af.mil/News/Photo...oto/2001869405)

"In-Between" Examples

The line between pure chemistry and applied chemistry is not always distinct. What may start out as simply asking a question may result is some very useful information. If scientists are studying the biochemistry of a microorganism that causes a disease, they may soon find information that would suggest a way to make a chemical that would inactivate the microorganism. The compound could be used to learn more about the biochemistry, but could also be used to cure the disease.

Hemoglobin is a protein in red blood cells that transports oxygen in the bloodstream. Scientists studied hemoglobin simply to learn how it worked. Out of this research came an understanding of how the protein changes shape when oxygen attaches to it. This information was then applied to help patients with sickle cell anemia, a disorder caused by an abnormal hemoglobin structure that makes hemoglobin molecules clump up when oxygen leaves the protein. Basic knowledge of protein structure led to an improved understanding of a wide-spread disease and opened the door for development of treatments.

Summary

- Pure research focuses on understanding basic properties and processes.
- Applied research focuses on the use of information to create useful materials.
- Sometimes there is no clear line between pure and applied research.

Review

1. What is pure research?
2. What is applied research?
3. Give one example of pure research.
4. Give on example of applied research.
5. Is it always easy to classify research as pure or applied? Explain your answer.

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1.5: Energy in Chemistry

What are some types of chemical reactions?

Some chemical reactions can be very calm and boring, while other reactions release a great deal of energy. Dynamite is a chemical that can explode violently. Here we see dynamite being used to move boulders to clear a path for a road. The chemical reactions involved here release tremendous amounts of energy very quickly.

Dynamite was invented by Alfred Nobel in 1866. Nitroglycerin, a very unstable explosive, was already known. Nobel mixed the nitroglycerin with silica to stabilize it and form a solid material. He made a fortune with this discovery and established the Nobel Foundation, which funds the Nobel Prizes every year.



Figure 1.5.1 (Mass Communication Specialist 2nd Class Aaron Burden, U.S. Navy; Source: http://commons.wikimedia.org/wiki/File:US_Navy_090422-N-7130B-318.jpg; License: Public Domain)

Energy in Chemical Bonds

Chemical reactions either require **energy** or release energy. The amount of energy needed or released depends upon the structure of the molecules that are involved in the reaction. Some reactions need to be heated for long periods of time in order for change to take place. Other reactions release energy, allowing heat to be given off to the surroundings. This energy can be used in a variety of ways.

Heating

Coal, natural gas, oil—these materials can be burned to produce heat. The reaction with oxygen releases a great deal of energy that can warm homes and offices. Wood is another example of a chemical (yes, a very complicated one) that will release energy when burned.

Transportation

A major use for petroleum products is fuel for cars, trucks, airplanes, trains, and other vehicles. The chemicals used are usually a mixture of compounds containing several carbon atoms in a chain. When the material is ignited, a massive amount of gas is created almost instantaneously. This increase in volume will move the pistons in an internal combustion engine to provide power. A jet plane works on a similar principle—air comes into the front of the engine and mixes with the jet fuel. This mixture is ignited and the gases formed create a lot of pressure to push the plane forward. The idea can be seen even more clearly in the case of a rocket launch. The ignition of the fuel (either solid-state or liquid) creates gases produced under great pressure that push the

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rocket up.



Figure 1.5.2: Rocket launch. (NASA; Source: http://commons.wikimedia.org/wiki/File:Apollo_15_launch.jpg; License: Public Domain)

Batteries

A major source of energy produced by chemical reactions involves batteries. There are many types of batteries that operate using a variety of chemical reactions. The general principle behind these reactions is the release of electrons that can then flow through a circuit, producing an electrical current.

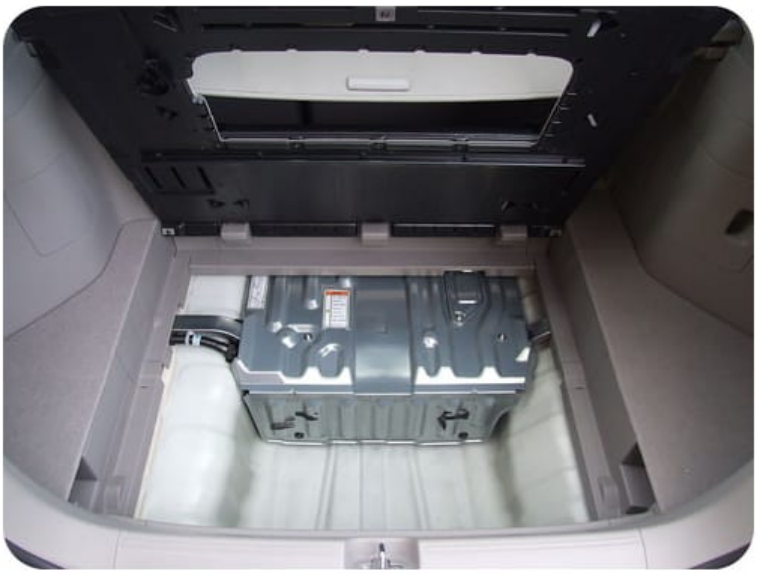


Figure 1.5.3: Car battery. (Public Domain; Toshinori Baba via Wikipedia)

Batteries are used in a wide variety of applications, among them:

1. flashlights
2. watches
3. computers
4. cars
5. hybrid vehicles (provide some power to wheels)
6. cell phones

Batteries in cars, computers, cell phones, and other devices are usually rechargeable. An electric current is passed through the battery to provide electrons that reverse (at least partially) the chemical reactions originally used to create the electric current. Eventually, however, the system can no longer be recharged and the battery has to be discarded.

Hand-Warmers

Hikers, campers, and other outdoor recreationists take advantage of chemical reactions to keep their hands warm. Small containers of chemicals can undergo reactions to generate heat that can be used to avoid frostbite. Some products contain iron filings that will react with air to release thermal energy. These types of warmers cannot be reused. Other systems rely on heat being released when certain chemicals crystallize. If the warmer is placed in very hot water after use, the system can be regenerated.

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Summary

- Some chemical reactions release energy, some require energy.
- Energy released by chemical reactions can be used in a variety of ways.

Review

1. Who invented dynamite?
2. How was the nitroglycerin made more stable?
3. What kind of energy is released when we burn natural gas?
4. How does burning gasoline power a car?
5. How do batteries create energy?
6. Explain how a hand-warmer works.

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1.6: Medicine

How has chemistry made a contributions to medicine?

Diabetes mellitus is a disease characterized by the body's inability to use glucose (a component of table sugar). Glucose is needed to provide biochemical energy for all the cells of the body. When the body cannot make energy using glucose, it begins to break down fat and protein to provide the needed energy, eventually leading to death. Diabetes is the result of the pancreas losing the ability to make insulin, a protein that helps glucose to enter the body's cells and be used for biochemical energy. A key piece of the puzzle surrounding our understanding of diabetes came when Frederick Sanger, a British biochemist, carried out experiments that gave him the structure of the **insulin** molecule. Sanger used basic chemistry techniques and reactions, and took twelve years to complete his research. Today, automated instruments based on his approach can perform the same analysis in a matter of days. Sanger was awarded the Nobel Prize in Chemistry in 1958 for his insulin research.

Chemistry in Medicine

Major contributions to health care have been made by chemistry. The development of new drugs involves chemical analysis and synthesis of new compounds. Many recent television programs advertise the large number of new drugs produced by chemists.

The development of a new drug for any disease is long and complicated. The chemistry of the disease must be studied, as well as how the drug affects the human body. A drug may work well in animals, but not in humans. Out of one hundred drugs that offer the possibility of treating disease, only a small handful actually turn out to be both safe and effective.

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Chemistry contributes to the preparation and use of materials for surgery (sutures, artificial skin, and sterile materials). The sutures used in many surgeries today do not have to be removed, because they simply dissolve in the body after a period of time. Replacement blood vessels for heart and other types of surgery are often made of chemicals that do not react with the tissues, so they will not be rejected by the body. Artificial skin can be used to replace human skin for burn patients.



Figure 1.6.2: Surgical procedure. (Mass Communication Specialist 3rd Class Matthew Jackson, U.S. Navy; Source: http://commons.wikimedia.org/wiki/File:Orif_surgery.jpg; License: Public Domain)

Clinical laboratory testing uses a wide variety of chemical techniques and instrumentation for analysis. Clinical laboratory testing allows us to answer commonly asked questions such as "is your cholesterol high?" and "do you have diabetes?" Some laboratory tests use simple techniques. Other processes involve complex equipment and computer analysis data, in order to perform measurements on large numbers of patient samples.



Figure 1.6.3: Blood samples for laboratory testing. (Photographer's Mate 3rd Class Jeremy L. Grisham, U.S. Navy; Source: http://commons.wikimedia.org/wiki/File:US_Navy_060105-N-8154G-010.jpg; License: Public Domain)

Laboratory testing has come to the local drug store or grocery store because of developments in chemistry. You can test your blood glucose using a simple portable device that runs a chemical test on the blood sample and tells you how much glucose is present, allowing a diabetic patient to regulate how much insulin to administer (chemistry is also used to produce the insulin and the disposable syringe that administers the drug).



Figure 1.6.4: Blood glucose testing device. (Courtesy Christopher Tidy (User:Christidy/Wikipedia); Source: http://commons.wikimedia.org/wiki/File:Glucose_meters.jpg; License: Public Domain)

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Science Friday: The Medical Wonders of Worm Spit

How useful is worm spit? It turns out that worm spit, also known as silk, is a very useful material in medicine. In this video by Science Friday, Dr. David Kaplan describes how silk is used in a variety of medical applications.

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Summary

- Chemistry finds many applications in the healthcare field.
- Development of medicines involves many complicated chemistry processes.
- Chemistry is used to create materials used in surgery.
- Much of laboratory testing is based on chemistry techniques.

Review

1. What chemical is missing in the diabetic patient?

2. Who discovered the structure of insulin?
3. What two things need to be studied to develop a new drug?
4. List two areas where chemistry has helped surgical patients
5. What blood test can be performed using material purchased from your local drugstore?

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1.7: Agriculture

How does chemistry aid the success of crops?

In the spring, many people begin to plant their gardens. They see ads in catalogs or shop the gardening section of a local store to get ideas. The right place in the garden is selected, seeds or plants are put in the soil, and then they wait. Whether it is a small home garden or a large thousand acre farm, chemistry contributes greatly to the success of the crop.

Agriculture

Crops need four things for good growth: sunlight, water, **nutrients** from the soil, and protection from predators such as insects. Chemistry has made major contributions within water usage, nutrient availability, and plant protection. Water purification uses a number of chemical and physical techniques to remove salts and contaminants that would pollute the soil. Chemical analysis of soil allows the grower to see what nutrients are lacking, so that they can be added. In the spring, grocery stores, hardware stores, and gardening centers have high stacks of bags containing fertilizers and weed killers that enrich the soil and keep down unwanted plants. These same stores also provide a number of sprays or solid treatments for insects that might otherwise eat the plants.

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Figure 1.7.1: Idaho wheat field. (the U.S. Department of Agriculture; Source: http://commons.wikimedia.org/wiki/File:Wheat_harvest.jpg; License: Public Domain)

Water Purification

Fresh water is essential for good crops. In some areas of the world, there is enough rain to accomplish this task. In other locales, water must be provided so that the crops will grow. Worldwide, irrigation covers about 18% of farm land and produces some 40% of crops. Obtaining clean water in many parts of the world is made possible by the process of **desalinization**.



Figure 1.7.2: Desalinization equipment. (Lance Cpl. Willard J. Lathrop, U.S. Marine Corps; Source: http://commons.wikimedia.org/wiki/File:US_Navy_060220-M-2061L-060_Hygiene_Equipment_Specialist,_Cpl._Xzavia_Granger_assigned_to_the_31st_Marine_Expeditionary_Unit_%2 License: Public Domain)

Desalinization involves to the treatment of sea water to remove salts; the resulting water can then be used for irrigation without contaminating the soil with materials that harm the growing plants.

Soil Nutrients

In many areas of the world, the soil is deficient in essential nutrients. A number of minerals such as phosphorus, potassium, calcium, and magnesium may not be present in large enough amounts for plants to grow well. Nitrogen is also extremely important

for good crops.



Figure 1.7.3: Man spreading chemicals on soil. (CDC/Dawn Arlotta; Source: http://commons.wikimedia.org/wiki/File:Man_applying_fertilizer.jpg; License: Public Domain)

Soil analysis is available from a variety of labs. Local university extension services can provide valuable information as to the composition of a soil, and will also make suggestions as to the types and amounts of needed nutrients. Fertilizers can be purchased and added to the soil to enrich it and ensure better yield of crops.

Insect Control

Even if the crop grows well, there is still the possibility of insect or pest damage. The insect or pest can consume the crop or can damage it to the point where it will not grow well. Infestations of army worms can do major damage to corn and grain crops. Aphids and boll weevils are major predators of cotton crops. Failure to control these pests results in widespread crop damage and financial loss for the farmer.



Figure 1.7.4: *Potato pests.* (Public Domain; John Curtis via Wikipedia)

A wide variety of **pesticides** have been developed by chemists and other scientists to deal with pests. The basic approach is to develop pesticides that interfere with some biochemical process in the pest. Ideally, the pesticide will not affect other living organisms, but this is not always the case. It is very important to read the labels and observe all precautions when using pesticides.

Summary

- Obtaining clean water in many parts of the world is made possible by the process of desalinization.
- Plant nutrients are very important for good plant growth.
- Chemical analysis of soil can tell he farmer or gardener what nutrients are needed.
- Chemists have developed many pesticides that will kill plant predators such as the army worm and the boll weevil.

Review

1. List three things crops need for good growth.
2. How much of the water used in farming is provided by irrigation?
3. What fraction of crops are grown using irrigation?
4. Why do nutrients need to be added to the soil?
5. How do pesticides work?

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1.8: Materials

How does chemistry affect the clothing that we wear?

Chemistry research is often full of surprises. One such surprise came to Stephanie Kwolek of the DuPont chemical company when she was working on materials known as polymers. These chemicals had been around for a while and were being used for new types of textiles. Kwolek was looking for a strong and rigid petroleum product. She came up with a material that did not look like your average polymer; she played a hunch and had it made into threads. This new material had stiffness about nine times that of any known polymers of the time. Further research and development led to the production of Kevlar, a material now widely used in body armor. Kevlar also found a range of application in racing sails, car tires, brakes, and fire-resistant clothing worn by firefighters.

Materials

Electronics

Chemists are involved in the design and production of new materials. Some of the materials that chemists have helped discover or develop in recent years include polymers, ceramics, adhesives, coatings, and liquid crystals. Liquid crystals are used in electronic displays, as in watches and calculators. The silicon-based computer chip has revolutionized modern society, and chemists have played a key role in the design and continued improvement of these chips. The calculator shown below uses both a liquid crystal display and chips inside the device.



Figure 1.8.1: Calculator with liquid crystal display. (Public Domain; Nom (User:Cjp24/Wikimedia Commons) via Wikipedia)

Superconductors

Many chemists are currently working in the field of superconductivity. Superconductors are materials that are able to conduct electricity with 100% efficiency, meaning that no energy is lost during the electrical transmission, as happens with conventional conduction materials like copper cable. The challenge is to design materials that can act as superconductors at normal temperatures, as opposed to only being able to superconduct at very low temperatures.

Clothing

The fibers that compose the materials for our clothing are either natural or human-made. Silk and cotton are examples of natural fibers. Silk is produced by the silkworm and cotton is grown as a plant. Human-made fabrics include nylon, orlon, and a number of other polymers. These materials are made from hydrocarbons found in petroleum products. Synthetic polymers are also used in shoes, rain gear, and camping items. Synthetic fabrics tend to be lighter than the natural ones, and can be treated to be more water-resistant and durable.

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Materials originally developed as textiles are finding a wide variety of other uses. Nylon is found in a number of plastic utensils. Taking advantage of its strength and light weight, nylon is a component of ropes, fishing nets, tents, and parachutes.



Figure 1.8.2: Nylon spatula. (Public Domain; User:Vanischenu/Wikimedia Commons via Wikipedia)

Summary

- Chemists produce materials for electronics, superconductors, textiles, and other applications.

Review

Use the link below to answer the following questions:

1. Who developed Kevlar?
2. Where are liquid crystals used?
3. What is a superconductor?
4. What are synthetic polymers made from?

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1.9: Chemistry and the Environment

What are the effects of waste within our environment?

Where to dump the garbage? Where to put the waste materials we don't want, after we have made what we do want? For many years we just dumped the waste in the nearest lake, river, empty spot of ground, and didn't worry about it. However, as we learn more about the effects of chemicals on living systems, we have become more concerned about the effects of all this dumping. In many cases, we just didn't know what the long-term effects would be. Today, dumping is generally illegal and offenders pay heavy penalties of fines and jail sentences.

Environment

Many chemicals that were once commonly used were later discovered to be harmful to the **environment**, to human health, or to both. The element **lead** was once a common additive to gasoline and to paint. Plumbing pipes were once typically made of lead. Only since the 1970s has the danger of lead become apparent. It causes brain damage; small children (who might chew on objects painted with lead-based paint) are particularly susceptible. The use of lead in gasoline, paint, and plumbing pipes is now banned; new materials are being developed to replace hazardous lead components.

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Lead is still widely used in car batteries. In recent years, battery **recycling** has become common—about 98% of car batteries are recycled today. The used batteries go to a processing plant where they are crushed and placed in a tank. The lead sinks to the bottom, to be separated and melted. The plastic can then be used in a number of applications. The battery acids are treated with chemicals to neutralize them before they are disposed of.



Figure 1.9.1: Batteries for recycling. (Public Domain; User:Lack Thereof/Wikipedia via Wikipedia)

How Chemists Help the Environment

Chemists are involved in all aspects of environmental protection. In the case of lead, chemists measured the amount of lead in soil, paint, plumbing, blood, and other materials. Chemists studied the chemical processes in the body to see how lead did its damage. Methods for removal of lead from the body were developed with the help of chemists. New gasoline products were developed that fostered fuel efficiency and smooth running of engines, without the presence of lead.

Chemists continue to look for threats to our health and the environment—and to search for alternatives—so that harmful chemicals can be replaced with others that will do the job as effectively, but without the harm.

Summary

- Environmental contamination is a major problem today.
- Lead contamination has serious harmful effects on the body.
- Chemists contribute in many ways to helping with issues of lead contamination.

Review

1. Where was lead found in products?
2. What is a hazard of lead?
3. What happens to the lead in used car batteries?
4. List two ways chemists help with issues related to lead contamination.

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1.10: Alchemy

A common area of technology to societies of the past was metallurgy. Useful tools could be made with metallurgy that would last for a long time. Weapons could stay sharp longer with improved metals. Precious metals, such as gold and silver, could be refined and used for jewelry or for money. Because it was fairly rare, gold was considered to be very valuable, and became a common means of paying for goods and services.

However, mining for gold is a slow, dirty, and dangerous process. Not everyone owns a gold mine—in both the ancient Egyptian society and during the Roman Empire, the gold mines were the property of the state, not an individual or group. There were few ways for most people to legally get any gold for themselves.

Alchemy

Chemistry is a subject that has its roots in the ancient tradition known as alchemy, from which it derives its name. Alchemy was a combination of philosophy and science that had both practical and mystical aspects. The goals of alchemy were varied and difficult to summarize. In many ways, the alchemists sought to achieve perfection, through such actions as the pursuit of the philosopher's stone and the elixir of life (Figure below). The philosopher's stone, it was believed, was a substance that was capable of being used to turn base metals (such as lead) into gold. It was also believed that it could be used to achieve rejuvenation and perhaps immortality. While alchemists did not ultimately succeed in these quests, their work provided the foundation for the modern study of chemistry.

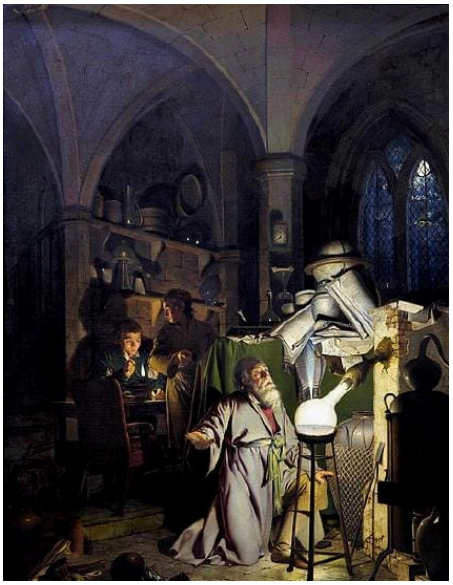
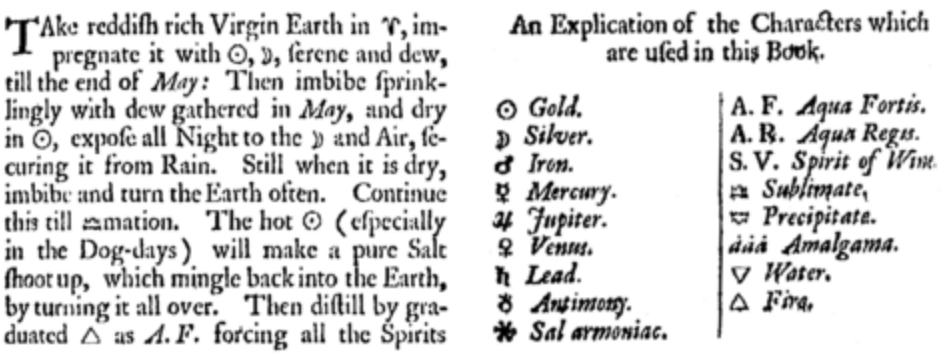


Figure 1.10.1: This painting by Joseph Wright (1771) is titled "The Alchymist, In Search of the Philosopher's Stone." (Public Domain; Joseph Wright of Derby / Derby Museum and Art Gallery via https://commons.wikimedia.org/wiki/F..._Alchemist.jpg)

Many of the specific approaches that alchemists used when they tried changing lead into gold are vague and unclear. Each alchemist had their own code for recording data. The processes were kept secret so that others could not profit from them. Different scholars developed their own set of symbols as they recorded the information they came up with. Many alchemists were not very honest, and sometimes took advantage of noblemen by taking money and claiming to be able to make gold from lead, then left town in the middle of the night. On occasion, the nobleman would detect the fraud and have the alchemist hung. By the 1300s, several European rules had declared alchemy to be illegal, and set out strict punishments for those practicing the alchemical arts.

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TAke reddish rich Virgin Earth in ♈, impregnate it with ☉, ☽, serene and dew, till the end of *May:* Then imbibe sprinklingly with dew gathered in *May*, and dry in ☉, expose all Night to the ☽ and Air, securing it from Rain. Still when it is dry, imbibe and turn the Earth often. Continue this till ♎mation. The hot ☉ (especially in the Dog-days) will make a pure Salt shoot up, which mingle back into the Earth, by turning it all over. Then distill by graduated △ as *A. F.* forcing all the Spirits

An Explication of the Characters which are used in this Book.

☉ *Gold.*
☽ *Silver.*
♂ *Iron.*
☿ *Mercury.*
♃ *Jupiter.*
♀ *Venus.*
♄ *Lead.*
♁ *Antimony.*
✻ *Sal armoniac.*

A. F. *Aqua Fortis.*
A. R. *Aqua Regis.*
S. V. *Spirit of Wine.*
♎ *Sublimate.*
♍ *Precipitate.*
ȧȧȧ *Amalgama.*
▽ *Water.*
△ *Fire.*

Figure 1.10.2: Alchemist recipe. (Public Domain; Kenelm Digby via Wikipedia)

Alchemist Contributions to Chemistry

Alchemists laid the groundwork for many chemical processes, such as the refining of ores, the production of gunpowder, the manufacture of glass and ceramics, leather tanning, and the production of inks, dyes, and paints. Alchemists also made the first attempts at organizing and classifying substances so that they could better understand their reactions and be able to predict the products of their experiments. This eventually led to the modern periodic table, which you will learn about in a later chapter. Alchemy began to fully evolve into chemistry in the 17th century, with a greater emphasis on rational thought and experimentation and less emphasis on spirituality and mysticism.

The alchemists were never successful in changing lead into gold, but modern nuclear physics can accomplish this task. Lead is subjected to nuclear bombardment in a particle accelerator. A small amount of gold can be obtained by this process. However, the cost of the procedure is far more than that of the amount of gold obtained. So, the dream of the alchemists has never (and will never) come true.

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Summary

- Gold has been considered valuable by all civilizations.
- The alchemists tried to find the philosopher's stone that would allow them to make gold from lead.
- The alchemists did not successfully transform lead into gold.
- The alchemists laid the ground work for many advances to the new science of chemistry.

Review

1. Why is gold considered to be valuable?
2. Who owned the gold mines during the ancient Egyptian and Roman civilizations?
3. What is the elixir of life?
4. What contributions to modern chemistry were made by the alchemists?

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1.11: Nature of Science

What can be accomplished through team work?

Most jobs to day involve a fair amount of meetings and team involvement. It doesn't matter what you do, you will spend a certain amount of time meeting with other people, sharing what you have done, planning projects, and organizing the work. Hardly anybody works by him or herself these days – everybody is part of a larger group.

The Nature of Science

Today's scientists rarely work alone. Rather, most scientists collaborate with one another as part of a group effort, no matter the setting. The majority of research scientists work either for a company such as DuPont Chemical Company in Wilmington, Delaware or for one of many universities, such as the California Institute of Technology. Working as part of a group has many advantages. Most scientific problems are so complex and time-consuming, that one person could not hope to address all of the issues by himself or herself. Instead, different members of a research group are each tasked with a particular small aspect of a larger research problem. Collaboration between members of the group is frequent. This occurs informally in the laboratory on an everyday basis. Research groups typically have regular meetings where one or more members of the group may give a presentation to the others on the status of the research that they are doing. Progress normally occurs in small steps rather than grand, sweeping discoveries, and that progress is helped along by the teamwork that comes from working as part of a group.

Modern scientific research is usually expensive. Lab equipment, chemicals, research space, and the upkeep of technical instrumentation all costs money. So research groups need to raise money in order to continue their research. Much of that money comes from government sources, such as the National Science Foundation or the National Institute of Health, especially in the case of research being done at universities. Private companies can fund their own research, but may also seek outside funding as well. Scientists write grants explaining the goals of their research along with projected costs, and funding agencies make decisions on which research projects they would like to fund. The long-term viability of most research labs depends on the ability to get and maintain funding.

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Communicating Results

Suppose that your research is a success. What now? Scientists communicate their results to one another and to the public at large in several ways. One is to publish their research findings in one of many publications called scientific **journals**. There are many hundreds of scientific journals covering every field of science imaginable. In chemistry, there is the "Journal of the American Chemical Society," the "Journal of Physical Chemistry," and the "Inorganic Chemistry," to name just a few. Some journals have a very narrow scope while others publish articles from many different sciences and appeal to a wider audience. Examples of the latter include the journals "Science" and "Nature." Journal articles are often very complex and detailed. They must be accurate, since the research field as a whole uses these journal articles as a way to make scientific progress. Therefore, journal articles are only published after having been extensively reviewed by other professional scientists in the same field. Reviewers have the power to make suggestions about the research or possibly question the validity of the author's conclusions. Only when the reviewers are satisfied that the research is correct, will the journal publish the article. In this way, all scientists can trust the research findings that they read about in journals.

Scientists also communicate with one another by presenting their findings at international conferences. Some scientists are

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chosen to give a lecture at a conference, typically about research that has already been published. Many other scientists at the same conference will present their work at **poster sessions**. These poster sessions are more informal and may often represent research that is still in progress.

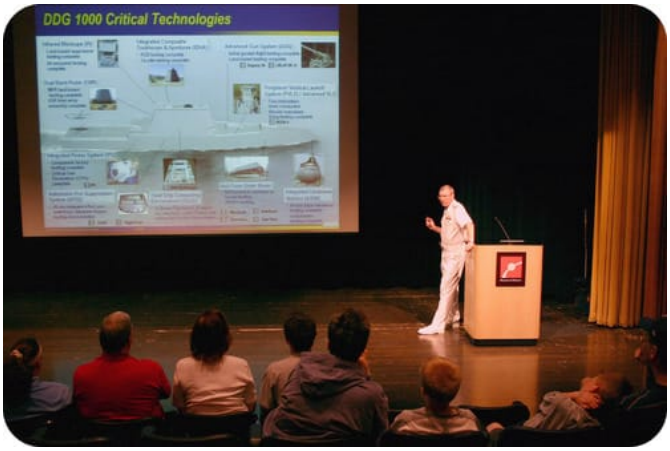


Figure 1.11.1: Scientific presentation. (Public Domain; Journalist 1st Class Dave Kaylor, U.S. Navy via Wikipedia)

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Science Friday

A Shot in the Dark: Alternative Uses for Squid Ink

Squid ink is commonly thought of as a defensive mechanism used by squids, but could there be other uses? In this video by Science Friday, Dr. Stephanie Bush talks about possible other functions of squid ink.

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Snowflake Safari

Next snowstorm, grab a magnifying glass and look carefully at snowflakes. Bullet rosettes, stellar plates and capped columns are just a few of the many varieties of snow crystals. In this video by Science Friday, physicist Kenneth Libbrecht shares secrets about snowflakes.

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Summary

- Scientific research today is a team effort.
- Support for research usually comes from grants.
- Findings are shared in meetings, conferences, and scientific publications.

Review

1. How is most scientific research carried out today?
2. How is research usually supported?
3. How do scientists share their findings?

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1.12: Scientific Problem Solving

How can we use problem solving in our everyday routines?

One day you wake up and realize your clock radio did not turn on to get you out of bed. You are puzzled, so you decide to find out what happened. You list three possible explanations:

1. There was a power failure and your radio cannot turn on.
2. Your little sister turned it off as a joke.
3. You did not set the alarm last night.

Upon investigation, you find that the clock is on, so there is no power failure. Your little sister was spending the night with a friend and could not have turned the alarm off. You notice that the alarm is not set—your forgetfulness made you late. You have used the scientific method to answer a question.

Scientific Problem Solving

Humans have always wondered about the world around them. One of the questions of interest was (and still is): *what is this world made of?* Chemistry has been defined in various ways as the study of matter. What matter consists of has been a source of debate over the centuries. One of the key areas for this debate in the Western world was Greek philosophy.

The basic approach of the Greek philosophers was to discuss and debate the questions they had about the world. There was no gathering of information to speak of, just talking. As a result, several ideas about matter were put forth, but never resolved. The first philosopher to carry out the gathering of data was Aristotle (384-322 B.C.). He recorded many observations on the weather, on plant and animal life and behavior, on physical motions, and a number of other topics. Aristotle could probably be considered the first "real" scientist, because he made systematic observations of nature and tried to understand what he was seeing.



Figure 1.12.1: Aristotle. (Public Domain; Raphael via Wikipedia)

Inductive and Deductive Reasoning

Two approaches to logical thinking developed over the centuries. These two methods are **inductive reasoning** and **deductive reasoning**. Inductive reasoning involves getting a collection of specific examples and drawing a general conclusion from them. Deductive reasoning takes a general principle and then draws a specific conclusion from the general concept. Both are used in the development of scientific ideas.

Inductive reasoning first involves the collection of data: "If I add sodium metal to water, I observe a very violent reaction. Every time I repeat the process, I see the same thing happen." A general conclusion is drawn from these observations: the addition of sodium to water results in a violent reaction.

In deductive reasoning, a specific prediction is made based on a general principle. One general principle is that acids turn blue

litmus paper red. Using the deductive reasoning process, one might predict: "If I have a bottle of liquid labeled 'acid', I expect the litmus paper to turn red when I immerse it in the liquid."

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The Idea of the Experiment

Inductive reasoning is at the heart of what is now called the "**scientific method**." In European culture, this approach was developed mainly by Francis Bacon (1561-1626), a British scholar. He advocated the use of inductive reasoning in every area of life, not just science. The scientific method, as developed by Bacon and others, involves several steps:

1. Ask a question - identify the problem to be considered.
2. Make observations - gather data that pertains to the question.
3. Propose an explanation (a hypothesis) for the observations.
4. Make new observations to test the hypothesis further.



Figure 1.12.2: Sir Francis Bacon. (Public Domain; Paul van Somer via Wikipedia)

Note that this should not be considered a "cookbook" for scientific research. Scientists do not sit down with their daily "to do" list and write down these steps. The steps may not necessarily be followed in order. But this does provide a general idea of how scientific research is usually done.

When a hypothesis is confirmed repeatedly, it eventually becomes a theory—a general principle that is offered to explain natural phenomena. Note a key word—*explain*, or *explanation*. A theory offers a description of why something happens. A law, on the other hand, is a statement that is always true, but offers no explanation as to why. The law of gravity says a rock will fall when dropped, but does not explain why (gravitational theory is very complex and incomplete at present). The kinetic molecular theory of gases, on the other hand, states what happens when a gas is heated in a closed container (the pressure increases), but also explains why (the motions of the gas molecules are increased due to the change in temperature). Theories do not get "promoted" to laws, because laws do not answer the "why" question.

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Summary

- The early Greek philosophers spent their time talking about nature, but did little or no actual exploration or investigation.
- Inductive reasoning - to develop a general conclusion from a collection of observations.
- Deductive reasoning - to make a specific statement based on a general principle.
- Scientific method - a process of observation, developing a hypothesis, and testing that hypothesis.

Review

1. What was the basic shortcoming of the Greek philosophers approach to studying the material world?
2. How did Aristotle improve the approach?
3. Define “inductive reasoning” and give an example.
4. Define “deductive reasoning” and give an example.
5. What is the difference between a hypothesis and a theory?
6. What is the difference between a theory and a law?

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2: MEASUREMENTS

Measurement is the assignment of a number to a characteristic of an object or event, which can be compared with other objects or events.

2.1: SI Base Units

2.2: Metric Prefixes

2.3: Scientific Notation in Chemistry

2.4: SI Length and Volume Units

2.5: SI Mass and Weight Units

2.6: SI Kinetic Energy Units

2.7: Temperature and Temperature Scales

2.8: Scientific Dimensional Analysis

2.9: Metric Unit Conversions

2.10: Derived Units

2.11: Density

2.12: Accuracy and Precision

2.13: Percent Error

2.14: Measurement Uncertainty

2.15: Significant Figures

2.16: Rounding

2.17: Uncertainty in Multiplication and Division

2.18: Uncertainty in Addition and Subtraction

2.19: Measurements and Problem-Solving (Exercises)

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2.1: SI Base Units

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 a football player moves the ball ten yards, they get a first down, and it doesn't matter whether they are playing in Los Angeles, Dallas, or Green Bay. At one time, however, 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 in that case is simple: new king, new distance (and then you have to re-mark all those football fields).

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, listed in Table 2.1.1.

Table 2.1.1: 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.



Figure 2.1.1: Meter Standard. (Public Domain; US Department of Commerce; National Institute of Standards and Technology via Wikipedia; http://commons.wikimedia.org/wiki/File:Standard_kilogram,_2.jpg)



Figure 2.1.2: Kilogram standard (Public Domain; National Institute of Standards and Technology via Wikipedia)

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The map below shows the adoption of the SI units in countries around the world. The United States has legally adopted 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.

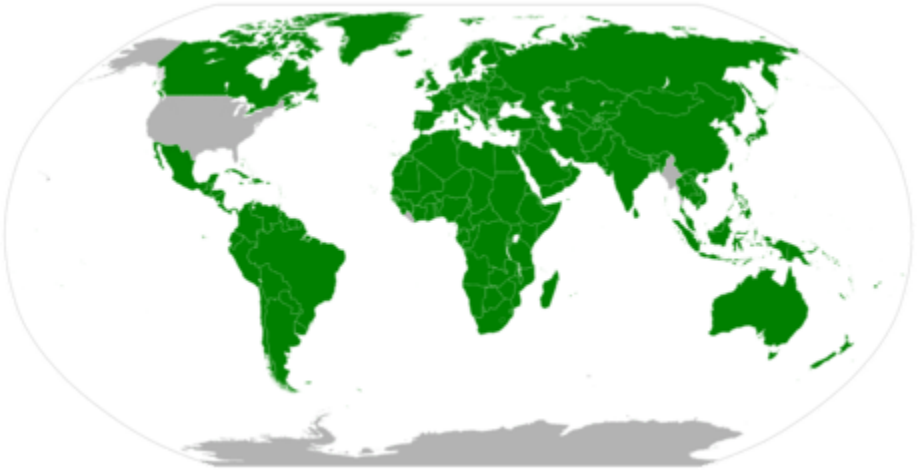


Figure 2.1.3: Areas of the world that use the metric system (green). (Public Domain; User:Canuckguy/Wikipedia and User:AzaToth/Wikimedia Commons via Wikipedia)

Summary

- The SI system is based on multiples of ten.
- There are seven basic units in the SI system.
- Five of these seven SI units are commonly used in chemistry.

Review

1. What does SI stand for?
2. When was this system adopted by the international community?
3. Which of the units are commonly used in chemistry?

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2.2: Metric Prefixes



Figure 2.2.1 (Public Domain; Joseph Wright of Derby via Wikipedia)

What do Latin and Greek have to do with modern science?

Isn't it hard enough to learn English terms? For hundreds of years, the languages of the educated class were Latin and Greek. In part, because the literature of philosophy was Latin and Greek. Even the medieval Bibles were written in those two languages – the first English translation was in the late 1380s. Using Latin and Greek allowed scholars from different countries to communicate more easily with one another. Today we still see many Latin phrases in legal communications ("pro bono" meaning to do something "for the good" and not charge legal fees), scientific naming of biological species, and Latin is used for the annual student speech at Harvard University graduations. Not bad for a "dead" language.

Metric Prefixes

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 2.2.1 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 2.2.1: SI Prefixes

Prefix	Unit Abbreviation	Meaning	Example
giga	G	1,000,000,000	1 gigameter $(\mathrm{Gm}) = 10^9\ \mathrm{m}$
mega	M	1,000,000	1 megameter $(\mathrm{Mm}) = 10^6\ \mathrm{m}$
kilo	k	1,000	1 kilometer $(\mathrm{km}) = 1,000\ \mathrm{m}$
hecto	h	100	1 hectometer $(\mathrm{hm}) = 100\ \mathrm{m}$
deka	da	10	1 dekameter $(\mathrm{dam}) = 10\ \mathrm{m}$
		1	**1 meter** (m)

Prefix	Unit Abbreviation	Meaning	Example
deci	d	1/10	1 decimeter $(\mathrm{dm}) = 0.1\ \mathrm{m}$
centi	c	1/100	1 centimeter $(\mathrm{cm}) = 0.01\ \mathrm{m}$
milli	m	1/1,000	1 millimeter $(\mathrm{mm}) = 0.001\ \mathrm{m}$
micro	μ	1/1,000,000	1 micrometer $(\mu\mathrm{m}) = 10^{-6}\ \mathrm{m}$
nano	n	1/1,000,000,000	1 nanometer $(\mathrm{nm}) = 10^{-9}\ \mathrm{m}$
pico	p	1/1,000,000,000,000	1 picometer $(\mathrm{pm}) = 10^{-12}\ \mathrm{m}$

There are more prefixes - some of them rarely used. Have you ever heard of a zeptometer?

There are a couple of odd little practices with the use of metric abbreviations. Most abbreviations are lowercase. The lowercase "m" is used for meter, instead of "M". However, when it comes to volume, the base unit "liter" is abbreviated as "L" and not "l". So, 3.5 milliliters is written 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 that 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.

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Simulations

Practice using metric prefixes with this simulation.

Summary

- Metric prefixes derive from Latin or Greek terms.
- Metric prefixes are used to make the units manageable.

Review

1. What is the prefix for “thousand”?
2. What is the prefix for 0.01?
3. How would you write 500 milliliters?
4. How many decimeters in one meter?
5. You have a mass that weighs 1.2 hectograms. How many grams does it weigh?

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2.3: Scientific Notation in Chemistry

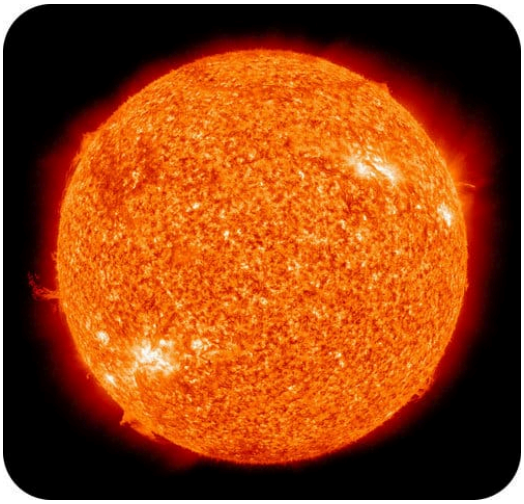


Figure 2.3.1 (Public Domain; NASA via Wikipedia)

How far is the Sun from Earth?

Astronomers are accustomed to very large numbers. While the moon is only $406,697\ \text{km}$ from earth at its maximum distance, the sun is much further away $(150\ \text{million km})$. Proxima Centauri, the star nearest the earth, is $39,900,000,000,000\ \text{km}$ away—and this is just the start of long distances. On the other end of the scale, some biologists deal with very small numbers: a typical fungus could be as small as 30 micrometers $(0.000030\ \text{meters})$ in length, and a virus might only be 0.03 micrometers $(0.00000003\ \text{meters})$ long.

Scientific Notation

Scientific notation is a way to express numbers as the product of two numbers: a coefficient and the number 10 raised to a power. It is a very useful tool for working with numbers that are either very large or very small. As an example, the distance from Earth to the Sun is about 150,000,000,000 meters—a very large distance indeed. In scientific notation, the distance is written as $1.5 \times 10^{11}\ \text{m}$. The coefficient is the 1.5 and must be a number greater than or equal to 1, and less than 10. The power of 10, or exponent, is 11 because you would have to multiply 1.5 by 10^{11} to get the correct number. Scientific notation is sometimes referred as exponential notation. A summary of SI units is given in the table below.

Table 2.3.1: SI Prefixes

Prefix	Unit Abbreviation	Exponential Factor	Meaning	Example
giga	G	10^9	1,000,000,000	1 gigameter $(\text{Gm}) = 10^9\ \text{m}$
mega	M	10^6	1,000,000	1 megameter $(\text{Mm}) = 10^6\ \text{m}$
kilo	k	10^3	1,000	1 kilometer $(\text{km}) = 1,000\ \text{m}$
hecto	h	10^2	100	1 hectometer $(\text{hm}) = 100\ \text{m}$
deka	da	10^1	10	1 dekameter $(\text{dam}) = 10\ \text{m}$
		10^0	**1**	**1 meter** (m)
deci	d	10^{-1}	1/10	1 decimeter $(\text{dm}) = 0.1\ \text{m}$
centi	c	10^{-2}	1/100	1 centimeter $(\text{cm}) = 0.01\ \text{m}$
milli	m	10^{-3}	1/1,000	1 millimeter $(\text{mm}) = 0.001\ \text{m}$

Prefix	Unit Abbreviation	Exponential Factor	Meaning	Example
micro	μ	10^{-6}	1/1,000,000	1 micrometer $(\mu\text{m}) = 10^{-6}$ m
nano	n	10^{-9}	1/1,000,000,000	1 nanometer $(\text{nm}) = 10^{-9}$ m
pico	p	10^{-12}	1/1,000,000,000,000	1 picometer $(\text{pm}) = 10^{-12}$ m

When working with small numbers, we use a negative exponent. So 0.1 meters is 1×10^{-1} meters, 0.01 is 1×10^{-2} and so forth. The table above gives examples of smaller units. Note the use of the **leading zero** (the zero to the left of the decimal point). That digit is there to help you see the decimal point more clearly. The figure 0.01 is less likely to be misunderstood than .01, where you may not see the decimal.

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Summary

- Scientific notation allows us to express very large or very small numbers in a convenient way.
- Scientific notation uses a coefficient (a number between 1 and 10) and a power of ten sufficient for the actual number.

Review

1. What is scientific notation?
2. What do we use scientific notation for?
3. What is a leading zero?
4. Express 150,000,000 in scientific notation.
5. Express 0.000043 in scientific notation.

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2.4: SI Length and Volume Units



Figure 2.4.1 (Public Domain; User:Żeglarz/Wikimedia Commons via Wikipedia)

How were sailors able to measure the depths of seas?

Back in the days before all the electronic gadgets for measuring depth and locating undersea objects existed, the "fathom" was the unit of measurement for depth. A rope was knotted every six feet and the end was dropped over the side of the ship. You could tell how deep the water was by how many knots went under the water before the rope hit bottom. Today we just turn on an instrument and read the depth to a high level of accuracy.

Length and Volume

Length is the measurement of the extent of something along its greatest dimension. The SI basic unit of length, or linear measure, is the **meter** (m). All measurements of length may be made in meters, though the prefixes listed in various tables will often be more convenient. The width of a room may be expressed as about 5 meters (m), whereas a large distance, such as the distance between New York City and Chicago, is better expressed as 1150 kilometers (km). Very small distances can be expressed in units such as the millimeter or the micrometer. The width of a typical human hair is about 20 micrometers $(\mu\mathrm{m})$.

Volume is the amount of space occupied by a sample of matter. The volume of a regular object can be calculated by multiplying its length by its width and height. Since each of those is a linear measurement, we say that units of volume are derived from units of length. The SI unit of volume is the cubic meter (m^3), which is the volume occupied by a cube that measures $1\ \mathrm{m}$ on each side. This very large volume is not convenient for typical use in a chemistry laboratory. A liter (L) is the volume of a cube that measures $10\ \mathrm{cm}\ (1\,\mathrm{dm})$ on each side. A liter is thus equal to both $1000\ \mathrm{cm}^3\ (10\,\mathrm{cm} \times 10\,\mathrm{cm} \times 10\,\mathrm{cm})$, and to $1\ \mathrm{dm}^3$. A smaller unit of volume that is commonly used is the milliliter (mL—note the capital L, which is a standard practice). A milliliter is the volume of a cube that measures $1\ \mathrm{cm}$ on each side. Therefore, a milliliter is equal to a cubic centimeter (cm^3). There are $1000\ \mathrm{mL}$ in $1\ \mathrm{L}$, which is the same as saying that there are $1000\ \mathrm{cm}^3$ in $1\ \mathrm{dm}^3$.



Figure 2.4.2: A typical water bottle is 1 liter in volume. (Public Domain; User:Kenyon/Wikimedia Commons via Wikipedia)

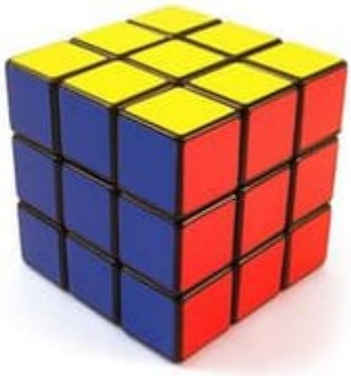


Figure 2.4.3: This Rubik's cube is $5.7\ \text{cm}$ on each side and has a volume of $185.2\ \text{cm}^3$ or $185.2\ \text{mL}$. (CC by 2.0; Bram Van Damme (Flickr: Bramus!) via Flickr)



Figure 2.4.4: A graduated cylinder is often used to measure volume in the laboratory, and comes in a variety of sizes. (Public Domain; User:Darrien/Wikipedia via Wikipedia)

Summary

- Length is the measurement of the extent of something along its greatest dimension.
- Volume is the amount of space occupied by a sample of matter.
- Volume can be determined by knowing the length of each side of an item.

Review

1. Define length.
2. Define volume.
3. An object measures 6.2 cm×13.7 cm×26.9 cm6.2 cm×13.7 cm×26.9 cm. Which value is the length of the object?

4. How big is a mL?

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2.5: SI Mass and Weight Units



Figure 2.5.1 (Public Domain; NASA via Wikipedia)

How is he floating?

One of the many interesting things about travel in outer space is the idea of weightlessness. If something is not fastened down, it will float in mid-air. Early astronauts learned that weightlessness had negative effects on bone structure. If there was no pressure on the legs, those bones would begin to lose mass. Weight provided by gravity is needed to maintain healthy bones. Specially designed equipment is now a part of every space mission, so the astronauts can maintain good body fitness.

Mass and Weight

Mass is a measure of the amount of matter that an object contains. The base SI unit of mass is the kilogram or kg, which was originally defined as the mass of $1\ \mathrm{L}$ of liquid water at $4^{\circ}\mathrm{C}$ (the volume of a liquid changes slightly with temperature). The kilogram is the only SI base unit that contains a prefix. This is because when the first physical standard for mass was created in 1799, the methods to measure the mass of something as small as a gram did not exist. In the laboratory, mass is measured with a balance (figure below), which must be calibrated with a standard mass so that its measurements are accurate.



Figure 2.5.2: An analytical balance makes very sensitive mass measurements in a laboratory, usually in grams. (Public Domain; US Drug Enforcement Administration via Wikipedia)

Other common units of mass are the gram and the milligram. A gram is 1/1000th of a kilogram, meaning that there are $1000\ \mathrm{g}$ in $1\ \mathrm{kg}$. A milligram is 1/1000th of a gram, so there are $1000\ \mathrm{mg}$ in $1\ \mathrm{g}$.

Mass is often confused with the term weight. **Weight** is a measure of force that is equal to the gravitational pull on an object. The weight of an object is dependent on its location. On the moon, the force due to **gravity** is about one sixth that of the gravitational

force on Earth. Therefore, a given object will weigh six times more on Earth than it does on the moon. Since mass is dependent only on the amount of matter present in an object, mass does not change with location. Weight measurements are often made with a spring scale by reading the distance that a certain object pulls down and stretches a spring.

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Summary

- Mass is a measure of the amount of matter that an object contains.
- Weight is a measure of force that is equal to the gravitational pull on an object.
- Mass is independent of location, while weight depends on location.

Review

1. Define mass.
2. Define weight.
3. If I weigh 180 pounds on Earth, what will I weigh on the moon?

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2.6: SI Kinetic Energy Units



Figure 2.6.1 (Public Domain; User:Dlls/Wikimedia Commons via Wikipedia)

Have you ever watched a cat in action?

When a cat is chasing something, its moves very fast. A person may comment, "that cat has a lot of energy"—which is more correct than they might realize! One form of energy is seen when an object is moving, and this type of energy is the basis for many chemical processes.

SI Kinetic Energy Units

An object's **kinetic energy** is the energy due to motion. Kinetic energy can be defined mathematically as:

$$KE = \frac{1}{2}mv^2$$

where KE is kinetic energy, m is mass, and v is velocity.

Energy is defined as the capacity to do work or to produce heat. As discussed previously, kinetic energy is one type of energy and is associated with motion. Another frequently encountered energy is **potential energy**, a type of energy that is stored in matter and released during a chemical reaction. The **joule** (J) is the SI unit of energy and is named after English physicist James Prescott Joule (1818-1889). If we go back to the equation for kinetic energy written above, we can put units in: kg for mass and $\mathrm{m}^2/\mathrm{s}^2$ for velocity squared. Then, in terms of SI base units, a joule is equal to a kilogram times meter squared, divided by a second squared $(\mathrm{kg} \cdot \mathrm{m}^2/\mathrm{s}^2)$. Another common unit of energy that is often used is the **calorie** (cal), which is equivalent to $4.184\ \mathrm{J}$.



Figure 2.6.2: James Prescott Joule. (Public Domain; John Collier via Wikipedia)

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Summary

- Energy is the capacity to do work or to produce heat.
- Kinetic energy is energy due to motion.
- Potential energy is energy stored in matter.
- The joule (J) is the SI unit of energy and equals $\mathrm{kg} \cdot \mathrm{m}^2/\mathrm{s}^2$.

Review

1. What is kinetic energy?
2. What is the mathematical equation for kinetic energy?
3. What is potential energy?
4. What is the SI unit for energy?

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2.7: Temperature and Temperature Scales



Figure 2.7.1 (CC0; Tantien3 via Pixabay / Public Domain)

An ice cube in your hand feels cold, but why?

The particles of matter in a hot object are moving much faster than the particles of matter in a cold object. An object's **kinetic energy** is the energy due to motion. The particles of matter that make up the hot stove have a greater amount of kinetic energy than those in the ice cube.

Temperature and Temperature Scales

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 an 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 2.7.2: The glowing charcoal on the left represents high kinetic energy, while the snow and ice on the right are of much lower kinetic energy. (Charcoal: CC-BY 2.0; Snow: Public Domain; Charcoal: Serge Melki; Snow: User:Haymanj/Wikimedia Commons via Charcoal: http://www.flickr.com/photos/sergemelki/3106924114; Snow: http://commons.wikimedia.org/wiki/File:Snow_Gum1.JPG)

Temperature Scales

The first thermometers were glass and contained alcohol, which expanded and contracted as the temperature changed. The

German scientist Daniel Gabriel Fahrenheit used mercury in the tube, an idea put forth by Ismael Boulliau. The Fahrenheit scale was first developed in 1724 and tinkered with for some time after that. The main problems with this scale are the arbitrary definitions of temperature. The freezing point of water was defined as $32^{\circ}\mathrm{F}$ and the boiling point as $212^{\circ}\mathrm{F}$. The Fahrenheit scale is typically not used for scientific purposes.



Figure 2.7.3: Daniel Gabriel Fahrenheit. (Public Domain; Anonymous via Wikipedia)

The Celsius scale of the metric system is named after Swedish astronomer Anders Celsius (1701-1744). The Celsius scale sets the freezing point and boiling point of water at $0^{\circ}\mathrm{C}$ and $100^{\circ}\mathrm{C}$, respectively. The distance between those two points is divided into 100 equal intervals, each of which is one degree. Another term sometimes used for the Celsius scale is "centigrade" because there are 100 degrees between the freezing and boiling points of water on this scale. However, the preferred term is "Celsius".



Figure 2.7.4: Anders Celsius. (Public Domain; Olof Arenius via Wikipedia)

The Kelvin temperature scale is named after Scottish physicist and mathematician Lord Kelvin (1824-1907). It is based on molecular motion, with the temperature of $0\ \mathrm{K}$, also known as absolute zero, being the point where all molecular motion ceases. The freezing point of water on the Kelvin scale is $273.15\ \mathrm{K}$, while the boiling point is $373.15\ \mathrm{K}$. Notice that there is no "degree" used in the temperature designation. Unlike the Fahrenheit and Celsius scales, where temperatures are referred to as "degrees F " or "degrees C", we simply designated temperatures in the Kelvin scale as kelvins.



Figure 2.7.5: Lord Kelvin. (Public Domain; Anonymous via Wikipedia)

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As can be seen by the 100 kelvin difference between the two, a change of one degree on the Celsius scale is equivalent to the change of one kelvin on the Kelvin scale. Converting from one scale to another is easy, as you simply add or subtract 273.

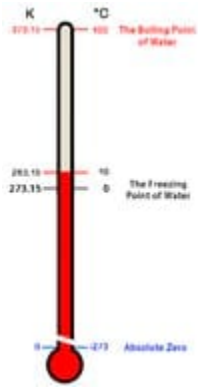


Figure 2.7.6: Comparison between Kelvin and Celsius temperature Scales. (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

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Simulation

No matter the temperature scale you use, the temperature at which a liquid boils depends on a few different variables. Try boiling water in a few places around the world to see what we mean.

Summary

- Temperature is a measure of the average kinetic energy of the particles in matter.
- The Fahrenheit scale defines the freezing point of water as 32°F and the boiling point as 212°F.
- The Celsius scale sets the freezing point and boiling point of water at 0°C and 100°C, respectively.
- The Kelvin scale is based on molecular motion, with the temperature of $0\ \text{K}$—also known as absolute zero—being the point where all molecular motion ceases.

Review

1. What is absolute zero on the Celsius temperature scale?
2. What are the freezing and boiling points of water in the Celsius scale?
3. Convert the following Kelvin temperatures to degrees Celsius.
   a. 188 K
   b. 631 K
4. Temperature in degrees Fahrenheit can be converted to Celsius by first subtracting 32, then dividing by 1.8. What is the Celsius temperature outside on a warm day (88°F)?
5. Why is the Celsius scale sometimes called “centigrade”?

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2.8: Scientific Dimensional Analysis

Conversion Factors

Many quantities can be expressed in several different ways. The English system measurement of 4 cups is also equal to 2 pints, 1 quart, and $\frac{1}{4}$ of a gallon.

$$4 \text{ cups} = 2 \text{ pints} = 1 \text{ quart} = 0.25 \text{ gallon}$$

Notice that the numerical component of each quantity is different, while the actual amount of material that it represents is the same. This is because the units are different. We can establish the same set of equalities for the metric system:

$$1 \text{ meter} = 10 \text{ decimeters} = 100 \text{ centimeters} = 1000 \text{ millimeters}$$

The metric system's use of powers of 10 for all conversions makes this quite simple.

Whenever two quantities are equal, a ratio can be written that is numerically equal to 1. Using the metric examples above:

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

The $\frac{1 \text{ m}}{100 \text{ cm}}$ is called a **conversion factor**. A conversion factor is a ratio of equivalent measurements. Because both 1 m and 100 cm represent the exact same length, the value of the conversion factor is 1. The conversion factor is read as "1 meter per 100 centimeters". Other conversion factors from the cup measurement example can be:

$$\frac{4 \text{ cups}}{2 \text{ pints}} = \frac{2 \text{ pints}}{1 \text{ quart}} = \frac{1 \text{ quart}}{0.25 \text{ gallon}} = 1$$

Since the numerator and denominator represent equal quantities in each case, all are valid conversion factors.

Scientific Dimensional Analysis

Conversion factors are used in solving problems in which a certain measurement must be expressed with different units. When a given measurement is multiplied by an appropriate conversion factor, the numerical value changes, but the actual size of the quantity measured remains the same. **Dimensional analysis** is a technique that uses the units (dimensions) of the measurement in order to correctly solve problems. Dimensional analysis is best illustrated with an example.

Example 2.8.1

How many seconds are in a day?

Solution

Step 1: List the known quantities and plan the problem.

Known

- 1 day = 24 hours
- 1 hour = 60 minutes
- 1 minute = 60 seconds

Unknown

The known quantities above represent the conversion factors that we will use. The first conversion factor will have day in the denominator so that the "day" unit will cancel. The second conversion factor will then have hours in the denominator, while the third conversion factor will have minutes in the denominator. As a result, the unit of the last numerator will be seconds and that will be the units for the answer.

Step 2: Calculate.

$$1\text{ d} \times \frac{24\text{ hr}}{1\text{ d}} \times \frac{60\text{ min}}{1\text{ hr}} \times \frac{60\text{ s}}{1\text{ min}} = 86,400\text{ s}$$

Applying the first conversion factor, the "d" unit cancels and $1 \times 24 = 24$. Applying the second conversion factor, the "hr" unit cancels and $24 \times 60 = 1440$. Applying the third conversion factor, the "min" unit cancels and $1440 \times 60 = 86,400$. The unit that remains is "s" for seconds.

Step 3: Think about your result.

Seconds is a much smaller unit of time than days, so it makes sense that there are a very large number of seconds in one day.

Summary

- A conversion factor is a ratio of equivalent measurements.
- Dimensional analysis is a technique that uses the units (dimensions) of the measurement to solve problems.

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Review

1. What is a conversion factor?
2. What is dimensional analysis?
3. How many meters are in 3.7 km?
4. How many kg in 12980 g?

Explore More

Use the link below to answer the following questions**:** http://www.felderbooks.com/papers/units.html

1. What do we always need to express measurements correctly?
2. What does dimensional analysis tell you?
3. How do you know that you have set the problem up incorrectly?
4. How do you know that you have set the problem up correctly?

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2.9: Metric Unit Conversions

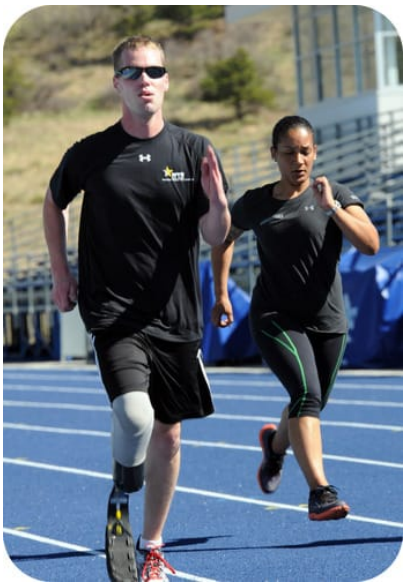


Figure 2.9.1 (Credit: Courtesy of the US Army; Source: http://commons.wikimedia.org/wiki/File:Flickr_-_The_U.S._Army_-_Track_practice.jpg; License: Public Domain)

How can a number of track laps be converted to a distance in meters?

You are training for a 10-kilometer run by doing laps on a 400-meter track. You ask yourself how many times you need to run around the track in order to cover ten kilometers (...more than you realize). By using dimensional analysis, you can easily determine the number of laps needed to cover the $10\ \text{km}$ distance.

Metric Unit Conversions

The metric system's many prefixes allow quantities to be expressed in many different units. Dimensional analysis is useful to convert from one metric system unit to another.

Example 2.9.1

A particular experiment for science class requires $120\ \text{mL}$ of a solution. The science teacher knows that he will need to make enough solution for 40 experiments to be performed throughout the day. How many liters of solution should he prepare?

Solution

Step 1: List the known quantities and plan the problem.

Known

- 1 experiment requires $120\ \text{mL}$
- $10^{-3}\ \text{L} = 1\ \text{mL}$

Unknown

Since each experiment requires $120\ \text{mL}$ of solution and the teacher needs to prepare enough for 40 experiments, multiply 120 by 40 to get $4800\ \text{mL}$ of solution needed. Now you must convert mL to L by using a conversion factor.

Step 2: Calculate.

$$4800\ \text{mL} \times \frac{10^{-3}\ \text{L}}{1\ \text{mL}} = 4.8\ \text{L} \tag{2.9.1}$$

Note that the conversion factor is arranged so that the mL unit is in the denominator and thus cancels out, leaving L as the remaining unit in the answer.

Step 3:Think about your result.

A liter is much larger than a milliliter, so it makes sense that the number of liters required is less than the number of milliliters.

Two-Step Metric Unit Conversions

Some metric conversion problems are most easily solved by breaking them down into more than one step. When both the given unit and the desired unit have prefixes, one can first convert to the simple (un-prefixed) unit, followed by a conversion to the desired unit. The following example will illustrate this method.

Example 2.9.2

Convert $4.3\ \text{cm}$ to μm.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $10^{-2}\ \text{m} = 1\ \text{cm}$
- $10^{-6}\ \text{m} = 1\ \mu\text{m}$

Unknown

You may need to consult a table for the multiplication factor represented by each metric prefix. First convert cm to m, then convert m to μm.

Step 2: Calculate.

$$4.3\ \text{cm} \times \frac{10^{-2}\ \text{m}}{1\ \text{cm}} \times \frac{1\ \mu\text{m}}{10^{-3}\ \text{m}} = 43,000\ \mu\text{m} \tag{2.9.2}$$

Each conversion factor is written so that the unit of the denominator cancels with the unit of the numerator of the previous factor.

Step 3: Think about your result.

A micrometer is a smaller unit of length than a centimeter, so the answer in micrometers is larger than the number of centimeters given.

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Summary

- Dimensional analysis can be used to carry out metric unit conversions.

Review

1. Perform the following conversions.
   a. 0.074 km to m
   b. 24,600 µg to g
   c. 4.9 × 107 µg to kg
   d. 84 dm to mm

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2.10: Derived Units



Figure 2.10.1 (Public Domain; User:Joegrimes/Wikipedia via Commons Wikimedia, Farm landly sander [commons.wikimedia.org])

How has farming evolved?

As farming becomes more expensive and less profitable (at least for small farms), farmers might decide to sell their land to builders who want to erect commercial or residential properties. In order to sell, an accurate property title is needed. The dimensions of the farm must be determined, and the acreage calculated from those dimensions.

Dimensional Analysis and Derived Units

Some units are combinations of SI base units. A **derived unit** is a unit that results from a mathematical combination of SI base units. We have already discussed volume and energy as two examples of derived units. Some others are listed in the table below:

Derived SI Units

Quantity	Symbol	Unit	Unit Abbreviation	Derivation
Area	A	square meter	m^2	$\text{length} \times \text{width}$
Volume	V	cubic meter	m^3	$\text{length} \times \text{width} \times \text{height}$
Density	D	kilograms/cubic meter	$\mathrm{kg/m}^3$	\(\frac{\text{mass}}{\text{volume}}\0
Concentration	c	moles/liter	$\mathrm{mol/L}$	$\frac{\text{amount}}{\text{volume}}$
Speed (velocity)	v	meters/second	$\mathrm{m/s}$	$\frac{\text{length}}{\text{time}}$
Acceleration	a	meters/second/second	$\mathrm{m/s}^2$	$\frac{\text{speed}}{\text{time}}$
Force	F	newton	N	$\text{mass} \times \text{acceleration}$
Energy	E	joule	J	$\text{force} \times \text{length}$

Using dimensional analysis with derived units requires special care. When units are squared or cubed as with area or volume, the conversion factors themselves must also be squared. Shown below is the conversion factor for cubic centimeters and cubic meters.

$$\left(\frac{1\ \mathrm{m}}{100\ \mathrm{cm}}\right)^3 = \frac{1\ \mathrm{m}^3}{10^6\ \mathrm{cm}^3} = 1$$

Because a cube has 3 sides, each side is subject to the conversion of $1\ \mathrm{m}$ to $100\ \mathrm{cm}$. Since 100 cubed is equal to 1 million (10^6), there are $10^6\ \mathrm{cm}^3$ in $1\ \mathrm{m}^3$. Two convenient volume units are the liter (which is equal to a cubic decimeter) and the milliliter, which

is equal to a cubic centimeter. The conversion factor would be:

$$\left(\frac{1\text{ dm}}{10\text{ cm}}\right)^3 = \frac{1\text{ dm}^3}{1000\text{ cm}^3} = 1$$

There are thus $1000\ \text{cm}^3$ in $1\ \text{dm}^3$, which is the same thing as saying there are $1000\ \text{mL}$ in $1\ \text{L}$.

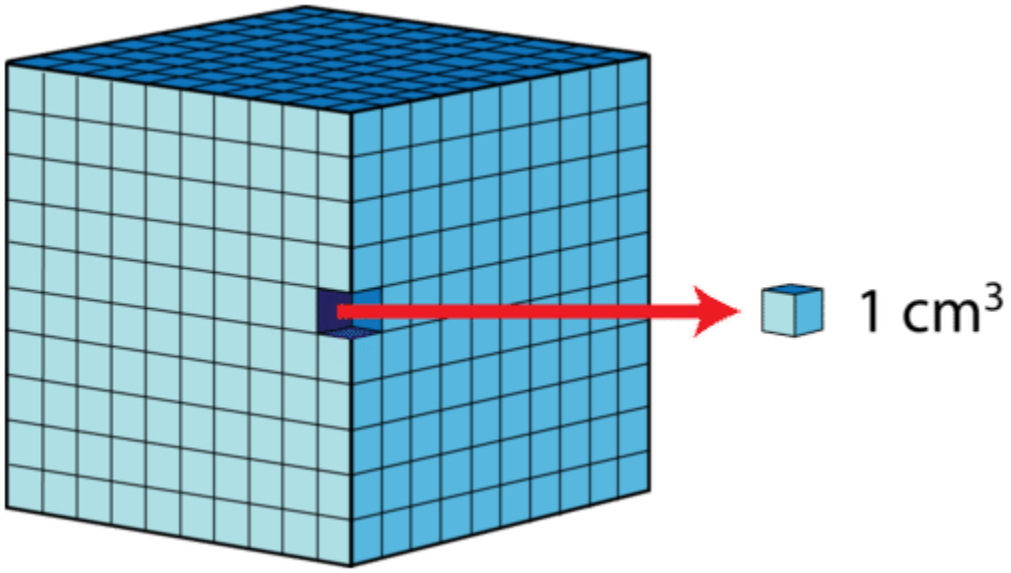


Figure 2.10.2: There are $1000\ \text{cm}^3$ in $1\ \text{dm}^3$. Since a cm^3 is equal to a mL and a dm^3 is equal to a L, we can say that there are $1000\ \text{mL}$ in $1\ \text{L}$. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Example 2.10.1

Convert $3.6 \times 10^8\ \text{mm}^3$ to mL.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $1\ \text{m} = 1000\ \text{mm}$
- $1\ \text{mL} = 1\ \text{cm}^3$
- $1\ \text{m} = 100\ \text{cm}$

Unknown

This problem requires multiple steps and the technique for converting with derived units. Simply proceed one step at a time: mm^3 to m^3 to $\text{cm}^3 = \text{mL}$.

Step 2: Calculate.

$$3.6\ \text{mm}^3 \times \left(\frac{1\text{ m}}{1000\text{ mm}}\right)^3 \times \left(\frac{100\text{ cm}}{1\text{ m}}\right)^3 \times \frac{1\text{ mL}}{1\text{ cm}^3} = 0.0036\text{ mL}$$

Numerically, the steps are to divide 3.6 by 10^9, followed by multiplying by 10^6. You may find that you can shorten the problem by a step by first determining the conversion factor from mm to cm, and using that instead of first converting to m. There are $10\ \text{mm}$ in $1\ \text{cm}$.

$$3.6\ \text{mm}^3 \times \left(\frac{1\text{ cm}}{10\text{ mm}}\right)^3 \times \frac{1\text{ mL}}{1\text{ cm}^3} = 0.0036\text{ mL}$$

In this case, 3.6/1000 gives the same result of 0.0036.

Step 3: Think about your result.

Cubic millimeters are much smaller than cubic centimeters, so the final answer is much less than the original number of mm^3.

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Summary

- A derived unit is a unit that results from a mathematical combination of SI base units.
- Calculations involving derived units follow the same principles as other unit conversion calculations.

Review

1. What is a derived unit?
2. Convert 0.00722 km^3 to m^3
3. Convert 129 cm^3 to L
4. Convert 4.9×10^5 μm^3 to mm^3

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2.11: Density



Figure 2.11.1 (CC BY 2.0; Tony Hisgett (Flickr:ahisgett) via Flickr)

How do logs stay afloat in water?

After trees are cut, logging companies often move the raw material down a river to a sawmill where it can be shaped into building materials or other products. The logs float on the water because they are less dense than the water they are in. Knowledge of density is important in the characterization and separation of materials. Information about density allows us to make predictions about the behavior of matter.

Density

A golf ball and a table tennis ball are about the same size. However, the golf ball is much heavier than the table tennis ball. Now imagine a similar size ball made out of lead. That would be very heavy indeed! What are we comparing? By comparing the mass of an object relative to its size, we are studying a property called **density**. Density is the ratio of the mass of an object to its volume.

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

Density is an intensive property, meaning that it does not depend on the amount of material present in the sample. Water has a density of $1.0\ \text{g/mL}$. That density is the same whether you have a small glass of water or a swimming pool full of water. Density is a property that is constant for the particular identity of the matter being studied.

The SI unit of density is kilograms per cubic meter $\left(\text{kg/m}^3\right)$, since the kg and the m are the SI units for mass and length respectively. In everyday usage in a laboratory, this unit is awkwardly large. Most solids and liquids have densities that are conveniently expressed in grams per cubic centimeter $\left(\text{g/cm}^3\right)$. Since a cubic centimeter is equal to a milliliter, density units can also be expressed as g/mL. Gases are much less dense than solids and liquids, so their densities are often reported in g/L. Densities of some common substances at 20°C are listed in the table below.

Table 2.11.1: Densities of Some Common Substances

Liquids and Solids	Density at 20°C (g/mL)	Gases	Density at 20°C (g/L)
Ethanol	0.79	Hydrogen	0.084

Liquids and Solids	Density at $20^{\circ}\mathrm{C}$ $(\mathrm{g/mL})$	Gases	Density at $20^{\circ}\mathrm{C}$ $(\mathrm{g/L})$
Ice $(0^{\circ}\mathrm{C})$	0.917	Helium	0.166
Corn oil	0.922	Air	1.20
Water	0.998	Oxygen	1.33
Water $(4^{\circ}\mathrm{C})$	1.000	Carbon dioxide	1.83
Corn syrup	1.36	Radon	9.23
Aluminum	2.70		
Copper	8.92		
Lead	11.35		
Mercury	13.6		
Gold	19.3		

Since most materials expand as temperature increases, the density of a substance is temperature dependent, and usually decreases as temperature increases.

Ice floats in water, and it can be seen from the table that ice is less dense. Corn syrup, being more dense, would sink if placed in water.

Example 2.11.1

An $18.2\ \mathrm{g}$ sample of zinc metal has a volume of $2.55\ \mathrm{cm}^3$. Calculate the density of zinc.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass $= 18.2\ \mathrm{g}$
- Volume $= 2.55\ \mathrm{cm}^3$

Unknown

Use the equation for density, $D = \frac{m}{V}$, to solve the problem.

Step 2: Calculate.

$$D = \frac{m}{V} = \frac{18.2\ \mathrm{g}}{2.55\ \mathrm{cm}^3} = 7.14\ \mathrm{g/cm}^3$$

Step 3: Think about your result.

If $1\ \mathrm{cm}^3$ of zinc has a mass of about 7 grams, then 2 and a half cm^3 will have a mass about 2 and a half times as great. Metals are expected to have a density greater than that of water, and zinc's density falls within the range of the other metals listed above.

Since density values are known for many substances, density can be used to determine an unknown mass or an unknown volume. Dimensional analysis will be used to ensure that units cancel appropriately.

Example 2.11.2

1. What is the mass of $2.49\ \mathrm{cm}^3$ of aluminum?
2. What is the volume of $50.0\ \mathrm{g}$ of aluminum?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Density $= 2.70\ \text{g/cm}^3$
- 1. Volume $= 2.49\ \text{cm}^3$
- 2. Mass $= 50.0\ \text{g}$

Unknown

- 1. Mass $=?\ \text{g}$
- 2. Volume $=?\ \text{cm}^3$

Use the equation for density, $D = \frac{m}{V}$, and dimensional analysis to solve each problem.

Step 2: Calculate.

$$1.\ 2.49\ \text{cm}^3 \times \frac{2.70\ \text{g}}{1\ \text{cm}^3} = 6.72\ \text{g}$$

$$2.\ 50.0\ \text{g} \times \frac{1\ \text{cm}^3}{2.70\ \text{g}} = 18.5\ \text{cm}^3$$

In problem 1, the mass is equal to the density multiplied by the volume. In problem 2, the volume is equal to the mass divided by the density.

Step 3: Think about your results.

Because a mass of $1\ \text{cm}^3$ of aluminum is $2.70\ \text{g}$, the mass of about $2.5\ \text{cm}^3$ should be about 2.5 times larger. The $50\ \text{g}$ of aluminum is substantially more than its density, so that amount should occupy a relatively large volume.

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Simulation

Go fishing and explore the variables that affect density in this simulation.

Summary

- Density is the ratio of the mass of an object to its volume.
- Gases are less dense than either solids or liquids.
- Both liquid and solid materials can have a variety of densities.
- For liquids and gases, the temperature will affect the density to some extent.

Review

1. Define “density.”
2. Are gases more or less dense that liquids or solids at room temperature?
3. How does temperature affect the density of a material?
4. A certain liquid sample has a volume of 14.7 mL and a mass of 22.8 grams. Calculate the density.
5. A material with a density of 2.7 grams/mL occupies 35.6 mL. How many grams of the material are there?
6. A certain material has a density of 19.3 g/mL. What is the material?

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2.12: Accuracy and Precision



Figure 2.12.1 (Public Domain; Daniel Arizpe via Commons Wikimedia, High School Basketball Game [commons.wikimedia.org])

How do professional basketball players improve their shooting accuracy?

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

Accuracy and Precision

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

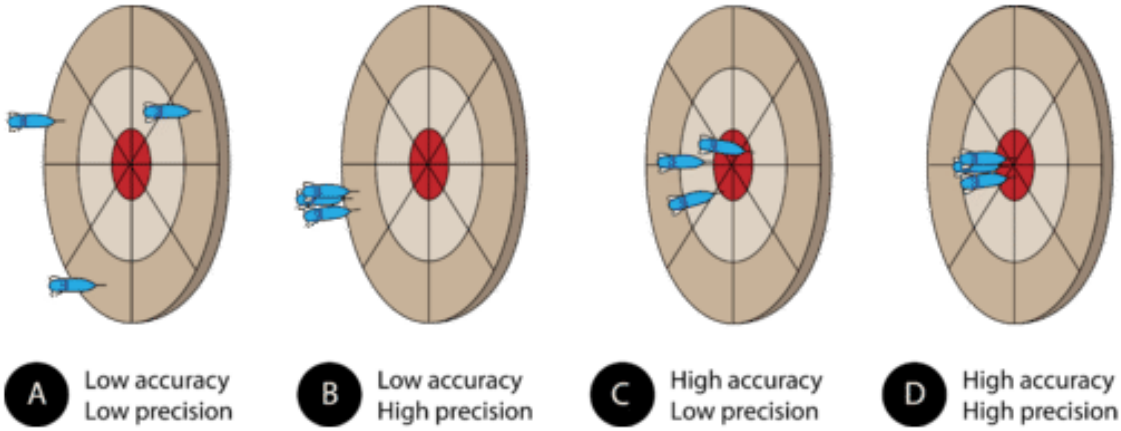


Figure 2.12.2: The distribution of darts on a dartboard shows the difference between accuracy and precision. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

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

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



Figure 2.12.3: Students in a chemistry lab are making careful measurements with a series of volumetric flasks. Accuracy and precision are critical in every experiment. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

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Simulation

How is your accuracy and precision? Try out your skills with this simulation.

Summary

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

Review

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

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2.13: Percent Error

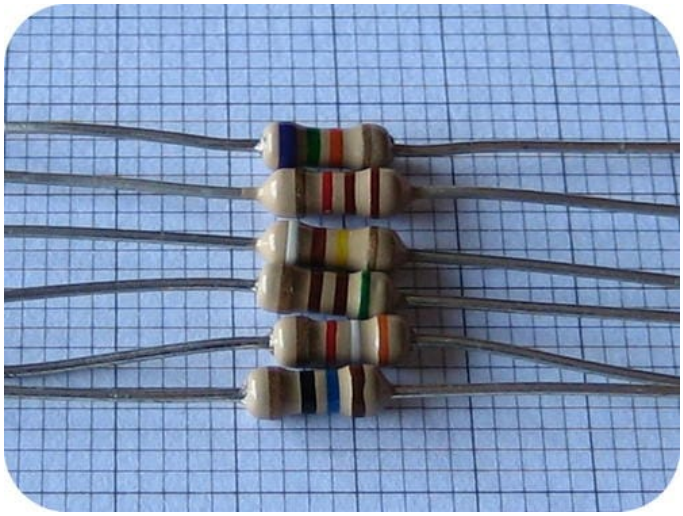


Figure 2.13.1 (Public Domain; Vald Kliper via Wikipedia)

How does an electrical circuit work?

A complicated piece of electronic equipment may contain several resistors that have the role of controlling the voltage and current in the electrical circuit. If there is too much current, the apparatus malfunctions. Conversely, too little current means the system simply will not perform. The resistor values are always given with an error range. A resistor may have a stated value of 200 ohms, but a 10% error range, meaning the resistance could be anywhere between 195 and 205 ohms. By knowing these values, an electronics person can design and service the equipment to make sure that it functions properly.

Percent Error

An individual measurement may be accurate or inaccurate, depending on how close it is to the true value. Suppose that you are doing an experiment to determine the density of a sample of aluminum metal. The **accepted value** of a measurement is the true or correct value based on general agreement with a reliable reference. For aluminum, the accepted density is $2.70\ \text{g/cm}^3$. The **experimental value** of a measurement is the value that is measured during the experiment. Suppose that in your experiment, you determine an experimental value for the aluminum density to be $2.42\ \text{g/cm}^3$. The **error** of an experiment is the difference between the experimental and accepted values.

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

If the experimental value is less than the accepted value, the error is negative. If the experimental value is larger than the accepted value, the error is positive. Often, error is reported as the absolute value of the difference in order to avoid the confusion of a negative error. The **percent error** is the absolute value of the error, divided by the accepted value, and multiplied by 100%.

$$\%\ \text{Error} = \frac{|\text{experimental value} - \text{accepted value}|}{\text{accepted value}} \times 100\%$$

To calculate the percent error for the aluminum density measurement, we can substitute the given values of $2.45\ \text{g/cm}^3$ for the experimental value and $2.70\ \text{g/cm}^3$ for the accepted value.

$$\%\ \text{Error} = \frac{\left|2.45\ \text{g/cm}^3 - 2.70\ \text{g/cm}^3\right|}{2.70\ \text{g/cm}^3} \times 100\% = 9.26\%$$

If the experimental value is equal to the accepted value, the percent error is equal to 0. As the accuracy of a measurement decreases, the percent error of that measurement rises.

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Summary

- Definitions of accepted value and experimental values are given.
- Calculations of error and percent error are demonstrated.

Review

1. Define accepted value.
2. Define experimental value
3. What happens as the accuracy of the measurement decreases?

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2.14: Measurement Uncertainty



Figure 2.14.1 (CC BY-SA 4.0; Tony Webster via Commons Wikimedia, Police Line Crime Scene [commons.wikimedia.org])

How do police officers identify criminals?

After a bank robbery has been committed, police will ask witnesses to describe the thieves. They will usually get an answer such as "medium height". Others may say something along the lines of "between 5 foot 8 inches and 5 foot 10 inches". In both cases, there is a significant amount of uncertainty about the height of the criminals.

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 $0.1\ \text{g}$, other highly sensitive balances are capable of measuring to the nearest $0.001\ \text{g}$, 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. The figure below shows two rulers making the same measurement of an object (indicated by the blue arrow).

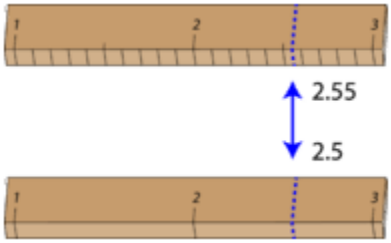


Figure 2.14.2: Uncertainty in measurement. (CC BY-NC-SA 3.0; Christopher Auyeung via CK-12 Foundation)

With either ruler, it is clear that the length of the object is between 2 and $3\ \text{cm}$. The bottom ruler contains no millimeter markings. With that ruler, the tenths digit can be estimated and the length may be reported as $2.5\ \text{cm}$. However, another person may judge that the measurement is $2.4\ \text{cm}$ or perhaps $2.6\ \text{cm}$. 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). With this ruler, the same object may be measured as $2.55\ \text{cm}$. 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 $2.54\ \text{cm}$ or $2.56\ \text{cm}$. 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.

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Summary

- Uncertainty exists in all measurements.
- The degree of uncertainty is affected in part by the quality of the measuring tool.

Review

1. What is uncertainty in measurements?
2. Why is the top ruler more reliable in measuring length than the bottom ruler?
3. How could the top ruler be made more accurate?

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2.15: Significant Figures



Figure 2.15.1 (Public Domain; User:WhisperToMe/Wikimedia Commons via Wikipedia)

How fast do you drive?

As you enter the town of Jacinto City, Texas, a sign tells you that the speed limit is 30 miles per hour. But what if you happen to be driving 31 miles an hour? Are you in trouble? Probably not, because there is a certain amount of leeway built into enforcing the regulation. Most speedometers do not measure the vehicle speed very accurately and could easily be off by a mile or so (on the other hand, radar measurements are much more accurate). So, a couple of miles per hour difference won't matter that much. However, if you stretch the limits any further, you may end up receiving a traffic ticket.

Significant Figures

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 $2.553\ \text{cm}$, as there is no possible way that the thousandths digit could be estimated. The 3 is not significant and would not be reported.

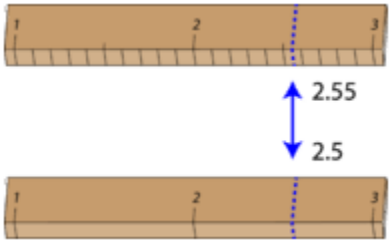


Figure 2.15.2: Measurement with two different rulers. (CC BY-NC-SA 3.0; Christopher Auyeung via CK-12 Foundation)

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.

Figure 2.15.1: 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 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 left-end zeros. Left-end 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 right-end zeros. Right-end zeros in a number that lacks a decimal point are not significant.	• 140 has two significant figures. • 75,210 has four significant figures.
5. Right-end 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.

It needs to be emphasized that to say a certain digit is not significant, does not mean that it is not important or can be left out. Though the zero in a measurement of 140 may not be significant, the value cannot simply be reported as 14. An insignificant zero functions as a placeholder for the decimal point. When numbers are written in scientific notation, this becomes more apparent. The measurement 140 can be written as 1.4×10^2 with two significant figures in the coefficient. For a number with left-end zeros, such as 0.000416, it can be written as 4.16×10^{-4} with 3 significant figures. In some cases, scientific notation is the only way to correctly indicate the correct number of significant figures. In order to report a value of 15,000,000 with four significant figures, it would need to be written as 1.500×10^7. The right-end zeros after the 5 are significant. The original number of 15,000,000 only has two significant figures.

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Interactive Element

Practice using significant figures in this simulation.

Summary

- 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.

Review

1. What does a significant figure tell us?
2. What is a left-end zero?
3. What is a right-end zero?
4. What does an insignificant zero do?

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2.16: Rounding



Figure 2.16.1 (Public Domain; Nancy Heise via Wikipedia)

Have you ever been fishing

People who fish often are a little unreliable when it comes to describing what fish they caught and how much it weighed. It's easier to say that the fish weighed ten pounds than it is to accurately describe the weight of 8 pounds 11 ounces. Ten pounds is "close enough" when a fisherman is talking about their catch.

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 $207.518\ \text{m}$. 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 2.16.1.

Table 2.16.1: 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.

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Summary

- Rounding involves the adjustment of a value to account for the proper number of significant digits.
- Rules exist for rounding numbers.

Review

1. Why do we round numbers?
2. What do we need to know before we round a number?
3. What is “rounding up”?
4. What is “rounding down”?

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2.17: Uncertainty in Multiplication and Division



Figure 2.17.1 (Public Domain; Adrian Pingstone (User:Arpingstone/Wikimedia Commons) via Wikipedia)

Who should report the numbers - you or your calculator?

Calculators do just what is asked of them—no more and no less. However, they sometimes can get a little out of hand. If you multiply 2.48 times 6.3, you get an answer of 15.687, a value that ignores the number of significant figures in either number. Division with a calculator is even worse. When you divide 12.2 by 1.7, the answer you obtain is 7.176470588. Neither piece of data is accurate to nine decimal places, but the calculator does not know that. The human being operating the instrument has to make the decision about how many places to report.

Uncertainty in Multiplication and Division

The density of a certain object is calculated by dividing the mass by the volume. Suppose that a mass of $37.46\ \text{g}$ is divided by a volume of $12.7\ \text{cm}^3$. The result on a calculator would be:

$$D = \frac{m}{V} = \frac{37.46\ \text{g}}{12.7\ \text{cm}^3} = 2.949606299\ \text{g/cm}^3$$

The value of the mass measurement has four significant figures, while the value of the volume measurement has only three significant figures. For multiplication and division problems, the answer should be rounded to the same number of significant figures as the measurement with the least number of significant figures. Applying this rule results in a density of $2.95\ \text{g/cm}^3$, with three significant figures—the same as the volume measurement.

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Example 2.17.1

Perform the following calculations, rounding the answers to the appropriate number of significant figures.

A. $0.048\ \text{m} \times 32.97\ \text{m}$

B. $14,570\ \text{kg} \div 5.81\ \text{L}$

Solution

Step 1: Plan the problem.

Analyze each of the measured values to determine how many significant figures should be in the result. Perform the calculation and round appropriately. Apply the correct units to the answer. When multiplying or dividing, the units are also multiplied or divided.

Step 2: Calculate.

A. $0.048\ \text{m} \times 32.97\ \text{m} = 1.6\ \text{m}^2$ Round to two significant figures because 0.048 has two.

B. $14,570\ \text{kg} \div 5.81\ \text{L} = 2510\ \text{kg/L}$ Round to three significant figures because 5.81 has three.

Exercise 2.17.1

How many significant figures should the answer contain?

$10.61 \times 12.133 \times 3.25$

Answer available at the end of this section.

Summary

- For multiplication and division problems, the answer should be rounded to the same number of significant figures as the measurement with the least number of significant figures.

Review

1. Perform the calculation and round your answer with the correct number of significant figures.
   a. $78.2 \div 32\ \text{cm}^3$
   b. 3.0 m/s × 9.21 s
2. What happens to units in multiplication and division problems?

Answers

Exercise 2.17.1

=418; 3 significant figures

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2.18: Uncertainty in Addition and Subtraction



Figure 2.18.1 (Public Domain; User:Septagram/Wikipedia via Wikipedia)

How old do you think this calculator is?

Calculators are great devices. Their invention has allowed for quick computation at work, school, and other places where manipulation of numbers needs to be done rapidly and accurately. But they are only as good as the numbers put into them. The calculator cannot determine how accurate each set of numbers is, and the answer given on the screen must be assessed by the user for reliability.

Uncertainty in Addition and Subtraction

Consider two separate mass measurements: $16.7\ \text{g}$ and $5.24\ \text{g}$. The first mass measurement, $(16.7\,\text{g})$, is known only to the tenths place, or to one digit after the decimal point. There is no information about its hundredth place and so that digit cannot be assumed to be zero. The second measurement, $(5.24\,\text{g})$, is known to the hundredths place, or to two digits after the decimal point.

When these masses are added together, the result on a calculator is $16.7 + 5.24 = 21.94\ \text{g}$. Reporting the answer as $21.94\ \text{g}$ suggests that the sum is known all the way to the hundredths place. However, that cannot be true because the hundredths place of the first mass was completely unknown. The calculated answer needs to be rounded in such a way as to reflect the certainty of each of the measured values that contribute to it. For addition and subtraction problems, the answer should be rounded to the same number of decimal places as the measurement with the least number of decimal places. The sum of the above masses would be properly rounded to a result of $21.9\ \text{g}$.

Example 2.18.1

Determine the combined molecular mass of a glucose molecule and a maltose molecule.

Glucose molecule = $180.156\frac{g}{mol}$

Maltose molecule = $342.3\frac{g}{mol}$

Solution

$180.156 + 342.4 = 522.456$

When adding and subtracting we know to look at the least number of decimals in our starting values; in this case 342.3 has only 1 digit after the decimal, so we need to round our answer to the same place.

\(522.456\to 522.5\frag{g}{mol}

When working with whole numbers, pay attention to the last significant digit that is to the left of the decimal point, and round your answer to that same point. For example, consider the subtraction: $78,500\ \text{m} - 362\ \text{m}$. The calculated result is $78,138\ \text{m}$. However, the first measurement is known only to the hundreds place, as the 5 is the last significant digit. Rounding the result to that same point means that the correct result is $78,100\ \text{m}$.

Example 2.18.2

What is $4200 + 540$ = ?

Solution

$4200 + 540 = 4740$

To determine where to round our answer, we look at our starting numbers to see which has the fewest number of decimal places. They both have 0 so we round to the nearest whole number, 4740.

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Summary

- For addition and subtraction problems, the answer should be rounded to the same number of decimal places as the measurement with the least number of decimal places.

Review

1. What is the basic principle to use in working with addition and subtraction?
2. What do you pay attention to when working with whole numbers?

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2.19: Measurements and Problem-Solving (Exercises)

These are homework exercises to accompany Chapter 1 of the University of Kentucky's LibreText for CHE 103 - Chemistry for Allied Health. Solutions are available below the questions.

Questions

1.1: Measurements Matter

(click here for solutions)

Q1.1.1

Express the following values in scientific notation.

a. 150,000,000
b. 0.000043
c. 332000
d. 0.0293
e. 932
f. 0.1873
g. 78,000
h. 0.0001
i. 4500
j. 0.00290
k. 6281
l. 0.00700

Q1.1.2

Express the following values in decimal notation.

a. 3.8×10^{-3}
b. 9.21×10^{5}
c. 7.91×10^{-2}
d. 2.5×10^{6}
e. 3.42×10^{-8}
f. 5.4×10^{5}
g. 3×10^{-3}
h. 7.34×10^{2}
i. 9.8×10^{-4}
j. 6×10^{7}
k. 4.20×10^{-6}
l. 4.20×10^{6}

Q1.1.3

What SI base unit would be appropriate for each measurement?

a. the length of a room
b. the amount of carbon in a diamond

c. the mass of NaCl in a bottle

Q1.1.4

List the meaning of each abbreviation of the base units.

a. m
b. K
c. kg
d. s
e. mol

Q1.1.5

What is the the derived unit from the SI base units for the relationship of each pair of quantities?

a. mass and volume
b. distance and time
c. amount of substance and volume
d. area

Q1.1.6

Give the meaning and name of each metrix prefix abbreviation.

a. M
b. m
c. n
d. d

Q1.1.7

Give the abbreviation and meaning of each metrix prefix.

a. kilo
b. centi
c. micro
d. giga

Q1.1.8

Name the prefix with the following numerical meaning.

a. 1/10
b. 1,000,000
c. 1/1,000,000
d. 1/100
e. 1

Q1.1.9

Convert each temperature to the missing one between Celsius and Fahrenheit.

a. 77°F
b. 212°F
c. 37°C
d. 22°C
e. 95°F
f. 15°C
g. 0°F

h. 0°C
i. –10°C
j. –10°F

1.2 Significant Figures

(click here for solutions)

Q1.2.1

Explain the similarities and differences between accuracy and precision.

Q1.2.2

The density of a copper sample was determined by three different students (shown below). Each performed the measurement three times and is reported below (all values in units of g/cm^3). The accepted value for the density of copper is 8.92 g/cm^3.

a. Determine if each student's data is accurate, precise, neither or both.
b. What is the average density based on Justin's data?
c. What is the average density based on Jane's data?
    - Jane: 8.94, 8.89, 8.91
    - Justin: 8.32, 8.31, 8.34
    - Julia: 8.64, 9.71, and 9.10

Q1.2.3

Determine the number of significant figures in each of the following values.

a. 406
b. 3.00
c. 3.20
d. 0.25
e. 0.0689
f. 0.002910
g. 3941
h. 46.250
i. 30.21
j. 0.10300

Q1.2.4

Write each value with three significant figures, use scientific notation if necessary.

a. 34500
b. 24
c. 0.0345
d. 0.012
e. 612.8
f. 98.22
g. 0.14928
h. 300

Q1.2.5

Give three examples of exact numbers.

Q1.2.6

Find the result of each of the following calculations and report the value with the correct number of significant figures.

a. 0.23 + 12.2 =
b. 13 - 1.03 =
c. 0.839 + 0.28925 =
d. 28 + 34.4 =
e. 0.8 + 2.3 =
f. 34.9 - 0.583 =
g. 21 - 0.132 =
h. 0.840 + 0.9334

Q1.2.7

Find the result of each of the following calculations and report the value with the correct number of significant figures.

a. 34 x 0.12 =
b. 68.2 / 0.78 =
c. $3.29 \times 10^4 \times 16.2$ =
d. 0.8449 x 29.7 =
e. $5.92 \times 10^3 / 0.628$ =
f. 3.00 x 2.6 =
g. 2.50 x 9.331 =
h. 3.20 / 12.75 =

1.3 Scientific Dimensional Analysis

(click here for solutions)

Q1.3.1

What is a conversion factor?

Q1.3.2

What is the conversion factor between each pair of units?

a. feet and inches
b. mL and cm^3
c. kg and g
d. cm and m
e. mm and cm
f. inches and centimeters
g. grams and pounds
h. g and µg (mcg)

Q1.3.3

Complete each of the following conversions.

a. 34 cm to m
b. 3.7 ft to in
c. 345 mg to Mg
d. 5.3 km to mm
e. 4.0 L to mL
f. 3.45×10^3 mm to km
g. 78 cm^3 to mL
h. 0.85 kg to dg

i. 13 pints to gallon
j. 0.35 L to cm^3

Q1.3.4

Complete each of the following conversions.

a. 342 cm^3 to dm^3
b. 2.70 g/cm^3 to kg/L
c. 34 mi/hr to km/min
d. 0.00722 km^2 to m^2
e. 4.9×10^5 mcm^3 to mm^3
f. 80. km/hr to mi/hr

1.4 Percentages

(click here for solutions)

Q1.4.1

Solve each of the following.

a. What percent of 35 is 8.2?
b. What percent of 56 is 12?
c. What percent of 923 is 38?
d. What percent of 342 is 118?

Q1.4.2

Solve each of the following?

a. What is 42% of 94?
b. What is 83% of 239?
c. What is 16% of 45?
d. What is 38% of 872?

Q1.4.3

Solve each of the following?

a. 42 is 34% of what number?
b. 73 is 82% of what number?
c. 13 is 57% of what number?
d. 75 is 25% of what number?
e. 25 is 15% of what number?
f. 98 is 76% of what number?

Q1.4.4

A patient originally weighs 182 pounds and loses 15.0% of their body weight. What is their final weight?

Q1.4.5

A patient's original weight was 135 pounds and they lose 12 pounds. What percent of their body weight did they lose?

Q1.4.6

A patient needs to increase their calcium supplement by 25% a week. If they are currently taking a 300. mg supplement, how much more will they need to take?

Q1.4.7

An infant's birth weight is 7 pounds, 1 ounce. Her discharge weight is 6 pounds, 13 ounces. What percent of her birth weight did she lose?

Q1.4.8

A patient needs a 20.% decrease in their medication dosage from 125 mg. What will his dosage be after the decrease?

Answers

1.1: Measurements Matter

Q1.1.1

a. 1.5×10^{8}
b. 4.3×10^{-5}
c. 3.32×10^{5}
d. 2.93×10^{-2}
e. 9.32×10^{2}
f. 1.873×10^{-1}
g. 7.8×10^{4}
h. 1×10^{-4}
i. 4.5×10^{3}
j. 2.9×10^{-3}
k. 6.281×10^{3}
l. 7×10^{-3}

Q1.1.2

a. 0.0038
b. 921000
c. 0.0791
d. 2500000
e. 0.0000000342
f. 540000
g. 0.003
h. 734
i. 0.00098
j. 60000000
k. 0.00000420
l. 4200000

Q1.1.3

a. meter
b. mole
c. kilogram

Q1.1.4

a. meter

b. Kelvin
c. kilogram
d. second
e. mole

Q1.1.5

a. kg/m^3
b. m/s
c. mol/m^3 is based on SI base units, but mol/L is also acceptable
d. m^2

Q1.1.6

a. Mega, 10^6
b. milli, 10^{-3}
c. nano, 10^{-9}
d. deci, 10^{-1}

Q1.1.7

a. k, 10^3
b. c, 10^{-2}
c. μ (or mc), 10^{-6}
d. G, 10^9

Q1.1.8

a. deci
b. mega
c. micro
d. centi
e. none (base unit)

Q1.1.9

a. 77°F = 25°C
b. 212°F = 100°C
c. 37°C = 98.6°F
d. 22°C = 72°F
e. 95°F = 35°C
f. 15°C = 59°F
g. 0°F = –18°C
h. 0°C = 32°F
i. –10°C = 14°F
j. –10°F = –23°C

1.2 Significant Figures

Q1.2.1

Accuracy is a measure of how close the values are close to the correct value while precision is a measure of how close values are to each other.

Q1.2.2

a.

- Jane: 8.94, 8.89, 8.91 - accurate and precise
- Justin: 8.32, 8.31, 8.34 - precise
- Julia: 8.64, 9.71, and 9.10 - neither accurate nor precise

a. 8.32 g/cm^3
b. 8.91 g/cm^3

Q1.2.3

a. 3
b. 3
c. 3
d. 2
e. 3
f. 4
g. 4
h. 5
i. 4
j. 5

Q1.2.4

a. 3.45×10^4
b. 2.40×10^1
c. 3.45×10^{-2}
d. 1.20×10^{-2}
e. 613 or 6.13×10^2
f. 9.82×10^1
g. 0.149 or 1.49×10^{-1}
h. 300. or 3.00×10^2

Q1.2.5

Answers will vary. 12 eggs, 100 cm = 1 m, 1 inch = 2.54 cm, 4 people

Q1.2.6

a. 0.23 + 12.2 = 12.43 = **12.4**
b. 13 - 1.03 = 11.97 = **12**
c. 0.839 + 0.28925 = 1.12825 = **1.128**
d. 28 + 34.4 = 62.4 = **62**
e. 0.8 + 2.3 = **3.1**
f. 34.9 - 0.583 = 34.317 = **34.3**
g. 21 - 0.132 = 20.868 = **21**
h. 0.840 + 0.9334 = 1.7734 = **1.773**

Q1.2.7

a. 34 x 0.12 = 4.08 = **4.1**
b. 68.2 / 0.78 = 87.4358974 = **87**
c. 3.29 x 10^4 x 16.2 = 5.32980×10^5 = 5.33×10^5
d. 0.8449 x 29.7 = 25.09353 = **25.1**
e. 5.92 x 10^3 / 0.628 = 9.4267515×10^3 = **9.43×10^3**
f. 3.00 x 2.6 = **7.8**

g. 2.50 x 9.331 = 23.3275 = **23.3**
h. 3.20 / 12.75 = 0.25098 = **0.251**

1.3 Scientific Dimensional Analysis

Q1.3.1

A conversion factor is a relationship between two units. The value in the numerator has some equivalence to the value in the denominator.

Q1.3.2

a. 1 foot = 12 inches
b. 1 mL = 1 cm^3
c. 1 kg = 1000 g or 1 × 10^{-3} kg = 1 g
d. 100 cm = 1 m or 1 cm = 1 × 10^{-2} m
e. 10 mm = 1 cm
f. 1 inch = 2.54 cm
g. 454 grams = 1 pound
h. 1 g = 1 × 10^{6} µg (mcg) or 1 × 10^{-6} g = 1 µg (mcg)

Q1.3.3

a. $34\ cm \times \frac{1\ m}{100\ cm} = 0.34\ m$
b. $3.7\ ft \times \frac{12\ in}{1\ ft} = 44.4\ in = 44\ in$
c. $345\ mg \times \frac{1\ g}{1000\ mg} \times \frac{1\ Mg}{1 \times 10^6\ g} = 3.45 \times 10^{-7}\ Mg$
d. $5.3\ km \times \frac{1000\ m}{1\ km} \times \frac{1000\ mm}{1\ m} = 5.3 \times 10^6\ mm$
e. $4.0\ L \times \frac{1000\ mL}{1\ L} = 4.0 \times 10^3\ mL$
f. $3.45 \times 10^3\ mm \times \frac{1\ m}{1000\ mm} \times \frac{1\ km}{1000\ m} = 3.45 \times 10^{-3}\ km$
g. $78\ cm^3 \times \frac{1\ mL}{cm^3} = 78\ mL$
h. $0.85\ kg \times \frac{1000\ g}{1\ kg} \times \frac{10\ dg}{1\ g} = 8.5 \times 10^3\ dg$
i. $13\ pints \times \frac{1\ quart}{2\ pints} \times \frac{1\ gallon}{4\ quarts} = 1.6\ gallons$
j. $0.35\ L \times \frac{1000\ mL}{1\ L} \times \frac{1\ mL}{1\ cm^3} = 3.5 \times 10^2\ cm^3$

Q1.3.4

a. $342\ cm^3 \times \frac{1\ dm}{10\ cm} \times \frac{1\ dm}{10\ cm} \times \frac{1\ dm}{10\ cm} = 0.342\ dm^3$ **or** $342\ cm^3 \times \left(\frac{1\ dm}{10\ cm}\right)^3 = 342\ cm^3 \times \frac{1^3\ dm^3}{10^3\ cm^3} = 0.342\ dm^3$
b. $\frac{2.70\ g}{cm^3} \times \frac{1\ kg}{1000\ g} \times \frac{1\ cm^3}{1\ mL} \times \frac{1000\ mL}{1\ L} = \frac{2.70\ kg}{L}$
c. $\frac{34\ mi}{hr} \times \frac{5280\ ft}{1\ mi} \times \frac{12\ in}{1\ ft} \times \frac{2.54\ cm}{1\ in} \times \frac{1\ m}{100\ cm} \times \frac{1\ km}{1000\ m} \times \frac{1\ hr}{60\ min} = \frac{0.91\ km}{min}$
d. $0.00722\ km^2 \times \frac{1000\ m}{1\ km} \times \frac{1000\ m}{1\ km} = 7.22 \times 10^3\ m^2$
e. $4.95 \times 10^5\ mcm^3 \times \frac{1\ mm}{1000\ mcm} \times \frac{1\ mm}{1000\ mcm} \times \frac{1\ mm}{1000\ mcm} = 4.95 \times 10^{-4}\ mm^3$
f. $\frac{80.\ km}{hr} \times \frac{1000\ m}{1\ km} \times \frac{100\ cm}{1\ m} \times \frac{1\ in}{2.54\ cm} \times \frac{1\ ft}{12\ in} \times \frac{1\ mi}{5280\ ft} = \frac{50.\ mi}{hr}$

1.4 Percentages

Q1.4.1

a. $\% = \frac{part}{whole} \times 100 = \frac{8.2}{35} \times 100 = 23\%$
b. $\% = \frac{part}{whole} \times 100 = \frac{12}{56} \times 100 = 21\%$
c. $\% = \frac{part}{whole} \times 100 = \frac{38}{923} \times 100 = 4.1\%$
d. $\% = \frac{part}{whole} \times 100 = \frac{118}{342} \times 100 = 34.5\%$

Q1.4.2

a.
$$\%=\frac{part}{whole}\times 100$$
$$42\%=\frac{part}{94}\times 100$$
$$part=39$$

b.
$$\%=\frac{part}{whole}\times 100$$
$$83\%=\frac{part}{239}\times 100$$
$$part=198=2.0\times 10^2$$

c.
$$\%=\frac{part}{whole}\times 100$$
$$16\%=\frac{part}{45}\times 100$$
$$part=7.2$$

d.
$$\%=\frac{part}{whole}\times 100$$
$$38\%=\frac{part}{872}\times 100$$
$$part=3.3\times 10^2$$

Q1.4.3

a.
$$\%=\frac{part}{whole}\times 100$$
$$34\%=\frac{42}{whole}\times 100$$
$$whole=1.2\times 10^2$$

b.
$$\%=\frac{part}{whole}\times 100$$
$$82\%=\frac{73}{whole}\times 100$$
$$whole=89$$

c.
$$\%=\frac{part}{whole}\times 100$$
$$57\%=\frac{13}{whole}\times 100$$
$$whole=23$$

d.
$$\%=\frac{part}{whole}\times 100$$
$$25\%=\frac{75}{whole}\times 100$$
$$whole=3.0\times 10^2$$

e.
$$\%=\frac{part}{whole}\times 100$$
$$15\%=\frac{25}{whole}\times 100$$
$$whole=1.7\times 10^2$$

f.
$$\%=\frac{part}{whole}\times 100$$
$$76\%=\frac{98}{whole}\times 100$$
$$whole=129$$

Q1.4.4

$$\%=\frac{part}{whole}\times 100$$
$$15.0\%=\frac{part}{182\ pounds}\times 100$$
$$part=27.3\ pounds\ lost$$

$$182\ pounds-27.3\ pounds=154.7\ pounds=155\ pounds$$

Q1.4.5

$$\% = \frac{part}{whole} \times 100$$
$$\% = \frac{12\ pounds}{135\ pounds} \times 100$$
$$\% = 8.9\%\ lost$$

Q1.4.6

$$\% = \frac{part}{whole} \times 100$$
$$25\% = \frac{part}{300.\ mg} \times 100$$
$$part = 75\ mg\ more$$

Q1.4.7

Convert both weights to ounces, find the ounces lost, and then find the percent lost.

Birth weight: $(7\ pounds \times 16) + 1\ \ ounce = 113\ \ ounces$

Discharge weight: $(6\ pounds \times 16) + 13\ \ ounces = 109\ \ ounces$

Weight lost: $113\ \ ounces - 109\ \ ounces = 4\ \ ounces$

Percent lost from original brith weight.

$$\% = \frac{part}{whole} \times 100$$
$$\% = \frac{4\ ounces}{113\ ounces} \times 100$$
$$\% = 3.5\% = 4\%$$

Q1.4.8

$$\% = \frac{part}{whole} \times 100$$
$$20\% = \frac{part}{125\ mg} \times 100$$
$$part = 25\ mg\ lost$$

$$125\ mg - 25\ mg = 100.\ mg$$

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3: MATTER AND CHANGE

3.1: Substances

3.2: Physical Properties

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3.1: Substances

Have you ever had the experience of opening the refrigerator and seeing something that you don't recognize?

Not an unfamiliar food, but something that has changed over the time period it was in the refrigerator. You are not sure just what it is, so you call it "stuff". Likely, after looking at it, you know it's not a good idea to eat it, so you throw it out. If you're not sure just what it is, you certainly do not want to put it in your body.

Substances

Silver and ordinary table salt are two examples of chemical **substances**. A substance is matter that has a uniform and definite composition. All samples of substances, sometimes called pure substances, have identical properties. When chemists run a chemical reaction, they want to use pure materials so they know exactly what they are dealing with. They know that the reaction involves a specific substance, so they expect the same reaction to give the same results each time that it is run.

One characteristic of a substance is that it cannot be separated into components by physical separation techniques. The **compound** acetylsalicylic acid (the active ingredient in aspirin) is a substance because it is a chemical compound that cannot be further reduced to other materials by physical means. However, the aspirin that contains the acetylsalicylic acid is a mixture. In addition to the active pain-relieving ingredient, aspirin also may contain waxes, corn starch, and cellulose, among other ingredients. These added materials bind the tablet components into a solid pill and influence how rapidly the material dissolves.



Figure 3.1.1: Aspirin tablets. (Pixabay License; Heung Soon via Pixabay)

Substances may be either **elements** or pure compounds. They may also be solid, liquid, or gas. Some of the elements (such as silver) exist in the solid form. Other elements (mercury) exist as liquids at room temperature. Elements such as helium and radon exist as gases at room temperature. A compound such as water will take the form of a solid, a liquid, or a gas, depending on the temperature.

Summary

- A substance is a pure material with a uniform and definite composition.
- A substance cannot be separated into components using physical means.
- A substance can be a solid, a liquid, or a gas, dependent on temperature.

Review

1. Define substance.
2. Can salt be a substance?

3. Is aspirin a substance?

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3.2: Physical Properties

Why are drag car standards constantly reinforced?

Drag racing is a highly competitive (and expensive) sport. There are a variety of classes of vehicles, ranging from stock classes (dependent on car weight, engine size, and degree of engine modification) all the way up to the Top Fuel class. Cars in the Top Fuel class may weigh over two thousand pounds, and are capable of top speeds of well over 300 miles per hour at the end of the quarter-mile. The standards for each class are well-defined, and frequent checks are made of engine dimensions and components to ensure that the rules are followed.

Physical Properties

A **physical property** is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Silver is a shiny metal that conducts electricity very well. It can be molded into thin sheets—a property called malleability. Salt is dull, brittle, and conducts electricity when it has been dissolved into water (which it does quite easily). Physical properties of matter include color, hardness, malleability, solubility, electrical conductivity, density, **melting point**, and **boiling point**.

For the elements, color does not vary much from one element to the next. The vast majority of elements are colorless, silver, or gray. Some elements do have distinctive colors: sulfur and chlorine are yellow, copper is (of course) copper-colored, and elemental bromine is red.

Density can be a very useful parameter for identifying an element. Of the materials that exist as solids at room temperature, iodine has a very low density compared to zinc, chromium, and tin. Gold has a very high density, as does platinum.

Hardness helps determine how an element (especially a metal) might be used. Many elements are fairly soft (silver and gold, for example) while others (such as titanium, tungsten, and chromium) are much harder. Carbon is an interesting example of hardness. In graphite (the "lead" found in pencils) the carbon is very soft, while the carbon in a diamond is roughly seven times as hard.



Figure 3.2.1: Pencil (Public Domain; Public Domain via Wikipedia)



Figure 3.2.2: Diamond ring (Public Domain; User:Ikkyu2/Wikimedia Commons via Wikipedia)

Melting point and boiling point are somewhat unique identifiers, especially in the case of compounds. In addition to giving some idea as to the identity of the compound, important information can be obtained about the purity of the material.

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Is that Really Gold? Archimedes and Density

How do you determine the density of an irregularly shaped object? Learn how Archimedes did this thousands of years ago with just a beaker, water, and some simple math.

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Summary

- A physical property is a characteristic of a substance that can be observed or measured without changing the identity of the substance.
- Physical properties include color, density, hardness, melting point, and boiling point.

Review

1. What is a physical property?
2. What color are most metals?
3. Is titanium harder or softer than gold?

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3.3: Extensive and Intensive Properties

How much is twenty dollars really worth?

You agree to mow someone's lawn for twenty dollars (it's a fairly large yard). When they pay you, they give you a $20 bill. It doesn't matter whether the bill is brand new or old, dirty, and wrinkled—all of these bills have the same value of $20. If you want more $20 bills, you have to mow more lawns. You can't say, "this particular bill is actually worth more than $20." To have more money, you have to put in more work.

Extensive Properties

Some properties of matter depend on the size of the sample, while some do not. An **extensive** property is a property that depends on the amount of matter in a sample. The mass of an object is a measure of the amount of matter that an object contains. A small sample of a certain type of matter will have a small mass, while a larger sample will have a greater mass. Another extensive property is volume. The volume of an object is a measure of the space that is occupied by that object.

The figure below illustrates the extensive property of volume. The pitcher and glass both contain milk. The pitcher holds approximately two quarts and the glass will hold about 8 ounces of milk. The same milk is in each container. The only difference is the amount of milk contained in the glass and in the pitcher.



Figure 3.3.1: Milk pitcher and glass. (Public Domain; Zenon Niewada (Wikimedia: Pitcherman) via Wikipedia)

Intensive Properties

The electrical conductivity of a substance is a property that depends only on the type of substance. Silver, gold, and copper are excellent conductors of electricity, while glass and plastic are poor conductors. A larger or smaller piece of glass will not change this property. An **intensive** property is a property of matter that depends only on the type of matter in a sample and not on the amount. Other intensive properties include color, temperature, density, and solubility.



Figure 3.3.1: Copper wire. (Credit: User:Inductiveload/Wikimedia Commons; Source: http://commons.wikimedia.org/wiki/File:Tinned_Copper_Wire_anaglyph.jpg)

The copper wire shown in the picture below has a certain electrical conductivity. You could cut off the small end that sticks out, and it would have the same conductivity as the entire long roll of wire shown here. The conductivity is a property of the copper metal itself, not of the length of the wire.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Summary

- An extensive property is a property that depends on the amount of matter in a sample.
- Mass and volume are examples of extensive properties.
- An intensive property is a property of matter that depends only on the type of matter in a sample and not on the amount.
- Color, temperature, and solubility are examples of intensive properties.

Review

1. Define extensive property.
2. Give two examples of extensive properties.
3. Define intensive property.
4. Give two examples of intensive properties.

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3.4: Solid, Liquid, and Gas

Why is the state of water different in each picture?

Water can take many forms. At low temperatures (below $0^{\circ}\mathrm{C}$), water is a solid. When at "normal" temperatures (between $0^{\circ}\mathrm{C}$ and $100^{\circ}\mathrm{C}$), it is a liquid. At temperatures above $100^{\circ}\mathrm{C}$, water is a gas (steam).



Figure 3.4.1 (Public Domain; Iceberg: Rear Admiral Harley D. Nygren, NOAA; Beach: User:Wicki/Wikimedia Commons; Volcano: E. Klett, US Fish and Wildlife Service via Iceberg: http://commons.wikimedia.org/wiki/File:Ice_berg.jpg; Beach: http://commons.wikimedia.org/wiki/File:Ocean_Spokojny.JPG; Volcano: http://commons.wikimedia.org/wiki/File:Dds40-097_large.jpeg)

The state of water depends on the temperature. Each state (solid, liquid, and gas) has its own unique set of physical properties.

Matter and its States

Matter typically exists in one of three states: **solid**, **liquid**, or **gas**. There is a fourth state of matter called **plasma**, which rarely exists on earth, but we will omit this from our current discussion. The state a given substance exhibits is also a physical property. Some substances exist as gases at room temperature (oxygen and carbon dioxide), while others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states.

Liquid

Liquids have the following characteristics:

- No definite shape (takes the shape of its container).
- Has definite volume.
- Particles are free to move over each other, but are still attracted to each other.

A familiar liquid is mercury metal. Mercury is an anomaly. It is the only metal we know of that is liquid at room temperature. Mercury also has an ability to stick to itself (surface tension), which is a property that all liquids exhibit. Mercury has a relatively high surface tension, and this makes it very unique. Here you can see mercury in its common liquid form.

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Figure 3.4.2: Mercury. (the EPA; Source: http://commons.wikimedia.org/wiki/File:Mercury-element.jpg; License: Public Domain)

If we heat liquid mercury to its boiling point of $357^{\circ}\mathrm{C}$, and contain it under the right pressure conditions, we would notice all particles in the liquid state go into the gas state.

Gas

Gases have the following characteristics:

- No definite shape (takes the shape of its container).
- No definite volume.
- Particles move in random motion with little or no attraction to each other.
- Highly compressible.

Solid

Solids are defined by the following characteristics:

- Definite shape (rigid).
- Definite volume.
- Particles vibrate around fixed axes.

If we were to cool a sample of liquid mercury to its freezing point of $-39^{\circ}\mathrm{C}$, and had it contained under the right pressure conditions, we would notice all of the liquid particles would go into the solid state.

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As you can see in the video, mercury can be solidified when its temperature is brought to its freezing point. However, when returned to room temperature conditions, mercury does not exist in solid state for long, and returns back to its more common liquid form.

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Plasma

Plasma is a state of matter that resembles a gas but has certain properties that gases do not have. Like a gas, plasma consists of particles of matter that can pull apart and spread out, so it lacks a fixed volume and a fixed shape. Unlike a gas, plasma can conduct electricity and respond to a magnetic field. That's because plasma consists of electrically charged particles called ions, instead of uncharged particles such as atoms or molecules.

Plasma are defined by the following characteristics:

- particles are charged ions and free electrons
- no definite shape
- no definite volume
- conducts electricity
- responds to magnetic field

Summary

- Three states of matter exist: solid, liquid, and gas.
- Solids have a definite shape and volume.
- Liquids have a definite volume, but take the shape of their container.
- Gases have no definite shape or volume.

Review

1. How many states of matter are there?
2. What is a solid?
3. What is a liquid?
4. What is a gas?

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3.5: Physical Change

How does a car go from new and shiny to old and beaten up?

Want to buy a car – cheap? Notice there is no specification such as “in good condition” or “needs a little work.” The car above is pretty beat up. The body is damaged, the windows are broken, and the interior is probably torn up. But this is still a car. It has all the components of a car, even though you would not want to buy it in the present condition. But change that condition and you have a (possibly) useable car.



Figure 3.5.1 (CC BY 2.0; User: Paul / Wikimedia Commons via Wikipedia)

Physical Change

As an ice cube melts, its shape changes as it acquires the ability to flow. However, its composition does not change. **Melting** is an example of a **physical change**. A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not. Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include **vaporization** (liquid to gas), **freezing** (liquid to solid), and **condensation** (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved into water, the salt is said to have entered the aqueous state. The salt may be regained by boiling off the water, leaving the salt behind.



Figure 3.5.2: Melting ice in the Beaufort Sea. (Public Domain; Rear Admiral Harley D. Nygren, NOAA via Wikipedia)

When a piece of wood is ground into sawdust, that change is irreversible since the sawdust cannot be reconstituted into the same piece of wood that it was before. Cutting grass or pulverizing a rock are examples of irreversible physical changes. Chopping wood for a fire also represents an irreversible physical change, since the pieces cannot be put back together to form the tree.



Figure 3.5.3: Chopping wood is an irreversible physical change. (Pixabay License; stevepb via Pixabay)

Physical Changes vs. Chemical Changes

What is the difference between physical and chemical changes? Watch the video to find out.

YouTube Video

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Practice

Explore a campout to see if you can determine which things are undergoing physical changes and which are undergoing chemical changes.

Summary

- A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not.
- A physical change can be reversible where the original form of the matter can be restored, or irreversible where the original form cannot be restored.
- Melting ice and grinding wood into sawdust are examples of physical changes.

Review

1. Define physical change.
2. Why is melting an ice cube a reversible physical change?
3. Give an example of an irreversible physical change.

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3.6: Homogeneous Mixture

How do you like your coffee?

Many people enjoy a cup of coffee at some point during the day. Some may drink it black, while others may put cream (or some dairy substitute) and sugar in their coffee. High-end coffee drinks can be purchased at espresso stands (either sit-down or drive-through). Whatever the coffee drinker's preference, they want the coffee to be the same at the beginning as it is at the end of the drink. They don't want the components to separate out, but want their drink to be uniform from top to bottom.

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 or large amount of salt 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.

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Homogenous Mixtures

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 would be considered homogeneous because the dissolved material is present in the same amount throughout the solution.

One characteristic of mixtures is that they can be separated into their components. Since each part of the mixture has not reacted with another part of the mixture, the identities of the different materials are unchanged.

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Simulation

Take a deep breath...you just breathed in a mixture. This simulation will help you explore the mixture that surrounds all of us.

Summary

- A homogeneous mixture is a mixture in which the composition is uniform throughout the mixture.
- All solutions would be considered homogeneous.

Review

1. What is a mixture?
2. What is a homogeneous mixture?
3. Are all solutions homogeneous mixtures?
4. Can homogeneous mixtures be separated into their components?

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3.7: Heterogeneous Mixtures

What is the best way to eat a bag of jelly beans? Many people open the bag and eat all the candy, no matter what flavor each piece is. Others pick through the collection. They may say, "I don't like the orange ones." Or maybe they only care for the lemon ones. There are different kinds of jelly beans in the mixture and people will eat what they want and get rid of the rest.

Heterogeneous Mixtures

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.

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.



Figure 3.7.1: Oil and water do not mix, instead forming two distinct layers called phases. The oil phase is less dense than the water phase, and so the oil floats on top of the water. (CC BY 2.0; Flickr: Yortw via Flickr)

In the case of vegetable soup, one phase would be the liquid soup itself. This phase has vitamins, minerals, and other components dissolved in the water. This phase would be homogeneous. The carrots, peas, corn, or other vegetables represent other phases of the soup.The various vegetables are not mixed evenly in the soup, but are spread around at random.

There are a large number of heterogeneous mixtures around us. Soil is composed of a variety of substances, and is often of different composition depending on the sample taken. One shovelful may come up with dirt and grass, while the next shovelful could contain an earthworm.

Smog is another example of a heterogeneous mixture. This murky collection of pollutants can be a mixture of water and contaminants from burning gasoline or plastics, mixed with nitric oxide derivatives and ozone. You can see that the smog distribution in the air illustrated below is not evenly spread out, but varies from one part of the atmosphere to another.



Figure 3.7.2: Smog in New York City. (Public Domain; Dr. Edwin P. Ewing, Jr./CDC via Wikipedia)

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Simulation

Whether the air is clean or smoggy, it's a collection of molecules. See what it's made of in this simulation.

Summary

- A heterogeneous mixture is a mixture in which the composition is not uniform throughout the mixture.
- A phase is a separate layer in a heterogeneous mixture.

Review

Use the link below to answer the following questions:

1. Define a heterogeneous mixture.
2. Why is vegetable soup a heterogeneous mixture?
3. How many phases are in a heterogeneous mixture?

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3.8: Methods for Separating Mixtures



Figure 3.8.1 (CC by 2.0; Marcin Chady via Flicker photo, Marcin Chady [www.flickr.com])

How did gold miners search for gold?

Beginning in the late 1840s, thousands of prospectors rushed to California in search of gold. One of the approaches taken to isolate the gold from the soil was called "panning". Dirt would be placed in the pan and covered with water. After thorough mixing, the pan is gently swirled to remove dissolved material while the heavier gold settles to the bottom of the pan. The gold is then separated from the mixture of soil and water.

Separation of Mixtures

Not everyone is out searching for gold (and not many of those searchers is going to get much gold, either). In a chemical reaction, it is important to isolate the component(s) of interest from all the other materials so they can be further characterized. Studies of biochemical systems, environmental analysis, pharmaceutical research – these and many other areas of research require reliable separation methods.

Here are a number of common separation techniques:

Chromatography

Chromatography is the separation of a mixture by passing it in solution or suspension, or as a vapor (as in gas chromatography), through a medium in which the components move at different rates. Thin-layer chromatography is a special type of chromatography used for separating and identifying mixtures that are or can be colored, especially pigments.

Distillation

Distillation is an effective method to separate mixtures that are comprised of two or more pure liquids. Distillation is a purification process where the components of a liquid mixture are vaporized and then condensed and isolated. In simple distillation, a mixture is heated, and the most volatile component vaporizes at the lowest temperature. The vapor passes through a cooled tube (a condenser), where it condenses back into its liquid state. The condensate that is collected is called distillate.

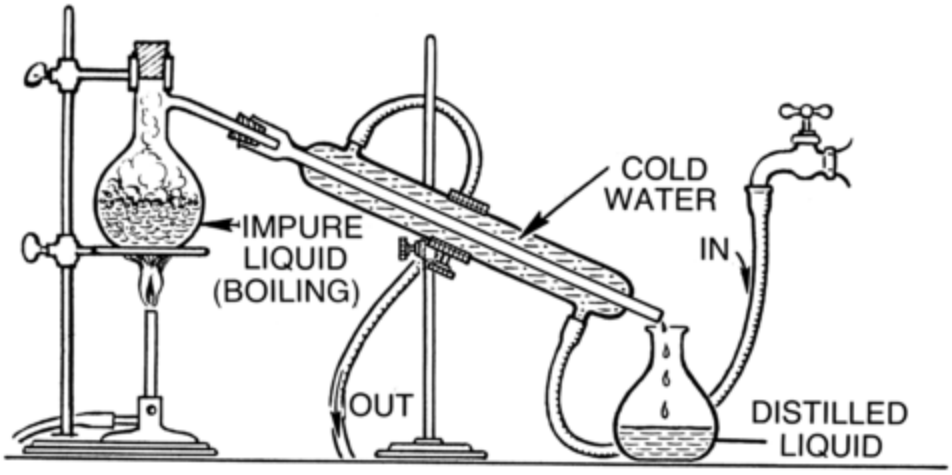


Figure 3.8.2: Distillation apparatus. (Public Domain; Pearson Scott Foresman via Commons Wikimedia, Distillation Apparatus [commons.wikimedia.org])

In the figure above, we see several important pieces of equipment. There is a heat source, a test tube with a one-hole stopper attached to a glass elbow, and rubber tubing. The rubber tubing is placed into a collection tube, which is submerged in cold water. There are other more complicated assemblies for distillation that can also be used, especially to separate mixtures which are comprised of pure liquids, with boiling points that are close to one another.

Evaporation

Evaporation is a technique used to separate out homogeneous mixtures that contain one or more dissolved salts. The method drives off the liquid components from the solid components. The process typically involves heating the mixture until no more liquid remains. Prior to using this method, the mixture should only contain one liquid component, unless it is not important to isolate the liquid components. This is because all liquid components will evaporate over time. The evaporation method is suitable to separate a soluble solid from a liquid.

In many parts of the world, table salt is obtained from the evaporation of sea water. The heat for the process comes from the sun.



Figure 3.8.3: Once the sea water in these evaporation ponds has evaporated, the salt can be harvested. (CC BY 3.0; Sandip Dey (Wikimedia: Dey.sandip) via Commons Wikimedia, Salt pans in Marakkanam, India [commons.wikimedia.org])

Filtration

Filtration is a separation method used to separate out pure substances in mixtures comprised of particles—some of which are large enough in size to be captured with a porous material. Particle size can vary considerably, given the type of mixture. For instance, stream water is a mixture that contains naturally occurring biological organisms like bacteria, viruses, and protozoa. Some water filters can filter out bacteria, the length of which are on the order of 1 micron. Other mixtures, like soil, have relatively large particle sizes, which can be filtered through something like a coffee filter.

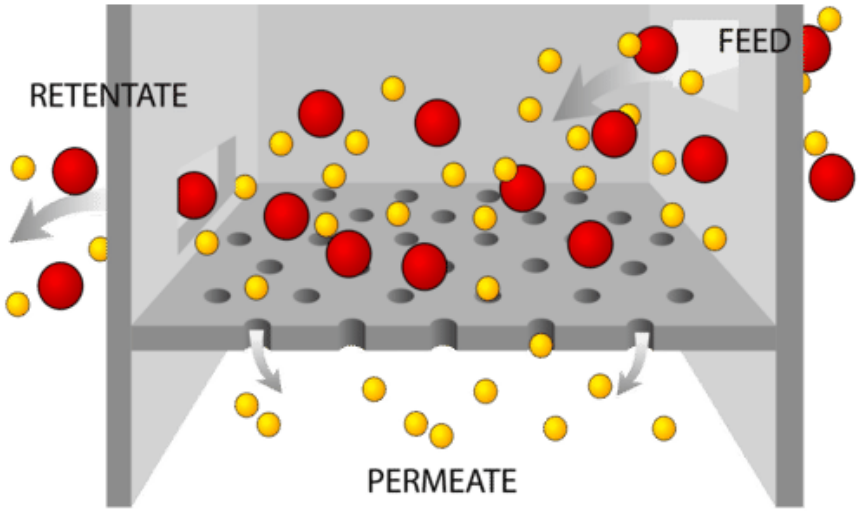


Figure 3.8.3: Filtration. (Public Domain; Mariana Ruiz Villarreal (Wikimedia: LadyofHats) via Commons Wikimedia, Filtration Diagram [commons.wikimedia.org])

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Summary

- Mixtures can be separated using a variety of techniques.
- Chromatography involves solvent separation on a solid medium.
- Distillation takes advantage of differences in boiling points.
- Evaporation removes a liquid from a solution to leave a solid material.
- Filtration separates solids of different sizes.

Review

1. Why is it important to separate materials from a mixture?
2. What is chromatography?
3. What is distillation
4. What is filtration?
5. What is evaporation?
6. What technique would you use to separate sand from water? There are two possibilities.
7. What technique would you use to separate alcohol from water?

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3.9: Elements



Figure 3.9.1 (Public Domain; Sidney Paget via Wikipedia)

Who is Sherlock Holmes?

The famous fictional British detective Sherlock Holmes supposedly often made the statement "elementary, my dear Watson". In reality, the closest he ever came to that line was an exchange with Watson in the short story "The Crooked Man". Holmes demonstrated shrewd insight into Watson's activities of the day. When asked how he knew what Watson was doing, Holmes simply replies "Elementary". Regardless of exactly how stated it, Sherlock Holmes was simply referring to what the Free Dictionary defines as "relating to, or constituting the basic, essential, or fundamental part".

Elements

An **element** is the simplest form of matter that has a unique set of properties. It is a pure substance that cannot be broken down into different types of substances. As of 2010, there are 118 known elements. Ninety eight of the elements are naturally occurring on earth, while the other 20 elements have been made synthetically in a lab. The names of all the elements can be found on the periodic table.

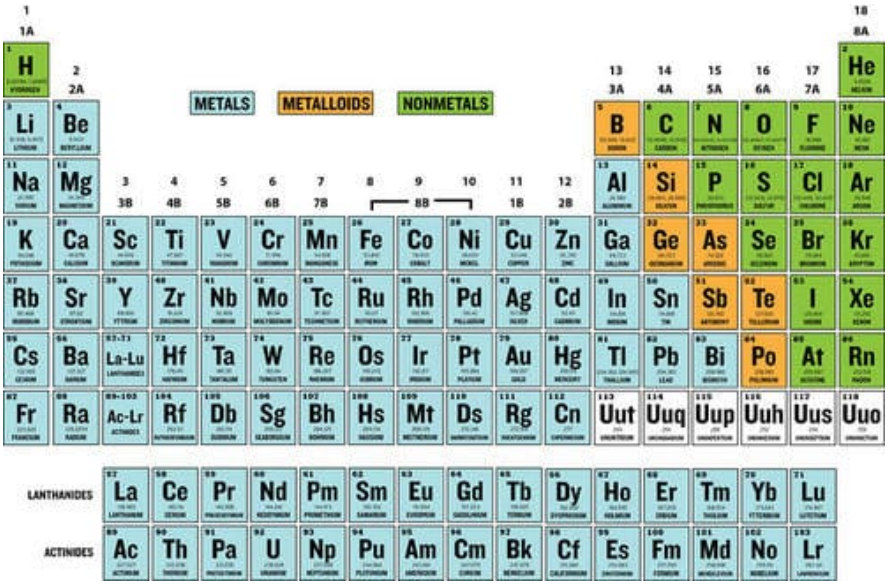


Figure 3.9.2: The Periodic Table. (CC BY-NC 3.0; Christopher AuYeung via CK-12 Foundation)

Examples of well-known elements include oxygen, iron, and gold (see the figure below). Elements cannot be broken down into a simpler substance. Likewise, one element cannot be chemically converted into a different element.



Figure 3.9.3: Chemical elements are the simplest of substances. (A) An oxygen tank like this is used by people who need breathing assistance. (B) A simple skillet can be made from cast iron. (C) Gold bars are formed and used for monetary purposes. (Credit: (A) Leon Brocard (Flickr: acme); (B) User:Evan-Amos/Wikimedia Commons; (C) Agnico-Eagle Mines Limited; Source: (A) http://www.flickr.com/photos/acme/8532536; (B) http://commons.wikimedia.org/wiki/File:Cast-Iron-Pan.jpg; (C) http://commons.wikimedia.org/wiki/File:Gold_Bars.jpg; http://commons.wikimedia.org/wiki/File:Coypel_Democritus.jpg License: (A) CC-BY 2.0; (B) Public Domain; (C) Public Domain)

Some elements have been known for centuries (gold, silver, iron, and copper, for example), while others have been created in a lab only within the last several years. Most elements do not exist as such in nature. They are so reactive that they can be found only in combination with other materials.

Some elements are very valuable, while others are quite inexpensive. Gold is currently worth almost $1700 per ounce. Aluminum, on the other hand, only sells for about 90 cents per pound, considerably lower than gold. Copper is worth somewhat more, selling for approximately $3.50 per pound. Platinum is very valuable at about $1650 an ounce, though not quite as expensive as gold.

Simulation

There are elements all around you, including in the air you breathe. See what elements are in the air in this simulation.

Video: The Chemical Elements Song

The Chemical Elements song, sung by Tom Lehrer. At the time, there were only 102 elements discovered.

YouTube Video

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Summary

- An element is the simplest form of matter that has a unique set of properties.
- One element cannot be chemically converted to another element.

Review

1. What is an element?
2. Give the names of any two elements.

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3. What is the name of the table that has the names of the elements?
4. How many elements are naturally occurring?

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3.10: Compounds



Figure 3.10.1 (Public Domain; User:JohnABerring27A via Wikipedia)

What do you need to start building a house?

Houses take on a variety of shapes. Some are large mansions while others are simpler dwellings. No matter what the house looks like there will be certain common components – wood, brick, plaster, nails, plumbing, wiring, etc. These components go together in numerous ways to make up the unique final product.

Compounds in chemistry are composed of different types and numbers of atoms. But atoms are common building blocks for a wide variety of compounds, just as brick, wood, and other materials are common building blocks for many types of houses.

Compounds

A **compound** is a substance that contains two or more elements chemically combined in a fixed proportion. The elements carbon and hydrogen combine to form many different compounds. One of the simplest is called methane, in which there are always four times as many hydrogen particles as carbon particles. Methane is a pure substance because it always has the same composition. However, it is not an element because it can be broken down into simpler substances—carbon and hydrogen.

Recall that the components of a mixture can be separated from one another by physical means. This is not true for a compound. Table salt is a compound consisting of equal parts of the elements sodium and chlorine. Salt cannot be separated into its two elements by filtering, distillation, or any other physical process. Salt and other compounds can only be decomposed into their elements by a chemical process. A **chemical change** is a change that produces matter with a different composition. Many compounds can be decomposed into their elements by heating. When sugar is heated, it decomposes into carbon and water. Water is still a compound, but one which cannot be broken down into hydrogen and oxygen by heating. Instead, the passage of an electrical current through water will produce hydrogen and oxygen gases.

The properties of compounds are generally very different than the properties of the elements from which the compound is formed. Sodium is an extremely reactive soft metal that cannot be exposed to air or water. Chlorine is a deadly gas. The compound sodium chloride is a white solid which is essential for all living things (see below).

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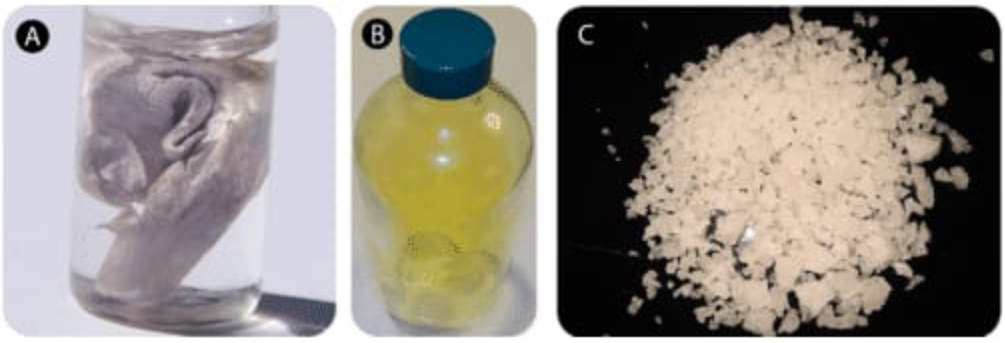


Figure 3.10.1: (A) Sodium is so reactive that it must be stored under oil. (B) Chlorine is a poisonous yellow-green gas. (C) Salt crystals, a compound of sodium and chlorine. (Credit: (A) Hi-Res Images of Chemical Elements; (B) User: Greenhorn1/Wikimedia Commons; (C) User:Miansari66/Wikimedia Commons; Source: (A) http://images-of-elements.com/sodium.php; (B) http://commons.wikimedia.org/wiki/File:Chlorine2.jpg; (C) http://commons.wikimedia.org/wiki/File:Salt_,_Lake_salt.JPG License: (A) CC by 3.0 (B) Public Domain; (C) Public Domain)

Science Friday: Celebrating Explosive Chemistry

July Fourth: A day for picnics, parades and chemistry. In this video by Science Friday, Professor Bassam Shakhashiri explains some of the science of fireworks.

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Summary

- A compound is a substance that contains two or more elements chemically combined in a fixed proportion.
- A chemical change is a change that produces matter with a different composition.

Review

1. What is a compound?
2. How is a compound different from an element?
3. What is a chemical change?

Explore More

Use the resource below to answer the following questions.

YouTube Video

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1. What lists all the elements we know about?
2. What compound is made of two hydrogen atoms and one oxygen atom?
3. What compound is made of one carbon atom and two oxygen atoms?

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3.11: Chemical Change



Figure 3.11.1 (Public Domain; Jon Sullivan via Wikipedia)

Do you like to cook?

Cooking is a valuable skill that can benefit everyone. Whether it is fixing a simple grilled cheese sandwich or preparing an elaborate meal, cooking demonstrates some basic concepts of chemistry. When you bake bread, you mix flour, sugar, yeast, and water together. After baking, this mixture has been changed to form bread, another substance that has different characteristics and qualities from the original materials. The process of baking has produced chemical changes in the ingredients that result in bread being made.

Chemical Change

Most of the elements we know about do not exist freely in nature. Sodium cannot be found by itself (unless we prepare it in the laboratory) because it interacts too easily with other materials. On the other hand, the element helium does not interact with other elements to any extent. We can isolate helium from natural gas during the process of drilling for oil.

A chemical change produces different materials than the ones we started with. One aspect of the science of chemistry is the study of the changes that matter undergoes. If chemistry was simply a study of elements that did nothing, life would be very boring (in fact, life would not exist since the elements are what make up our bodies and sustain us). But the processes of change that take place when different chemicals are combined produce all the materials that we use daily.

One type of chemical change (already mentioned) is when two elements combine to form a compound. Another type involves the breakdown of a compound to produce the elements that make it up. If we pass an electric current through bauxite (aluminum oxide, the raw material for aluminum metal), we get metallic aluminum as a product.



Figure 3.11.1: Electrolytic production of aluminum. (Public Domain; Ronny Jacques via Wikipedia)

However, the vast majority of chemical changes involve one compound being transformed into another compound. There are literally millions of possibilities when we take this approach to chemical change. New compounds can be made to produce better fabrics that are easier to clean and maintain; they can help preserve food so it doesn't spoil as quickly; we can make new medicines to treat diseases; even products, such as airbags (as seen in the simulation below), use a chemical change to improve the world around us.

Simulation

How does a chemical reaction inflate an airbag?

Summary

A chemical change produces different materials than the ones we started with.

Review

1. What is a chemical change?
2. List three types of chemical changes.

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3.12: Chemical Symbols and Formulas

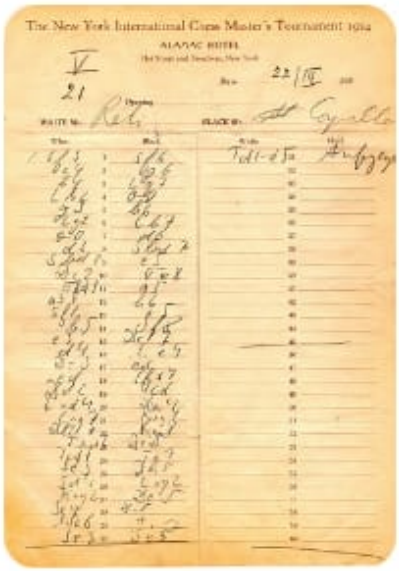


Figure 3.12.1 (Author Unknown; Source: Commons Wikimedia, Capablanca [commons.wikimedia.org]; License: Public Domain)

How do chess players monitor their moves in a game?

Suppose you were walking along and noticed a piece of paper on the ground with markings on it. You pick it up and see the paper in the picture above.

To most people, these notes are meaningless (maybe they're a secret spy code). But to a chess player, these symbols tell the story of a chess game. Each abbreviation describes a chess piece or a move during the game. The use of special symbols allows chess players to "see" the game without having to read a wordy and possibly incomplete description of what happened.

Chemical Symbols and Formulas

Chemical Formula

In order to illustrate chemical reactions and the elements and **compounds** involved in them, chemists use symbols and formulas. 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 below.

Table 1: 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
Sn	Tin	Stannum

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

Watch the following video to learn more about the origin of different element's symbols.

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Chemical Formulas

Compounds are combinations of two or more elements. A **chemical formula** is an expression that shows the elements in a compound and the relative proportions of those elements. Water is composed of hydrogen and oxygen in a two to one ratio. The chemical formula for water is $\mathrm{H_2O}$. Sulfuric acid is one of the most widely produced chemicals in the United States and is composed of the elements hydrogen, sulfur, and oxygen. The chemical formula for sulfuric acid is $\mathrm{H_2SO_4}$.

Summary

- A chemical symbol is a one- or two-letter designation of an element.
- Compounds are combinations of two or more elements.
- A chemical formula is an expression that shows the elements in a compound and the relative proportions of those elements.
- Some elements have symbols that derive from the Latin name for the element.

Review

1. What is a chemical symbol?
2. What is a chemical formula?
3. How many hydrogen atoms are in one molecule of the compound $C_{12}H_{22}O_{11}$?
4. What is the Latin name for the element potassium?

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3.13: Chemical Properties and Chemical Reactions

Have you ever left your bicycle out in the rain?

It's not a good idea because the rain can cause the bicycle to rust. You start to get a reddish-orange build-up of a rough deposit on the metal. It may start with the chain, but can spread to other parts of the bicycle, especially if there are scratches that create a bare metal surface. The formation of rust is a chemical process that takes place when iron is exposed to water and oxygen. It is estimated that damage due to rust costs U.S. businesses, military, and government over 276 billion dollars a year—a very expensive chemical process.

Chemical Properties

A **chemical property** describes the ability of a substance to undergo a specific chemical change. A chemical property of iron is that it is capable of combining with oxygen to form iron oxide, the chemical name of rust. The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used to describe chemical changes are: burn, rot, explode, decompose, and ferment. Chemical properties are very useful in identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

Chemical Change

A **chemical change** is also called a chemical reaction. A chemical reaction is a process that occurs when one or more substances are changed into one or more new substances. Zinc (Zn) is a silver gray element that can be ground into a powder. If zinc is mixed at room temperature with powdered sulfur (S), a bright yellow element, the result will simply be a mixture of zinc and sulfur. No chemical reaction occurs. However, if energy is provided to the mixture in the form of heat, the zinc will chemically react with the sulfur to form the compound zinc sulfide (ZnS). The figure below shows the substances involved in this reaction.

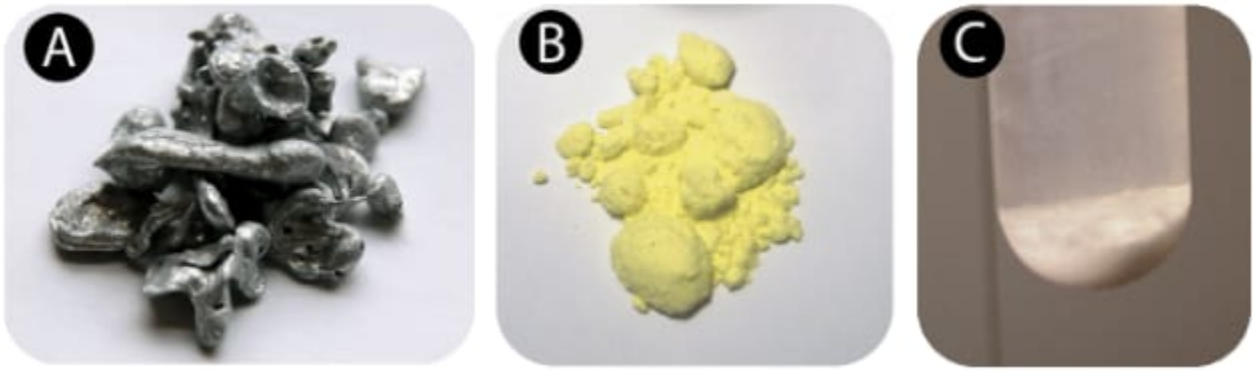


Figure 3.13.1: Zinc and sulfur are two elements that undergo a chemical reaction when heated to form the compound zinc sulfide. ((A; (A) Ben Mills (Benjah-bmm27); (B) User:Vineyard/Wikimedia Commons; (C) A. Eisen/NOAA via (A) http://commons.wikimedia.org/wiki/File:Zinc-sample.jpg; (B) http://commons.wikimedia.org/wiki/File:SulfurReagent.jpg; (C) http://oceanexplorer.noaa.gov/explorations/03windows/logs/jul30/media/zincsulfideprecipitate.html) Public Domain; (B) CC BY 3.0; (C) Public Domain)

Alkaline metals are highly reactive with water. Watch the video to see what happens when different alkaline metals are exposed to water.

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Summary

- A chemical property describes the ability of a substance to undergo a specific chemical change.
- A chemical reaction is a process that occurs when one or more substances are changed into one or more new substances.

Review

1. What is a chemical property?
2. What is a chemical reaction?
3. When can we observe a chemical property?
4. Is freezing water a chemical property?

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3.14: Reactants and Products



Figure 3.14.1 (Public Domain; Jon Sullivan via Wikipedia)

What would you do without a computer?

Over the last few decades, computers have proven to be extremely useful tools for organizing and processing information. Data, such as the structural parameters of a compound, is put into the computer; the output can be a detailed diagram of how that molecule looks in three dimensions. The computer and its programs transform input data into a useful final product.

Chemical Equation

The reaction between zinc and sulfur can be shown in what is called a **chemical equation**. In words, we could write the reaction as:

$$\text{zinc} + \text{sulfur} \rightarrow \text{zinc sulfide}$$

The more convenient way to express a chemical reaction is to use the symbols and formulas of the substances involved:

$$\text{Zn} + \text{S} \rightarrow \text{ZnS}$$

The substance(s) to the left of the arrow in a chemical equation are called reactants. A **reactant** is a substance that is present at the start of a chemical reaction. The substance(s) to the right of the arrow are called **products**. A product is a substance that is present at the end of a chemical reaction. In the equation above, the zinc and sulfur are the reactants that chemically combine to form the zinc sulfide product.

There is a standard way of writing chemical equations. The reactants are all written on the left-hand side of the equation, with the products on the right-hand side. An arrow points from the reactants to the products to indicate the direction of the reaction:

$$\text{reactants} \rightarrow \text{products}$$

When appropriate, a symbol may be written above or below the arrow to indicate some special circumstance. The symbol "Δ" is often used to indicate that the reaction is to be heated.

The presence of the arrow also indicates that the reaction goes in one direction under the conditions indicated. There are reactions which can be easily reversed, but we will not take those up right now.

There are a wide variety of reactions possible: elements may form compounds (as seen in the reaction above), compounds may form elements (water will break down in the presence of an electric current to form hydrogen gas and oxygen gas) or compounds may combine, break apart, or rearrange to form new materials.

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Summary

- A chemical equation describes a chemical reaction.
- Reactants are starting materials and are written on the left-hand side of the equation.
- Products are the end result of the reaction and are written on the right-hand side of the equation.

Review

1. What is a reactant?
2. What is a product?
3. What does a chemical equation do?
4. In the reaction sodium + water → sodium hydroxide + hydrogen,
   a. what are the reactants?
   b. what are the products?

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3.15: Recognizing Chemical Reactions



Figure 3.15.1 (Public Domain; Jon Sullivan via Wikipedia)

Have you ever cooked a pizza?

Making a pizza can be as easy as buying a "take and bake" from a store and putting it in the oven, or as complicated as mixing the dough and loading it up with your favorite toppings before baking it. How do you know when it is done? The most obvious sign is that the crust turns light brown. The dough is no longer flexible, but much more rigid. Maybe the cheese has melted. You certainly want the pizza to be cooked, not half-raw.

Recognizing Chemical Reactions

How can a scientist tell if a **chemical reaction** is taking place? There are four visual clues that indicate that a chemical reaction is likely occurring:

1. A change in color occurs during the reaction.
2. A **gas** is produced during the reaction.
3. A solid product called a **precipitate** is produced in the reaction.
4. A transfer of energy occurs as a result of the reaction.

Mercury (II) oxide is a red solid. When it is heated to a temperature above $500^{\circ}\mathrm{C}$, it easily decomposes into mercury and oxygen gas. The red color of the mercury oxide reactant becomes the silver color of mercury. The color change is a sign that the reaction is occurring.



Figure 3.15.1: Mercuric oxide. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)



Figure 3.15.2: Mercury metal. (Public Domain; User:A/Wikimedia Commons via Wikipedia)

When zinc reacts with hydrochloric acid, the reaction bubbles vigorously as hydrogen gas is produced. The production of a gas is also an indication that a chemical reaction is occurring.



Figure 3.15.3: Zinc reacting with hydrochloric acid produces bubbles of hydrogen gas. (Public Domain; User:Chemicalinterest/Wikimedia Commons via Wikipedia)

When a colorless solution of lead (II) nitrate is added to a colorless solution of potassium iodide, a yellow solid called a precipitate is instantly produced. A precipitate is a solid product that forms from a reaction and settles out of a liquid solution. The formation of a precipitate is an indication of a chemical reaction.



Figure 3.15.4: A yellow precipitate of solid lead (II) iodide forms immediately when solutions of lead (II) nitrate and potassium iodide are mixed. (CC by 2.0; Paige Powers (Flickr: paigggeyy) via Flickr)

All chemical changes involve a transfer of energy. When zinc reacts with hydrochloric acid, the test tube becomes very warm as energy is released during the reaction. Some other reactions absorb energy. While energy changes are a potential sign of a chemical reaction, care must be taken to ensure that a chemical reaction is indeed taking place. Physical changes also involve a transfer of energy. The melting of a solid absorbs energy, while the condensation of a gas releases energy. The only way to be certain that a chemical reaction has taken place is to test the composition of the substances after the change has taken place, to see if they are different from the starting substances.

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Summary

- There are four visual clues that indicate that a chemical reaction is likely occurring:
  - A change of color occurs during the reaction.
  - A gas is produced during the reaction.
  - A solid product called a precipitate is produced in the reaction.
  - A transfer of energy occurs as a result of the reaction.

Review

1. What was the color change when mercury (II) oxide was heated?
2. What happened when zinc metal was mixed with hydrochloric acid?
3. What happens when lead nitrate and potassium iodide are mixed?

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4: ATOMIC STRUCTURE

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4.1: Democritus' Idea of the Atom



Figure 4.1.1 (CC by 2.0; Justin Norris (Flickr:JustinMN) via Flickr)

What would the philosophers do?

People enjoy getting together to discuss things, whether it is how your favorite sports team is doing, what the best new movie is, the current politics, or any number of other topics. Often the question is raised about who is right and who is wrong. If the football game is to be played this coming weekend, all we can do is offer opinions as to its outcome. The game has not been played yet, so we don't know who will actually win.

The ancient Greek philosophers did a lot of discussing, with part of their conversations concerning the physical world and its composition. There were different opinions about what made up matter. Some felt one thing was true while others believed another set of ideas. Since these scholars did not have laboratories and had not developed the idea of the experiment, they were left to debate. Whoever could offer the best argument was considered right. However, often the best argument had little to do with reality.

One of the on-going debates had to do with sand. The question posed was: into how small of pieces can you divide a grain of sand? The prevailing thought at the time, pushed by Aristotle, was that the grain of sand could be divided indefinitely, that you could always get a smaller particle by dividing a larger one and there was no limit to how small the resulting particle could be.



Figure 4.1.2: Into how small of pieces can you divide a grain of sand? (Public Domain; User:(WT-shared) Shoestring/WTS.Wikivoyage via Wikipedia)

Since Aristotle was such an influential philosopher, very few people disagreed with him. However, there were some philosophers who believed that there was a limit to how small a grain of sand could be divided. One of these philosophers was Democritus (~460-370 B.C.E.), often referred to as the "laughing philosopher" because of his emphasis on cheerfulness. He taught that there were substances called **atoms** and that these atoms made up all material things. The atoms were unchangeable, indestructible, and always existed.



Figure 4.1.3: Democritus. (Public Domain; Antoine Coypel via Wikipedia)

The word "**atom**" comes from the Greek $\alpha\tau o\mu o\sigma$ and means "indivisible". The atomists of the time (Democritus being one of the leading atomists) believed there were two realities that made up the physical world: atoms and void. There were an infinite number of atoms, but different types of atoms had different sizes and shapes. The void was the empty space in which the atoms moved and collided with one another. When these atoms collided with one another, they might repel each other or they might connect in clusters, held together by tiny hooks and barbs on the surfaces of the atoms.

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Aristotle disagreed with Democritus and offered his own idea of the composition of matter. According to Aristotle, everything

was composed of four elements: earth, air, fire, and water. Deomcritus' theory better explained things, but Aristotle was more influential, so his ideas prevailed. It took almost two thousand years before scientists came around to seeing the atom as Democritus did.

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How right was Democritus?

It is very interesting that Democritus had the basic idea of atoms, even though he had no experimental evidence to support his thinking. We now know more about how atoms hold together in "clusters" (compounds), but the basic concept existed over two thousand years ago. We also know that atoms can be further subdivided, but there is still a lower limit to how small we can break up that grain of sand.

Summary

- Greek philosophers debated many ideas.
- Aristotle and others believed that a grain of sand could be divided indefinitely.
- Democritus believed there was a lower limit to the division of a grain of sand.

Review

1. How did the ancient Greek philosophers spend their time?
2. What approach did they not have for studying nature?
3. Who was the most influential philosopher of that time?
4. What was the major contribution Democritus made to the thinking of his day?
5. List characteristics of atoms according to Democritus.

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4.2: Conservation of Mass



Figure 4.2.1 (CC BY 2.0; Douglas Heriot via Flickr)

Have you ever lost a screw?

The following situation happens all too often. You have taken apart a piece of equipment to clean it up. When you put the equipment back together, somehow you have an extra screw or two. Or you find out that a screw is missing that was a part of the original equipment. In either case, you know something is wrong. You expect to end up with the same amount of material that you started with, not with more or less than what you had originally.

Law of Conservation of Mass

By the late 1700s, chemists accepted the definition of an element as a substance that cannot be broken down into a simpler substance by ordinary chemical means. It was also clear that elements combine with one another to form more complex substances called compounds. The chemical and physical properties of these compounds are different than the properties of the elements from which they were formed. There were questions about the details of these processes.

In the 1790s, a greater emphasis began to be placed on the quantitative analysis of chemical reactions. Accurate and reproducible measurements of the masses of reacting elements and the compounds they form led to the formulation of several basic **laws**. One of these is called the law of **conservation of mass**, which states that during a chemical reaction, the total mass of the **products** must be equal to the total mass of the **reactants**. In other words, mass cannot be created or destroyed during a chemical reaction, but is always conserved.

As an example, consider the reaction between silver nitrate and sodium chloride. These two compounds will dissolve in water to form silver chloride and sodium nitrate. The silver chloride does not dissolve in water, so it forms a solid that we can filter off. When we evaporate the water, we can recover the sodium nitrate formed. If we react 58.5 grams of sodium chloride with 169.9 grams of silver nitrate, we start with 228.4 grams of materials. After the reaction is complete and the materials separated, we find that we have formed 143.4 grams of silver chloride and 85.0 grams of sodium nitrate, giving us a total mass of 228.4 grams for the products. So, the total mass of reactants equals the total mass of products, a proof of the law of conservation of mass.

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Summary

- The law of conservation of mass states that, during a chemical reaction, the total mass of the products must be equal to the total mass of the reactants.

Review

1. The law of conservation of mass states that, during a chemical reaction, the total _____ of the products must be equal to the total _____ of the reactants.
2. Describe an example of the law of conservation of mass.

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4.3: Law of Definite Proportions



Figure 4.3.1 (Public Domain; Staff Sgt. Christopher Flurry, U.S. Marine Corps via Commons Wikimedia, File:USMC-110609-M-FK922-004.jpg [commons.wikimedia.org])

We use electricity for many purposes—from cooking, to powering our televisions, to charging our cell phones. Wherever we travel in the United States, we want electricity to be available. What we also want (although we usually don't think about it) is for the electricity supply to be the same wherever we go. We want the same voltage (110 volts for the U.S.) to come from the outlet to whatever we plug in. If the voltage is less, the system will not work. If it is more, the equipment will be damaged. We want a definite amount of voltage—no more and no less.

Law of Definite Proportions

The discovery that mass was always conserved in chemical reactions was soon followed by the **law of definite proportions**, which states that a given chemical compound always contains the same elements in the exact same proportions by mass. As an example, any sample of pure water contains 11.19% hydrogen and 88.81% oxygen by mass. It does not matter where the sample of water came from or how it was prepared. Its composition, like that of every other compound, is fixed.



Figure 4.3.2: Water. (CC BY 2.0; Jenn Durfey via Flickr Photos, Water Faucet [www.flickr.com])

Another example is carbon dioxide. This gas is produced from a variety of reactions, often by the burning of materials. The structure of the gas consists of one atom of carbon and two atoms of oxygen. Carbon dioxide production is of interest in many areas, from the amount we breathe to the amount of the gas produced by burning wood or fossil fuels. By knowing the exact composition of carbon dioxide, we can make predictions as to the effects of different chemical processes.



Figure 4.3.3: Carbon dioxide is produced during the burning process. (Public Domain; Patsy Lynch, FEMA via Commons Wikimedia, FEMA Photograph by Patsy Lynch taken on 04-17-2006 in Missouri [commons.wikimedia.org])

Summary

- The law of definite proportions states that a given chemical compound always contains the same elements in the exact same proportions by mass.

Review

1. State the law of definite proportions.
2. Will the composition of water molecules vary depending on its source?
3. Why is this law important?

Explore More

Use the resource below to answer the questions that follow.

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1. How many hydrogen atoms are there in a molecule of water?
2. How many oxygen atoms are there in a molecule of water?
3. What is the mass ratio of hydrogen to oxygen in a molecule of water?
4. Will the mass ratio of hydrogen to oxygen change depending on the size of the sample? For example if the sample size was 10 L or 10,000 L.

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4.4: Law of Multiple Proportions

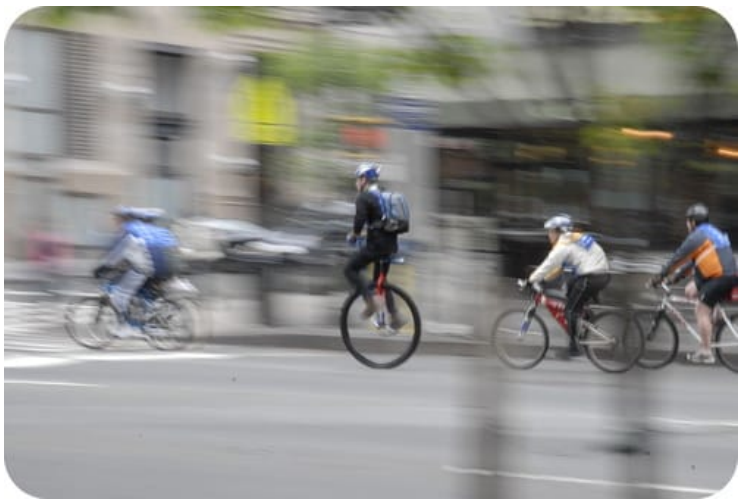


Figure 4.4.1 (CC by 2.0; Flickr: Sutibu via Commons WIkimedia, Five Boro Bike Tour Unicycle [commons.wikimedia.org])

What are the similarities and differences between a unicycle and a bicycle?

Just from the words themselves, the astute Latin-speaking scholar can tell that, whatever it is made of, the unicycle has one of them (*uni* = "one") and the bicycle has two (*bi* = "two"). From pictures, we get additional information that helps us tell the two apart. The unicycle has one wheel and the bicycle has two. In particular, they are made up of the same materials, and the only significant difference is the number of wheels on the two vehicles. Now—how many wheels are on a tricycle?

Law of Multiple Proportions

Once the idea that elements combined in definite proportions to form compounds was established, experiments also began to demonstrate that the same pairs of certain elements could combine to form more than one compound. Consider the elements carbon and oxygen. Combined in one way, they form the familiar compound carbon dioxide. In every sample of carbon dioxide, there are $32.0\ \text{g}$ of oxygen present for every $12.0\ \text{g}$ of carbon. By dividing 32.0 by 12.0, this simplifies to a mass ratio of oxygen to carbon of 2.66 to 1. There is another compound that forms from the combination of carbon and oxygen called carbon monoxide. Every sample of carbon monoxide contains $16.0\ \text{g}$ of oxygen for every $12.0\ \text{g}$ of carbon. This is a mass ratio of oxygen to carbon of 1.33 to 1. In the carbon dioxide, there is exactly twice as much oxygen present as there is in the carbon monoxide. This example illustrates the **law of multiple proportions**: whenever the same two elements form more than one compound, the different masses of one element that combine with the same mass of the other element are in the ratio of small whole numbers.

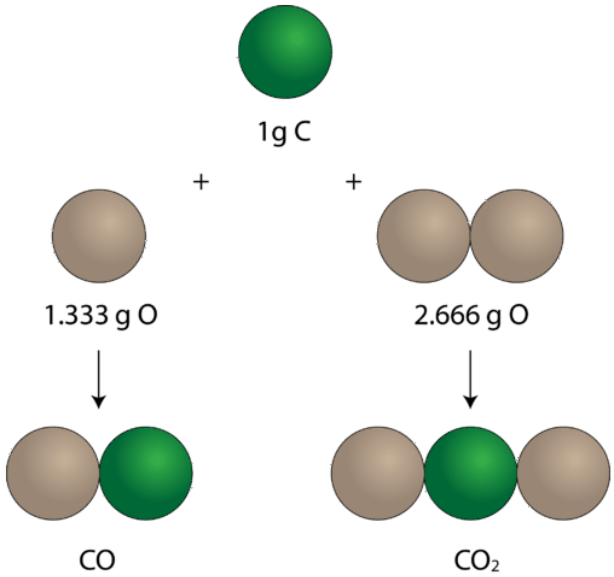


Figure 4.4.2: Carbon can form two different compounds with oxygen. (CC BY-NC-SA 3.0; Christopher Auyeung via CK-12 Foundation)

In carbon monoxide, on the left, there is $1.333\ \text{g}$ of oxygen for every $1\ \text{g}$ of carbon. In carbon dioxide, on the right, there is $2.666\ \text{g}$ of oxygen for every gram of carbon. So the ratio of oxygen in the two compounds is 1:2, a small whole number ratio.

The difference between carbon monoxide and carbon dioxide is significant. Carbon monoxide is a deadly gas, formed from the incomplete combustion of some carbon-containing materials (such as wood and gasoline). This compound will attach to hemoglobin in the red blood cells and block the binding of oxygen to those cells. If oxygen does not bind, it cannot be carried to the cells of the body where it is needed, and death can occur. Carbon dioxide, on the other hand, is not toxic like carbon monoxide is. However, it can displace oxygen in systems since it is heavier. Carbon dioxide fire extinguishers cut off the flow of oxygen in a fire, putting out the fire.

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Summary

- The law of multiple proportions states that whenever the same two elements form more than one compound, the different masses of one element that combine with the same mass of the other element are in the ratio of small whole numbers.

Review

1. State the law of multiple proportions.
2. In carbon dioxide (CO2), how many grams of oxygen (O) would there be if there are 24 grams of carbon (C)?
3. How many grams of carbon (C) would be present in carbon monoxide (CO) that contains 2.666 grams of oxygen (O)?

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4.5: Mass Ratio Calculation

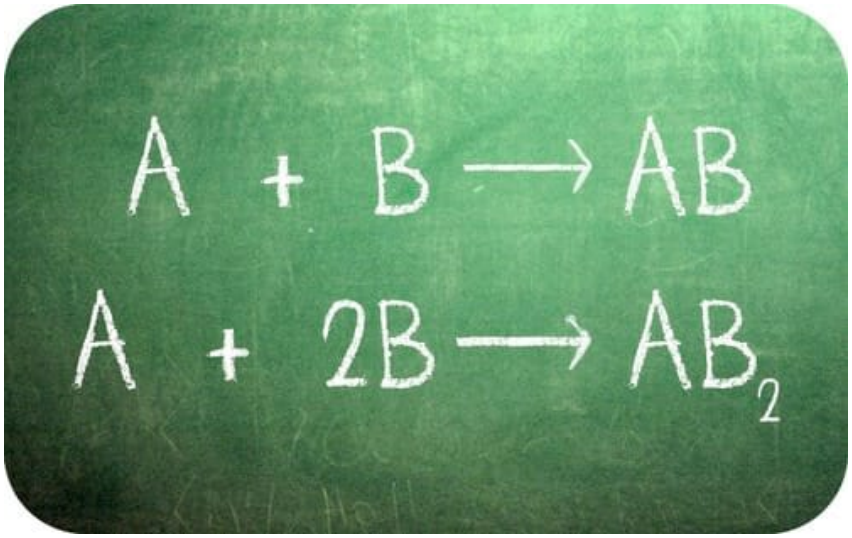


Figure 4.5.1 (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

What are the similarities and differences between these two equations?

One fundamental law of chemistry deals with the fact that we cannot create or destroy matter (using chemical means). When a reaction is run, the number of atoms of each specific type must be the same on both sides of the equation. For some materials, it turns out that one element can combine with a second element in more than one ratio. Carrying out mass ratio calculations helped establish the law of multiple proportions.

Mass Ratio Calculations

Example 4.5.1

Copper reacts with chlorine to form two compounds. Compound A consists of $4.08\ \text{g}$ of copper for every $2.28\ \text{g}$ of chlorine. Compound B consists of $7.53\ \text{g}$ of copper for every $8.40\ \text{g}$ of chlorine. What is the lowest whole number mass ratio of copper that combines with a given mass of chlorine?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Compound A $= 4.08\ \text{g Cu}$ and $2.28\ \text{g Cl}$
- Compound B $= 7.53\ \text{g Cu}$ and $8.40\ \text{g Cl}$

Apply the law of multiple proportions to the two compounds. For each compound, find the grams of copper that combine with $1.00\ \text{g}$ of chlorine by dividing the mass of copper by the mass of chlorine. Then find the ratio of the masses of copper in the two compounds by dividing the larger value by the smaller value.

Step 2: Calculate

$$\text{Compound A}\ \frac{4.08\ \text{g Cu}}{2.28\ \text{g Cl}} = \frac{1.79\ \text{g Cu}}{1.00\ \text{g Cl}}$$

$$\text{Compound B}\ \frac{7.53\ \text{g Cu}}{8.40\ \text{g Cl}} = \frac{0.896\ \text{g Cu}}{1.00\ \text{g Cl}}$$

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Compare the masses of copper per gram of chlorine in the two samples.

$$\frac{1.79\text{ g Cu (in compound A)}}{0.896\text{ g Cu (in compound B)}} = \frac{2.00}{1} = 2:1$$

The mass ratio of copper per gram of chlorine in the two compounds is 2:1.

Step 3: Think about your result.

The ratio is a small whole-number ratio. For a given mass of chlorine, compound A contains twice the mass of copper as does compound B.



Figure 4.5.2: $CuCl_2$. (Public Domain; User:Chemicalinterest/Wikimedia Commons via Commons Wikimedia, Copper Chloride Crystals [commons.wikimedia.org])

Summary

- The mass ratio gives the mass of an element that is found in combination with another element.

Review

1. What does the mass ratio between two elements tell us?
2. If we compare the mass ratio of elements in one compound to that in an second compound what can we learn? Give an example from the lesson above.
3. In compound A, there is 6.3 g of hydrogen and 18.7 g of carbon, while in compound B there is 6.9 g of hydrogen and 41.0 g of carbon, what is the carbon to hydrogen mass ratio in each compound and how do these ratios compare?
4. What are lowest ratios of hydrogen and carbon in compounds A and B? Predict their formulas

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4.6: Dalton's Atomic Theory



Figure 4.6.1 (Public Domain; Luca della Robbia, photographed by Jastrow via Wikipedia)

"*Pick a little, talk a little, pick a little, talk a little,*

Cheep cheep cheep, talk a lot, pick a little more"

These lyrics from the musical "Music Man" sum up the way science was done for centuries. The lyrics refer to a group of gossiping ladies, but the outcome was the same. The Greek and Roman philosophers debated, discussed, and sometimes even attacked one another. But the mode of discovery was talk. There was no experimentation—the idea had not yet been thought of. So science did not develop very far, and there was no reliable way to establish what was true and what was false.

John Dalton

While it must be assumed that many more scientists, philosophers, and others studied the composition of matter after Democritus, a major leap forward in our understanding of the composition of matter took place in the 1800s with the work of the British scientist John Dalton. He started teaching school at age twelve, and was primarily known as a teacher. In his twenties, he moved to the growing city of Manchester, where he was able to pursue some scientific studies. His work in several areas of science brought him a number of honors. When he died, over 40,000 people in Manchester marched at his funeral.

Dalton studied the weights of various elements and compounds. He noticed that matter always combined in fixed ratios based on weight, or volume in the case of gases. Chemical compounds always contain the same proportion of elements by mass, regardless of amount, which provided further support for Proust's law of definite proportions. Dalton also observed that there could be more than one combination of two elements.



Figure 4.6.2: John Dalton. (Public Domain; Henry Roscoe (author), William Henry Worthington (engraver), and Joseph Allen (painter) via Wikipedia)

Dalton's Atomic Theory (1804)

From his own experiments and observations, as well as the work of his peers, Dalton proposed a new theory of the **atom**. This later became known as Dalton's **atomic theory**. The general tenets of this theory are as follows:

- 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.

Dalton's atomic theory has been largely accepted by the scientific community, with the exception of three changes. We now know that (1) an atom can be further subdivided, (2) all atoms of an element are not identical in mass, and (3) using nuclear fission and fusion techniques, we can create or destroy atoms by changing them into other atoms.

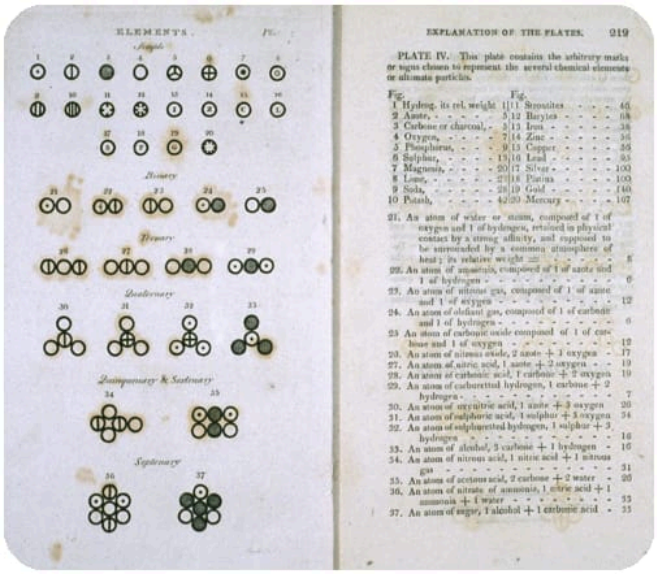


Figure 4.6.3: Dalton's symbols. (Public Domain; John Dalton via Wikipedia)

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Summary

- Dalton proposed his atomic theory in 1804.
- The general tenets of this theory are:
  - All matter is composed of extremely small particles called atoms
  - Atoms cannot be subdivided, created, or destroyed.
  - 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 of different elements can combine in whole number ratios to form chemical compounds;
  - In chemical reaction, atoms are rearranged, combined, or separated.

Review

1. Did Dalton believe that atoms could be created or destroyed?
2. According to Dalton's theory, all atoms of the same _____ are identical in size, mass, and other properties.
3. What parts of the theory are not considered valid any more?

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4.7: Cathode Ray Tube

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Figure 4.7.1 (Public Domain; User:Mahlum/Wikimedia Commons via Commons Wikimedia, Radionette TV Set [commons.wikimedia.org])

How old do you think this TV is?

The TV set seen above is becoming harder and harder to find these days. The main reason is because they are older and based on outdated technology. The new TV sets are flat screen technology that take up less space and give better picture quality, especially with the advent of high-definition broadcasting. The technology used in the older TV sets used cathode ray tubes. A beam of electrons was sprayed to a picture tube which was treated to react with the electrons to produce an image. Similar CRT devices were used in computer monitors, now also replaced by flat screen monitors.

Discovery of the Electron

The first discovery of a subatomic particle was a result of experiments into the nature of the relationship between electricity and matter.

Cathode Rays

The first **cathode ray tube** prototype was developed by Heinrich Geissler, a German glassblower and physicist. He used a mercury pump to create a vacuum in a tube. Geissler explored a number of techniques to remove air from the tube and to prevent leaks, as well as ways to get good connections of the wires in the tubes.

In 1878, Sir William **Crookes**, a British scientist, displayed the first cathode rays using a modification of the Geissler apparatus. His major contribution to construction of the tube was to develop ways to evacuate almost all the air from the tube. Crookes also carried out many experiments using more reliable equipment to confirm earlier findings about the properties of cathode rays. He made two discoveries which supported the hypothesis that the cathode ray consisted of a stream of particles:

- When an object was placed between the cathode and the opposite end of the tube, it cast a shadow on the glass. The shadow caused by the object indicates that particles were being blocked on their way from the cathode to the anode.

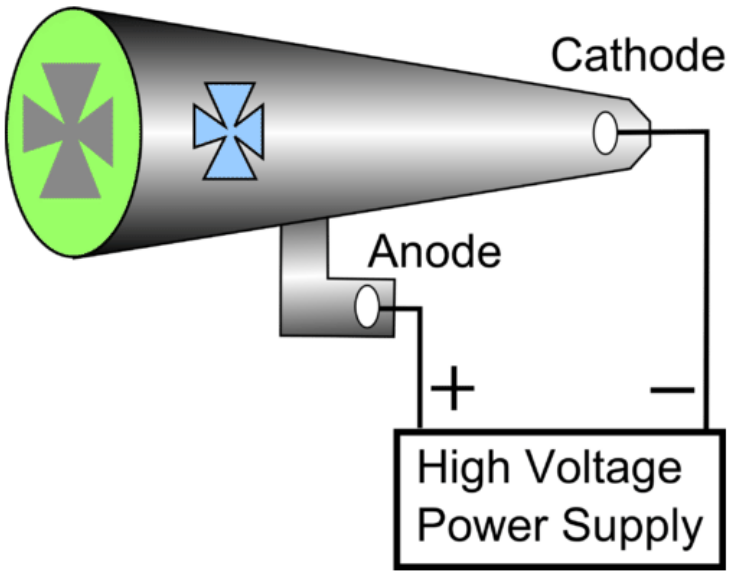


Figure 4.7.2 (Public Domain; User:Chetvornoa/Wikipedia and User:Drondent/Wikipedia via Commons Wikimedia, Crookes tube2 Diagram [commons.wikimedia.org])

- A cathode ray tube was constructed with a small metal rail between the two electrodes. Attached to the rail was a paddle wheel capable of rotating along the rail. Upon starting up the cathode ray tube, the wheel rotated from the cathode towards the anode. Notice that the cathode and anode are positioned so that the rays will strike the top of the paddle wheel. Crookes concluded that the cathode ray was made of particles which must have mass.

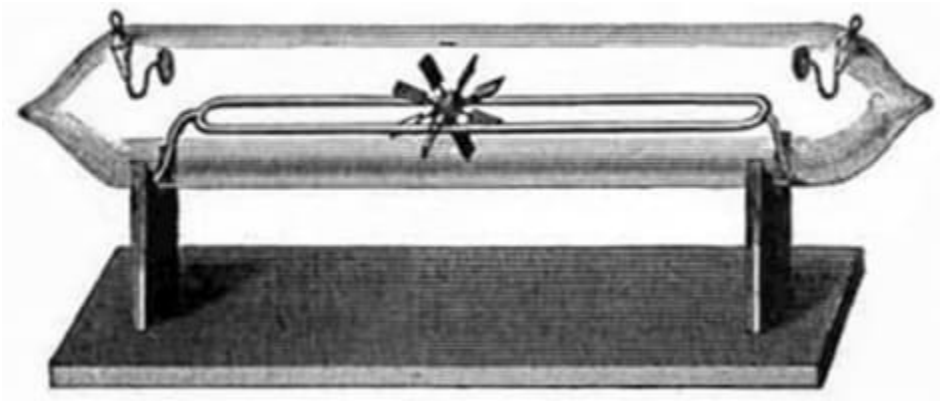


Figure 4.7.3: The cathode ray tube was first invented by Sir William Crookes. (Public Domain; William Crookes via Commons Wikimedia, Crookes Paddle Wheel tube [commons.wikimedia.org])

Further Research with the Crookes Tube

Crookes' work opened the door to a number of important discoveries. Other scientists were able to demonstrate that the "cathode ray" was actually a stream of **electrons**. In 1897, Karl Ferdinand Braun developed the first oscilloscope, using a cathode ray tube to see an electrical pulse as it passed through the instrument. The invention of television would not have been possible without the cathode ray tube. Work with a modified system led to the discovery of x-rays in 1895 by the German physicist Wilhelm Roentgen. This simple device has led to major advances in science and technology.

Summary

- The cathode ray tube was first invented by Sir William Crookes.
- Experiments showed that the rays had mass.

Review

- Who developed the first cathode ray tube?
- What improvement did Crookes make to the cathode ray tube?
- How did Crookes show there were particles being emitted?
- What did Karl Ferdinand Braun invent?
- What did Wilhelm Roentgen invent?

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4.8: Electrons

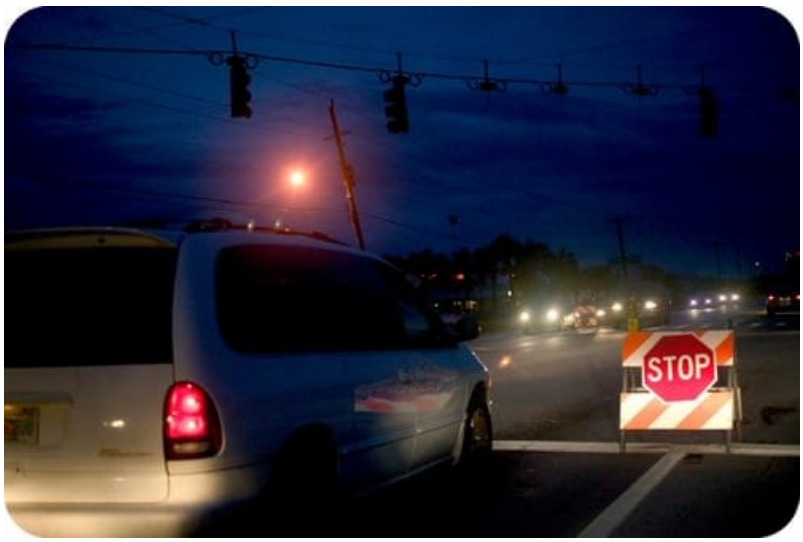


Figure 4.8.1 (Public Domain; Andrea Booher, FEMA via Wikipedia)

What causes a power outage?

In a power outage, all your electrical equipment suddenly stops working. The radio was on just a minute ago and now it is silent. What happened? Somewhere between a power generator and your electrical device was an interruption. Power stopped flowing through the wires and into your radio. That “power” turns out to be electrons that move through the wires and cause an electrical current to flow.

Is There Anything Inside an Atom?

As the nineteenth century began to draw to a close, the concept of atoms was well-established. We could determine the mass of different atoms and had some good ideas about the atomic composition of many compounds. Dalton’s atomic theory held that atoms were indivisible, so scientists did not ask questions about what was inside the atom – it was solid and could not be broken down further. But then things began to change.

The Electron

In 1897, English physicist J.J. Thomson (1856-1940) experimented with a device called a cathode ray tube, in which an electric current was passed through gases at low pressure. A cathode ray tube consists of a sealed glass tube fitted at both ends with metal disks called electrodes. The electrodes are then connected to a source of electricity. One electrode, called the anode, becomes positively charged while the other electrode, called the cathode, becomes negatively charged. A glowing beam (the cathode ray) travels from the cathode to the anode.

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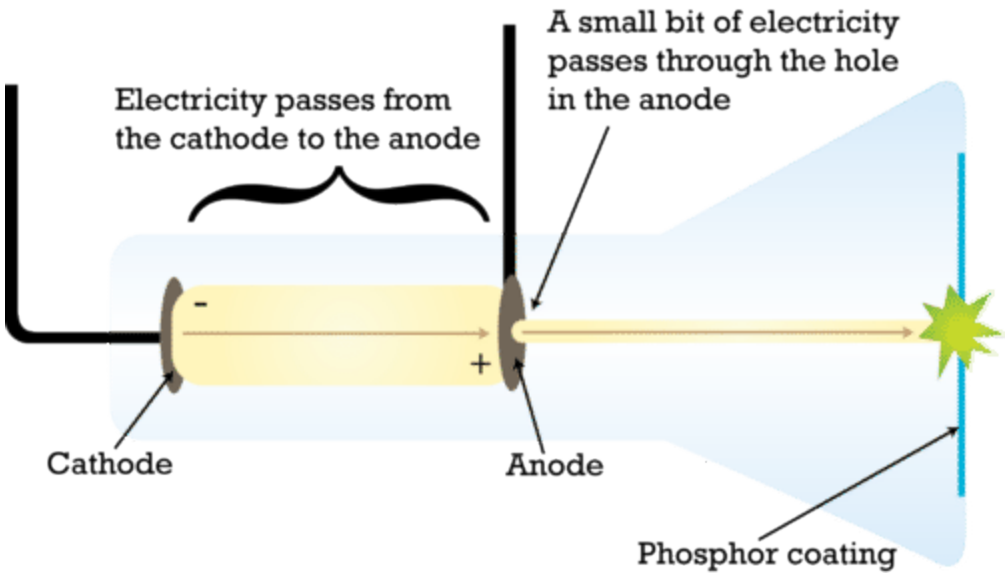


Figure 4.8.2 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Earlier investigations by Sir William Crookes and others had been carried out to determine the nature of the cathode ray. Thomson modified and extended these experiments in an effort to learn about these mysterious rays. He discovered two things, which supported the hypothesis that the cathode ray consisted of a stream of particles.

- When an object was placed between the cathode and the opposite end of the tube, it cast a shadow on the glass.
- A cathode ray tube was constructed with a small metal rail between the two electrodes. Attached to the rail was a paddle wheel capable of rotating along the rail. Upon starting up the cathode ray tube, the wheel rotated from the cathode towards the anode. This proved that the cathode ray was made of particles which must have mass. Crooke had first observed this phenomenon and attributed it to pressure by these particles on the wheel. Thomson correctly surmised that these particles were producing heat, which caused the wheel to turn.

In order to determine if the cathode ray consisted of charged particles, Thomson used magnets and charged plates to deflect the cathode ray. He observed that cathode rays were deflected by a **magnetic field** in the same manner as a wire carrying an electric current, which was known to be negatively charged. In addition, the cathode ray was deflected away from a negatively charged metal plate and towards a positively charged plate.

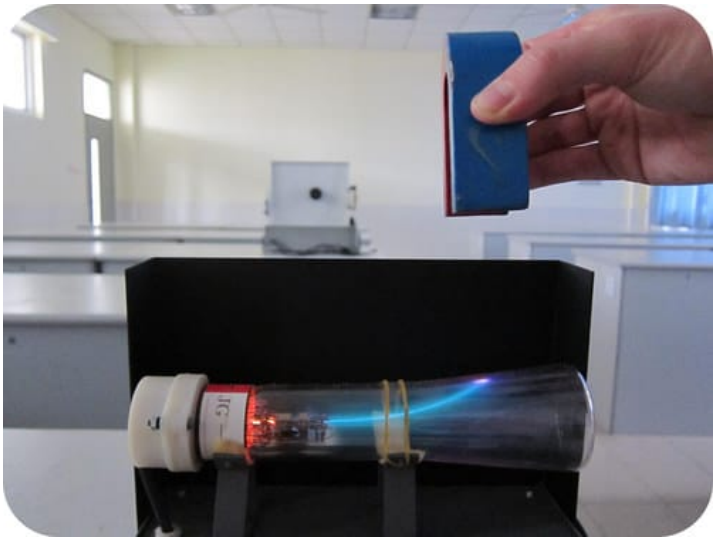


Figure 4.8.3 (CC BY 2.0; Micah Sittig via Flickr)

Thomson knew that opposite charges attract one another, while like charges repel one another. Together, the results of the cathode ray tube experiments showed that cathode rays are actually streams of tiny negatively charged particles moving at very

high speeds. While Thomson originally called these particles corpuscles, they were later named electrons.

Thomson conducted further experiments, which allowed him to calculate the charge-to-mass ratio $\frac{e}{m_e}$ of the electron. In units of coulombs to grams, this value is 1.8×10^8 Coulombs/gram. He found that this value was a constant and did not depend on the gas used in the cathode ray tube or on the metal used as the electrodes. He concluded that electrons were negatively charged subatomic particles present in atoms of all elements.

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Summary

- Cathode rays are deflected by a magnetic field.
- The rays are deflected away from a negatively charged electrical field and toward a positively charge field.
- The charge/mass ratio for the electron is 1.8×10^8 Coulombs/gram.

Review

1. What subatomic particle creates electric power, and how does it do it?
2. Whose work did Thomson repeat and revise?
3. What experiment did Thomson perform that showed cathode rays to be particles?
4. How did he show that these particles had a charge on them?
5. Did the cathode ray have positive or negative charge?

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4.9: Oil Drop Experiment

"How tall are you? How much do you weigh?" Questions like these are easy to answer because we have tools to make the measurements. A yard stick or tape measure will suffice to measure height. You can stand on a bathroom scale and determine your weight.

It is a very different matter to measure properties of objects that we cannot see with the naked eye. If we want to measure the size of a germ, we have to use a microscope. To learn the size of a single molecule, we have to use even more sophisticated instruments. So how would we measure something even smaller than a molecule, even smaller than an atom?

Charge and Mass of the Electron

The man who measured properties of the **electron** was Robert Millikan (1868-1953). He taught himself physics while a student at Oberlin College' since there was nobody on the faculty to instruct him in this field. Millikan completed postgraduate research training in the U.S. and in Germany. His studies on the properties of the electron proved to be of great value in many areas of physics and chemistry.



Figure 4.9.1: Robert Millikan. (Public Domain; Robert Millikan via Wikipedia)

Oil Drop Experiment

Millikan carried out a series of experiments between 1908 and 1917 that allowed him to determine the charge of a single electron, famously known as the oil drop experiment.

Millikan sprayed tiny drops of oil into a chamber. In his first experiment, he simply measured how fast the drops fell under the force of gravity. He could then calculate the mass of the individual drops. Then, he sprayed oil drops and applied an electrical charge to them by shining x-rays up through the bottom of the apparatus. The x-rays ionized the air, causing electrons to attach to the oil drops. The oil drops picked up static charge and were suspended between two charged plates. Millikan was able to observe the motion of the oil drops with a microscope and found that the drops lined up in a specific way between the plates, based on the number of electric charges they had acquired.

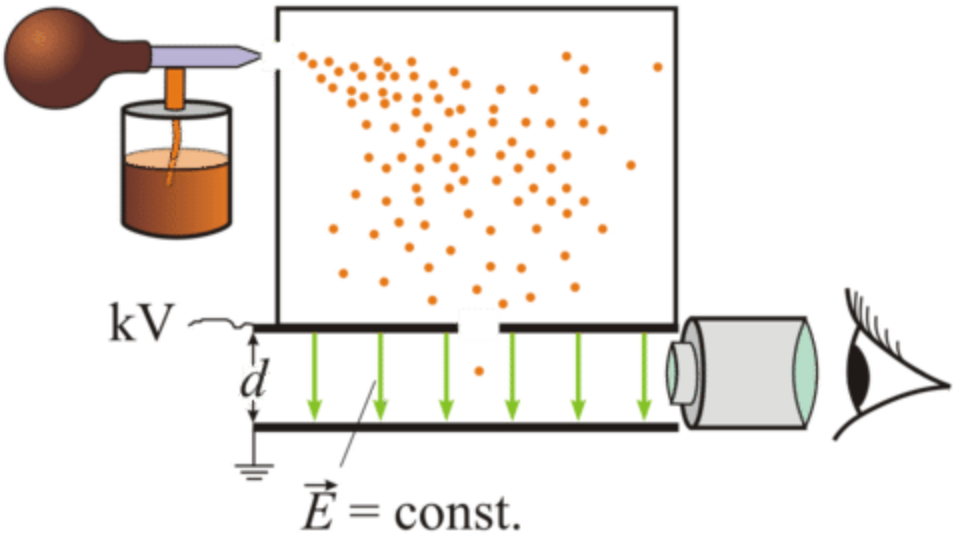


Figure 4.9.2: Oil Drop experiment. (Public Domain; User:Mpfiz/Wikimedia Commons via Wikipedia)

Millikan used the information to calculate the charge of an electron. He determined the charge to be 1.5924×10^{-19} C, where C stands for **coulomb**, which is one ampere*second. Today, the accepted value of the charge of an electron is $1.602176487 \times 10^{-19}$ C. Millikan's experimental value proved very accurate; it is within 1% of the currently accepted value. Millikan later used the information from his oil drop experiment to calculate the mass of an electron. The accepted value today is $9.10938215 \times 10^{-31}$ kg. The incredibly small mass of the electron was found to be approximately 1/1840 the mass of a hydrogen atom. Therefore, Scientists realized that atoms must contain another particle that carries a positive charge and is far more massive than the electron.

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Summary

- The oil drop experiment allowed Millikan to determine the charge on the electron.
- He later used this data to determine the mass of the electron.

Review

- How did Millikan learn physics in college?
- What did Millikan use to pick up static charge?
- Where did the oil drops go to be measured?

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4.10: Protons



Figure 4.10.1 (Public Domain; Adrian Pingstone (User:Arpingstone/Wikimedia Commons) via Wikipedia)

Can you name this car?

Describing what we can see is a fairly easy matter. If we are asked to describe the sports car illustrated below, we could all quickly come up with a fairly accurate description. A person knowledgeable about cars would include more details, but everyone would have the basic information in their description.

What makes the description easy to come up with? We can see it, we have a common language to describe it (size, color, construction), and we have a basic idea of what it is (a car, not a house or a tree). Scientists have far more difficulty in describing things they cannot see. There is no way to look directly at an atom and see its detailed structure. When a discovery is first made, there is often no language to use to tell others exactly what it is. This was the problem in talking about the atom and its structure.

Putting the Puzzle Pieces Together

Research builds upon itself – one piece connects to another. Sometimes the puzzle doesn't seem to make sense because some of the pieces are missing at the moment. Each finding gives a clearer picture of the whole and also raises new questions. The detective work that led to the discovery of the proton was built upon finding pieces to the puzzle and putting them together in the right way.

The electron was discovered using a cathode ray tube. An electric current was passed from the cathode (the negative pole) to the anode (positive pole). Several experiments showed that particles were emitted at the cathode and that these particles had a negative charge. These experiments demonstrated the presence of electrons.

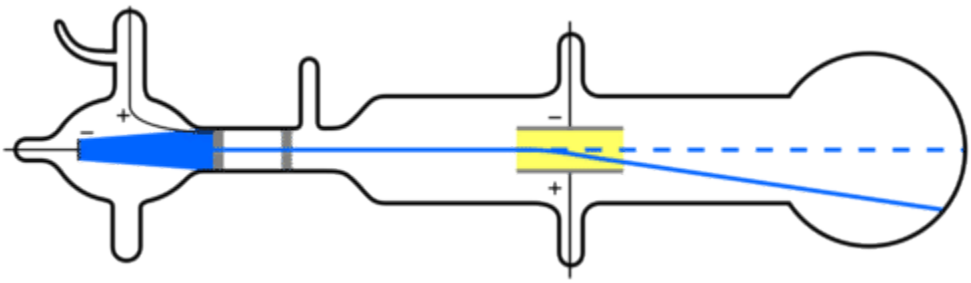


Figure 4.10.2: JJ Thomson's experiment with cathode rays. (Public Domain; User:Kurzon/Wikimedia Commons via Wikipedia)

If cathode rays are electrons that are given off by the metal atoms of the cathode, then what remains of the atoms that have

lost those electrons? We know several basic things about electrical charges. They are carried by particles of matter. Millikan's experiment showed that they exist as whole-number multiples of a single basic unit. Atoms have no overall electrical charge, meaning that each and every atom contains an exactly equal number of positively and negatively charged particles. A hydrogen atom is the simplest kind of atom with only one electron. When that electron is removed, a positively charged particle should remain.

Discovery of the Proton

In 1886, Eugene Goldstein (1850-1930) discovered evidence for the existence of this positively charged particle. Using a cathode ray tube with holes in the cathode, he noticed that there were rays traveling in the opposite direction from the cathode rays. He called these **canal rays** and showed that they were composed of positively charged particles. The **proton** is the positively charged subatomic particle present in all atoms. The mass of the proton is about 1840 times the mass of the electron.

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Summary

- When an electron is removed from a hydrogen atom, a proton remains.
- Goldstein observed rays traveling in the opposite direction of the cathode rays in a cathode ray tube.
- He demonstrated that these rays were positive particles and called the canal rays.

Review

1. Why is it easy to describe things we can see?
2. Why did researchers believe that the particle left after electrons were emitted as cathode rays had to be positive?
3. Atoms, which are always neutral in electric charge, contain electrons as well as protons and neutrons. An electron has an electrical charge of -1. If an atom has three electrons, infer how many protons it has.
4. How many electrons does it take to weight the same as one proton?

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4.11: Neutrons



Figure 4.11.1 (Public Domain; Sidney Paget via Wikipedia)

Was Sherlock Holmes real or the product of someone's imagination?

The most famous detective in history and literature never existed. Sherlock Holmes was the creation of the British author Sir Arthur Conan Doyle. This mythical person had capabilities far beyond those of mere mortals. Holmes was capable of spotting the tiniest clue, the smallest piece of evidence to solve the crime. He could link all sorts of seemingly irrelevant data into a coherent whole to clear up whatever mystery he was dealing with.

The Quest for the Neutron

Clues are generally considered to involve the presence of something – a footprint, a piece of fabric, a bloodstain, something tangible that we can measure directly. The discoveries of the electron and the proton were accomplished with the help of those kinds of clues. Cathode ray tube experiments showed both the negatively charged electrons emitted by the cathode and the positively charged proton (also emitted by the cathode). The **neutron** was initially found not by a direct observation, but by noting what was not found.

Research had shown the properties of the electron and the proton. Scientists learned that approximately 1,837 electrons weighed the same as one proton. There was evidence to suggest that electrons went around the heavy nucleus composed of protons. Charge was balanced with equal numbers of electrons and protons which made up an electrically neutral atom. But there was a problem with this model – the atomic number (number of protons) did not match the atomic weight. In fact, the atomic number was usually about half the atomic weight. This indicated that something else must be present. That something must weigh about the same as a proton, but could not have a charge – this new particle had to be electrically neutral.

In 1920, Ernest Rutherford tried to explain this phenomenon through the presence of another particle in the nucleus. He proposed that the "extra" particles were protons and electrons that had combined into a new particle in the nucleus (this did not turn out to be the case). These new particles would have a mass very similar to a proton, but would be electrically neutral since the positive charge of the proton and the negative charge of the electron would cancel each other out.

In 1930, German researchers bombarded the element beryllium with alpha particles (helium nuclei containing two protons and two neutrons with a charge of +2). The particles produced in this process had strong penetrating power, which suggested they were fairly large. In addition, they were not affected by a magnetic field, so they were electrically neutral. The French husband-wife research team of Frederic and Irene Joliot-Curie used these new "rays" to bombard paraffin, which was rich in protons. The unknown particles produced a large emission of protons from the paraffin.

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The English physicist James Chadwick (1891-1974) repeated these experiments and studied the energy of these particles. By measuring velocities, he was able to show that the new particle has essentially the same mass as a proton. So we now have a third subatomic particle with a mass equal to that of a proton, but with no charge. This particle is called the neutron. Chadwick won the Nobel Prize in Physics in 1935 for his research.

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Neutron Applications

Neutrons can be used in a variety of ways. One important use is in nuclear **fission** to produce new isotopes. A neutron will collide with a large atom (such as uranium) and cause it to split into smaller atoms, such as in figure below.

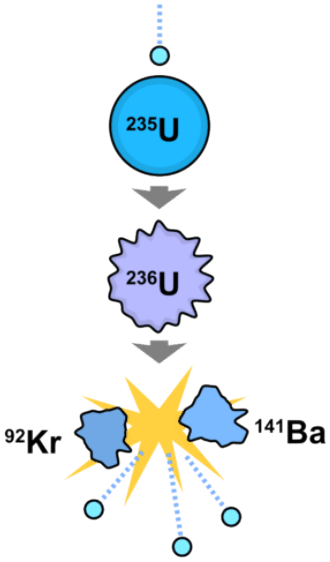


Figure 4.11.2: A neutron collides with a large atom, splitting it into smaller atoms. (Public Domain; User:Fastfission/Wikimedia Commons via Wikipedia)

Nuclear reactors utilize **chain reactions** involving neutrons to heat water which drive turbines for the generation of electricity. When a neutron collides with a large atom, the atom splits with the release of more neutrons and also a large amount of energy. The energy converts water to steam for the turbine, while the neutrons serve to continue the chain reaction (see figure below).

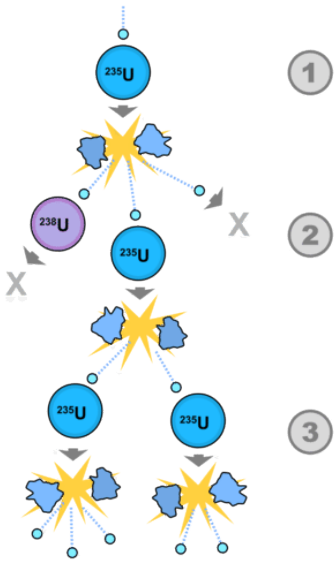


Figure 4.11.3: How nuclear fission produces new isotopes. (Public Domain; User: Fastfission/Wikimedia Commons via Wikipedia)

Summary

- Rutherford proposed that "extra" particles in nucleus were combinations of protons and electrons.
- Bombardment of beryllium with alpha particles produced large, neutral particles.
- Chadwick determined the mass of the neutron.
- Nuclear fission produces new elements.
- Nuclear reactors use chain reactions to produce heat.

Review

1. How did Rutherford try to explain the differences between the number of protons in the nucleus and the atomic weight?
2. What did German researchers find when they bombarded beryllium with alpha particles?
3. What did Chadwick determine about these new particle (observed by the German scientist and the Curies)?

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4.12: Thomson's Atomic Model



Figure 4.12.1 (Public Domain; User:Fokker/Wikimedia Commons via Wikipedia)

What is this model airplane composed of?

Millions of children over the years have enjoyed building models - this model airplane is one example of the types of models that can be constructed. Perhaps sixty years ago the models were made of balsa wood, a very light material. Parts would be cut by hand, carefully glued together, and then covered with paper or other fabric. The development of plastics made the construction of model aircraft much simpler in many respects. And, the end-product is more durable and damage-proof.

A model serves a useful purpose – it gives us an idea of what the real thing is like. The model plane seen above has wings, a tail, and an engine just like the real thing. This model also has a propeller, as is the case with most small planes and some smaller passenger planes. However, the model is not the real thing. We certainly cannot fly people or cargo in the model (besides maybe a tiny mouse), but we can get some idea of what a real plane looks like and how it works.

Science uses many models to explain ideas. We model the electron as a very small particle with a negative charge. That gives us a picture, but a very incomplete one. This picture works fine for most chemists, but is inadequate for a physicist. Models give us a start toward understanding structures and processes, but certainly are not a complete representation of the entity we are examining.

Atomic Models

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 similar to a thick soup. The electrons were considered 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 toward the center region of the atom.

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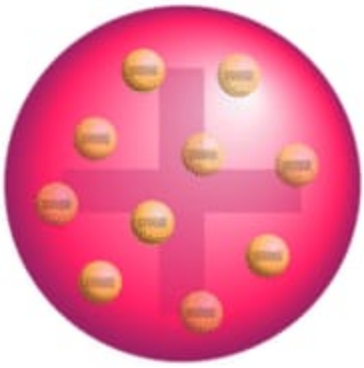


Figure 4.12.1: The "plum pudding" model (Public Domain; User: Fastfission/Wikimedia Commons via Wikipedia)

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.

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Summary

- A model gives an idea of what something looks like, but is not the real thing.
- The "plum pudding" model of the atom consisted of a uniform sphere of positive charge with negative electrons embedded in the sphere.

Review

- What is a model?
- Why are models useful in science?
- In Thomson's model of the atom, where were the electrons?
- What was the positive charge in this model?
- What kept the electrons in the atom?
- Whose model replaced Thomson's?
- What awards did Thomson receive?

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4.13: Rutherford's Atomic Model

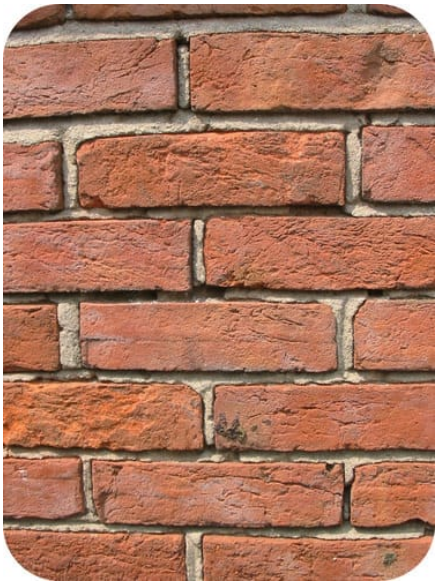


Figure 4.13.1 (Public Domain; Man vyi via https://commons.wikimedia.org/wiki/F...k_wall_old.jpg)

How much space do bricks occupy?

As we look at the world around us, it looks pretty solid. We hit a wall with our hand and the hand stops – it does not (normally) go through the wall. We think of matter as occupying space. But there is a lot of empty space in matter. In fact, most of the matter *is* empty space.

The Gold Foil Experiment

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.

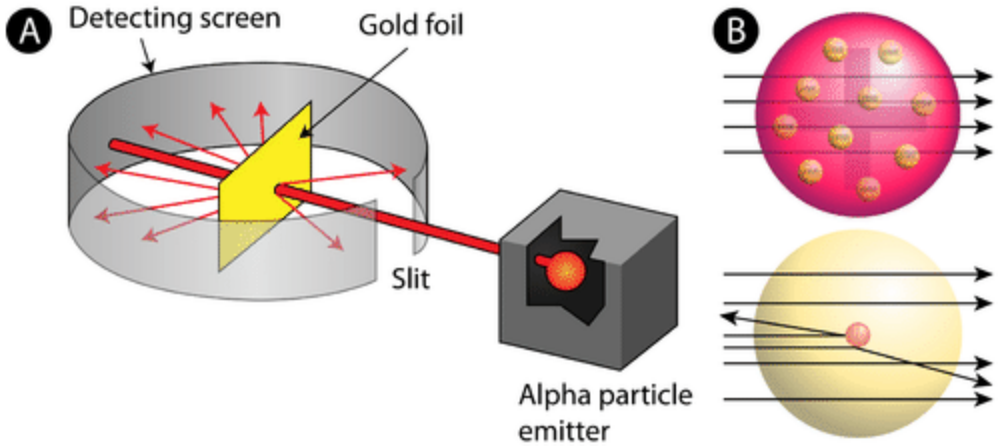


Figure 4.13.2: (A) The experimental setup for Rutherford's gold foil experiment: A radioactive element that emitted alpha particles was directed toward a thin sheet of gold foil that was surrounded by a screen which would allow detection of the deflected particles. (B) According to the plum pudding model (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). ((A; (A) Christopher Auyeung; (B) User:Fastfission/Wikimedia Commons via (A) CK-12 Foundation; (B) http://commons.wikimedia.org/wiki/File:Rutherford_gold_foil_experiment_results.svg) CC BY-NC 3.0; (B) Public Domain)

According to the accepted atomic model—in which an atom's mass and charge are uniformly distributed throughout the atom—the scientists hypothesized 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 not deflected, 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.

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. This proved to be the key to understanding the chemical properties of elements.

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Summary

- Bombardment of gold foil with alpha particles showed that a very small percentage of alpha particles were deflected.
- The nuclear model of the atom consists of a small and dense positively charged interior surrounded by a cloud of electrons.

Review

1. What is an alpha particle?
2. What did Rutherford observe from shooting thousands and thousands of alpha particles at a thin piece of gold foil?
3. How did Rutherford explain the observation that most alpha particles went straight through the gold foil?
4. What did he say about the particles that were deflected?
5. Describe Rutherford's nuclear model.

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4.14: Atomic Nucleus



Figure 4.14.1 (CC by 3.0; Charles Hamm via Wikipedia)

How is science like a jigsaw puzzle?

Many people enjoy putting jigsaw puzzles together. As the different pieces go together, the picture begins to become clearer. When the puzzle is completed, you see that what had been a confused collection of individual components that made little or no sense by themselves fit together to give a clear picture.

Science works the same way as a jigsaw puzzle. Different researchers make individual discoveries that answer a specific question or questions. When enough data are gathered, we get a better understanding of a process or a structure. The experiments of Crooke, Millikan, Rutherford, and many others gave us pieces of the puzzle that is the atomic nucleus.

Different Models of the Nucleus

After the electron and proton were discovered, people began to try to build a picture of the atom. The Thomson model had electrons mixed in with a positive core of some sort that provided electrical neutrality. Rutherford showed that this model was incomplete. His picture of the atom involved a small solid core that alpha particles could zoom past with very few collisions. So the picture of the atom became a little clearer – electrons surrounded a very small core nucleus. The discovery of the neutron helped fill out the picture even more. We now have protons and neutrons in a concentrated center of the atom with electrons surrounding the nucleus.

One problem still existed. We have a number of positively charged protons in the nucleus. Why don't they push each other apart? Physicists postulate a strong nuclear force that acts at very short distances. At these distances the attraction between protons is greater than the force causing the positive charges to push each other away. Neutrons are also involved in this process somehow. So the strong force holds protons together, it holds neutrons together, and it causes protons and neutrons to be attracted to one another.

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Summary

- Rutherford proposed a model of the atomic nucleus which had a solid core.
- Physicists postulate a strong nuclear force holds the protons and neutrons together in the nucleus of the atom.

Review

1. How did Rutherford change our thinking about atomic structure?
2. What is our current picture of the atom?
3. Why is the presence of positively charged protons a problem with current models of the atom?
4. How do we explain why the nucleus does not fall apart?

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4.15: Atomic Number



Figure 4.15.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

What is unique about each one of us?

For the vast majority of people, it is not their name, because it is quite possible for others in the world to have the same name (check it out by doing an internet search for your name and see how many other of "you" there are). It is not your physical description. Eye-witnesses to crime scenes often pick the wrong person when trying to identify the criminal.

There may be some unique identifiers for us. If you have a cell phone in your name, nobody else in the world has that number. Email addresses are different for each of us, which is a good thing since we can email almost anywhere in the world. Our DNA is unique, but getting a DNA analysis is expensive and time-consuming, so we really don't want to have to explore that.

Organizing the Elements

One of the goals of science is to discover the order in the universe and to organize information that reflects that order. As information about the different elements was made known, efforts were made to see if there were patterns in all of the data. An early attempt to organize data was made by Mendeleev, who developed the first periodic table. His data set was based on **atomic weights** and was instrumental in providing clues as to the possible identity of new elements. Once we learned the details of the atomic nucleus, the table was based on the number of protons in the nucleus, called the **atomic number** of the element.

Atomic Number

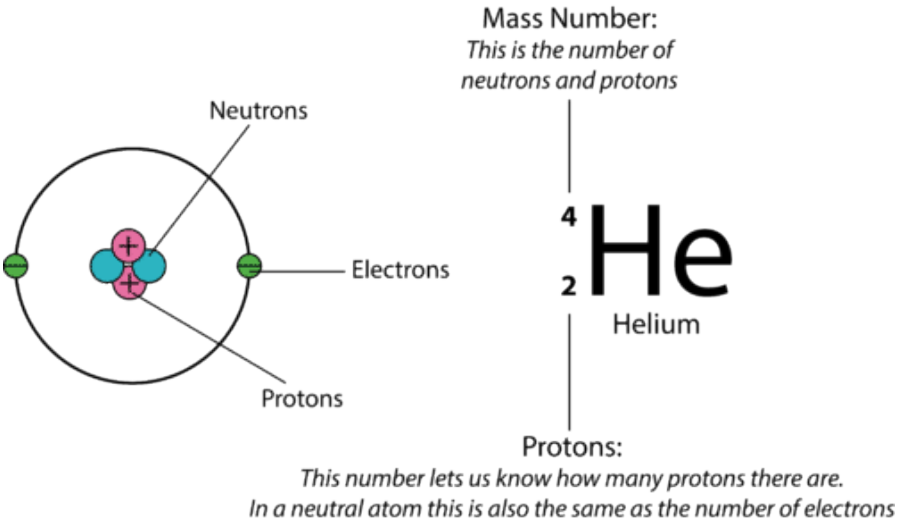


Figure 4.15.2: How can you determine the atomic number of an element. (Public Domain; User:Materialscientist/Wikimedia Commons via Wikipedia)

The atomic number (Z) of an element is the number of protons in the nucleus of each atom of that element. This means that the number of protons is the characteristic which makes each element unique compared to all other elements. Elements are different because of their atomic number. The **periodic table** displays all of the known elements and is arranged in order of increasing atomic number. In this table, an element's atomic number is indicated above the elemental symbol. Hydrogen, at the upper left of the table, has an atomic number of 1. Every hydrogen atom has one proton in its nucleus. Following on the table is helium, whose atoms have two protons in the nucleus. Lithium atoms have three protons, and so forth.

Since atoms are neutral, the number of electrons is equal to the number of protons. Hydrogen atoms all have one electron occupying the space outside of the nucleus. Manganese (atomic number 25) would have twenty-five protons and twenty-five electrons.

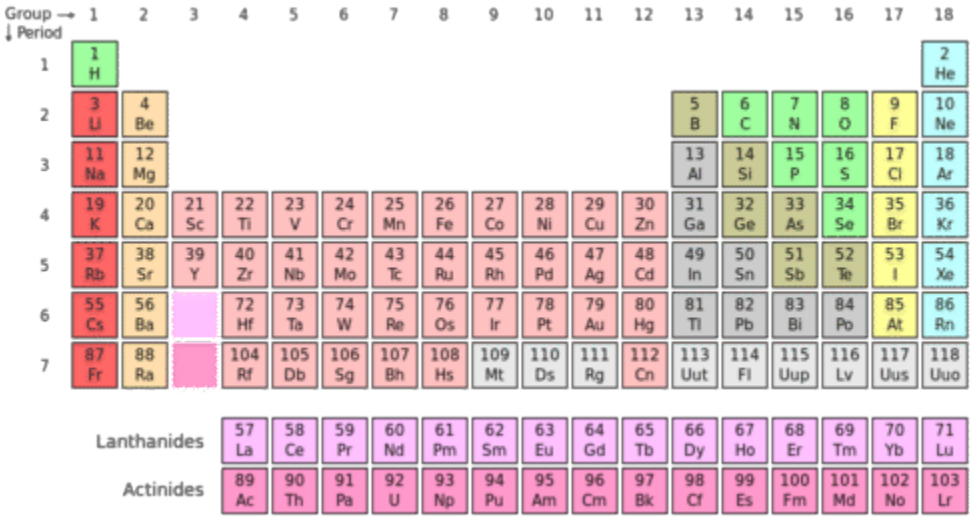


Figure 4.15.3: The periodic table classifies elements by atomic number. (Public Domain; User:Cepheus/Wikimedia Commons via Wikipedia)

The classification of elements by atomic number allows us to understand many properties of the atom and makes it possible to predict behaviors instead of just having to memorize everything.

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Summary

- The atomic number (Z) of an element is the number of protons in the nucleus of each atom of that element
- The number of electrons is equal to the number of protons in an atom of an element.

Review

1. What is the atomic number of an atom? Why is this number important?
2. Using a periodic table, what is the atomic number of helium have
3. How many protons are in the following elements:?
   a. Ne
   b. Ca
   c. Pt
4. Write the symbol for the element with the following atomic number:
   a. 18
   b. 41
   c. 82
   d. 12

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4.16: Mass Number

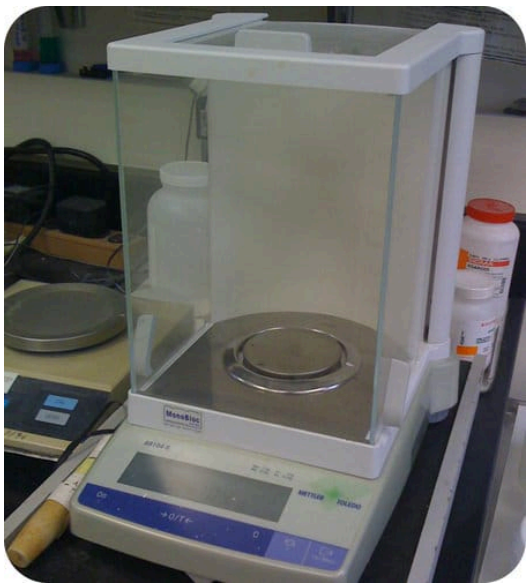


Figure 4.16.1 (Public Domain; User:Ebultoof/Wikipedia via Wikipedia)

How can you determine the mass of a chemical?

Often a student will need to weigh out a chemical for an experiment. If he or she uses a watch glass (a small, round piece that will hold the solid chemical), the weight of the watch glass must be determined first. Then the solid is added to the glass and the weight of the glass plus the solid is measured. The balance reading will be the total of the glass plus the chemical.

History of Atomic Weight Determinations

As a part of his research on atoms, John Dalton determined a number of atomic weights of elements in the early 1800s. Atomic weights were the basis for the periodic table that Mendeleev developed. Originally all atomic weights were based on a comparison to hydrogen, which has an atomic weight of one. After the discovery of the proton, scientists assumed that the weight of an atom was essentially that of the protons—electrons were known to contribute almost nothing to the atomic weight of the element.

This approach worked until we learned how to determine the number of protons in an element. We then saw that the atomic weight for an element was often twice the number of protons (or more). The discovery of the neutron provided the missing part of the picture. The atomic mass is now known to be the sum of the protons and neutrons in the nucleus.

Mass Number

Rutherford showed that the vast majority of the mass of an atom is concentrated in its nucleus, which is composed of protons and neutrons. The **mass number** is defined as the total number of protons and neutrons in an atom. It can be calculated by adding the number of neutrons and the number of protons (atomic number) together.

Mass number = atomic number + number of neutrons

Consider Table 4.16.1 below that shows data from the first six elements of the periodic table.

Table 4.16.1: Atoms of the First Six Elements

Name	Symbol	Protons	Neutrons	Electrons	Atomic Number	Mass Number
Hydrogen	H	1	0	1	1	1

Name	Symbol	Protons	Neutrons	Electrons	Atomic Number	Mass Number
Helium	He	2	2	2	2	4
Lithium	Li	3	4	3	3	7
Beryllium	Be	4	5	4	4	9
Boron	B	5	6	5	5	11
Carbon	C	6	6	6	6	12

Consider the element helium. Its atomic number is 2, so it has two protons in its nucleus. Its nucleus also contains two neutrons. Since $2 + 2 = 4$, we know that the mass number of the helium atom is 4. Finally, the helium atom also contains two electrons since the number of electrons must equal the number of protons. This example may lead you to believe that atoms have the same number of protons and neutrons, but further examination of the table above will show that this is not the case. Lithium, for example, has three protons and four neutrons, leaving it with a mass number of 7.

Knowing the mass number and the atomic number of an atom allows you to determine the number of neutrons present in that atom by subtraction.

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number}$$

Example 4.16.1

Atoms of the element chromium (Cr) have an atomic number of 24 and a mass number of 52. How many neutrons are in the nucleus of a chromium atom?

Solution

To determine this, you subtract the atomic number from the mass number, as shown above:

$$52 - 24 = 28 \text{ neutrons in a chromium atom}$$

The composition of any atom can be illustrated with a shorthand notation using the atomic number and the mass number. Both are written before the chemical symbol, with the mass number written as a superscript and the atomic number written as a subscript. The chromium atom discussed above would be written as:

$$^{52}_{24}\mathrm{Cr}$$

Another way to refer to a specific atom is to write the mass number of the atom after the name, separated by a hyphen. The above atom would be written as chromium-52.

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Summary

- The mass number is defined as the total number of protons and neutrons in an atom.
- The mass number = number of neutrons + atomic number.

Review

1. Who first determined atomic weights for elements?
2. What were the original atomic weights based on?
3. Why were calculations based on numbers of protons not valid for determining atomic weights?
4. A tin atom has an atomic number of 50 and a mass number of 118. How many neutrons are present in this atom?
5. What is the mass number of a cobalt atom that has 27 protons and 30 neutrons?

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4.17: Isotopes



Figure 4.17.1 (Public Domain; Chief Photographer's Mate Johnny Bivera, U.S. Navy via Wikipedia)

Are all the members of the football team shown above identical?

They are on the same team and are all known by the same team name, but there are individual differences among the players. We do not expect the kicker to be as big as the quarterback. The tight end is very likely to weigh less than the defensive tackle on the other side of the ball. They play as a unit, but they have different weights and heights.

What are Isotopes?

The history of the atom is full of some of these differences. Although John Dalton stated in his atomic theory of 1804 that all atoms of an element are identical, the discovery of the neutron began to show that this assumption was not correct. The study of radioactive materials (elements that spontaneously give off particles to form new elements) by Frederick Soddy (1877-1956) gave important clues about the internal structure of atoms. His work showed that some substances with different radioactive properties and different atomic masses were in fact the same element. He coined the term **isotope** from the Greek roots isos (ίσος "equal") and topos (τόπος "place"). He described isotopes as, "Put colloquially, their atoms have identical outsides but different insides." Soddy won the Nobel Prize in Chemistry in 1921 for his work.

As stated earlier, not all atoms of a given element are identical. Specifically, the number of neutrons can be variable for many elements. As an example, naturally occurring carbon exists in three forms. Each carbon atom has the same number of protons (6), which is its atomic number. Each carbon atom also contains six electrons in order to maintain electrical neutrality. However the number of neutrons varies as six, seven, or eight. Isotopes are atoms that have the same number atomic number, but different mass numbers due to a change in the number of neutrons.

The three isotopes of carbon can be referred to as carbon-12 ($^{12}_{6}C$), carbon-13 ($^{13}_{6}C$), and carbon-14 ($^{14}_{6}C$) refers to the nucleus of a given isotope of an element. A carbon atom is one of three different nuclides. Most elements naturally consist of mixtures of isotopes. Carbon has three natural isotopes, while some heavier elements can have many more. Tin has ten stable isotopes, the most of any element.

While the presence of isotopes affects the mass of an atom, it does not affect its chemical reactivity. Chemical behavior is governed by the number of electrons and the number of protons. Carbon-13 behaves chemically in exactly the same way as the more plentiful carbon-12.

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Summary

- Isotopes are atoms that have the same atomic number, but different mass numbers due to a change in the number of neutrons.
- The term nuclide refers to the nucleus of a given isotope of an element.
- The atomic mass of an atom equals the sum of the protons and the neutrons.

Review

1. What are isotopes?
2. Why do different isotopes of an element generally have the same physical and chemical properties?
3. How would the nucleus of the hydrogen-1 and hydrogen-2 differ?
4. Relate the concepts of isotope and mass number.
5. All oxygen atoms have eight protons, and most have eight neutrons as well. What is the mass number of an oxygen isotope that has nine neutrons? What is the name of this isotope?
6. An isotope of yttrium has 39 protons and 59 neutrons. What is the mass number of that isotope?
7. An isotope with a mass number of 193 has 116 neutrons. What is the atomic number of this isotope?
8. An isotope of barium (atomic number 56) has an mass of 138. How many neutrons are in the nucleus of this isotope?
9. Relate the concepts of isotope and mass number.

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4.18: Atomic Mass Unit



Figure 4.18.1 (Public Domain; Derek Jensen (User:Tysto/Wikimedia Commons) via Wikipedia)

The current system of measurement in the United States is a hodge-podge of different units, many of which are hard to interconvert. We at least have standardized units these days, unlike centuries past. At one time, measurements of length were defined as the distance from the end of one appendage to another. For example, the yard was defined as the distance from the king's nose to the tip of his thumb, when his arm was stretched out.

Standardized measurements make it possible for people everywhere to get the same amount of something. Note the red and white labels on the gas pump pictured below. These labels certify that the gas pump has been checked, and is pumping an accurate gallon of gas. Standard measurements in science are very important so that we can compare experimental data from one lab to another, and make sure that we are all talking about the same thing.

Atomic Mass

Masses of individual atoms are very, very small. Using a modern device called a mass spectrometer, it is possible to measure such miniscule masses. An atom of oxygen-16, for example, has a mass of 2.66×10^{-23} g. While comparisons of masses measured in grams would have some usefulness, it is far more practical to have a system that will allow us to more easily compare relative atomic masses. Scientists decided on using the carbon-12 nuclide as the reference standard by which all other masses would be compared. By definition, one atom of carbon-12 is assigned a mass of 12 **atomic mass units** (amu). An atomic mass unit is defined as a mass equal to one twelfth the mass of an atom of carbon-12. The mass of any isotope of any element is expressed in relation to the carbon-12 standard. For example, one atom of helium-4 has a mass of 4.0026 amu. An atom of sulfur-32 has a mass of 31.972 amu.

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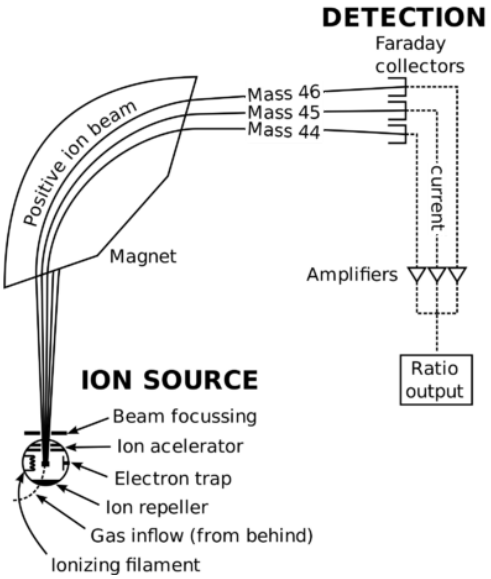


Figure 4.18.2: Mass spectrometer schematic. (Public Domain; Devon Fyson, based on image created by the U.S. Geological Survey via Wikipedia)

The carbon-12 atom has six protons and six neutrons in its nucleus for a mass number of 12. Since the nucleus accounts for nearly all of the mass of the atom, a single proton or single neutron has a mass of approximately $1\ \text{amu}$. However, as seen by the helium and sulfur examples, the masses of individual atoms are not whole numbers. This is because an atom's mass is affected very slightly by the interactions of the various particles within the nucleus, and the small mass of the electron is taken into account.

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Summary

- Carbon-12 is the reference for all atomic mass calculations.
- An atomic mass unit is defined as a mass equal to one twelfth the mass of an atom of carbon-12.
- The mass of an atom is affected by the interactions of particles within the nucleus.

Review

1. What instrument is used to measure the mass of atoms?
2. How much does a single oxygen-16 atom weigh in grams?
3. What is the reference standard for atomic mass units?
4. How is an atomic mass unit defined?
5. Why are the numbers for atomic mass of individual atoms not whole numbers?

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4.19: Calculating Atomic Mass



Figure 4.19.1 (CC BY 2.0; George M. Groutas via Flickr Boulders beach, Cape Peninsula [www.flickr.com])

Have you ever tried to move a boulder?

You have a pile of rocks to move and need to decide what equipment you want to rent to move them. If the rocks are fairly small, you can get a shovel to pick them up. Larger rocks could be moved by hand, but big boulders will need some sort of mechanical scoop. The amount of each kind of rock will also determine how much time you will need to get the job done. Knowing the relative amounts of large, medium, and small rocks can be very useful in deciding how to approach the job.

Percent Natural Abundance

Most elements occur naturally as a mixture of two or more isotopes. The table below shows the natural isotopes of several elements, along with the **percent natural abundance** of each.

Table 4.19.1: Atomic Masses and Percents of Abundance of Some Natural Isotopes

Element	Isotope (Symbol)	Percent Natural Abundance	Atomic Mass (amu)	Average Atomic Mass (amu)
Hydrogen	$^{1}_{1}\mathrm{H}$	99.985	1.0078	1.0080
	$^{2}_{1}\mathrm{H}$	0.015	2.0141	
	$^{3}_{1}\mathrm{H}$	negligible	3.0160	
Carbon	$^{12}_{6}\mathrm{C}$	98.89	12.000	12.011
	$^{13}_{6}\mathrm{C}$	1.11	13.003	
	$^{14}_{6}\mathrm{C}$	trace	14.003	
Oxygen	$^{16}_{8}\mathrm{O}$	99.759	15.995	15.999
	$^{17}_{8}\mathrm{O}$	0.037	16.995	
	$^{18}_{8}\mathrm{O}$	0.204	17.999	
Chlorine	$^{35}_{17}\mathrm{Cl}$	75.77	34.969	35.453
	$^{37}_{17}\mathrm{Cl}$	24.23	36.966	
Copper	$^{63}_{29}\mathrm{Cu}$	69.17	62.930	63.546
	$^{65}_{29}\mathrm{Cu}$	30.83	64.928	

For some elements, one particular isotope predominates greatly over the other isotopes. Naturally occurring hydrogen is nearly all hydrogen-1 and naturally occurring oxygen is nearly all oxygen-16. For many other elements, however, more than one isotope may exist in more substantial quantities. Chlorine (atomic number 17) is a yellowish-green toxic gas. About three quarters of all chlorine atoms have 18 neutrons, giving those atoms a mass number of 35. About one quarter of all chlorine atoms have 20 neutrons, giving those atoms a mass number of 37. Were you to simply calculate the arithmetic average of the precise **atomic masses**, you would get 36.

$$\frac{(34.969 + 36.966)}{2} = 35.968 \text{ amu}$$

Clearly the actual average atomic mass from the last column of the table is significantly lower. Why? We need to take into account the percent natural abundance of each isotope, in order to calculate the weighted average. The atomic mass of an element is the weighted average of the atomic masses of the naturally occurring isotopes of that element. The sample problem below demonstrates how to calculate the atomic mass of chlorine.

Example 4.19.1

Use the atomic masses of each of the two isotopes of chlorine along with their respective percent abundances to calculate the average atomic mass of chlorine.

Solution

Step 1: List the known and unknown quantities and plan the problem.

Known

- Chlorine-35: atomic mass $= 34.969$ amu and percent abundance $= 75.77\%$
- Chlorine-37: atomic mass $= 36.966$ amu and percent abundance $= 24.23\%$

Unknown

- Average atomic mass of chlorine

Change each percent abundance into decimal form by dividing by 100. Multiply this value by the atomic mass of that isotope. Add together for each isotope to get the average atomic mass.

Step 2: Calculate.

chlorine-35	$0.7577 \times 34.969 = 26.50$ amu
chlorine-37	$0.2423 \times 36.966 = 8.957$ amu
average atomic mass	$26.50 + 8.957 = 35.46$ amu

Note: Applying significant figure rules results in the 35.45 amu result without excessive rounding error. In one step:

$$(0.7577 \times 34.969) + (0.2423 \times 36.966) = 35.46 \text{ amu}$$

Step 3: Think about your result.

The calculated average atomic mass is closer to 35 than to 37 because a greater percentage of naturally occurring chlorine atoms have the mass number of 35. It agrees with the value from the table above.

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Summary

- The atomic mass of an element is the weighted average of the atomic masses of the naturally occurring isotopes of that element.
- Calculations of atomic mass use the percent abundance of each isotope.

Review

1. Define atomic mass.
2. What information do you need to calculate atomic mass for an element?
3. Calculate the atomic mass for carbon using the data provided in the table below.

Isotope	Atomic Mass	Percent Abundance
carbon-12	12.000000	98.90
carbon-13	13.003355	1.100

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4.20: Atomic Structure Practice

Sacramento City College
CHEM 309: General, Organic and Biochemistry
Prof. Dianne Bennett
Agenda (Fall) • **Agenda (Spring)** • **Homework Problems**

1. Chapter 2: The Classical Wave Equation
2. 2.1: The One-Dimensional Wave Equation
3. 2.2: The Method of Separation of Variables
4. 2.3: Oscillatory Solutions to Differential Equations
5. 2.4: The General Solution is a Superposition of Normal Modes
6. 2.5: A Vibrating Membrane
7. 2.E: The Classical Wave Equation (Exercises)

- **Homework Solutions**

1. Chapter 11: Computational Quantum Chemistry
2. 11.1: Gaussian Basis Sets
3. 11.2: Extended Basis Sets
4. 11.3: Orbital Polarization Terms in Basis Sets
5. 11.4: The Ground-State Energy of H2H2
6. 11.5: Quantum Calculations
7. 11.E: Computational Quantum Chemistry (Exercises)

Chapters

1. Chapter 12: Group Theory: The Exploitation of Symmetry
2. 12.1: The Exploitation of Symmetry
3. 12.2: Symmetry Elements
4. 12.3: Symmetry Operations Define Groups
5. 12.4: Symmetry Operations as Matrices
6. 12.5: The C3VC3V Point Group
7. 12.6: Character Tables
8. 12.7: Characters of Irreducible Representations
9. 12.8: Using Symmetry to Solve Secular Determinants
10. 12.9: Generating Operators
11. 12.E: Group Theory: The Exploitation of Symmetry (Exercises)

- **Video Tutorials**

1. Chapter 13: Molecular Spectroscopy
2. 13.1: The Electromagnetic Spectrum
3. 13.2: Rotations Accompany Vibrational Transitions
4. 13.3: Unequal Spacings in Vibration-Rotation Spectra
5. 13.4: Unequal Spacings in Pure Rotational Spectra
6. 13.5: Vibrational Overtones
7. 13.6: Electronic Spectra Contain Electronic, Vibrational, and Rotational Information
8. 13.7: The Franck-Condon Principle
9. 13.8: Rotational Spectra of Polyatomic Molecules
10. 13.9: Normal Modes in Polyatomic Molecules
11. 13.10: Irreducible Representation of Point Groups
12. 13.11: Time-Dependent Perturbation Theory
13. 13.12: The Selection Rule for the Rigid Rotator
14. 13.13: The Harmonic Oscillator Selection Rule
15. 13.14: Group Theory Determines Infrared Activity
16. 13.E: Molecular Spectroscopy (Exercises)

- **Basic Knowledge Practice Quizzes**

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There are select answers to these problems here.

ALL questions listed must be answered to earn credit for completing the text homework. Yes, even the Additional Exercises.

Elements and Atomic Structure

(1) Of the three subatomic particles, which is the least massive? Which of the subatomic particles are located inside the nucleus of the atom?

(2) Neon has 10 electrons, and Argon has 18 electrons. Of the two Noble gases, which would you expect to have the greater atomic radius? Justify your answer.

(3) Using the periodic table, identify the atomic symbols and names of the elements that correspond to the following atomic numbers:

a. 8
b. 27
c. 33
d. 59
e. 90

(4) How many electrons do each of the following elements contain? Hint: These are not ions, but neutral elements.

a. Fluorine
b. Magnesium
c. Iodine
d. Copper
e. Silicon

(5) Which elements contain the following numbers of protons and electrons?

a. 15 protons and 15 electrons
b. 77 protons and 77 electrons
c. 42 protons and 42 electrons
d. 4 protons and 4 electrons
e. 99 protons and 99 electrons

(6) Between elements in the same column, would you expect elements that are higher up in the column to have a larger atomic radius, or ones that are lower in the column? Justify your answer.

(7) Of the three subatomic particles that make up an element, which one can always be used to identify the element? Why can the others not be used in this way? Explain.

(8) Define the mass number of an element. Is this different from the atomic number? Explain.

(9) What is an isotope? How does one isotope differ from another within the same element?

(10) In a sample of different isotopes of Carbon-11, Carbon-12, Carbon-13 and Carbon-14, fill in the blanks for the atomic number, how many protons, neutrons, electrons and the mass number of each isotope.

(11) What do all isotopes of nitrogen have in common? Is this true for isotopes of all elements?

(12) How would you determine which isotope in a group is the smallest? The largest?

(13) Two of the most well-known isotopes of carbon are carbon-12 and carbon-14. Answer the following:

a. What is the atomic mass of each isotope?
b. What is the atomic number of each isotope?
c. What is the difference between the two isotopes?
d. Based on the atomic mass listed on your periodic table, which of the two is more abundant in nature, carbon-12 or carbon-14? The mass of carbon on the periodic table is 12.01.

(14) Starting with polonium, atomic number 84, there are no stable, non-radioactive isotopes of any subsequent elements on the periodic table. Using the atomic symbol and the mass and atomic numbers, express the isotope of polonium-84.

Periodic Table of Elements

(15) One of the main groups on the periodic tables is the Metals. How many different subgroups of metals are there on the periodic table? Name each of them. Where in the periodic table can each subgroup be located? Be specific.

(16) Apart from Metals, there are Metalloids and Non-Metals as groups on the periodic table. Name at least two subgroups within the Non-Metals. Where on the periodic table are these subgroups located? Be specific.

(17) How do the transition metals differ from the alkali and alkaline earth metals?

(18) Determine if the following descriptions are examples of a physical or chemical property.

a. Iron nails rust when left out in the rain
b. Glass is transparent and rigid
c. Table salt creates an electrical current when dissolved in water
d. Diamond is virtually unbreakable
e. Metallic sodium explodes when exposed to water

(19) Of the following elements, which of the three categories, metal, non-metal and metalloid, do they each belong to?

a. Manganese
b. Bromine
c. Silicon
d. Uranium
e. Phosphorus

(20) Using the given clues, determine the identity of the following elements:

a. A transition metal in period 4 with 3 valence electrons
b. A noble gas with atomic number less than 10
c. A metalloid from period 2
d. Group 7A element that does NOT exist as a diatomic molecule
e. Group 5A element directly adjacent to a metalloid

(21) For periods 1-6, how many electrons are in each energy level? How many valence electrons do the following elements from levels 2 and 3 have?

a. Fluorine
b. Aluminum
c. Nitrogen
d. Magnesium
e. Silicon

(22) Which family of elements has the least number of valence electrons? Which has the most?

Additional Exercises

(23) If the atomic radius of sodium atoms is 1.86×10^{-10} m, how many sodium atoms are needed to make a line that is 1.00 cm in length?

(24) If the atomic radius of osmium atoms is 1.34×10^{-10} m, how many osmium atoms are needed to make a line that is 5.85 m in length?

(25) The compound that sodium makes with chlorine has sodium and chlorine atoms in a 1:1 ratio. Name two other elements that should make a compound having a 1:1 ratio of atoms with sodium.

(26) The compound that magnesium makes with oxygen has magnesium to oxygen atoms in a 1:1 ratio. Name two other elements that should make a compound having a 1:1 ratio of atoms with magnesium.

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5: ELECTRONS IN ATOMS

5.1: Properties of Light

5.2: Atomic Emission Spectra

5.3: Bohr's Atomic Model

5.4: Spectral Lines of Atomic Hydrogen

5.5: Quantum Mechanical Atomic Model

5.6: Quantum Numbers

5.7: Orbitals

5.8: Aufbau Principle

5.9: Pauli Exclusion Principle

5.10: Electron Configurations

5.11: Valence Electrons

5.12: Noble Gas Configuration

5.13: Hund's Rule and Orbital Filling Diagrams

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5.1: Properties of Light



Figure 5.1.1 (Public Domain; Jon Sullivan / pdphoto.org via Wikipedia)

What are waves?

Waves can come in many sizes. Here we see a large wave crashing on the beach. Other waves can be very small and regular. We normally think of waves as being made of water, but there are forms of energy that take on the characteristics of waves. The idea of a wave has played a major role in our understanding of how the atom is put together and why it behaves the way it does.

Properties of Light

The nuclear atomic model proposed by Rutherford was a great improvement over previous models, but was still not complete. It did not fully explain the location and behavior of the electrons in the vast space outside of the nucleus. In fact, it was well known that oppositely charged particles attract one another. Rutherford's model did not explain why the electrons don't simply move toward, and eventually collide with, the nucleus. Experiments in the early twentieth century began to focus on the absorption and emission of light by matter. These studies showed how certain phenomena associated with light reveal insight into the nature of matter, energy, and atomic structure.

Wave Nature of Light

In order to begin to understand the nature of the electron, we first need to look at the properties of light. Prior to 1900, scientists thought light behaved solely as a wave. As we will see later, this began to change as new experiments demonstrated that light also has some of the characteristics of a particle. First, we will examine the wavelike properties of light.

Visible light is one type of **electromagnetic radiation**, which is a form of energy that exhibits wavelike behavior as it moves through space. Other types of electromagnetic radiation include gamma rays, x-rays, ultraviolet light, infrared light, microwaves, and radio waves. The figure below shows the **electromagnetic spectrum**, which is all forms of electromagnetic radiation. Notice that visible light makes up only a very, very small portion of the entire electromagnetic spectrum. All electromagnetic radiation moves through a vacuum at a constant speed of $2.998 \times 10^8 \text{ m/s}$. While the presence of air molecules slows the speed of light by a very small amount, we will use the value of $3.00 \times 10^8 \text{ m/s}$ as the speed of light in air.

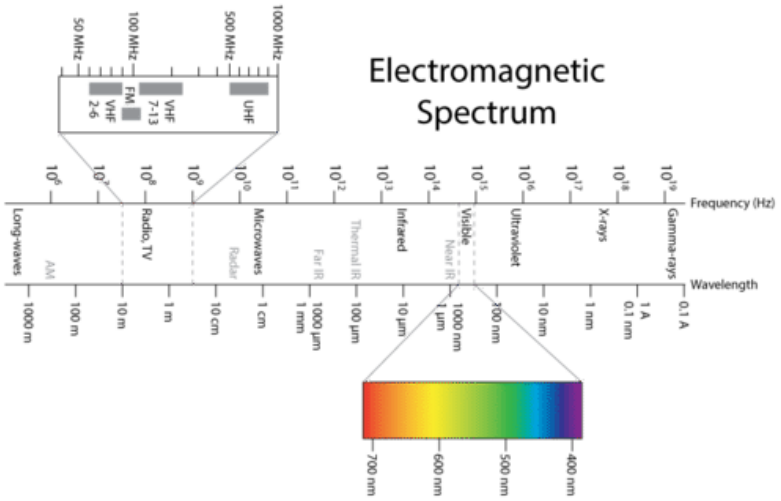


Figure 5.1.2: The electromagnetic spectrum encompasses a very wide range of wavelengths and frequencies. Visible light is only a very small portion of the spectrum with wavelengths from 400-700 nm. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

The figure above shows how the electromagnetic spectrum displays a wide variation in wavelength and frequency. Radio waves have wavelengths of as long as hundreds of meters, while the wavelength of gamma rays are on the order of 10^{-12} m. The corresponding frequencies range from 10^{6} to 10^{21} Hz. Visible light can be split into colors with the use of a prism (see below), yielding the visible spectrum of light. Red light has the longest wavelength and lowest frequency, while violet light has the shortest wavelength and highest frequency. Visible light wavelength ranges from about 400 to 700 nm with frequencies in the range of 10^{14} Hz.

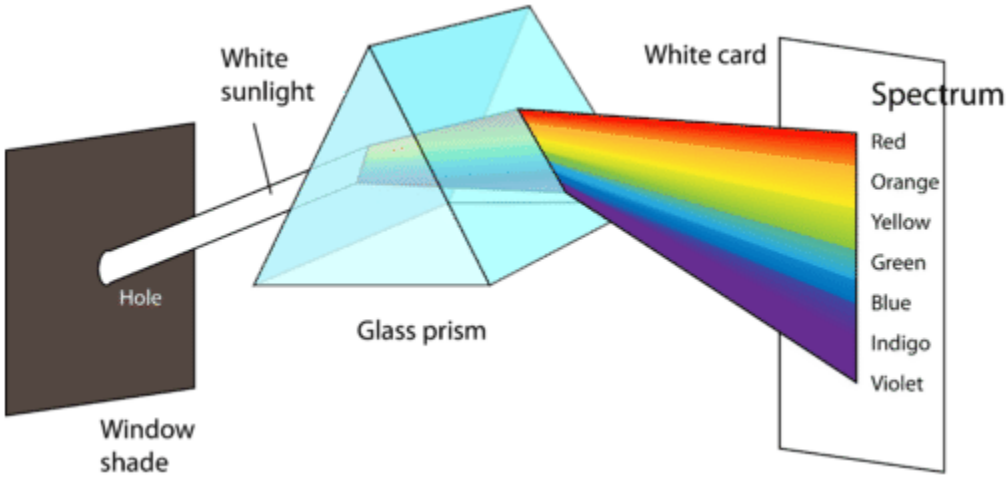


Figure 5.1.3: A small beam of white light is refracted (bent) as it passes through a glass prism. The shorter the wavelength of light, the greater is the refraction, so the light is separated into all of its colors. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

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Simulation

If we could see other wavelengths of energy, would they appear as colors? Use this simulation to explore the electromagnetic spectrum.

Summary

- Electromagnetic radiation is a form of energy.
- Visible light has wavelengths from $400\text{-}700\ \text{nm}$.
- The speed of light in air is $3.00 \times 10^8\ \text{m/s}$.

Review

1. What did Rutherford's nuclear atomic model not explain?
2. Prior to 1900, what did scientists believe about the nature of light?
3. What is visible light?
4. What is the range of wavelengths for visible light?

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5.2: Atomic Emission Spectra



Figure 5.2.1 (Public Domain; Department of Defense Warrior Games via Commons Wikimedia, Archery at the 2017 Warrior Games [commons.wikimedia.org])

How much energy does it take to shoot an arrow?

Archery as a sport or a means of defense has existed for centuries. At rest, there is no tension on the bowstring and no force on the arrow. When the string and arrow are pulled back, we now have a situation where kinetic energy (pulling of the string) has been converted to potential energy (the tension on the string). The archer releases the arrow and the potential energy is translated into kinetic energy as the arrow moves. It turns out that electrons behave the same way when energy is put into the system or released from the system.

Atomic Emission Spectra

The electrons in an atom tend to be arranged in such a way that the energy of the atom is as low as possible. The **ground state** of an atom is the lowest energy state of the atom. When those atoms are given energy, the electrons absorb the energy and move to a higher energy level. These energy levels of the electrons in atoms are quantized, meaning again that the electron must move from one energy level to another in discrete steps, rather than continuously. An **excited state** of an atom is a state where its potential energy is higher than the ground state. An atom in the excited state is not stable. When it returns back to the ground state, it releases the energy that it had previously gained in the form of electromagnetic radiation.

So how do atoms gain energy in the first place? One way is to pass an electric current through an enclosed sample of a gas at low pressure. Since the electron energy levels are unique for each element, every gas discharge tube will glow with a distinctive color, depending on the identity of the gas (see below).

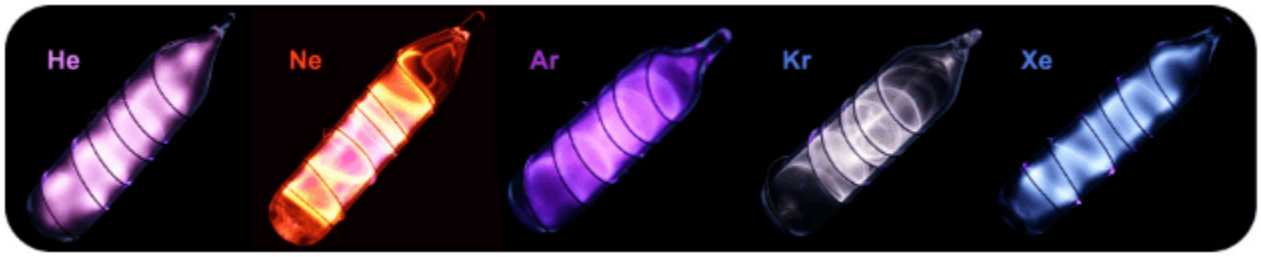


Figure 5.2.2: Gas discharge tubes are enclosed glass tubes filled with a gas at low pressure, through which an electric current is passed. Electrons in the gaseous atoms first become excited, and then fall back to lower energy levels, emitting light of a distinctive color in the process. Shown are gas discharge tubes of helium, neon, argon, krypton, and xenon. (CC by 3.0; User:Jurii/Wikimedia Commons, Heinrich Pniok (Wikimedia: Alchemist-hp) via Commons Wikimedia, Glowing noble gases [commons.wikimedia.org])

"Neon" signs are familiar examples of gas discharge tubes. However, only signs that glow with the red-orange color seen in the figure are actually filled with neon. Signs of other colors contain different gases or mixtures of gases.

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Scientists studied the distinctive pink color of the gas discharge created by hydrogen gas. When a narrow beam of this light was viewed through a prism, the light was separated into four lines of very specific wavelengths (and frequencies since λ and ν are inversely related). An **atomic emission spectrum** is the pattern of lines formed when light passes through a prism to separate it into the different frequencies of light it contains. The figure below shows the atomic emission spectrum of hydrogen.

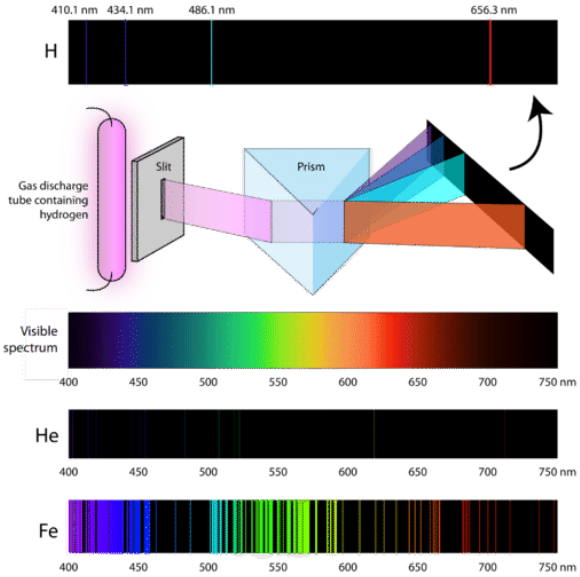


Figure 5.2.3: When light from a hydrogen gas discharge tube is passed through a prism, the light is split into four visible lines. Each of these spectral lines corresponds to a different electron transition from a higher energy state to a lower energy state. Every element has a unique atomic emission spectrum, as shown by the examples of helium (He) and iron (Fe). (CC BY-NC 3.0; Christopher Auyeung, using emission spectra available in the public domain via CK-12 Foundation; H spectrum: Commons Wikimedia, Emission Spectrum- H [commons.wikimedia.org]; visible spectrum: Commons Wikimedia, Linear Visible Spectrum [commons.wikimedia.org]; He spectrum: Commons Wikimedia, Helium Emission Spectrum [commons.wikimedia.org]; Fe spectrum: Commons Wikimedia, Emission Spectrum-Fe [commons.wikimedia.org])

Classical theory was unable to explain the existence of atomic emission spectra, also known as line-emission spectra. According to classical physics, a ground state atom would be able to absorb any amount of energy rather than only discrete amounts. Likewise, when the atoms relaxed back to a lower energy state, any amount of energy could be released. This would result in what is known as a **continuous spectrum**, where all wavelengths and frequencies are represented. White light viewed through a prism and a rainbow are examples of continuous spectra. Atomic emission spectra were more proof of the quantized nature of light and led to a new model of the atom based on quantum theory.

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Simulation

Look up at the sky at night and you'll see stars. How do we know what they are made of? Check out this simulation explore atomic colors and the ingredients that make up our universe.

Summary

- Atomic emission spectra are produced when excited electrons return to the ground state.
- The emitted light of electrons corresponds to energies of the specific electrons.

Review

1. What is the ground state of an atom?
2. What is an excited state?
3. Why do we see emission lines when electrons return to the ground state?

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5.3: Bohr's Atomic Model

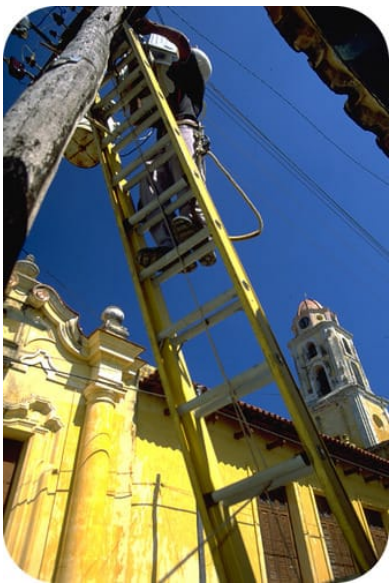


Figure 5.3.1 (Public Domain; the US Department of Agriculture via Wikipedia)

How does this worker's energy change as he climbs up and down the ladder?

Climbing a ladder takes energy. At every step, you push yourself up against gravity, and accumulate potential energy. Coming back down releases that potential energy as you descend step by step. If you are not careful, you can release that potential energy all at once by falling off the ladder (never a good thing). In addition, you take the climb or descent in steps. There is no "in-between" position on the ladder—your foot either hits a rung, or it hits empty space, and you are in trouble until you find a rung to stand on.

Bohr's Atomic Model

Following the discoveries of hydrogen emission spectra and the photoelectric effect, the Danish physicist Niels Bohr (1885-1962) proposed a new model of the atom in 1915. Bohr proposed that electrons do not radiate energy as they orbit the nucleus, but exist in states of constant energy that he called **stationary states**. This means that the electrons orbit at fixed distances from the nucleus (see below). Bohr's work was primarily based on the emission spectra of hydrogen. This is also referred to as the planetary model of the atom. It explained the inner workings of the hydrogen atom. Bohr was awarded the Nobel Prize in physics in 1922 for his work.

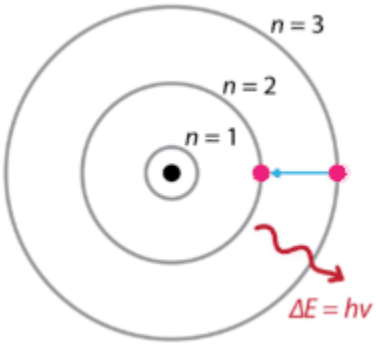


Figure 5.3.2: Bohr's atomic model hydrogen emission spectra. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Bohr explained that electrons can be moved into different orbits with the addition of energy. When the energy is removed, the electrons return back to their ground state, emitting a corresponding amount of energy—a quantum of light, or photon. This

was the basis for what later became known as **quantum theory**. This is a theory based on the principle that matter and energy have the properties of both particles and waves. It accounts for a wide range of physical phenomena, including the existence of discrete packets of energy and matter, the uncertainty principle, and the exclusion principle.

According to the Bohr model, often referred to as a **planetary model**, the electrons encircle the nucleus of the atom in specific allowable paths called orbits. When the electron is in one of these orbits, its energy is fixed. The ground state of the hydrogen atom, where its energy is lowest, is when the electron is in the orbit that is closest to the nucleus. The orbits that are further from the nucleus are all of successively greater energy. The electron is not allowed to occupy any of the spaces in between the orbits. An everyday analogy to the Bohr model is the rungs of a ladder. As you move up or down a ladder, you can only occupy specific rungs and cannot be in the spaces in between rungs. Moving up the ladder increases your potential energy, while moving down the ladder decreases your energy.

Bohr's work had a strong influence on our modern understanding of the inner workings of the atom. However, his model worked well as an explanation for the emissions of the hydrogen atom, but was seriously limited when applied to other atoms. Shortly after Bohr published his planetary model of the atom, several new discoveries were made, which resulted in, yet again, a revised view of the atom.

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Summary

- The Bohr model postulates that electrons orbit the nucleus at fixed energy levels.
- Orbits further from the nucleus exist at higher energy levels.
- When electrons return to a lower energy level, they emit energy in the form of light.

Review

1. When did Bohr propose his model of the atom?
2. What is a stationary state?
3. What is the ground state?
4. Can the electron occupy any space between the orbits?

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5.4: Spectral Lines of Atomic Hydrogen

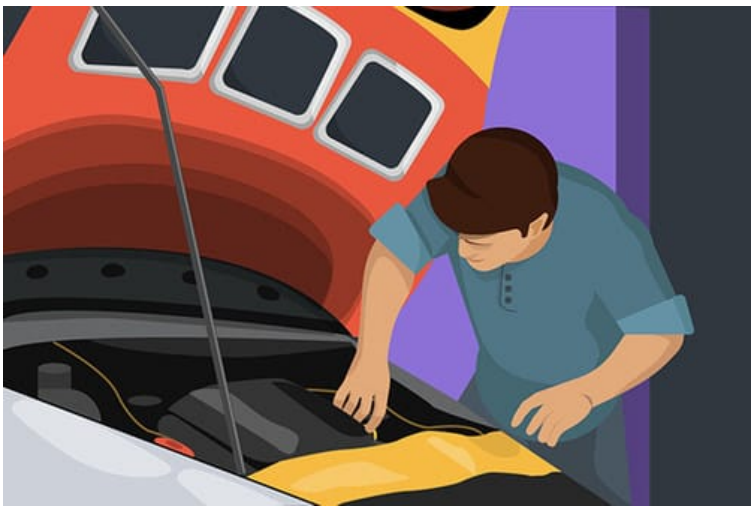


Figure 5.4.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Can you fix a car engine?

While less common in this day and age, there was a time when many people could work on their own cars if there was a problem. Today, engines are computerized and require specialized training and tools in order to be fixed. When people did their own repairs, it was sometimes a trial and error process. Maybe the spark plugs needed to be replaced. No, that didn't fix the problem completely, but it was a start in the right direction. Science operates the same way. A theory that is developed may work for a while, but then there are data that the theory cannot explain. This means that it's time for a newer and more inclusive theory.

Spectral Lines of Hydrogen

Bohr's model explains the spectral lines of the hydrogen atomic emission spectrum. While the electron of the atom remains in the ground state, its energy is unchanged. When the atom absorbs one or more quanta of energy, the electron moves from the ground state orbit to an excited state orbit that is further away. Energy levels are designated with the variable n. The ground state is $n = 1$, the first excited state is $n = 2$, and so on. The energy that is gained by the atom is equal to the difference in energy between the two energy levels. When the atom relaxes back to a lower energy state, it releases energy that is again equal to the difference in energy of the two orbits (see below).

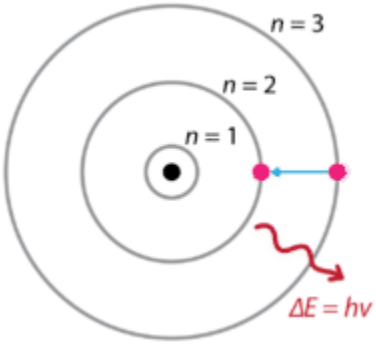


Figure 5.4.2: Bohr model of the atom: electron is shown transitioning from the $n = 3$ energy level to the $n = 2$ energy level. The photon of light that is emitted has a frequency that corresponds to the difference in energy between the two levels. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

The change in energy, ΔE, then translates to light of a particular frequency being emitted according to the equation $E = h\nu$. Recall that the atomic emission spectrum of hydrogen had spectral lines consisting of four different frequencies. This is explained

in the Bohr model by the realization that the electron orbits are not equally spaced. As the energy increases further and further from the nucleus, the spacing between the levels gets smaller and smaller.

Based on the wavelengths of the spectral lines, Bohr was able to calculate the energies that the hydrogen electron would have in each of its allowed energy levels. He then mathematically showed which energy level transitions correspond to the spectral lines in the atomic emission spectrum (see below).

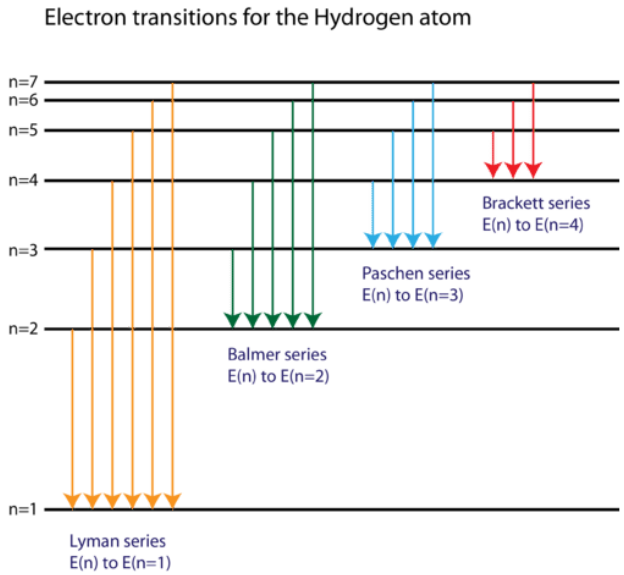


Figure 5.4.3: The electron energy level diagram for the hydrogen atom. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

He found that the four visible spectral lines corresponded to transitions from higher energy levels down to the second energy level $(n = 2)$. This is called the Balmer series. Transitions ending in the ground state $(n = 1)$ are called the Lyman series, but the energies released are so large that the spectral lines are all in the ultraviolet region of the spectrum. The transitions called the Paschen series and the Brackett series both result in spectral lines in the infrared region, because the energies are too small.

Bohr's model was a tremendous success in explaining the spectrum of the hydrogen atom. Unfortunately, when the mathematics of the model were applied to atoms with more than one electron, it was not able to correctly predict the frequencies of the spectral lines. While Bohr's model represented a great advancement in the atomic model and the concept of **electron transitions** between energy levels is valid, improvements were needed in order to fully understand all atoms and their chemical behavior.

Summary

- Emission lines for hydrogen correspond to energy changes related to electron transitions.
- The Bohr model works only for the hydrogen atom.

Review

1. What happens when a hydrogen atom absorbs one or more quanta of energy?
2. How do we detect the change in energy?
3. What electron transitions are presented by the lines of the Paschen series?
4. Does the Bohr model work for atoms other than hydrogen?

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5.5: Quantum Mechanical Atomic Model



Figure 5.5.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Imagine that a news flash interrupts your favorite TV program... "There has been a hold-up at the First National Bank. The suspect fled in a car and is believed to be somewhere in the downtown district. Everyone is asked to be on the alert." The robber can be located only within a certain area - the police do not have an exact location, just a general idea as to the whereabouts of the thief.

Quantum Mechanical Atomic Model

In 1926, Austrian physicist Erwin Schrödinger (1887-1961) used the wave-particle duality of the electron to develop and solve a complex mathematical equation that accurately described the behavior of the electron in a hydrogen atom. **The quantum mechanical model** of the atom comes from the solution to Schrödinger's equation. Quantization of electron energies is a requirement in order to solve the equation. This is unlike the Bohr model, in which quantization was simply assumed with no mathematical basis.

Recall that in the Bohr model, the exact path of the electron was restricted to very well-defined circular orbits around the nucleus. The quantum mechanical model is a radical departure from that. Solutions to the Schrödinger wave equation, called **wave functions**, give only the probability of finding an electron at a given point around the nucleus. Electrons do not travel around the nucleus in simple circular orbits.

The location of the electrons in the quantum mechanical model of the atom is often referred to as an **electron cloud**. The electron cloud can be thought of in the following way: Imagine placing a square piece of paper on the floor with a dot in the circle representing the nucleus. Now take a marker and drop it onto the paper repeatedly, making small marks at each point the marker hits. If you drop the marker many, many times, the overall pattern of dots will be roughly circular. If you aim toward the center reasonably well, there will be more dots near the nucleus and progressively fewer dots as you move away from it. Each dot represents a location where the electron could be at any given moment. Because of the uncertainty principle, there is no way to know exactly where the electron is. An electron cloud has variable densities: a high density where the electron is most likely to be and a low density where the electron is least likely to be (see below).

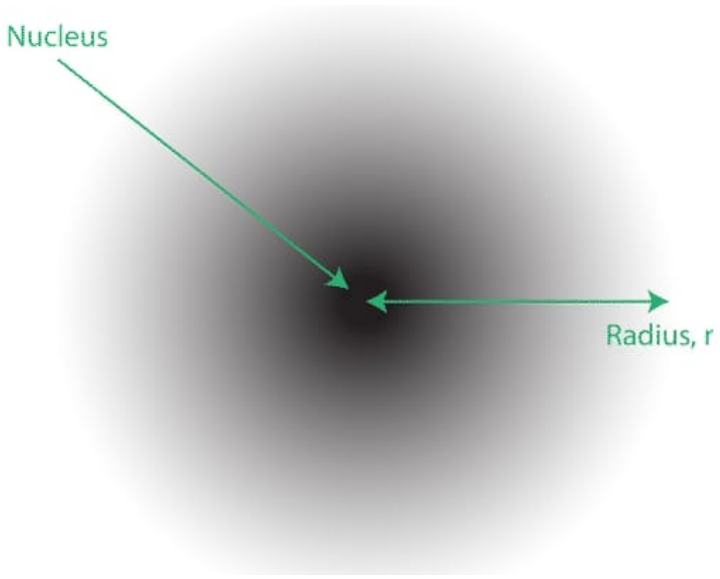


Figure 5.5.2: An electron cloud: the darker region nearer the nucleus indicates a high probability of finding the electron, while the lighter region further from the nucleus indicates a lower probability of finding the electron. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

In order to specifically define the shape of the cloud, it is customary to refer to the area within which there is a 90 chance of finding the electron. This is called an **orbital**, the three-dimensional region of space that indicates where there is a high probability of finding an electron.

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Summary

- The Schrödinger wave equation replaced Bohr's ideas about electron location with an uncertainty factor.
- The location of the electron can only be given as a probability that the electron is somewhere in a certain area.

Review

1. What does the quantum mechanical view of the atom require?
2. What is a wave function?
3. What does a high density electron cloud suggest?

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5.6: Quantum Numbers

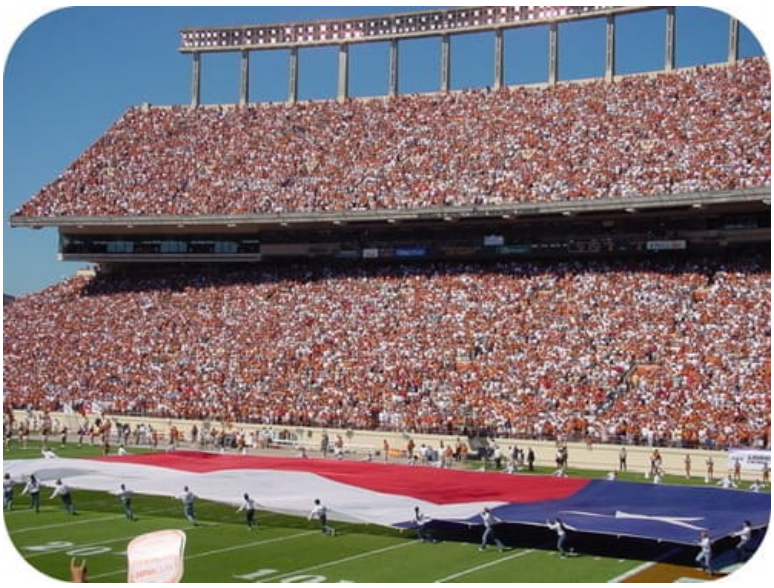


Figure $5.6.1$ (Public Domain; Enoch Lai via Wikipedia)

Can you guess how many people are in this stadium?

If you attend a college or professional football game, you need a ticket to get in. It is very likely that your ticket may specify a gate number, section number, row, and seat number. No other ticket can have the same four parts to it. It may even have the same gate, section, and row, but it would still have to include a different seat number. Each seat is unique and allows only one occupant to fill it.

Quantum Numbers

We use a series of specific numbers, called **quantum numbers**, to describe the location of an electron in an associated atom. Quantum numbers specify the properties of the atomic orbitals and the electrons in those orbitals. An electron in an atom or ion has four quantum numbers to describe its state. Think of them as important variables in an equation that describe the three-dimensional position of electrons in a given atom.

Principal Quantum Number (n)

The **principal quantum number**, signified by n, is the main energy level occupied by the electron. Energy levels are fixed distances from the nucleus of a given atom. They are described in whole number increments (e.g., 1, 2, 3, 4, 5, 6, ...). At location $n = 1$, an electron would be closest to the nucleus, while at $n = 2$ the electron would be farther, and at $n = 3$ farther yet. As we will see, the principal quantum number corresponds to the row number for an atom on the periodic table.

Angular Momentum Quantum Number (l)

The **angular momentum quantum number**, signified by l, describes the general shape or region an electron occupies—its orbital shape. The value of l depends on the value of the principal quantum number, n. The angular momentum quantum number can have positive values of zero to $(n - 1)$. If $n = 2$, l could be either 0 or 1.

Magnetic Quantum Number (m_l)

The **magnetic quantum number**, signified as m_l, describes the orbital orientation in space. Electrons can be situated in one of

three planes in three dimensional space around a given nucleus (x, y, and z). For a given value of the angular momentum quantum number, l, there can be $(2l + 1)$ values for m_l. As an example:

$n = 2$

$l = 0$ or 1

for $l = 0$, $m_l = 0$

for $l = 1$, $m_l = -1, 0, +1$

Table 5.6.1: Principal Energy Levels and Sublevels

Principal Energy Level	Number of Possible Sublevels	Possible Angular Momentum Quantum Numbers	Orbital Designation by Principal Energy Level and Sublevel
$n = 1$	1	$l = 0$	$1s$
$n = 2$	2	$l = 0$	$2s$
		$l = 1$	$2p$
$n = 3$	3	$l = 0$	$3s$
		$l = 1$	$3p$
		$l = 2$	$3d$
$n = 4$	4	$l = 0$	$4s$
		$l = 1$	$4p$
		$l = 2$	$4d$
		$l = 3$	$4f$

The table above shows the possible angular momentum quantum number values (l) for the corresponding principal quantum numbers $(n)\, of \backslash (n = 1, n = 2, n = 3,$ and $n = 4$.

Spin Quantum Number (m_s)

The **spin quantum number** describes the spin for a given electron. An electron can have one of two associated spins, $(+\frac{1}{2})$ spin, or $(-\frac{1}{2})$ spin. An electron cannot have zero spin. We also represent spin with arrows ↑ or ↓. A single orbital can hold a maximum of two electrons, and each must have opposite spin.

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Summary

- Quantum numbers specify the arrangements of electrons in orbitals.
- There are four quantum numbers that provide information about various aspects of electron behavior.

Review

1. What do quantum numbers do?
2. What is the principal quantum number?
3. What does the spin quantum number represent?

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5.7: Orbitals



Figure 5.7.1 (Public Domain; User:Baseball Bugs/Wikimedia Commons via Commons Wikimedia, Plane over Lake Harriet [commons.wikimedia.org])

How is it that so many planes are able to fly without running into each other?

The flight path of a commercial airliner is carefully regulated by the Federal Aviation Administration. Each airplane must maintain a distance of five miles from another plane flying at the same altitude, and be 2,000 feet above and below another aircraft (1,000 feet if the altitude is less than 29,000 feet). So, each aircraft only has certain positions it is allowed to maintain while it flies. Quantum mechanics demonstrates that electrons have similar restrictions on their locations.

Orbitals

We can apply our knowledge of quantum numbers to describe the arrangement of electrons for a given atom. We do this with something called **electron configurations**. They are effectively a map of the electrons for a given atom. We look at the four quantum numbers for a given electron and then assign that electron to a specific orbital.

s Orbitals

For any value of n, a value of $l = 0$ places that electron in an s **orbital**. This orbital is spherical in shape:

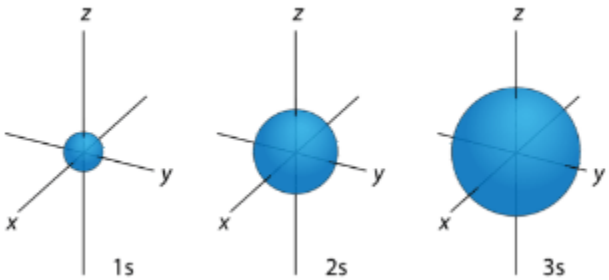


Figure 5.7.2 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

p Orbitals

From the table below, it is evident that there are three possible orbitals when $l = 1$. These are designated as p **orbitals** and have

dumbbell shapes. Each of the p orbitals has a different orientation in three-dimensional space.

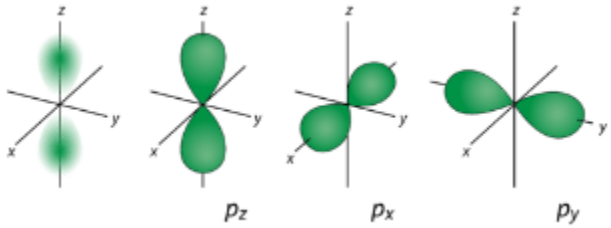


Figure 5.7.3 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

d Orbitals

When $l = 2$, m_l values can be $-2,\ -1,\ 0,\ +1,\ +2$ for a total of five d **orbitals**. Note that all five of the orbitals have specific three-dimensional orientations.

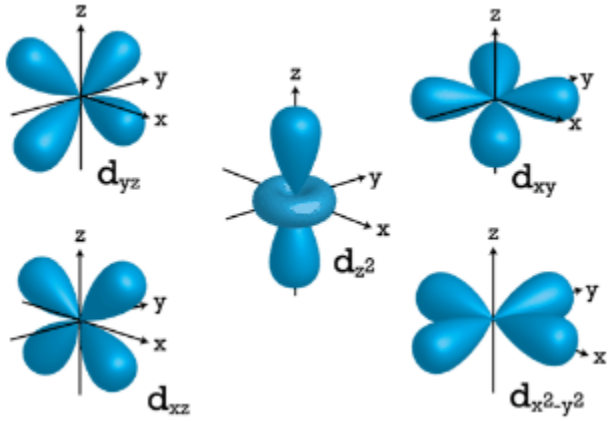


Figure 5.7.4 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

f Orbitals

The most complex set of orbitals are the f **orbitals**. When $l = 3$, m_l values can be $-3,\ -2,\ -1,\ 0,\ +1,\ +2,\ +3$ for a total of seven different orbital shapes. Again, note the specific orientations of the different f orbitals.

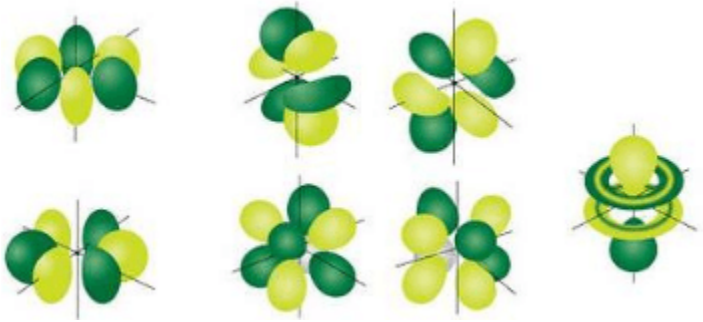


Figure 5.7.5 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Table 5.7.1: Electron Arrangement Within Energy Levels

Principal Quantum Number (n)	Allowable Sublevels	Number of Orbitals per Sublevel	Number of Orbitals per Principal Energy Level	Number of Electrons per Sublevel	Number of Electrons per Principal Energy Level
1	s	1	1	2	2

Principal Quantum Number (n)	Allowable Sublevels	Number of Orbitals per Sublevel	Number of Orbitals per Principal Energy Level	Number of Electrons per Sublevel	Number of Electrons per Principal Energy Level
2	s	1	4	2	8
	p	3		6	
3	s	1	9	2	18
	p	3		6	
	d	5		10	
4	s	1	16	2	32
	p	3		6	
	d	5		10	
	f	7		14	

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Summary

- There are four different classes of electron orbitals.
- Electron orbitals are determined by the value of the angular momentum quantum number l.

Review

1. What is an electron configuration?
2. How many electrons are in the $n = 1$ orbital?
3. What is the total number of electrons in a p orbital?
4. How many electrons does it take to completely fill a d orbital?

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5.8: Aufbau Principle



Figure 5.8.1 (Public Domain; Gary Minnaert via Commons Wikimedia, File:LACMA BCAM02.jpg [commons.wikimedia.org])

How are buildings constructed?

Construction of a building begins at the bottom. The foundation is laid, and the building goes up step by step. You obviously cannot start with the roof, since there is no place to hang it. The building goes from the lowest level to the highest level in a systematic way.

Aufbau Principle

In order to create ground state electron configurations for any element, it is necessary to know the way in which the atomic sublevels are organized in order of increasing energy. Figure 5.8.2 shows the order of increasing energy of the sublevels.

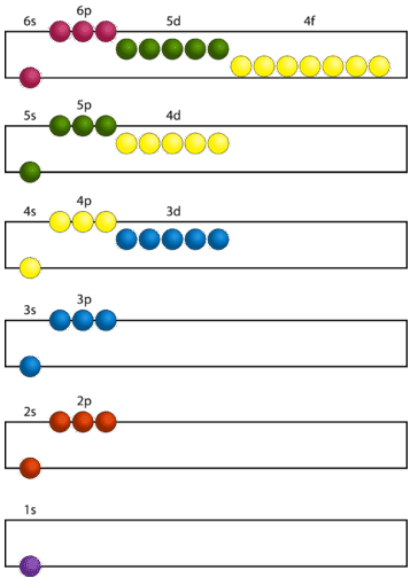


Figure 5.8.2: Electrons are added to atomic orbitals in order from low energy (bottom of the graph) to high (top of the graph), according to the Aufbau principle. Principle energy levels are color coded, while sublevels are grouped together, and each circle represents an orbital capable of holding two electrons. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The lowest energy sublevel is always the $1s$ sublevel, which consists of one orbital. The single electron of the hydrogen atom will occupy the $1s$ orbital when the atom is in its ground state. As we proceed to atoms with multiple electrons, those electrons are added to the next lowest sublevel: $2s$, $2p$, $3s$, and so on. The **Aufbau principle** states that an electron occupies orbitals in order from lowest energy to highest. The Aufbau (German for *building up, construction*) principle is sometimes referred to as the "building up" principle. It is worth noting that in reality, atoms are not built by adding protons and electrons one at a time, and that this method is merely an aid to understand the end result.

As seen in the figure above, the energies of the sublevels in different principal energy levels eventually begin to overlap. After the $3p$ sublevel, it would seem logical that the $3d$ sublevel should be the next lowest in energy. However, the $4s$ sublevel is slightly lower in energy than the $3d$ sublevel and thus fills first. Following the filling of the $3d$ sublevel is the $4p$, then the $5s$ and the $4d$. Note that the $4f$ sublevel does not fill until just after the $6s$ sublevel. Figure $5.8.2$ is a useful and simple aid for keeping track of the order of fill of the atomic sublevels.

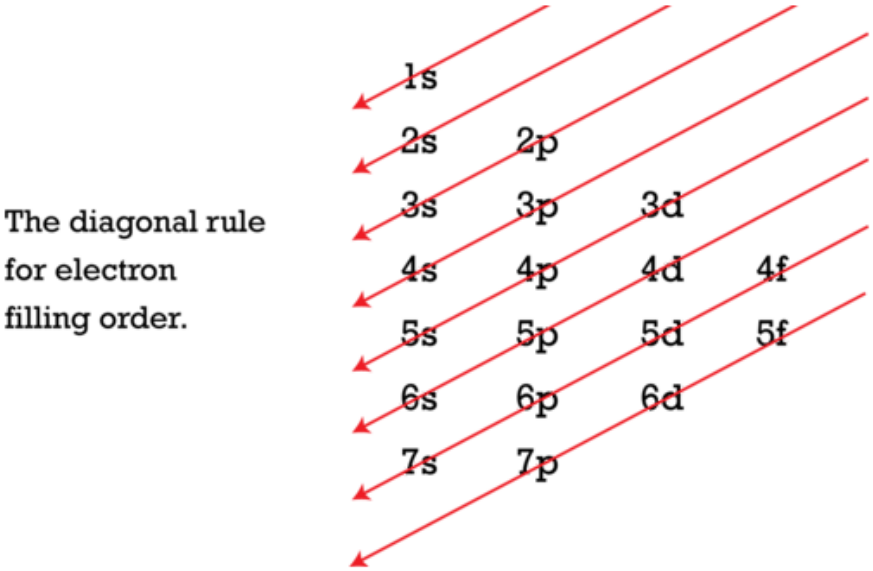


Figure $5.8.3$: The Aufbau principle is illustrated in the diagram by following each red arrow in order from top to bottom: $1s$, $2s$, $2p$, $3s$, etc. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

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Summary

- The Aufbau principle gives the order of electron filling in an atom.
- It can be used to describe the locations and energy levels of every electron in a given atom.

Review

1. What is the Aufbau principle?
2. Which orbital is filled after the $2p$?
3. Which orbital is filled after $4s$?
4. Which orbital is filled after $6s$?

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5.9: Pauli Exclusion Principle



Figure 5.9.1 (Pixelbay License; Pixabay:Nemo via http://pixabay.com/en/mail-internet-...ectronic-35636)

Can you name one thing that easily distinguishes you from the rest of the world?

And we're not talking about DNA—that's a little expensive to sequence. For many people, it is their email address. My email address allows people all over the world to contact me. It does not belong to anyone else, but serves to identify me. Electrons also have a unique set of identifiers in the quantum numbers that describe their location and spin.

Pauli Exclusion Principle

When we look at the orbital possibilities for a given atom, we see that there are different arrangements of electrons for each different type of atom. Since each electron must maintain its unique identity, we intuitively sense that the four quantum numbers for any given electron must not match up exactly with the four quantum numbers for any other electron in that atom.

For the hydrogen atom, there is no problem since there is only one electron in the H atom. However, for helium, we see that the first three quantum numbers for the two electrons are the same: same energy level, same spherical shape. What differentiates the two helium electrons is their spin. One of the electrons has a $+\frac{1}{2}$ spin while the other electron has a $-\frac{1}{2}$ spin. So, the two electrons in the $1s$ orbital are each unique and distinct from one another because their spins are different. This observation leads to the **Pauli exclusion principle**, which states that no two electrons in an atom can have the same set of four quantum numbers. The energy of the electron is specified by the principal, angular momentum, and magnetic quantum numbers. If those three numbers are identical for two electrons, the spin numbers must be different in order for the two electrons to be differentiated from one another. The two values of the spin quantum number allow each orbital to hold two electrons. Figure 5.9.2 shows how the electrons are indicated in a diagram.



Figure 5.9.2: In an orbital filling diagram, a square represents an orbital, while arrows represent electrons. An arrow pointing upward represents one spin direction, while an arrow pointing downward represents the other spin direction. (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)

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Summary

- The Pauli exclusion principle specifies limits on how identical quantum numbers can be for two electrons in the same atom.

Review

1. What is the difference between the two helium electrons?
2. What does the Pauli exclusion principle state?
3. What does the two values for the spin quantum number allow?

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5.10: Electron Configurations

```
C:\Temp> dir
 Volume in drive C is C
 Volume Serial Number is 74F5-B93C

 Directory of C:\Temp

2009-08-25  11:59    <DIR>          .
2009-08-25  11:59    <DIR>          ..
2007-03-01  11:37         2,321,600 AdobeUpdater12345.exe
2009-04-03  10:01            27,988 dd_depcheckdotnetfx30.txt
2009-04-03  10:01               764 dd_dotnetfx3error.txt
2009-04-03  10:01            32,572 dd_dotnetfx3install.txt
2009-06-09  13:46            35,145 GenProfile.log
2009-08-05  12:11               155 KB969856.log
2009-04-20  08:37               402 MSI29e0b.LOG
2009-04-09  16:34            38,895 offcln11.log
2009-04-03  16:02    <DIR>          OfficePatches
2009-07-14  14:30    <DIR>          OHotfix
2009-08-25  10:52            16,384 Perflib_Perfdata_c30.dat
2009-04-03  10:01             1,744 uxeventlog.txt
2009-08-25  11:42        50,245,632 WFV2F.tmp
2009-04-20  10:07             1,397 {AC76BA86-7AD7-1033-7B44-A81200000003}.ini
2009-04-20  10:13               617 {AC76BA86-7AD7-1033-7B44-A81300000003}.ini
              13 File(s)     52,723,295 bytes
               4 Dir(s)  83,570,208,768 bytes free
```

Figure $5.10.1$ (Public Domain; David R. Tribble (User:Loadmaster/Wikimedia Commons) via Commons wikimedia, Dir Command in Windows Command Prompt [commons.wikimedia.org])

How big is a file?

If you keep your papers in manila folders, you can pick up a folder and see how much it weighs. If you want to know how many different papers (articles, bank records, or whatever else you keep in a folder), you have to take everything out and count. A computer directory, on the other hand, tells you exactly how much you have in each file. We can get the same information on atoms. If we use an orbital filling diagram, we have to count arrows. When we look at electron configuration data, we simply add up the numbers.

Electron Configurations

Electron configuration notation eliminates the boxes and arrows of orbital filling diagrams. Each occupied sublevel designation is written followed by a superscript that is the number of electrons in that sublevel. For example, the hydrogen configuration is $1s^1$, while the helium configuration is $1s^2$. Multiple occupied sublevels are written one after another. The electron configuration of lithium is $1s^2 2s^1$. The sum of the superscripts in an electron configuration is equal to the number of electrons in that atom, which is in turn equal to its atomic number.

Example $5.10.1$

Draw the orbital filling diagram for carbon and write its electron configuration.

Solution

Step 1: List the known quantities and plan the problem.

Known

- atomic number of carbon, Z=6

Use the order of fill diagram to draw an orbital filling diagram with a total of six electrons. Follow Hund's rule. Write the electron configuration.

Step 2: Construct diagram.

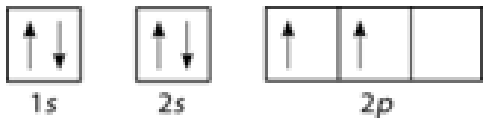


Figure 5.10.2: Orbital filling diagram for carbon. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Electron configuration $1s^2 2s^2 2p^2$

Step 3: Think about your result.

Following the $2s$ sublevel is the $2p$, and p sublevels always consist of three orbitals. All three orbitals need to be drawn even if one or more is unoccupied. According to Hund's rule, the sixth electron enters the second of those p orbitals, with the same spin as the fifth electron.

Second Period Elements

Periods refer to the horizontal rows of the periodic table. The first period of the periodic table contains only the elements hydrogen and helium. This is because the first principal energy level consists of only the s sublevel, and so only two electrons are required in order to fill the entire principal energy level. Each time a new principal energy level begins, as with the third element lithium, a new period is started on the periodic table. As one moves across the second period, electrons are successively added. With beryllium $(Z = 4)$, the $2s$ sublevel is complete and the $2p$ sublevel begins with boron $(Z = 5)$. Since there are three $2p$ orbitals and each orbital holds two electrons, the $2p$ sublevel is filled after six elements. Table 5.10.1 shows the electron configurations of the elements in the second period.

Table 5.10.1: Electron Configurations of Second-Period Elements

Element Name	Symbol	Atomic Number	Electron Configuration
Lithium	Li	3	$1s^2 2s^1$
Beryllium	Be	4	$1s^2 2s^2$
Boron	B	5	$1s^2 2s^2 2p^1$
Carbon	C	6	$1s^2 2s^2 2p^2$
Nitrogen	N	7	$1s^2 2s^2 2p^3$
Oxygen	O	8	$1s^2 2s^2 2p^4$
Fluorine	F	9	$1s^2 2s^2 2p^5$
Neon	Ne	10	$1s^2 2s^2 2p^6$

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Summary

- Electron configuration notation simplifies the indication of where electrons are located in a specific atom.
- Superscripts are used to indicate the number of electrons in a given sublevel.

Review

1. What does electron configuration notation eliminate?
2. How do we know how many electrons are in each sublevel?
3. An atom has the electron configuration of $1s^2 2s^2 2p^5$. How many electrons are in that atom?
4. Which element has the electron configuration of $1s^2 2s^2 2p^6 3s^2$?

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5.11: Valence Electrons



Figure 5.11.1 (Public Domain; User:Chemicalinterest/Wikipedia via Wikipedia)

What makes a particular element very reactive and another element non-reactive?

A chemical reaction involves either electron removal, electron addition, or electron sharing. The path that a specific element will take in a reaction depends on where the electrons are in the atom and how many there are.

Valence Electrons

In the study of chemical reactivity, electrons in the outermost principal energy level are very important and so are given a special name. **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? 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 $2s^2 2p^6$, has eight valence electrons.

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Summary

- Valence electrons are the outer-shell electrons of an atom.
- Valence electrons determine the reactivity of an atom.

Review

1. Define valence electron.
2. Define inner shell electron.
3. How many valence electrons are there in fluorine?
4. What are the 2s electrons in nitrogen?
5. How many inner shell electrons are there in beryllium?

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5.12: Noble Gas Configuration



Figure 5.12.1 (Public Domain; User:Danbold/Wikipedia via Wikipedia)

How does it feel to be full after a meal?

Envision that you have nearly finished a great meal, but cannot put another bite in your mouth because there is no place for it to go. The noble gases have the same problem—there is no room for any more electrons in their outer shells. They are completely full and cannot handle any more.

Noble Gas Configuration

Sodium, element number 11, is the first element in the third period of the periodic table. Its electron configuration is $1s^2 2s^2 2p^6 3s^1$. The first ten electrons of the sodium atom are the inner-shell electrons and the configuration of just those ten electrons is exactly the same as the configuration of the element neon $(Z = 10)$. This provides the basis for a shorthand notation for electron configurations called the noble gas configuration. The elements that are found in the last column of the periodic table are an important group of elements called the noble gases. They are helium, neon, argon, krypton, xenon, and radon. A **noble gas configuration** of an atom consists of the elemental symbol of the last noble gas prior to that atom, followed by the configuration of the remaining electrons. So for sodium, we make the substitution of $[\text{Ne}]$ for the $1s^2 2s^2 2p^6$ part of the configuration. Sodium's noble gas configuration becomes $[\text{Ne}]\, 3s^1$. Table 5.12.1 shows the noble gas configurations of the third period elements.

Table 5.12.1: Electron Configurations of Third-Period Elements

Element Name	Symbol	Atomic Number	Noble Gas Electron Configuration
Sodium	Na	11	$[\text{Ne}]\, 3s^1$
Magnesium	Mg	12	$[\text{Ne}]\, 3s^2$
Aluminum	Al	13	$[\text{Ne}]\, 3s^2 3p^1$
Silicon	Si	14	$[\text{Ne}]\, 3s^2 3p^2$
Phosphorus	P	15	$[\text{Ne}]\, 3s^2 3p^3$
Sulfur	S	16	$[\text{Ne}]\, 3s^2 3p^4$
Chlorine	Cl	17	$[\text{Ne}]\, 3s^2 3p^5$
Argon	Ar	18	$[\text{Ne}]\, 3s^2 3p^6$

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Again, the number of valence electrons increases from one to eight across the third period.

The fourth and subsequent periods follow the same pattern, except for the use of a different noble gas. Potassium has nineteen electrons, one more than the noble gas argon, so its configuration could be written as $[\mathrm{Ar}]\,4s^1$. In a similar fashion, strontium has two more electrons than the noble gas krypton, which would allow us to write its electron configuration as $[\mathrm{Kr}]\,5s^2$. All elements can be represented in this fashion.

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Summary

- The noble gas configuration system allows some shortening of the total electron configuration by using the symbol for the noble gas of the previous period as part of the pattern of electrons.

Review

1. What is the element represented by $[\mathrm{Ne}]\,3s^2 3p^2$?
2. What element has this electron configuration $[\mathrm{Ar}]\,3d^7 4s^2$?
3. What noble gas would be part of the electron configuration notation for Mn?
4. How would you write the electron configuration for Ba?

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5.13: Hund's Rule and Orbital Filling Diagrams



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Have you ever wondered what those load limit signs mean on a bridge?

The sign above signifies that nothing over five tons is allowed because it will do damage to the structure. There are limits to the amount of weight that a bridge can support, there are limits to the number of people that can safely occupy a room, and there are limits to what can go into an electron orbital.

Hund's Rule

The last of the three rules for constructing electron arrangements requires electrons to be placed one at a time in a set of orbitals within the same sublevel. This minimizes the natural repulsive forces that one electron has for another. **Hund's rule** states that orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and that each of the single electrons must have the same spin. The figure below shows how a set of three p orbitals is filled with one, two, three, and four electrons.

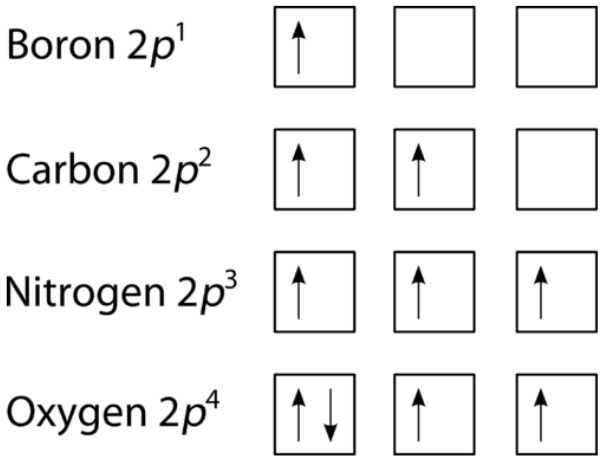


Figure 5.13.2: The $2p$ sublevel, for the elements boron $(Z = 5)$, carbon $(Z = 6)$, nitrogen $(Z = 7)$, and oxygen $(Z = 8)$. According to Hund's rule, as electrons are added to a set of orbitals of equal energy, one electron enters each orbital before any orbital receives a second electron. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Orbital Filling Diagrams

An **orbital filling diagram** is the more visual way to represent the arrangement of all the electrons in a particular atom. In an orbital filling diagram, the individual orbitals are shown as circles (or squares) and orbitals within a sublevel are drawn next to each other horizontally. Each sublevel is labeled by its principal energy level and sublevel. Electrons are indicated by arrows inside the circles. An arrow pointing upwards indicates one spin direction, while a downward pointing arrow indicates the other direction. The orbital filling diagrams for hydrogen, helium, and lithium are shown in the figure below.

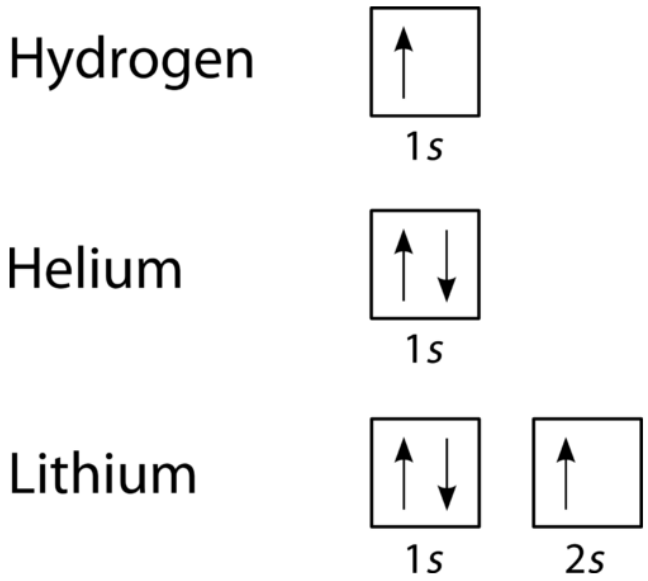


Figure 5.13.3: Orbital filling diagrams for hydrogen, helium, and lithium. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

According to the Aufbau process, sublevels and orbitals are filled with electrons in order of increasing energy. Since the s sublevel consists of just one orbital, the second electron simply pairs up with the first electron as in helium. The next element is lithium and necessitates the use of the next available sublevel, the $2s$.

The filling diagram for carbon is shown in the figure below. There are two $2p$ electrons for carbon and each occupies its own $2p$ orbital.

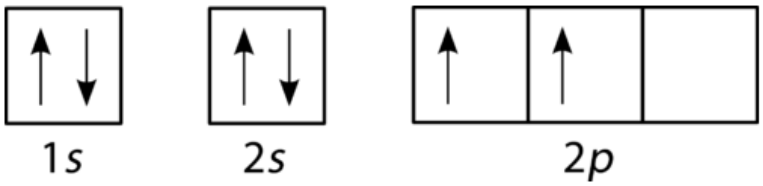


Figure 5.13.4: Orbital filling diagram for carbon. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Oxygen has four $2p$ electrons. After each $2p$ orbital has one electron in it, the fourth electron can be placed in the first $2p$ orbital with a spin opposite that of the other electron in that orbital.

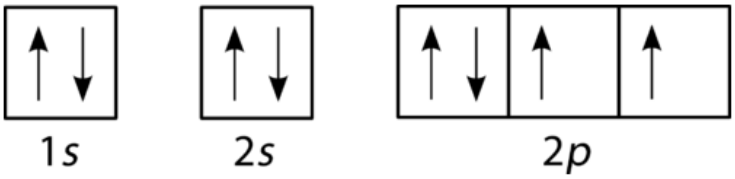


Figure 5.13.5: Orbital filling diagram for oxygen. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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Summary

- Hund's rule specifies the order of electron filling within a set of orbitals.
- Orbital filling diagrams are a way of indicating electron locales in orbitals.

Review

1. State Hund's rule.
2. What is an orbital filling diagram?
3. Is the diagram in figure below correct? Explain your answer.

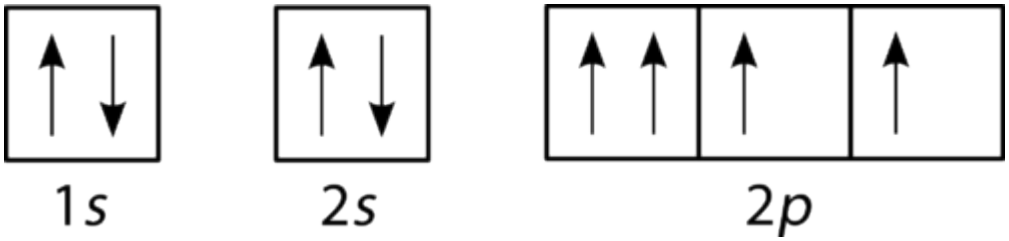


Figure 5.13.6 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

4. Is the diagram in figure below correct? Explain your answer.

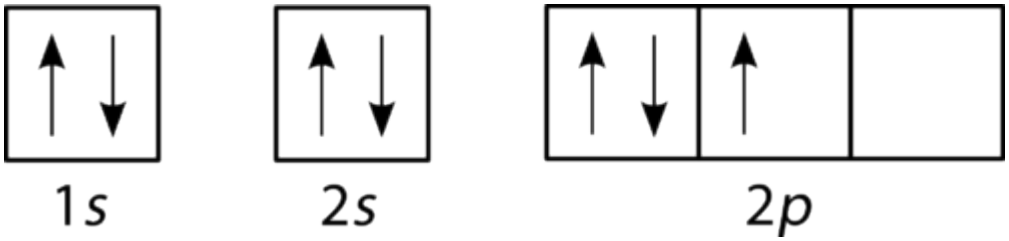


Figure 5.13.7 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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6: THE PERIODIC TABLE

The modern periodic table emphasizes the electronic structure of atoms. The original periodic table emphasized the reactivity of the elements. In this chapter we will learn about the connection between the two, and how the electronic structure is related to the macroscopic properties of the elements.

6.1: Early History of the Periodic Table

6.2: Mendeleev's Periodic Table

6.3: Periodic Law

6.4: Modern Periodic Table- Periods and Groups

6.5: Metals

6.6: Nonmetals

6.7: Metalloids

6.8: Blocks of the Periodic Table

6.9: Hydrogen and Alkali Metals

6.10: Alkaline Earth Metals

6.11: Noble Gases

6.12: Halogens

6.13: Transition Elements

6.14: Periodic Trends- Atomic Radius

6.15: Periodic Trends- Ionization Energy

6.16: Ionic Radii

6.17: Metallic and Nonmetallic Character

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6.1: Early History of the Periodic Table



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When you go to the library to find a book, how do you locate it?

If it is a fiction book, you look by author since the fiction materials are filed by the author's last name. If you are looking for a non-fiction publication, you look in a catalog (most likely on a computer these days). The book you are looking for will have a number by the title. This number refers to the Dewey Decimal system, developed by Melvil Dewey in 1876 and used in over 200,000 libraries throughout the world. Another system in wide use is the Library of Congress approach, developed in the late 1800s-early 1900s to organize the materials in the federal Library of Congress. This method is one of the most widely used ways to organize libraries in the world. Both approaches organize information so that people can easily find what they are looking for. Chemistry information also needs to be organized so we can see patterns of properties in elements.

Early Attempts to Organize Elements

By the year 1700, only a handful of elements had been identified and isolated. Several of these, such as copper and lead, had been known since ancient times. As scientific methods improved, the rate of discovery dramatically increased. With the ever-increasing number of elements, chemists recognized that there may be some kind of systematic way to organize the elements. The question was: how?

A logical way to begin grouping elements together was by their chemical properties. (In other words, putting elements in separate groups based on how they reacted with other elements.) In 1829, a German chemist, Johann Dobereiner (1780-1849), placed various groups of three elements into groups called **triads**. One such triad was lithium, sodium, and potassium. Triads were based on both physical, as well as chemical, properties. Dobereiner found that the atomic masses of these three elements, as well as other triads, formed a pattern. When the atomic masses of lithium and potassium were averaged together $\left(\frac{(6.94+39.10)}{2} = 23.02\right)$, it was approximately equal to the atomic mass of sodium (22.99). These three elements also displayed similar chemical reactions, such as vigorously reacting with the members of another triad: chlorine, bromine, and iodine. While Dobereiner's system would pave the way for future ideas, a limitation of the triad system was that not all of the known elements could be classified in this way.

English chemist John Newlands (1838-1898) ordered the elements in increasing order of atomic mass and noticed that every eighth element exhibited similar properties. He called this relationship the "Law of **Octaves**". Unfortunately, there were some elements that were missing and the law did not seem to hold for elements that were heavier than calcium. Newlands' work was largely ignored and even ridiculed by the scientific community in his day. It was not until years later that another more extensive periodic table effort would gain much greater acceptance, and that the pioneering work of John Newlands would be appreciated.

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No.	No.	No.	No.	No.	No.	No.	No.
H 1	F 8	Cl 15	Co & Ni 22	Br 29	Pd 36	I 42	Pt & Ir 50
Li 2	Na 9	K 16	Cu 23	Rb 30	Ag 37	Cs 44	Os 51
G 3	Mg 10	Ca 17	Zn 24	Sr 31	Cd 38	Ba & V 45	Hg 52
Bo 4	Al 11	Cr 19	Y 25	Ce & La 33	U 40	Ta 46	Tl 53
C 5	Si 12	Ti 18	In 26	Zr 32	Sn 39	W 47	Pb 54
N 6	P 13	Mn 20	As 27	Di & Mo 34	Sb 41	Nb 48	Bi 55
O 7	S 14	Fe 21	Se 28	Ro & Ru 35	Te 43	Au 49	Th 56

Figure 6.1.2 (Public Domain; John Alexander Reina Newlands via Commons Wikimedia, Newlands Periodiska System 1866 [commons.wikimedia.org])

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Summary

- Johann Dobereiner organized elements into groups called triads.
- John Newlands proposed the "Law of Octaves" for organizing the elements.

Review

1. List some elements known since ancient times?
2. What properties were the basis of the triad system?
3. Why did Dobereiner believe that lithium, sodium, and potassium belonged in a triad?
4. What was a shortcoming of the triad system?
5. How did Newlands arrange the element?
6. What was a problem with the “Law of Octaves"?

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6.2: Mendeleev's Periodic Table



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When you study for a test, how do you approach the task?

One useful way to study for a test is to use flashcards. To make them, write down vocabulary words, foreign language terms, math formulas, chemistry reactions—anything you want to learn. Then sort these cards into categories and topics that go together. This organization of information helps you see patterns in the material so that you can tie different ideas together and make better sense of them.

The periodic table was first built using a set of cards. With this strategy, Dmitri Mendeleev, the creator of the periodic table, could organize and rearrange material until patterns emerged.

Mendeleev's Periodic Table

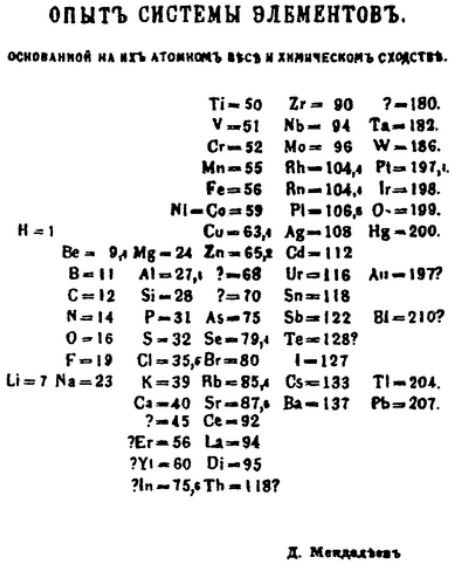
In 1869, Russian chemist and teacher Dmitri Mendeleev (1836-1907) published a periodic table of the elements. The following year, German chemist Lothar Meyer independently published a very similar table. Mendeleev is generally given more credit than Meyer because his table was published first, and because of several key insights that he made regarding the table.



Figure 6.2.2: Dmitri Mendeleev. (Public Domain; Unknown via Wikipedia)

Mendeleev was writing a chemistry textbook for his students and wanted to organize all of the known elements at that time

according to their chemical properties. He famously organized the information for each element onto separate note cards that were then easy to rearrange as needed. He discovered that when he placed them in order of increasing atomic mass, certain similarities in chemical behavior repeated at regular intervals. This type of a repeating pattern is called "periodic". A pendulum that swings back and forth in a given time interval is periodic, as is the movement of the moon around the Earth.



ОПЫТЪ СИСТЕМЫ ЭЛЕМЕНТОВЪ.

ОСНОВАННОЙ НА ИХЪ АТОМНОМЪ ВѢСѢ И ХИМИЧЕСКОМЪ СХОДСТВѢ.

		Ti = 50	Zr = 90	? = 180.
		V = 51	Nb = 94	Ta = 182.
		Cr = 52	Mo = 96	W = 186.
		Mn = 55	Rh = 104,4	Pt = 197,1.
		Fe = 56	Ru = 104,4	Ir = 198.
		Ni = Co = 59	Pl = 106,6	O = 199.
H = 1		Cu = 63,4	Ag = 108	Hg = 200.
	Be = 9,4	Mg = 24	Zn = 65,2	Cd = 112	
	B = 11	Al = 27,4	? = 68	Ur = 116	Au = 197?
	C = 12	Si = 28	? = 70	Sn = 118	
	N = 14	P = 31	As = 75	Sb = 122	Bi = 210?
	O = 16	S = 32	Se = 79,4	Te = 128?	
	F = 19	Cl = 35,6	Br = 80	I = 127	
Li = 7	Na = 23	K = 39	Rb = 85,4	Cs = 133	Tl = 204.
		Ca = 40	Sr = 87,6	Ba = 137	Pb = 207.
		? = 45	Ce = 92		
		?Er = 56	La = 94		
		?Yt = 60	Di = 95		
		?In = 75,6	Th = 118?		

Д. Менделѣевъ

Figure 6.2.3: Mendeleev's 1869 periodic table. (Public Domain; Dmitri Mendeleev via Wikipedia)

In the figure above, atomic mass increases from top to bottom of vertical columns, with successive columns going left to right. As a result, elements that are in the same horizontal row are groups of elements that were known to exhibit similar chemical properties. One of Mendeleev's insights is illustrated by the elements tellurium (Te) and iodine (I). Notice that tellurium is listed before iodine even though its atomic mass is higher. Mendeleev reversed the order because he knew that the properties of iodine were much more similar to those of fluorine (F), chlorine (Cl), and bromine (Br) than they were to oxygen (O), sulfur (S), and selenium (Se). He simply assumed that there was an error in the determination of one or both of the atomic masses. This turned out not to be the case, but Mendeleev was indeed correct to group these two elements as he did.

Notice that there are several places in the table that have no chemical symbol, but are instead labeled with a question mark. Between zinc (Zn) and arsenic (As) are two such missing elements. Mendeleev believed that elements with atomic masses of 68 and 70 would eventually be discovered and that they would fit chemically into each of those spaces. Listed in the table below are other properties that Mendeleev predicted for the first of these two missing elements, which he called "eka-aluminum", compared with the element gallium.

Mendeleev's predicted properties for Eka-aluminum and gallium.

	Eka-Aluminum (Ea)	**Gallium** (Ga)
Atomic Mass	$68\ \mathrm{amu}$	$69.9\ \mathrm{amu}$
Melting Point	Low	$30.15^{\circ}\mathrm{C}$
Density	$5.9\ \mathrm{g/cm^3}$	$5.94\ \mathrm{g/cm^3}$
Formula of Oxide	$\mathrm{Ea_2O_3}$	$\mathrm{Ga_2O_3}$

The element gallium was discovered four years after the publication of Mendeleev's table, and its properties matched up remarkably well with eka-aluminum, fitting into the table exactly where he had predicted. This was also the case with the element that followed gallium, which was eventually named germanium.

Mendeleev's periodic table gained wide acceptance with the scientific community and earned him credit as the discoverer of the periodic law. Element number 101, synthesized in 1955, is named **mendelevium** after the founder of the periodic table. It was, however, several years after Mendeleev died before the several discrepancies with the atomic masses could be explained, and before the reasons behind the repetition of chemical properties could be fully explained.

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Summary

- Mendeleev published his periodic table in 1869.
- His organization of elements was based on atomic mass.
- Mendeleev's periodic table made it possible to predict properties of elements that had not yet been discovered.

Review

1. When did Mendeleev publish his periodic table?
2. Who else came out with a periodic table at about the same time?
3. Why was Mendeleev's table considered to be superior?
4. What element did Mendeleev predict to exist?
5. What element was named after Mendeleev?

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6.3: Periodic Law



Figure 6.3.1: Copy and Paste Caption here. (CC-BY 2.0; Christopher Fynn via Flickr, Mysore - Fruit vendor [www.flickr.com])

How are these items related to one another?

We have all enjoyed looking around a market for delicious foods to eat later at home. When you get to the market you know you need to get fruits, vegetables and grains to ensure you eat a balanced diet. In the market, these items are all grouped together to make it easier to find the type of fruit, vegetable or grain you are looking for; it would be inconvenient if the different fruits were all scattered in different places. The periodic table is organized in a similar way, ensuring similar elements are found in the same group or period.

The Periodic Law

When Mendeleev put his periodic table together, nobody knew about the existence of the nucleus. It was not until 1911 that Rutherford conducted his gold foil experiment that demonstrated the presence of the nucleus in the atom. Just two years later, in 1913, English physicist Henry Moseley (1887-1915) examined x-ray spectra of a number of chemical elements. He would shoot x-rays through crystals of the element and study the wavelengths of the radiation that he detected. Moseley found that there was a relationship between wavelength and atomic number. His results led to the definition of atomic number as the number of protons contained in the nucleus of each atom. He then realized that the elements of the periodic table should be arranged in order of increasing atomic number, rather than increasing atomic mass.

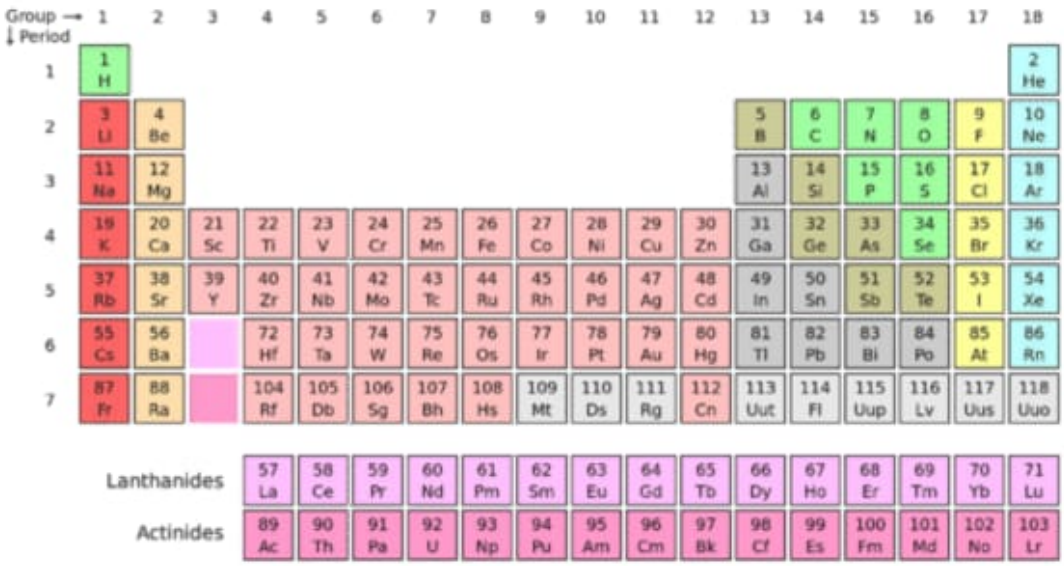


Figure 6.3.2 (Public Domain; User:Cepheus/Wikimedia Commons via Commons Wikimedia, Periodic Table [commons.wikimedia.org])

When ordered by atomic number, the discrepancies within Mendeleev's table disappeared. Tellurium has an atomic number of 52, while iodine has an atomic number of 53. So even though tellurium does indeed have a greater atomic mass than iodine, it is properly placed before iodine in the periodic table. Mendeleev and Moseley are credited with being most responsible for the modern **periodic law**: When elements are arranged in order of increasing atomic number, there is a periodic repetition of their chemical and physical properties. The result is the periodic table as we know it today. Each new horizontal row of the periodic table corresponds to the beginning of a new **period** because a new principal energy level is being filled with electrons. Elements with similar chemical properties appear at regular intervals, within the vertical columns called **groups**.

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Summary

- Elements of the periodic table are arranged in order of increasing atomic number.
- The periodic law states: "When elements are arranged in order of increasing atomic number, there is a periodic repetition of their chemical and physical properties."

Review

1. Did Mendeleev know about the nucleus of an atom?
2. Who discovered the relationship between wavelength of X-rays and atomic number?
3. What did Moseley conclude from his research?
4. What is the “periodic law”?
5. What do the vertical columns (groups) in the periodic table represent?

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6.4: Modern Periodic Table- Periods and Groups

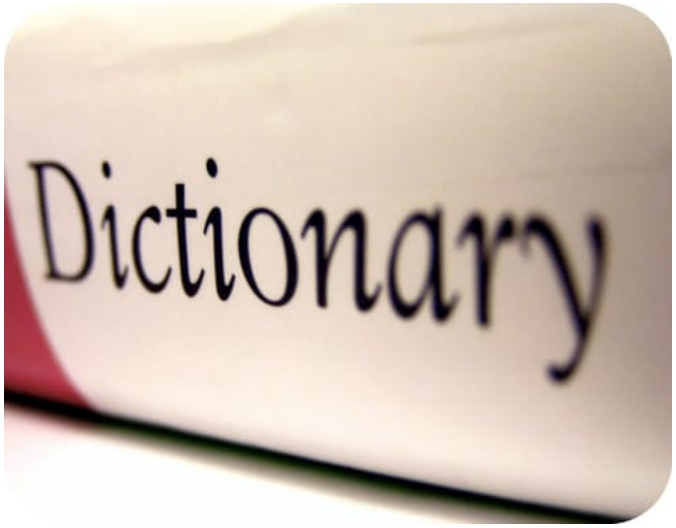


Figure 6.4.1 (Copyright; author via source)

How has the English dictionary evolved over time?

Language changes with time. New words enter the language and old words often disappear from lack of use. Dictionaries are published so that people can keep up with changes in language and know how to use words properly. These publications may be in print, as is the law dictionary below, or they may be electronic. Dictionaries can be found on the internet and apps are available for smartphones. Dictionaries are invaluable for good, reliable communication.

The Modern Periodic Table

The periodic table has undergone extensive changes in the time since it was originally developed by Mendeleev and Moseley. Many new elements have been discovered, while others have been artificially synthesized. Each fits properly into a **group** of elements with similar properties. The periodic table is an arrangement of the elements in order of their atomic numbers, so that elements with similar properties appear in the same vertical column or group.

The figure below shows the most commonly used form of the periodic table. Each square shows the chemical symbol of the element along with its name. Notice that several of the symbols seem to be unrelated to the name of the element: Fe for iron, Pb for lead, etc. Most of these are the elements that have been known since ancient times and have symbols based on their Latin names. The atomic number of each element is written above the symbol.

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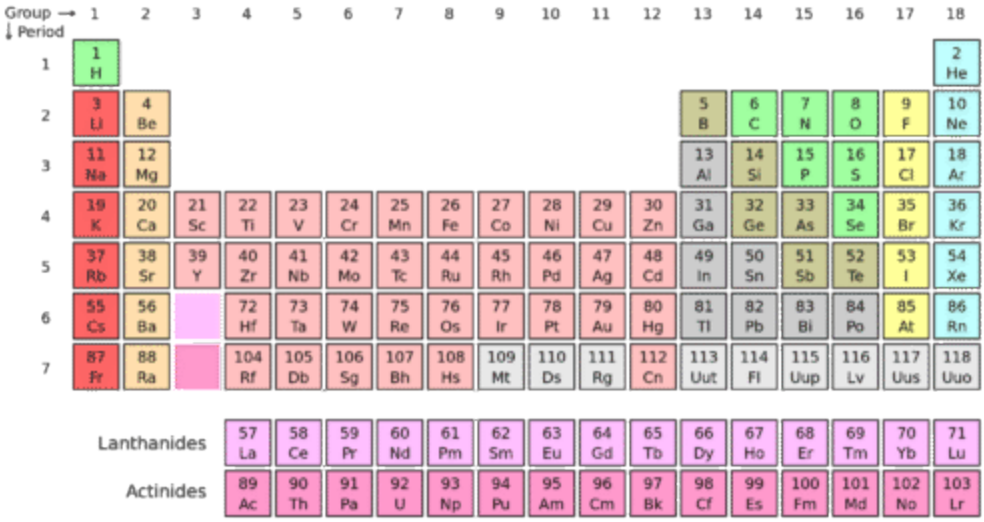


Figure 6.4.2 (Public Domain; User:Cepheus/Wikimedia Commons via Wikipedia)

A **period** is a horizontal row of the periodic table. There are seven periods in the periodic table, with each one beginning at the far left. A new period begins when a new principal energy level begins filling with electrons. Period 1 has only two elements (hydrogen and helium), while periods 2 and 3 have 8 elements. Periods 4 and 5 have 18 elements. Periods 6 and 7 have 32 elements, because the two bottom rows that are separate from the rest of the table belong to those periods. These two rows are pulled out in order to make the table itself fit more easily onto a single page.

A group is a vertical column of the periodic table, based on the organization of the outer shell electrons. There are a total of 18 groups. There are two different numbering systems that are commonly used to designate groups, and you should be familiar with both. The traditional system used in the United States involves the use of the letters A and B. The first two groups are 1A and 2A, while the last six groups are 3A through 8A. The middle groups use B in their titles. Unfortunately, there was a slightly different system in place in Europe. To eliminate confusion, the International Union of Pure and Applied Chemistry (IUPAC) decided that the official system for numbering groups would be a simple 1 through 18 from left to right. Many periodic tables show both systems simultaneously.

Most recent depictions of the periodic table show an incomplete seventh period. In fact, many of those seventh period elements were not known before the early twentieth century. Many elements have been synthesized by bombarding known elements with sub-atomic particles such as neutrons or alpha particles. Uranium has been used in this manner to produce elements 93-100. Larger elements (atomic numbers 101 and above) are formed by fusing nuclei of smaller elements together.

These synthetic elements tend to be very unstable, often existing for less than a second, so little is known about them. Elements up to atomic number 112 are known (not in any detail, however), and some evidence for elements 113 and above had been put forth, with elements 114 and 116 being added to the table in 2011. In December 2015, the IUPAC verified the existence of four new elements 113, 115, 117, and 118 and approved their addition to the periodic table.

The following are the new element names and their origins:

- Element 113 was named Nihonium, symbol Nh, proposed by Japanese researchers after the Japanese word Nihon, which means Japan.
- A team of scientists from Russia and the United States named element 115, Moscovium, symbol Mc, after Moscow and element 117, Tennessine, symbol Ts, after Tennessee.
- The Russian team that discovered element 118 named it Oganesson, symbol Og, after Yuri Oganessian, a prolific element hunter.

These elements complete the seventh period or row of the periodic table.

The video below reviews groups and trends in the periodic table:

YouTube Video

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Watch the additional videos below to learn more about the newly identified and named chemical elements:

Element Name and Symbol: Nihonium (Nh)

Atomic Number: 113

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Element Name and Symbol: Moscovium (Mc)

Atomic Number: 115

YouTube Video

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Element Name and Symbol: Tennessine (Ts)

Atomic Number: 117

YouTube Video

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Element Name and Symbol: Oganesson (Og)

Atomic Number: 118

YouTube Video

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Summary

- The periodic table is arranged in order of atomic number.
- A period is a horizontal row of the periodic table.
- A group is a vertical row of the periodic table.

Review

1. How is today's periodic table different from the one that Mendeleev published?
2. Are all the elements in today's periodic table naturally occurring? Explain your answer.
3. What is a "period"? What does it represent?
4. What is a "group"? What does it represent?
5. Why are there two different numbering systems for groups?

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6.5: Metals



Figure 6.5.1 (Public Domain; User:Brianiac/Wikipedia via Wikipedia)

Can you guess what types of metals these screws are made of?

Screws come in all sizes and shapes. They are all (well, almost all) made of some kind of metal. But they have differences in size, shape, and type of metal. Physical characteristics also differ. Some screws are long, and others are short. One screw may have a flat-head slot while another screw may have a Phillips-head. Some of the screws in the picture below are used to fasten things together, and others are used to hang heavy objects on a wall.

Chemists classify materials in many ways. We can sort elements on the basis of their electron arrangements. The way the electrons are distributed determines the chemical properties of the element. Another way is to classify elements based on physical properties. Some common physical properties are color, volume, and density. Other properties that allow us to sort on the basis of behavior are conduction of heat and electricity, malleability (the ability to be hammered into very thin sheets), ductility (the ability to be pulled into thin wires), melting point, and boiling point. Three broad classes of elements based on physical properties are metals, nonmetals, and metalloids.

Metals

A metal is an element that is a good conductor of heat and electricity. Metals are also malleable, which means that they can be hammered into very thin sheets without breaking. They are ductile, which means that they can be drawn into wires. When a fresh surface of any metal is exposed, it will be very shiny because it reflects light well. This is called luster. All metals are solid at room temperature with the exception of mercury (Hg), which is a liquid. Melting points of metals display a very wide variance. The melting point of mercury is -39°C, while the highest melting metal is tungsten (W), with a melting point of 3422°C. The elements in blue in the periodic table below are metals. About 80 percent of the elements are metals.

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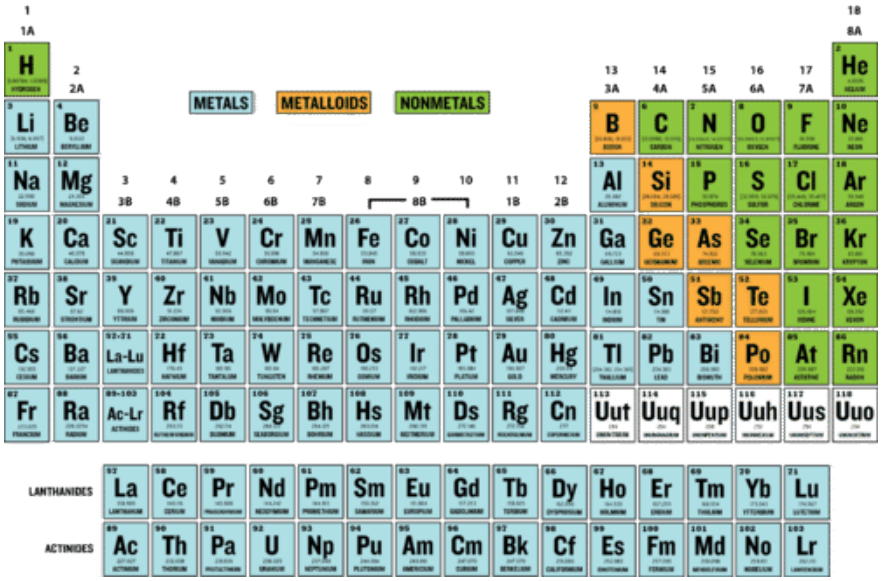


Figure 6.5.2 (CC BY-NC-SA 3.0; Christopher Auyeung via CK-12 Foundation)

Gold has been used by many civilizations for making jewelry (see figure below). This metal is soft and easily shaped into a variety of items. Since gold is very valuable and often used as currency, gold jewelry has also often represented wealth.



Figure 6.5.3: Gold jewelry. (Public Domain; User:Dorieo21/Wikimedia Commons via Wikipedia)

Copper is a good conductor of electricity and is very flexible and ductile. This metal is widely used to conduct electric current in a variety of appliances, from lamps to stereo systems to complex electronic devices (see figure below).



Figure 6.5.4: Copper wire exposed. (Public Domain; Scott Ehardt via Wikipedia)

Mercury is the only metal to exist as a liquid at room temperature (see figure below). This metal was extensively used in thermometers for decades until information about its toxicity became known. Mercury switches were once common, but are no longer used. However, new federally-mandated energy-efficient light bulbs that are now used contain trace amounts of mercury and represent a hazardous waste.



Figure 6.5.5: Pouring Mercury. (CC BY 3.0; User:Bionerd/Wikimedia Commons and User:Materialscientist/Wikimedia Commons via Wikipedia)

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Summary

- Metals are good conductors of heat and electricity.
- Metals are malleable and ductile
- All metals are solids at room temperature with the exception of mercury
- Gold, silver, iron, and mercury are typical metals.

Review

1. What properties of an element are affected by electron distribution?
2. Define malleability.
3. Define ductility.
4. State one reason gold is used in jewelry.
5. Why is mercury no longer used in many devices?

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6.6: Nonmetals



Figure 6.6.1 (Public Domain; Neville Street via Wikipedia)

When we sort parts in our shop or garage, we often classify them in terms of common properties. One container might hold all the screws (possibly sub-divided by size and type). Another container would be for nails. Maybe there is a set of drawers for plumbing parts.

When you get finished, you could also have a collection of things that don't nicely fit a category. You define them in terms of what they are not. They are not electrical components, or sprinkler heads for the yard, or parts for the car. These parts may have some common properties, but are a variety of items.

Nonmetals

In the chemical world, these "spare parts" would be considered nonmetals, loosely defined as not having the properties of metals. A **nonmetal** is an element that is generally a poor conductor of heat and electricity. Most properties of nonmetals are the opposite of metals. There is a wider variation in properties among the nonmetals than among the metals. Nonmetals exist in all three states of matter. The majority are gases, such as nitrogen and oxygen. Bromine is a liquid. A few are solids, such as carbon and sulfur. In the solid state, nonmetals are **brittle**, meaning that they will shatter if struck with a hammer. The solids are not lustrous. Melting points are generally much lower than those of metals. The green elements in the table below are nonmetals.

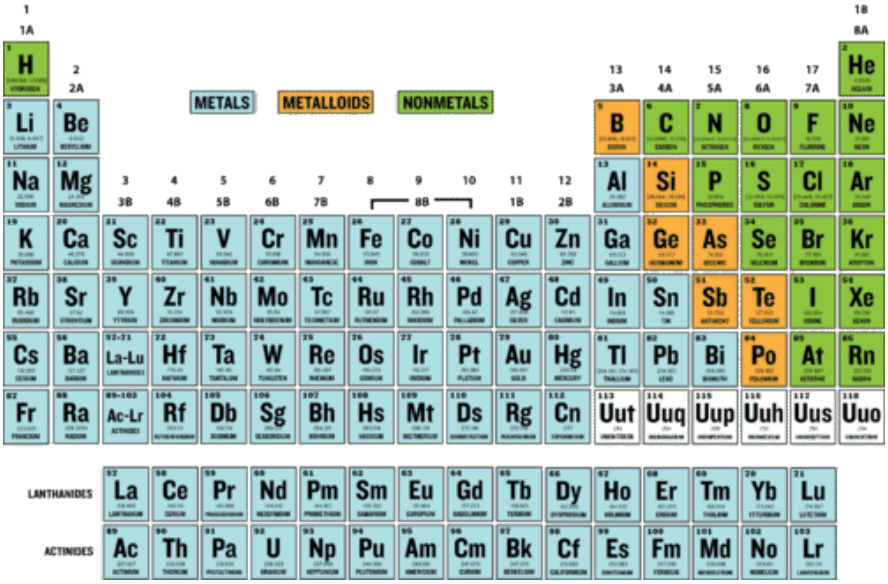


Figure 6.6.2: Copy and Paste Caption here. (CC BY-NC-SA 3.0; Christopher Auyeung via CK-12 Foundation)

Nonmetals have a wide variety of uses. Sulfur can be employed in gunpowder, fireworks, and matches to facilitate ignition (see

Figure below). This element is also widely used as an insecticide, a fumigant, or a means of eliminating certain types of fungus. An important role for sulfur is the manufacture of rubber for tires and other materials. First discovered in 1839 by Charles Goodyear, the process of vulcanization makes the rubber more flexible and elastic as well as being more resistant to changes in temperature. A major use of sulfur is for the preparation of sulfur-containing compounds such as sulfuric acid.



Figure 6.6.3: Sulfur. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

Bromine is a versatile compound, used mainly in manufacture of flame-retardant materials, especially important for children's clothing (see Figure below). For treatment of water in swimming pools and hot tubs, bromine is beginning to replace chlorine as a disinfectant because of its higher effectiveness. When incorporated into compounds, bromine atoms play important roles in pharmaceuticals for treatment of pain, cancer, and Alzheimer's disease.



Figure 6.6.4: Bromine. (CC BY 3.0; User:Jurii/Wikimedia Commons via Wikipedia)

Helium is one of the many nonmetals that is a gas. Other nonmetal gases include hydrogen, fluorine, chlorine, and all the group eighteen noble (or inert) gases. Helium is chemically non-reactive, so it is useful for applications such as balloons (see figure below) and lasers, where non-flammability is extremely important. Liquid helium exists at an extremely low temperature and can be used to cool superconducting magnets for imaging studies (MRI, magnetic resonance imaging). Leaks in vessels and many types of high-vacuum apparatus can be detected using helium. Inhaling helium changes the speed of sound, producing a higher pitch in your voice. This is definitely an unsafe practice and can lead to physical harm and death.



Figure 6.6.5: Blimp. (Public Domain; Derek Jensen (Wikimedia: Tysto) via Wikipedia)

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Summary

- Nonmetals are generally poor conductors of heat and electricity.
- Properties of nonmetals are usually the opposite of properties of metals
- Nonmetals can be solid, liquid, or gas at room temperature depending upon the element.
- Sulfur, bromine, and helium are typical nonmetals.

Review

1. What are the properties of nonmetals?
2. List the states of matter in which nonmetals can exist and give one example of each state.
3. What are the physical properties and uses of sulfur?
4. What are the physical properties and uses of bromine?
5. What are the physical properties and uses of helium?

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6.7: Metalloids

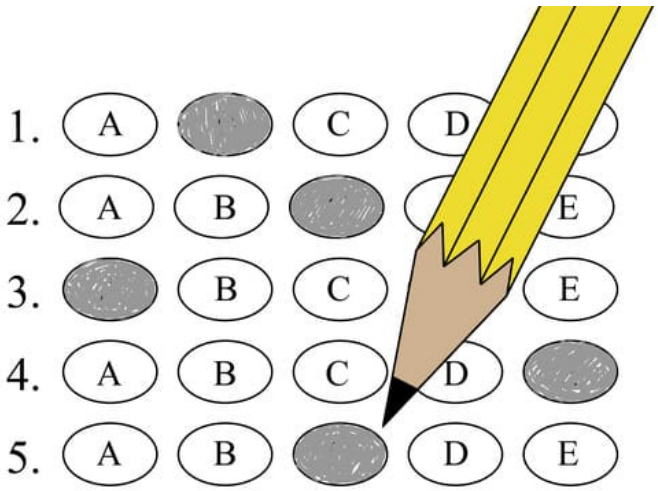


Figure 6.7.1 (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

Have you ever taken a multiple-choice test?

Most of the time the answers are specific choices – is the answer possibility a or possibility b? Quite often you can "think through" the choices to come up with the correct answer. More frustrating is the choice "none of the above." You feel very uncertain checking that possibility.

Metalloids

Some elements are "none of the above." They don't fit neatly into the categories of metal or non-metal because of their characteristics. A **metalloid** is an element that has properties that are intermediate between those of metals and nonmetals. Metalloids can also be called semimetals. On the periodic table, the elements colored yellow, which generally border the stair-step line, are considered to be metalloids. Notice that aluminum borders the line, but it is considered to be a metal since all of its properties are like those of metals.

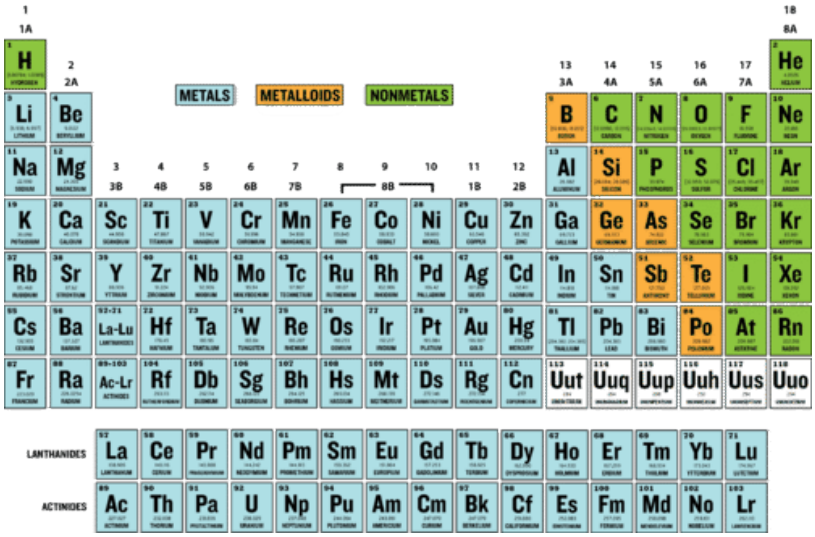


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Examples of Metalloids

Silicon is a typical metalloid (see figure below). It has luster like a metal, but is brittle like a nonmetal. Silicon is used extensively in computer chips and other electronics because its electrical conductivity is in between that of a metal and a nonmetal.



Figure 6.7.3: Silicon. (Public Domain; User:Enricoros/Wikipedia via Wikipedia)

Boron is a versatile element that can be incorporated into a number of compounds (see figure below). Borosilicate glass is extremely resistance to thermal shock. Extreme changes in the temperature of objects containing borosilicates will not create any damage to the material, unlike other glass compositions, which would crack or shatter. Because of their strength, boron filaments are used as light, high-strength materials for airplanes, golf clubs, and fishing rods. Sodium tetraborate is widely used in fiberglass as insulation and also is employed in many detergents and cleaners.

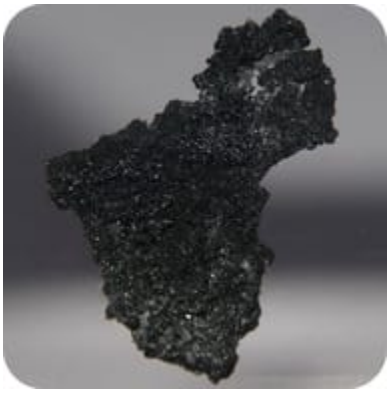


Figure 6.7.4: Boron. (CCSA;CC-BY 3.0; User:Jurii/Wikimedia Commons via Wikipedia)

Arsenic has long played a role in murder mysteries, being used to commit the foul deed (see figure below). This use of the material is not very smart since arsenic can be easily detected on autopsy. We find arsenic in pesticides, herbicides, and insecticides, but the use of arsenic for these applications is decreasing due to the toxicity of the metal. Its effectiveness as an insecticide has led arsenic to be used as a wood preservative.

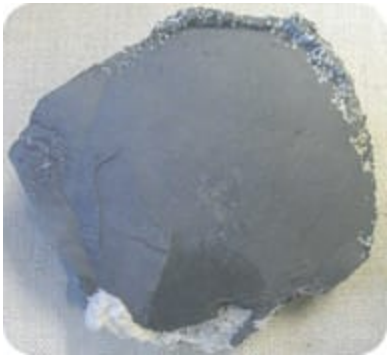


Figure 6.7.5: Arsenic. (Public Domain; Aram Dulyan (Wikimedia: Aramgutang) via Wikipedia)

Antimony is a brittle, bluish-white metallic material that is a poor conductor of electricity (see figure below). Used with lead, antimony increases the hardness and strength of the mixture. This material plays an important role in the fabrication of electronic and semiconductor devices. About half of the antimony used industrially is employed in the production of batteries, bullets, and alloys.



Figure 6.7.6: Antimony. (Public Domain; Aram Dulyan (User:Aramgutang/Wikimedia Commons) via Wikipedia)

Summary

- Metalloids are elements with properties intermediate between those of metals and non-metals
- Silicon is a metalloid because it has luster, but is brittle.
- Boron, arsenic, and antimony are metalloids with a variety of uses.

Review

1. Define “metalloid.”
2. Why would it be difficult to decide whether or not an element was a metalloid based on its properties?
3. Why is silicon used extensively in electronics?
4. What are borosilicates used for?
5. Why is the use of arsenic as an insecticide decreasing?
6. What is a main application of antimony?

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6.8: Blocks of the Periodic Table
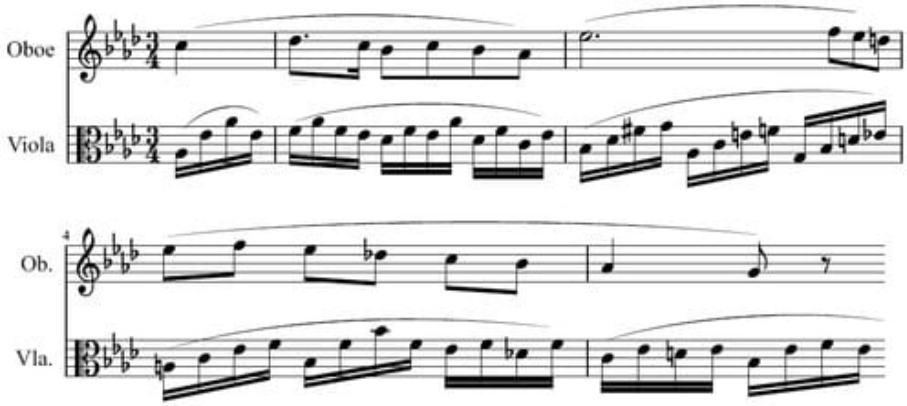


Figure 6.8.1 (Public Domain; User:Szalax/Wikimedia Commons via Wikipedia)

What makes these music notes unique?

We all enjoy music of some sort. Some people like classical music, others like jazz or country. Music styles change from one period of time to the next, and from one region to another. Each type of music has its language that describes it. Classical music has a certain structure, style, and content. There are different expressions of classical music – the symphony, concerto, sonata. We have ballet and opera as well as choral music. Jazz has a different set of characteristics from classical and different styles of performance. Each type of music can be described and compared to other types on the basis of certain common qualities like notes, chords, and melodic styles.

The elements in the periodic table could be considered to be similar to types of music. Each set of elements has its unique set of properties, with different sets of elements having some common characteristics in terms of electron arrangements. We can see patterns of electronic structure and reactivity in the periodic table that allow us to understand better the behavior of individual elements.

Periods and Blocks

There are seven horizontal rows of the periodic table, called **periods**. The length of each period is determined by the number of electrons that are capable of occupying the **sublevels** that fill during that period, as seen in the table below.

Table 6.8.1: Period Length and Sublevels in the Periodic Table

Period	Number of Elements in Period	Sublevels in Order of Fill
1	2	$1s$
2	8	$2s\ 2p$
3	8	$3s\ 3p$
4	18	$4s\ 3d\ 4p$
5	18	$5s\ 4d\ 5p$
6	32	$6s\ 4f\ 5d\ 6p$
7	32	$7s\ 5f\ 6d\ 7s$

Recall that the four different sublevels each consist of a different number of orbitals. The s sublevel has one orbital, the p sublevel has three orbitals, the d sublevel has five orbitals, and the f sublevel has seven orbitals. In the first period, only the $1s$ sublevel is being filled. Since all orbitals can hold two electrons, the entire first period consists of just two elements. In the second period, the $2s$ sublevel, with two electrons, and the $2p$ sublevel, with six electrons, are being filled. Consequently, the second period contains

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eight elements. The third period is similar to the second, filling the $3s$ and $3p$ sublevels. Notice that the $3d$ sublevel does not actually fill until after the $4s$ sublevel. This results in the fourth period containing 18 elements due to the additional 10 electrons that are contributed by the d sublevel. The fifth period is similar to the fourth. After the $6s$ sublevel fills, the $4f$ sublevel with its 14 electrons fills. This is followed by the $5d$ and the $6p$. The total number of elements in the sixth period is 32. The later elements in the seventh period are still being created. So while there is a possible number of 32 elements in the period, the current number is slightly less.

The period to which a given element belongs can be easily determined by its electron configuration. For example, consider the element nickel (Ni). Its electron configuration is $[\mathrm{Ar}]\ 3d^8\ 4s^2$. The highest occupied principal energy level is the fourth, indicated by the 4 in the $4s^2$ portion of the configuration. Therefore, nickel can be found in the fourth period of the periodic table.

Based on electron configurations, the periodic table can be divided into **blocks** denoting which sublevel is in the process of being filled. The s, p, d, and f blocks are illustrated below.

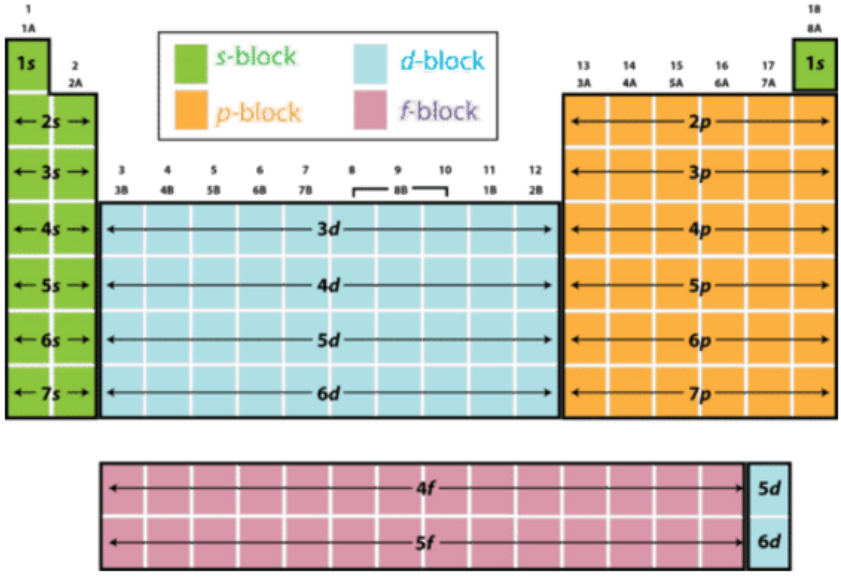


Figure $6.8.2$ (CC BY-NC 3.0; Christopher Auyeung and Joy Sheng via CK-12 Foundation)

The figure also illustrates how the d sublevel is always one principal level behind the period in which that sublevel occurs. In other words, the $3d$ sublevel fills during the fourth period. The f sublevel is always two levels behind. The $4f$ sublevel belongs to the sixth period.

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Summary

- The horizontal rows of the periodic table are called periods.
- The length of a period depends on how many electrons are needed to occupy the sublevels that fill the period.
- Blocks indicate which sublevel is being filled.

Review

1. What are the horizontal rows of the periodic table called?
2. Which sublevel is being filled in period 1?
3. Which sublevel is being filled in period 7?
4. How does the electron configuration of an element give information about the period it is in?
5. What block of elements has the d sublevels being filled?

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6.9: Hydrogen and Alkali Metals



Figure 6.9.1 (Public Domain; User:Ajhalls/Wikimedia Commons via Commons Wikimedia, Large Sodium Explosion [commons.wikimedia.org])

Can you guess what kind of reaction is taking place in this picture?

Some chemistry students just enjoy learning about the science, while others are intrigued by the violent reactions that sometimes can occur. Many chemistry classes have been enlivened by the demonstration of how reactive sodium is with water. In some instances, the demonstration has gone off safely. Unfortunately, in other situations students and instructors have incurred serious injury due to their failure to observe proper safety precautions.

One value of the periodic table is the ability to make predictions about the behavior of individual elements. By knowing which group an element is in, we can determine the number of reactive electrons and say something about how that element will behave.

Hydrogen and Alkali Metals

The periodic table is arranged on the basis of atomic numbers (number of protons in the nucleus). One of the valuable consequences of this arrangement is that we can learn a lot about the electron distribution in these atoms. The colors in the table below indicate the different groupings of atoms based on the location and number of electrons in the atom.



Figure 6.9.2 (Public Domain; User:Cepheus/Wikimedia Commons, modified by CK-12 Foundation via Commons Wikimedia, Periodic Table [commons.wikimedia.org])

If we look at Group I (red column), we see that it is labeled alkali metals. Also note the green H above the alkali metals. All of these elements have a similar configuration of outer-shell electrons (see table below). In each case, there is one electron in the outer orbital and that is an s-orbital electron. Hydrogen is not an alkali metal itself, but has some similar properties due to its simple one proton (located in the nucleus), one electron arrangement. The lone electron exists in a s-orbital around the nucleus. For lithium, there are two 1s electrons in an inner orbit and one 2s electron in the outer orbit. The same pattern holds for sodium and potassium.

Element, Symbol, Electron Configuration

Element	Symbol	Electron Configuration
hydrogen	H	$\backslash 1s^1$
lithium	Li	$[\mathrm{He}]\, 2s^1$
sodium	Na	$[\mathrm{Ne}]\, 3s^1$
potassium	K	$[\mathrm{Ar}]\, 4s^1$
rubidium	Rb	$[\mathrm{Kr}]\, 5s^1$
cesium	Cs	$[\mathrm{Xe}]\, 6s^1$
francium	Fr	$[\mathrm{Rn}]\, 7s^1$

Even an atom with a very complex electron composition such as cesium still has the single s electron in its outer orbital (see figure below).

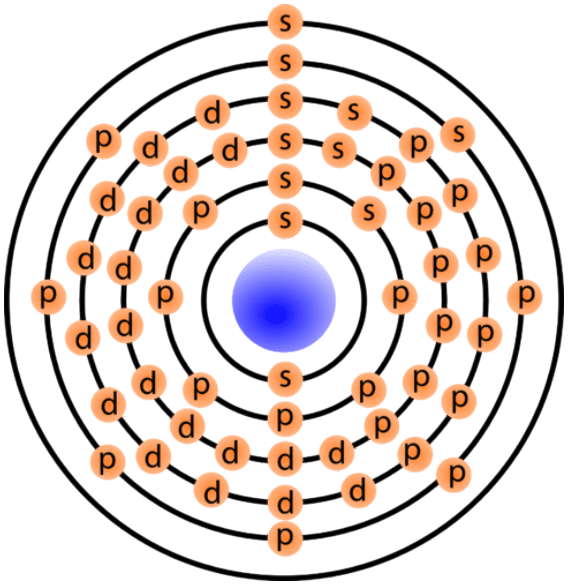


Figure 6.9.3: Celsium Orbitals. (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)

This one electron is very easily removed during chemical reactions. The group I elements react rapidly with oxygen to produce metal oxides. They are very soft metals, which become liquid just above room temperature.

Li reacts with water to produce hydrogen gas. Sodium also reacts the same way, just more rapidly. Potassium reacts rapidly with water producing hydrogen gas and heat which ignites the hydrogen gas. Rubidium and cesium react yet more vigorously and explode on contact with water.

YouTube Video

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Summary

- Group I (alkali metals and hydrogen) elements all have one electron in their outer shell. This electron is in a s orbital.
- The Group I metals are all very reactive with water.

Review

1. What group are the alkali metals and hydrogen in?
2. What is the outer shell electron configuration in this group?
3. How reactive are the alkali metals with oxygen?
4. How reactive are these metals with water?

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6.10: Alkaline Earth Metals



Figure 6.10.1 (Public Domain; U.S. Marine Corps via Wikipedia)

How are oyster shells and chemistry related?

We take a lot of chemistry for granted. Very few of us think about the chemistry of bone or oyster shells. Both of these materials have large amounts of calcium compounds in them and play important roles in maintaining the structure of the organism. The shell provides a solid surrounding for the oyster. Bones give support to the body so the person can move around and not just be a soft mass of tissue.



Figure 6.10.2 (Public Domain; Jossifresco via Wikipedia)

Alkaline Earth Metals

Group 2 elements are referred to as "**alkaline earth**" metals (tan column below). The name "**alkaline**" comes from the fact that compounds of these elements form **basic** (pH greater than 7) or alkaline solutions when dissolved in water. If the Group 1 elements all have one s electron in their outer orbital, we can predict that the Group 2 elements will have two electrons in that outer shell.

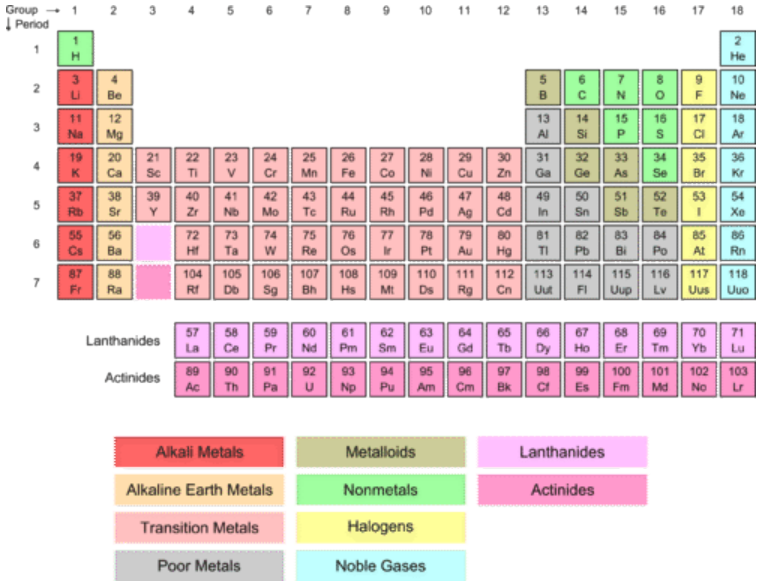


Figure 6.10.3 (Public Domain; User:Cepheus/Wikimedia Commons, modified by CK-12 Foundation via Wikipedia)

The beryllium atom, the first element of Group 2, has an atomic number of four. The atom has the 1s shell filled as well as the 2s shell, giving a total of four electrons ($1s^2 2s^2$). Note that there are two s electrons in the outer shell, a structure that is characteristic of the Group 2 elements. Barium (atomic number 56) has the same outer shell structure of two electrons in the s orbital, even though the internal electron structure for barium is quite complicated.

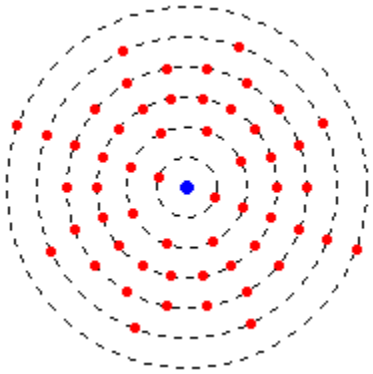


Figure 6.10.4 (Public Domain; Ingmar Runge via Wikipedia)

Radium (atomic number 88) has similar properties to barium and is also in the Group 2 category. However, radium is a radioactive element and is generally under the category of radioisotopes in addition to being an alkaline earth metal, because it is not a stable element.

The Group 2 elements tend to be less reactive than their Group 1 counterparts. The need to remove two electrons in order for the material to react means more energy is needed for electron removal. However, these elements are reactive enough that they do not exist in their elemental forms in nature, but are present as compounds.

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Uses of Alkaline Earth Compounds

Since magnesium burns brightly, it is used in flares and fireworks. Magnesium alloys with aluminum provide light weight and sturdy materials for airplanes, missiles, and rockets. Several antacids use magnesium hydroxide to neutralize excess stomach acid.

Calcium compounds are widely found in limestone, marble, and chalk. Calcium is an important constituent of cement. Other uses include calcium chloride as a deicer and limestone as a white pigment in paints and toothpaste.

Strontium is widely used in fireworks and magnets. Barium compounds can be used in paints, filler for rubber, plastic, and resins, and as a contrast medium for X-rays. Many beryllium compounds are toxic, but these materials have been employed in metal alloys.

Summary

- The alkaline earth elements are in Group 2 of the periodic table.
- These elements each have two s electrons in their outer shell.
- The alkaline earth elements are less reactive than the alkali metals.

Review

1. Why are these elements known as “alkaline earth” elements?
2. How many electrons are in the outer shell of the alkaline earth elements?
3. Are the alkaline earth elements more or less reactive than the alkali metals? Explain your answer.
4. Is radium usually considered as part of the alkaline earth category in terms of chemistry? Explain your answer.

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6.11: Noble Gases



Figure 6.11.1 (Public Domain; User:Ingolfson/Wikimedia Commons via Wikipedia)

What gives these lights their color?

Cities at night would be rather boring without all the bright lights. They provide colorful Illuminations and help make things much more visible. We call these lights “neon lights,” but they use several gases to make the different colors.

Noble Gases

The reactivity of an element can give us important clues as to the electron configuration of that material. If an element is extremely unreactive, this suggests that the electron configuration is such that adding or removing electrons is very unlikely. There must be a stable electron configuration that resists further reaction.

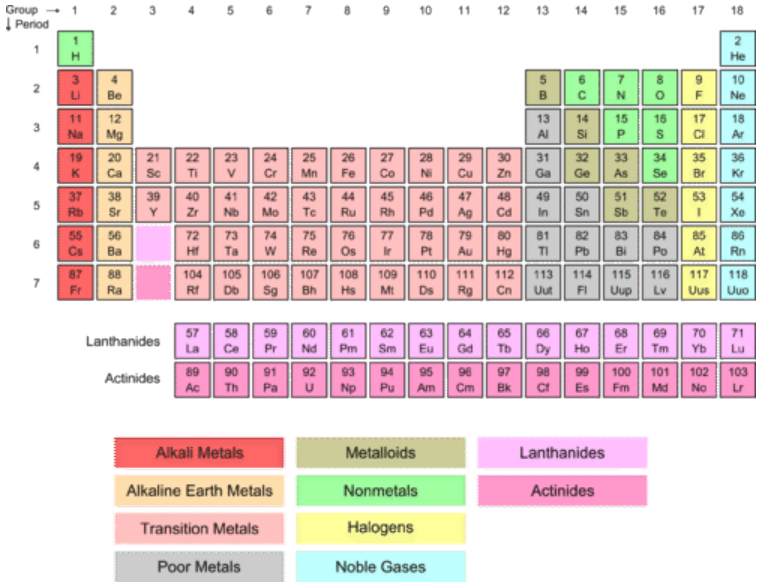


Figure 6.11.2 (Public Domain; User:Cepheus/Wikimedia Commons, modified by CK-12 Foundation via Wikipedia)

The Group 18 (once known as group VIII) elements are essentially chemically **inert** (light blue column on the right). All these elements exist as **monatomic** gases at room temperature. If we look at the electron configurations, we see that helium (atomic

number 2) has a full shell of two s electrons. Since there are no electrons shielding this shell from the nucleus, these two electrons will be very difficult to remove, making helium unreactive.

The remaining elements in the group have full outer shells consisting of two s electrons and six p electrons for an outer shell content of eight electrons. This particular arrangement renders the atoms fairly unreactive. This group has been referred to as the "inert" gases, indicating that they are chemically inert, or unreactive. Another popular term is "**noble gases**," suggesting that these gases do not like to have much to do with the other, more common materials (or that they don't do a lot of work).

Noble Gas Compounds

In more recent years, a number of reactions using the noble gas elements have been discovered. Although the conventional wisdom was that the complete outer shells of these elements would not allow them to react, some scientists believed that the outer electrons of the larger elements (such as Rn, Xe, and Kr) were far enough away from the nucleus that they should be able to be displaced under the right set of conditions. The first compound ($XePtF_6$) was made with xenon in 1962. Since then, several compounds have been made with radon, xenon, krypton, and argon. Only helium and neon have not formed compounds at this time.

Colors of Noble Gases

The different gases glow when an electric current is passed through them. Many of these gases are used in displays because of their chemical inertness. They are stable and will not react with other materials in the system. Radon also will give a reddish glow, but is not used because it is radioactive and will not retain its structure as radon for any significant length of time.



Figure 6.11.3 (CC BY 3.0; User:Jurii/Wikimedia Commons via Wikipedia)

Royal Society of Chemistry - Colors of the Noble Gases

YouTube Video

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Science Friday: 4850 Feet Below: The Hunt for Dark Matter

Why is there a laboratory 4850 feet below the surface? It turns out the depth is a necessary component of the extremely sensitive detector called the Large Underground Xenon experiment. In this video by Science Friday, several scientists explain how this detector helps physicists look for extremely rare particles.

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View this YouTube Video online

Summary

- The noble gases are in Group VIII of the periodic table.
- Helium has a full outer shell of two s electrons.
- The other gases have full outer shells of two s and six p electrons.
- Compounds have been formed with Rn, Xe, Kr, and Ar.

Review

1. What elements comprise the noble elements?
2. What state are they in at room temperature?
3. Why is helium non-reactive?
4. Why were the other noble gases believed to be non-reactive?
5. When was the first compound formed from xenon?
6. What happens when an electric current is passed through these gases?

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6.12: Halogens

How do you study a gas that does not exist as such in nature?

It's not as easy as you think. Fluorine is so reactive that we cannot find it free in nature. None of the halogens exist free in nature (unlike some of the metals such as gold and silver) because they are very reactive. The video below shows how violently elemental fluorine reacts with other materials.

YouTube Video

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Halogens

Some elements are much more reactive than others. The Group I (red) and Group II (tan) elements can easily lose electrons during a reaction. Elements of other groups are much more likely to accept electrons as they react.

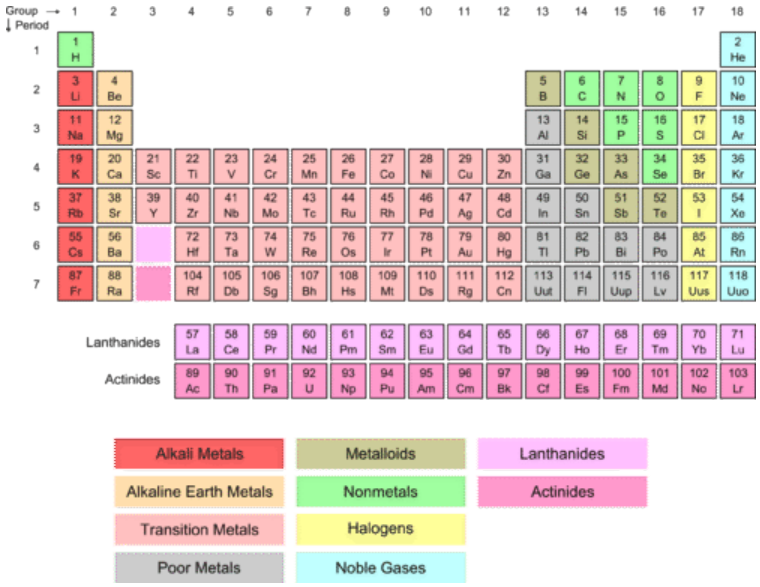


Figure 6.12.1 (Public Domain; User:Cepheus/Wikimedia Commons, modified by CK-12 Foundation via Wikipedia)

The elements of Group VIIA (new Group 17 - fluorine, chlorine, bromine, iodine, and astatine) are called the **halogens** (yellow column). The term "halogen" means "salt-former" because these elements will readily react with alkali metal and alkaline earth metals to form halide salts. The halogens all have the general electron configuration ns^2np^5, giving them seven valence electrons. They are one electron short of having the full outer s and p sublevel, which makes them very reactive.

Physical Properties of Halogens

As elements, chlorine and fluorine are gases at room temperature, bromine is a dark orange liquid, and iodine is a dark purple-gray solid. Astatine is so rare that its properties are mostly unknown. In the picture below we see chlorine gas on the left (green),

bromine solid and vapor in the middle (orange), and solid iodine (grey) on the right. Fluorine is not shown in the picture below because it is too corrosive and will destroy the glass container.



Figure 6.12.2 (Chlorine and iodine: Public Domain; Bromine: The copyright holder of this work allows anyone to use it for any purpose including unrestricted redistribution, commercial use, and modification; Chlorine: User:Greenhorn1/Wikimedia Commons Bromine: User:Tomihahndorf/De.Wikipedia; Iodine: Ben Mills (Wikimedia: Benjah-bmm27) via Chlorine: http://commons.wikimedia.org/wiki/File:Chlorinegas.jpg; Bromine: http://commons.wikimedia.org/wiki/File:Brom_amp.jpg; Iodine: http://commons.wikimedia.org/wiki/File:Iodine-sample.jpg)

None of these elements are found free in nature because of their reactivity. Fluorine is found in combination with cations in several minerals. Chlorine is found in table salt, in the oceans (which are about 2% chloride ion by weight) and in lakes such as the Great Salt Lake in Utah. Small amounts of bromide and iodide salts can be found in the oceans and in brine wells in several states.



Figure 6.12.3 (Public Domain; NASA via Wikipedia)

Watch the following two video experiments of p block elements:

This first video is of bromine reacting with aluminum.

YouTube Video

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View this YouTube Video online

This second video discusses the properties of halogens and shows a few more reactions they can participate in.

YouTube Video

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Summary

- The halogens all have seven electrons in their outer shells.
- The electron configuration in the outer shell is ns^2np^5.
- As the atomic number increases, the reactivity of the halogens decreases.
- Fluorine and chlorine exist as gases at room temperature, while bromine is a liquid, and iodine is a solid.

Review

1. Pick two elements that are halogens. For each, write the name, chemical symbol, and atomic weight.
2. What does the term “halogen” mean?
3. What is the outer shell electron configuration for the halogens?
4. What is the physical state of each halogen at room temperature?
5. Where are the halogens found?

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6.13: Transition Elements



Figure 6.13.1 (Public Domain; Stock model: Bull-Doser; Race model: Dana60Cummins via Stock model: Commons Wikimedia, Ford Focus ZX5 [commons.wikimedia.org]; Race model: Commons Wikimedia, Ford Focus Ecoboost Race Car [commons.wikimedia.org])

What are the similarities and differences between these two cars?

From the outside, the two cars below look the same (except for the flashy paint job on the racing model). They are the same model of the car, but one is a stock edition for regular driving, while the other one is built for high-speed racing. We really can't tell much from the external view. To see the differences, we need to go under the hood, take the engines apart, and look at the braking and suspension systems in order to see how the two cars differ.

Many electron configurations of elements are simple and straightforward. We can look at the outer shell and easily understand how that set of elements will react in terms of electron gain or loss. However, there are sets of elements that are more complex in their behavior. One such group is the **transition elements**.



Figure 6.13.2 (Public Domain; User:Cepheus/Wikimedia Commons, modified by CK-12 Foundation via Commons Wikimedia, Periodic Table [commons.wikimedia.org])

The Transition Elements

Transition elements are the elements that are found in Groups 3-12 (old groups IIA-IIB) on the periodic table (salmon-colored block in the middle of the table). The term refers to the fact that the d **sublevel**, which is in the process of being filled, is in a lower principal energy level than the s sublevel filled before it. For example, the electron configuration of scandium, the first transition element, is $[\mathrm{Ar}]\ \ 3d^1\ \ 4s^2$. Remember that the configuration is reversed from the fill order—the $4s$ fills before the $3d$ begins. Because they are all metals, the transition elements are often called the transition metals. As a group, they display typical

metallic properties and are less reactive than the metals in Groups 1 and 2. Some of the more familiar ones are so unreactive that they can be found in nature in their free, or uncombined, state. These include platinum, gold, and silver. Because of this unique filling order, the transition elements are often referred to as "d-block" elements.



Figure $6.13.3$: Piece of silver. (Public Domain; US Geological Survey via Commons Wikimedia, Silver [commons.wikimedia.org])

Compounds of many transition elements are distinctive for being widely and vividly colored. As visible light passes through a transition metal compound dissolved in water, the d-orbitals absorb light of various energies. The visible light of a given energy level that is not absorbed produces a distinctly colored solution.



Figure $6.13.4$: Transition metal compounds dissolved in water exhibit a wide variety of bright colors. From left to right are shown solutions of cobalt (II) nitrate, potassium dichromate, potassium chromate, nickel (II) chloride, copper (II) sulfate, and potassium permanganate. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Commons Wikimedia, Coloured Transition Metal Solutions [commons.wikimedia.org])

YouTube Video

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Summary

- The transition elements are found in groups IIA-IIB (new groups 3-12).
- These elements are characterized by having unfilled d sublevels.
- In general, the next higher s sublevel is already filled or has one electron missing.
- Many transition element compounds are brightly colored due to the inner-level d electron transitions.

Review

1. List five different transition elements, giving their name, chemical symbol, and atomic number.
2. What is unique about the transition elements in terms of electron configurations?
3. Why are these elements often referred to as "d-block" elements?
4. Why do many transition element compounds have bright colors?

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6.14: Periodic Trends- Atomic Radius



Figure 6.14.1 (CC by 2.0; James Cridland via Flickr)

How can all of these people fit in such a small space?

Events draw large numbers of people to them. Even an outdoor event can fill up so that there is no room for more people. The crowd capacity depends on the amount of space in the venue, and the amount of space depends on the size of the objects filling it. We can get more people into a given space than can elephants, because elephants are larger than people. We can get more squirrels into that same space than we can people for the same reason. Knowing the sizes of objects to be dealt with can be important in deciding how much space is needed.

Atomic Radius

The size of atoms is important to explanations of the behavior of atoms or compounds. One way to express the size of atoms is by use of **atomic radius**. This data helps us understand why some molecules fit together and why other molecules have parts that get too crowded under certain conditions.

The size of an atom is defined by the edge of its orbital. However, orbital boundaries are fuzzy, and variable under different conditions. In order to standardize the measurement of atomic radii, the distance between the nuclei of two identical atoms bonded together is measured. The atomic radius is defined as one-half the distance between the nuclei of identical atoms that are bonded together.

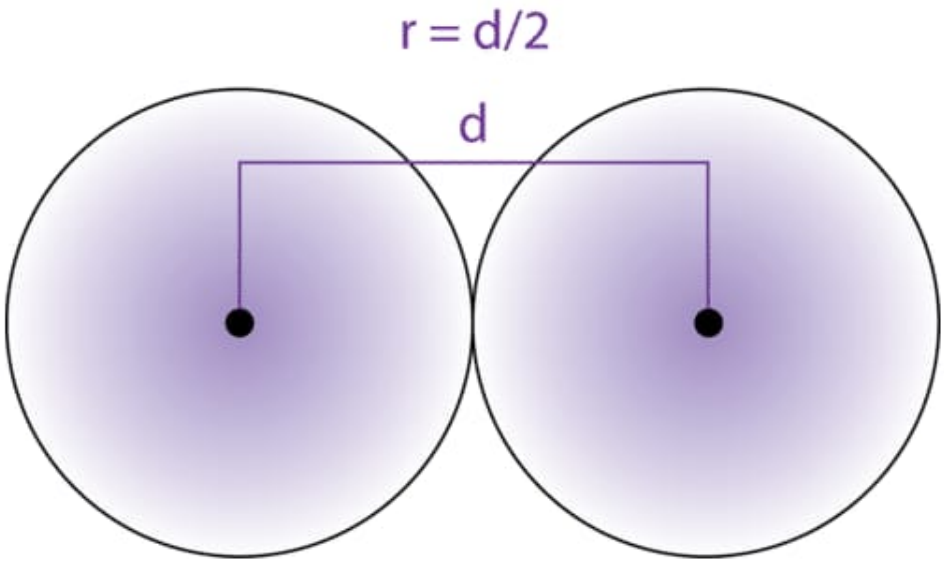


Figure 6.14.2: The atomic radius (r) of an atom can be defined as one half the distance (d) between two nuclei in a diatomic molecule. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Atomic radii have been measured for elements. The units for atomic radii are picometers, equal to 10^{-12} meters. As an example, the internuclear distance between the two hydrogen atoms in an H_2 molecule is measured to be $74\ \text{pm}$. Therefore, the atomic radius of a hydrogen atom is $\frac{74}{2} = 37\ \text{pm}$.

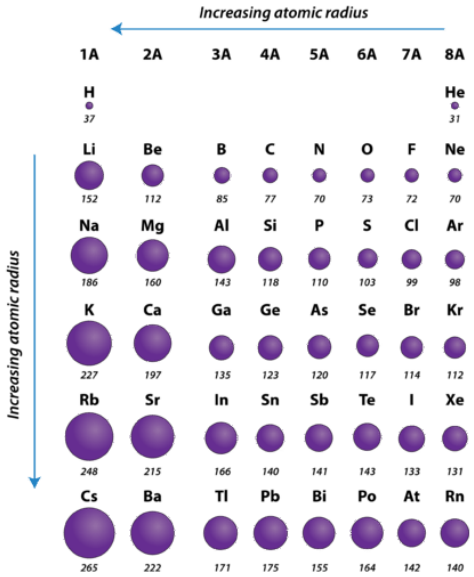


Figure 6.14.3: Atomic radii of the representative elements measured in picometers. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Periodic Trend

The atomic radius of atoms generally decreases from left to right across a period. There are some small exceptions, such as the oxygen radius being slightly greater than the nitrogen radius. Within a period, protons are added to the nucleus as electrons are being added to the same principal energy level. These electrons are gradually pulled closer to the nucleus because of its increased positive charge. Since the force of attraction between nuclei and electrons increases, the size of the atoms decreases. The effect lessens as one moves further to the right in a period, because of electron-electron repulsions that would otherwise cause the atom's size to increase.

Group Trend

The atomic radius of atoms generally increases from top to bottom within a group. As the atomic number increases down a group, there is again an increase in the positive nuclear charge. However, there is also an increase in the number of occupied principal energy levels. Higher principal energy levels consist of orbitals which are larger in size than the orbitals from lower energy levels. The effect of the greater number of principal energy levels outweighs the increase in nuclear charge, and so atomic radius increases down a group.

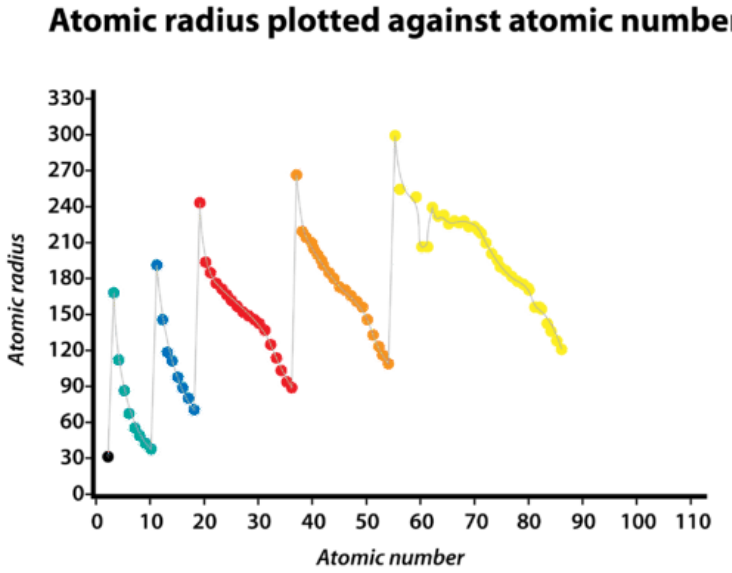


Figure 6.14.4: A graph of atomic radius plotted versus atomic number. Each successive period is shown in a different color. As the atomic number increases within a period, the atomic radius decreases. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Interactive Content

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Summary

- Atomic radius is determined as half the distance between the nuclei of two identical atoms bonded together.
- The atomic radius of atoms generally decreases from left to right across a period.
- The atomic radius of atoms generally increases from top to bottom within a group.

Review

1. Define “atomic radius.”
2. What are the units of measurement for atomic radius?
3. How does the atomic radius of different elements change across a period?
4. How does atomic radius change from top to bottom within a group?
5. Explain why the atomic radius of hydrogen is so much smaller than the atomic radius of potassium.

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6.15: Periodic Trends- Ionization Energy



Figure 6.15.1 (CC by 2.0; Nick Russill via Flickr, Northern Lights, Greenland [www.flickr.com])

The incredible green lights in this cold northern sky consist of charged particles known as ions. Their swirling pattern is caused by the pull of Earth's magnetic field. Called the northern lights, this phenomenon of nature shows that ions respond to a magnetic field. Do you know what ions are? Read on to find out.

Atoms Are Neutral

The northern lights aren't caused by atoms, because atoms are not charged particles. An atom always has the same number of electrons as protons. Electrons have an electric charge of -1 and protons have an electric charge of +1. Therefore, the charges of an atom's electrons and protons "cancel out." This explains why atoms are neutral in electric charge.

Q: What would happen to an atom's charge if it were to gain extra electrons?

A: If an atom were to gain extra electrons, it would have more electrons than protons. This would give it a negative charge, so it would no longer be neutral.

Atoms to Ions

Atoms cannot only gain extra electrons. They can also lose electrons. In either case, they become **ions**. Ions are atoms that have a positive or negative charge because they have unequal numbers of protons and electrons. If atoms lose electrons, they become positive ions, or cations. If atoms gain electrons, they become negative ions, or anions. Consider the example of fluorine (see figure below). A fluorine atom has nine protons and nine electrons, so it is electrically neutral. If a fluorine atom gains an electron, it becomes a fluoride ion with an electric charge of -1.

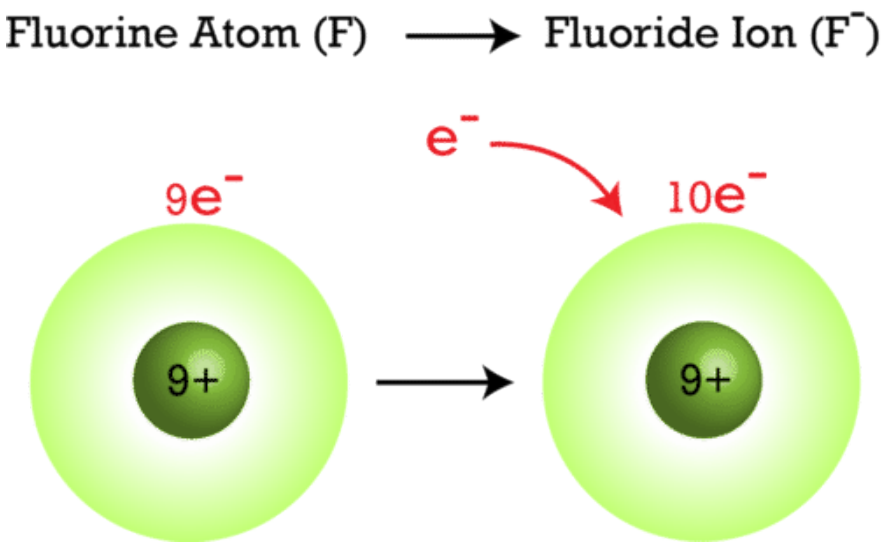


Figure 6.15.2 (CC BY-NC 3.0; Christopher AuYeung via CK-12 Foundation)

Names and Symbols

Like fluoride, other negative ions usually have names ending in *–ide*. Positive ions, on the other hand, are just given the element name followed by the word *ion*. For example, when a sodium atom loses an electron, it becomes a positive sodium ion. The charge of an ion is indicated by a plus (+) or minus sign (-), which is written to the right of and just above the ion's chemical symbol. For example, the fluoride ion is represented by the symbol F^-, and the sodium ion is represented by the symbol Na^+. If the charge is greater than one, a number is used to indicate it. For example, iron (Fe) may lose two electrons to form an ion with a charge of plus two. This ion would be represented by the symbol Fe^{2+}. This and some other common ions are listed with their symbols in the table below.

Table $PageIndex1$: Some Common Ions

Cations		Anions	
Name of Ion	Chemical Symbol	Name of Ion	Chemical Symbol
Calcium ion	Ca^{2+}	Chloride	Cl^-
Hydrogen ion	H^+	Fluoride	F^-
Iron(II) ion	Fe^{2+}	Bromide	Br^-
Iron(III) ion	Fe^{3+}	Oxide	O^{2-}

Q: How does the iron(III) ion differ from the iron(II) ion?

A: The iron(III) ion has a charge of +3, so it has one less electron than the iron(II) ion, which has a charge of +2.

Q: What is the charge of an oxide ion? How does its number of electrons compare to its number of protons?

A: An oxide ion has a charge of -2. It has two more electrons than protons.

How Ions Form

The process in which an atom becomes an ion is called ionization. It may occur when atoms are exposed to high levels of radiation. The radiation may give their outer electrons enough energy to escape from the attraction of the positive nucleus. However, most ions form when atoms transfer electrons to or from other atoms or molecules. For example, sodium atoms may transfer electrons to chlorine atoms. This forms positive sodium ions (Na^+) and negative chloride ions (Cl^-).

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Q: Why do you think atoms lose electrons to, or gain electrons from, other atoms?

A: Atoms form ions by losing or gaining electrons because it makes them more stable and this state takes less energy to maintain. The most stable state for an atom is to have its outermost energy level filled with the maximum possible number of electrons. In the case of metals such as lithium, with just one electron in the outermost energy level, a more stable state can be achieved by losing that one outer electron. In the case of nonmetals such as fluorine, which has seven electrons in the outermost energy level, a more stable state can be achieved by gaining one electron and filling up the outer energy level.

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Properties of Ions

Ions are highly reactive, especially as gases. They usually react with ions of opposite charge to form neutral compounds. For example, positive sodium ions and negative chloride ions react to form the neutral compound sodium chloride, commonly known as table salt. This occurs because oppositely charged ions attract each other. Ions with the same charge, on the other hand, repel each other. Ions are also deflected by a magnetic field, as you saw in the opening image of the northern lights.

Summary

- Atoms have equal numbers of positive protons and negative electrons, so they are neutral in electric charge.
- Atoms can gain or lose electrons and become ions, which are atoms that have a positive or negative charge because they have unequal numbers of protons and electrons.
- The process in which an atom becomes an ion is called ionization. It may occur when atoms are exposed to high levels of radiation or when atoms transfer electrons to or from other atoms.
- Ions are reactive, attracted or repulsed by other charged particles, and deflected by a magnetic field.

Review

1. Why are atoms neutral in electric charge?
2. Define ion.
3. Compare and contrast cations and anions, and give an example of each.
4. Describe how ions form.
5. List properties of ions.
6. The model in the illustration below represents an atom of lithium (Li). If the lithium atom becomes an ion, which type of ion will it be, a cation or an anion? What will be the electric charge of this ion? What will the ion be named? What symbol will be used to represent it?

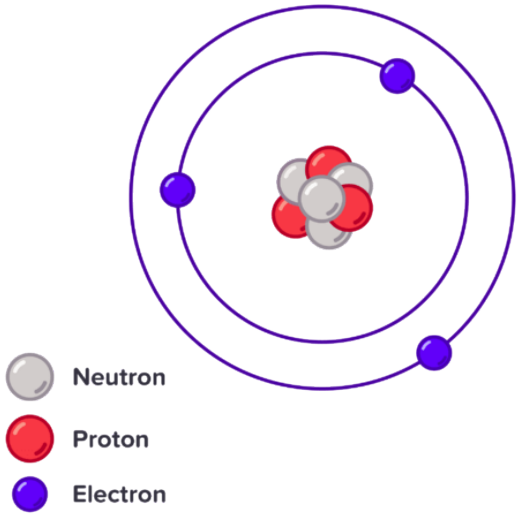


Figure 6.15.3 (CC BY-NC 3.0; CK-12 Foundation Employee via CK-12 Foundation)

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6.16: Ionic Radii



Figure 6.16.1 (Public Domain; the US Department of Agriculture via Commons Wikimedia, Peanuts [commons.wikimedia.org])

How are peanuts sold?

Peanuts can be sold two ways. The bare peanut without the shell (brown inner portion of peanut) can be purchased in jars and packages for casual munching or for cooking. The size of the peanut in this situation is smaller than the peanut plus shell since the outer portion is missing. If we add the shell to the peanut, we have a larger size for the combination.

Electrons and protons are strongly attracted to one another. The strength of that attraction and the relative numbers of the two particles in a given atom or ion have a significant influence on the size of that species. When an atom loses one or more electrons, the resulting ion becomes smaller. If electrons are added to the atom, the ion becomes larger.

Ionic Radius

The **ionic radius** for an atom is measured in a **crystal lattice**, requiring a solid form for the compound. These radii will differ somewhat depending upon the technique used. Usually, x-ray crystallography is employed to determine the radius for an ion.

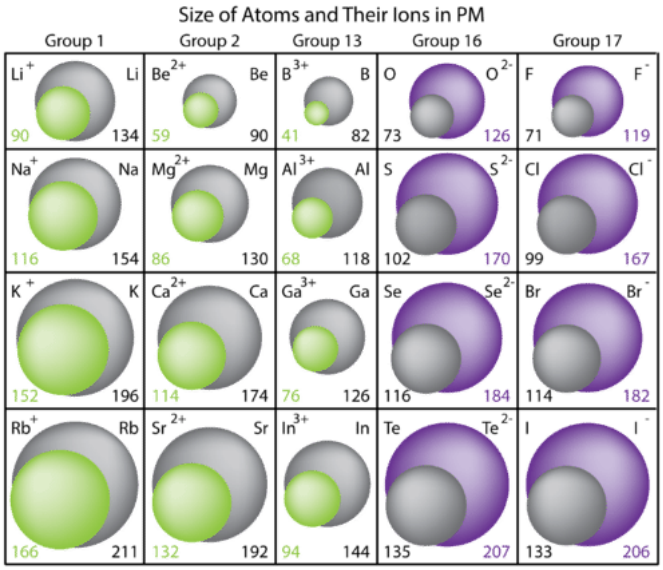


Figure 6.16.2: Comparison of ion sizes for Groups 1, 2, 13, 16, and 17. The atoms are shown in gray. Groups 1, 2, and 13 are metals and form cations, shown in green. Groups 16 and 17 are nonmetals and form anions, shown in purple. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

The removal of electrons always results in a cation that is considerably smaller than the parent atom. When the valence electron(s) are removed, the resulting ion has one fewer occupied principal energy level, so the electron cloud that remains is smaller. Another reason is that the remaining electrons are drawn closer to the nucleus because the protons now outnumber the electrons. One other factor is the number of electrons removed. The potassium atom has one electron removed to form the corresponding ion, while calcium loses two electrons.

The addition of electrons always results in an anion that is larger than the parent atom. When the electrons outnumber the protons, the overall attractive force that the protons have for the electrons is decreased. The electron cloud also spreads out because more electrons results in greater electron-electron repulsion. Notice that the group 16 ions are larger than the group 17 ions. The group 16 elements each add two electrons, while the group 17 elements add one electron per atom, to form the anions.

Interactive Content

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Summary

- Ionic radius is determined by measuring the atom in a crystal lattice.
- Removal of electrons results in an ion that is smaller than the parent element.
- Addition of electrons results in an ion that is larger than the parent atom.

Review

1. How are ionic radii measured?
2. Explain why the radius of the rubidium ion is smaller than the radius of the rubidium atom.
3. Explain why the radius of the tellurium ion is larger than the radius of the tellurium atom.
4. Why is the oxygen anion larger than the fluoride anion?

5. Why is the sodium cation larger than the magnesium cation?

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6.17: Metallic and Nonmetallic Character

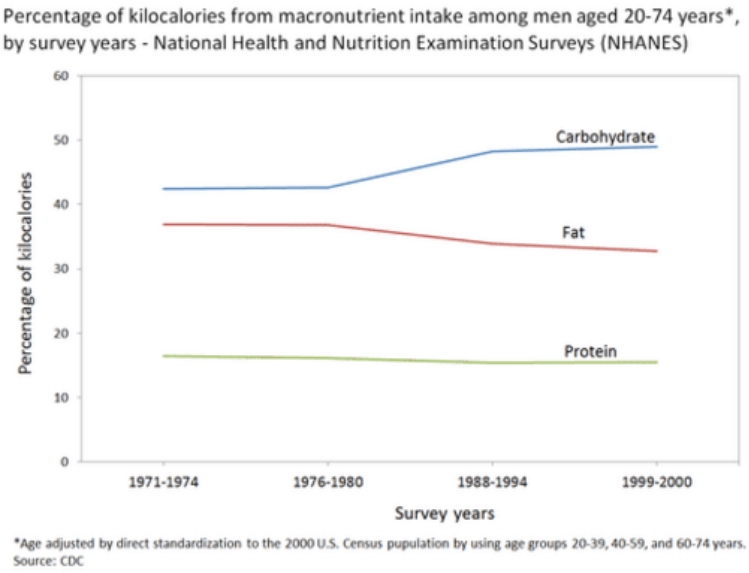


Figure 6.17.1: Graph indicating eating habits of American men from 1971 to 2000. (Public Domain; Recreated by CK-12 Foundation based on data from the CDC via CDC GOV Preview [www.cdc.gov]; Flickr, Venn Diagram [www.flickr.com])

What are we eating as a nation?

The graph above indicates some trends in the U.S. diet over a thirty-year period. By observing the direction our eating habits are going, steps can be taken to help prevent bad eating habits and decrease problems such as high blood pressure and heart attacks.

Development of the periodic table has helped organize chemical information in many ways. We can now see trends among properties of different atoms and make predictions about the behavior of specific materials.

Metallic and Nonmetallic Character

Metallic character refers to the level of reactivity of a metal. Metals tend to lose electrons in chemical reactions, as indicated by their low ionization energies. Within a compound, metal atoms have relatively low attraction for electrons, as indicated by their low electronegativities. By following the trend summary in the figure below, you can see that the most reactive metals would reside in the lower left portion of the periodic table. The most reactive metal is cesium, which is not found in nature as a free element. It reacts explosively with water and will ignite spontaneously in air. Francium is below cesium in the alkali metal group, but is so rare that most of its properties have never been observed.

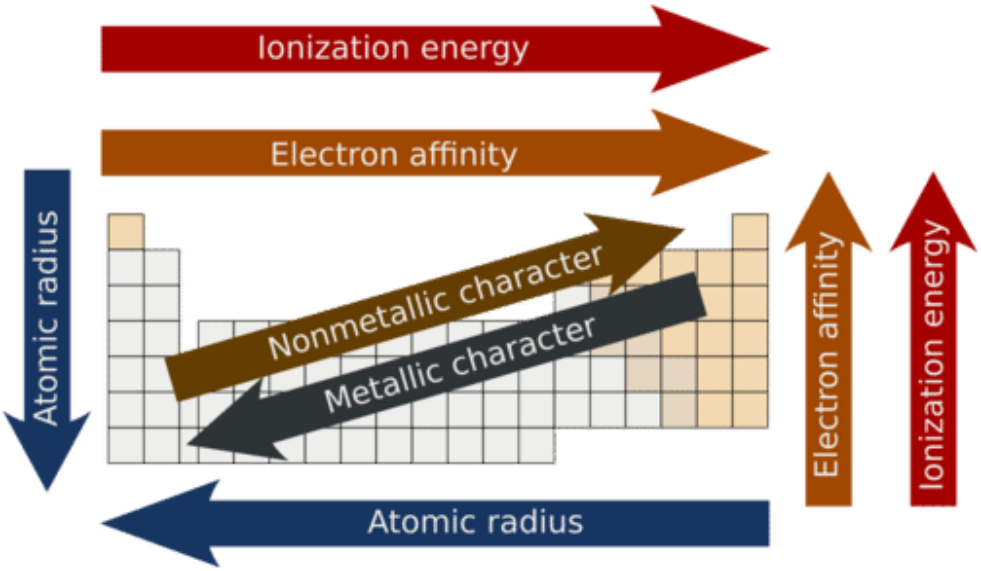


Figure 6.17.2: Trends in behaviors of elements. (Public Domain; User:Mirek2/Wikimedia Commons via Commons Wikimedia, Periodic Trends [commons.wikimedia.org])

Reactivity of metals is based on processes such as the formation of halide compounds with halogens, and how easily the element displaces hydrogen from dilute acids.

The metallic character increases as you go down a group. Since the ionization energy decreases going down a group (or increases going up a group), the increased ability for metals lower in a group to lose electrons makes them more reactive. In addition, the atomic radius increases going down a group, placing the outer electrons further away from the nucleus and making that electron less attracted by the nucleus.

Nonmetals tend to gain electrons in chemical reactions, and have a high attraction for electrons within a compound. The most reactive nonmetals reside in the upper right portion of the periodic table. Since the noble gases are a special group because of their lack of reactivity, the element fluorine is the most reactive nonmetal. It is not found in nature as a free element. Fluorine gas reacts explosively with many other elements and compounds, and is considered to be one of the most dangerous known substances.

Note that there is no clear division between metallic and **non-metallic** character. As we move across the periodic table, there is an increasing tendency to accept electrons (nonmetallic) and a decrease in the possibility that an atom will give up one or more electrons.

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Summary

- Metallic character refers to the level of reactivity of a metal.
- Non-metallic character relates to the tendency to accept electrons during chemical reactions.
- Metallic tendency increases going down a group.
- Non-metallic tendency increases going from left to right across the periodic table.

Review

1. Define “metallic character.”
2. Define “non-metallic character.”
3. Describe the trend in metallic character going down a group.
4. Describe the trend in non-metallic character going across the periodic table.
5. Why does the metallic character increase as you go down a group?

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7: CHEMICAL NOMENCLATURE

The International Union of Pure and Applied Chemistry (IUPAC) is an international federation of organizations that represents chemists in individual countries. IUPAC is best known for its works standardizing nomenclature in chemistry and other fields of science.

7.1: Molecular Formula

7.2: Empirical Formula

7.3: Cations

7.4: Anions

7.5: Transition Metal Ions

7.6: Naming Binary Ionic Compounds

7.7: Stock System Naming

7.8: Formulas for Binary Ionic Compounds

7.9: Polyatomic Ions

7.10: Names and Formulas of Ternary Ionic Compounds

7.11: Naming Binary Molecular Compounds

7.12: Naming Acids

7.13: Names and Formulas of Bases

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7.1: Molecular Formula
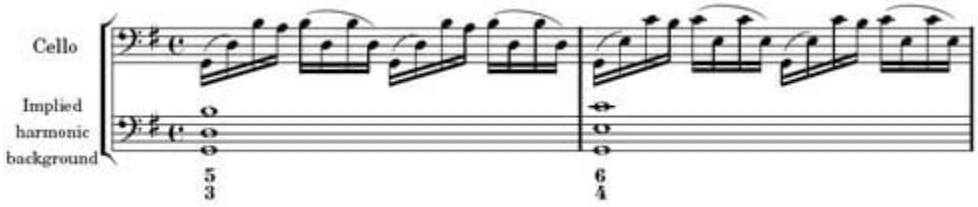


Figure 7.1.1 (Public Domain; J. S. Bach via Commons Wikimedia, Bach cello harmony [commons.wikimedia.org])

Why are music notes a unique "language"?

There are many "universal languages" in the world. Musicians of every culture recognize music embodied in a series of notes on a staff.

This passage from a Bach cello suite could be played by any trained musician from any country, because there is an agreement as to what the symbols on the page mean. In the same way, molecules are represented using symbols that all chemists agree upon.

Molecular Formula

A **molecule** is comprised of two or more atoms that have been chemically combined. A **molecular formula** is a chemical formula of a molecular compound that shows the kinds and numbers of atoms present in a molecule of the compound. Ammonia is a compound of nitrogen and hydrogen as shown below:

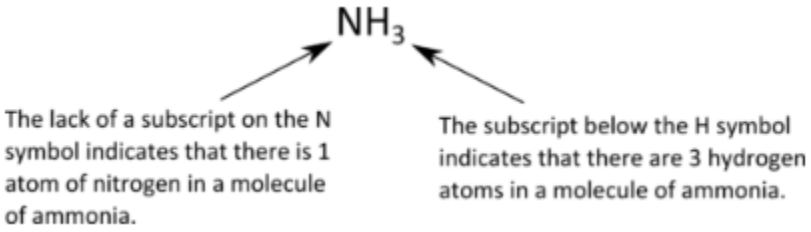


Figure 7.1.2: The molecular formula for ammonia. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Note from the example that there are some standard rules to follow in writing molecular formulas. The arrangements of the elements depend on the particular structure, so we will not concern ourselves with that point right now. The number of atoms of each kind is indicated by a subscript following the atom. If there is only one atom, no number is written. If there is more than one atom of a specific kind, the number is written as a subscript following the atom. We would not write $\mathrm{N_3H}$ for ammonia, because that would mean that there are three nitrogen atoms and one hydrogen atom in the molecule, which is incorrect.

The molecular formula does not tell us anything about the shape of the molecule or where the different atoms are. The molecular formula for sucrose (table sugar) is $\mathrm{C_{12}H_{22}O_{11}}$. This simply tells us the number of carbon, hydrogen, and oxygen atoms in the molecule. There is nothing said about where the individual atoms are located. We need a much more complicated formula (shown below) to communicate that information.

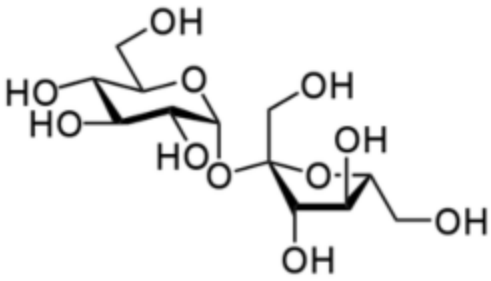


Figure 7.1.3: The molecular structure of sucrose, $C_{12}H_{22}O_{11}$. (Public Domain; User:Calvero/Wikimedia Commons via Commons Wikimedia, Sucrose [commons.wikimedia.org])

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Summary

- A molecular formula tells us which atoms and how many of each type of atom are present in a molecule.
- If only one atom of a specific type is present, no subscript is used.
- For atoms that have two or more of a specific type of atom present, a subscript is written after the symbol for that atom.
- Molecular formulas do not indicate how the atoms are arranged in a molecule.

Review

1. What does a molecular formula tell us?
2. What does a molecular formula not tell us?
3. What do the subscripts mean in a molecular formula?
4. If you wrote $C_6H_{11}O_5C_6H_{11}O_6$ as the molecular formula for sucrose, would that be correct? Explain your answer.
5. Sometimes the formula for acetic acid is written CH_3COOH. Is this a true molecular formula?

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7.2: Empirical Formula

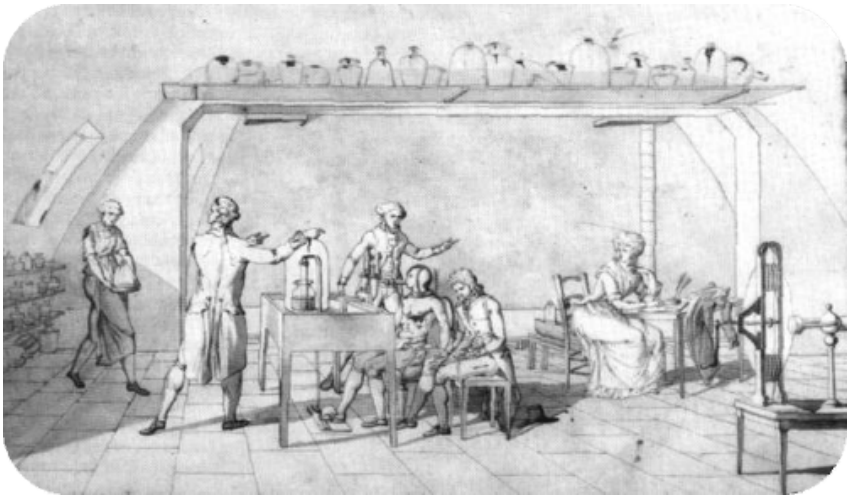


Figure 7.2.1 (Public Domain; Madame Lavoisier via Wikipedia)

What might the scientists in the picture be discussing?

When the French scientist Antoine Lavoisier conducted his experiments, he did not know what the products of each reaction would be. He had to isolate the material (whether he was heating mercury or measuring gases from breathing) and then study its elemental composition before he could understand the processes that were occurring.

Discovering that a new compound exists is the start of a long research project. In order to make this new compound in the lab, we need to know a lot about its structure. Often, the place to start is to determine the elements in the material. Then, we can find out the relative amounts of each element to continue our evaluation of this new material.

Empirical Formula

An **empirical formula** is a formula that shows the elements in a compound in their lowest whole-number ratio. Glucose is an important simple sugar that cells use as their primary source of energy. Its molecular formula is $\mathrm{C_6H_{12}O_6}$. Since each of the subscripts is divisible by 6, the empirical formula for glucose is $\mathrm{CH_2O}$. When chemists analyze an unknown compound, often the first step is to determine its empirical formula. There are a great many compounds whose molecular and empirical formulas are the same. If the molecular formula cannot be simplified into a smaller whole-number ratio, as in the case of $\mathrm{H_2O}$ or $\mathrm{P_2O_5}$, then the empirical formula is also the molecular formula.

How do we determine an empirical formula for a compound? Consider a compound composed of carbon, hydrogen, and oxygen. We can analyze the relative amounts of each element in the compound. When we get a percent figure for each element, we now know how many grams of each are in 100 grams of the original material. This allows us to determine the number of moles for each element. The ratios can then be reduced to small whole numbers to give the empirical formula. If we wanted a molecular formula, we would need to determine the molecular weight of the compound.

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Summary

- The empirical formula gives the lowest whole-number ratio of elements in a compound.
- The empirical formula does not show the actual number of atoms.

Review

1. Define “empirical formula.”
2. Why is $C_6H_{12}O_6$ not considered to be an empirical formula for glucose?
3. Can the empirical formula for a compound be the same as the molecular formula?
4. What do we need to know in order to determine a molecular formula from an empirical formula?
5. Give three examples of compounds whose empirical formulas are the same as their molecular formulas.

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7.3: Cations



Figure 7.3.1 (Public Domain; User:Gold Guru/Wikipedia via Wikipedia)

Have you ever gone digging for gold?

When the prospectors during the California Gold Rush (1848-1855) searched for gold nuggets in the earth, they were able to find the nuggets because gold is an unreactive material that exists in its elemental state in many places. Not everyone was fortunate enough to find significantly large gold nuggets, but a number of miners did become very wealthy (of course, a large number of others went back home broke).

Many of the elements we know about do not exist in their native form. They are so reactive that they are found only in compounds. These non-elemental forms are known as ions. Their properties are very different from those of the elements they come from. The term comes from a Greek word meaning "move" and was first coined by Michael Faraday, who studied the movement of materials in an electrical field.

Cations

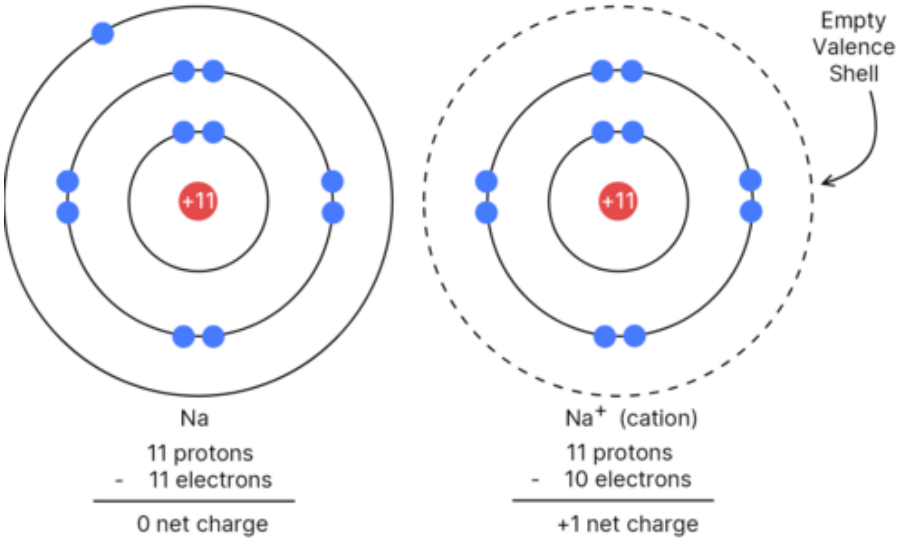


Figure 7.3.2: Sodium loses an electron to become a cation. (CK-12 Curriculum Materials License; CK-12 Foundation via CK-12 Foundation)

Some elements lose one or more electrons in forming ions. These ions are known as "**cations**" because they are positively charged and migrate toward the negative electrode (**cathode**) in an electrical field. Looking at the periodic table below, we know that the group 1 elements are all characterized by having one s electron in the outer orbit; group 2 elements have two s electrons in the outer orbit. These electrons are loosely attached to the atom and can easily be removed, leaving more protons in the atom than there are electrons, so that the resulting ion has a positive charge. Cations can also be formed from electron loss to many of the transition elements.

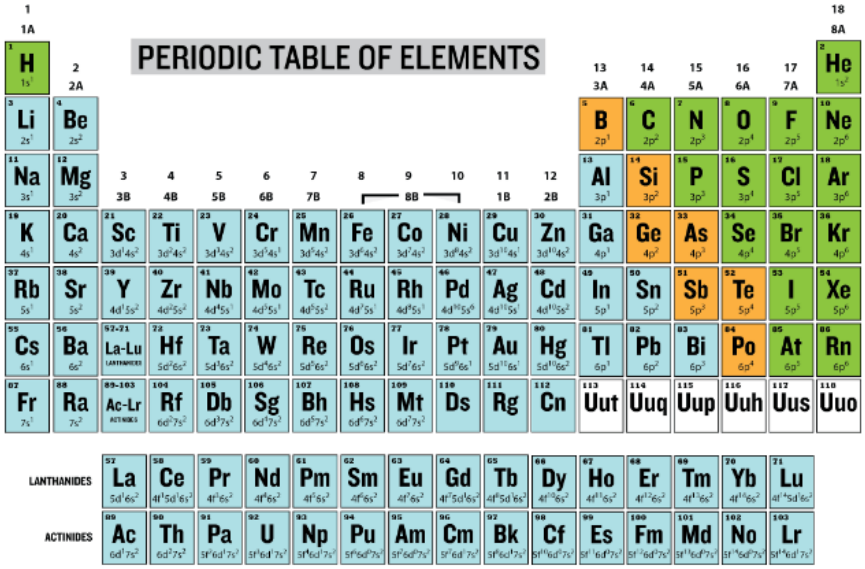


Figure 7.3.3: Periodic Table of Elements, notated with group numbers. (CC BY-NC 3.0; Christopher Auyeung via CK-12)

The cations are designated by the symbol for the parent element and a plus charge as a superscript after the element symbol—the potassium cation would be indicated as K^+. Note that the charge is placed after the symbol and not before it. The potassium ion is **monovalent**, meaning that it has lost one electron and has a $+1$ charge. The symbol for the magnesium cation would be Mg^{2+} or Mg^{++} to indicate that it has lost two electrons and has a $+2$ charge, so the magnesium cation would be referred to as a **divalent** cation.

The cations are simply named as the parent element. The sodium cation is still called "sodium". Often, the charge would be attached for clarity, so the sodium cation might be referred to as "sodium one plus".

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Applications of Cations

Cations play important roles in our daily lives. Sodium, potassium, and magnesium ions are essential for such processes as blood pressure regulation and muscle contraction. Calcium ions are an important part of bone structure. Sodium ions can be used in water softeners to remove other harmful elements. We put sodium chloride (table salt) on our food and use it as a preservative.

Summary

- Cations are formed by the loss of one or two electrons from an element.
- Groups 1 and 2 elements form cations.
- Cations are named according to the parent element.
- Cation charges are indicated with a superscript following the chemical symbol.

Review

1. What is a cation?
2. Write the symbol for the barium cation.
3. Write the symbol for the cesium cation.
4. List three ways cations are useful.

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7.4: Anions



Figure 7.4.1 (Public Domain; Shane Anderson/NOAA via Commons Wikimedia, Kelp 300 [commons.wikimedia.org])

What does the amount of salt in seaweed tell us?

Before iodized salt was developed, some people experienced a number of developmental difficulties, including problems with thyroid gland function and intellectual disabilities. In the 1920s, we learned that these conditions could usually be treated easily with the addition of iodide anion to the diet. One easy way to increase iodide intake was to add the anion to table salt. This simple step greatly enhanced health and development. Large amounts of iodide ion are also found in seaweed such as kelp (see picture above) and saltwater fish.

When a metal loses an electron, energy is needed to remove that electron. The other part of this process involves the addition of the electron to another element. The electron adds to the outer shell of the new element. Just as the loss of the electron from the metal produces a full shell, when the electron or electrons are added to the new element, it also results in a full shell.

Anions

Anions are negative ions that are formed when a **nonmetal** atom gains one or more electrons. Anions are so named because they are attracted to the **anode** (positive field) in an electrical field. Atoms typically gain electrons so that they will have the electron configuration of a noble gas. All the elements in Group 17 have seven valence electrons due to the outer $ns^2\ np^5$ configuration. Therefore, each of these elements would gain one electron and become an anion with a -1 charge. Likewise, Group 16 elements form ions with a -2 charge, and the Group 15 nonmetals form ions with a -3 charge.

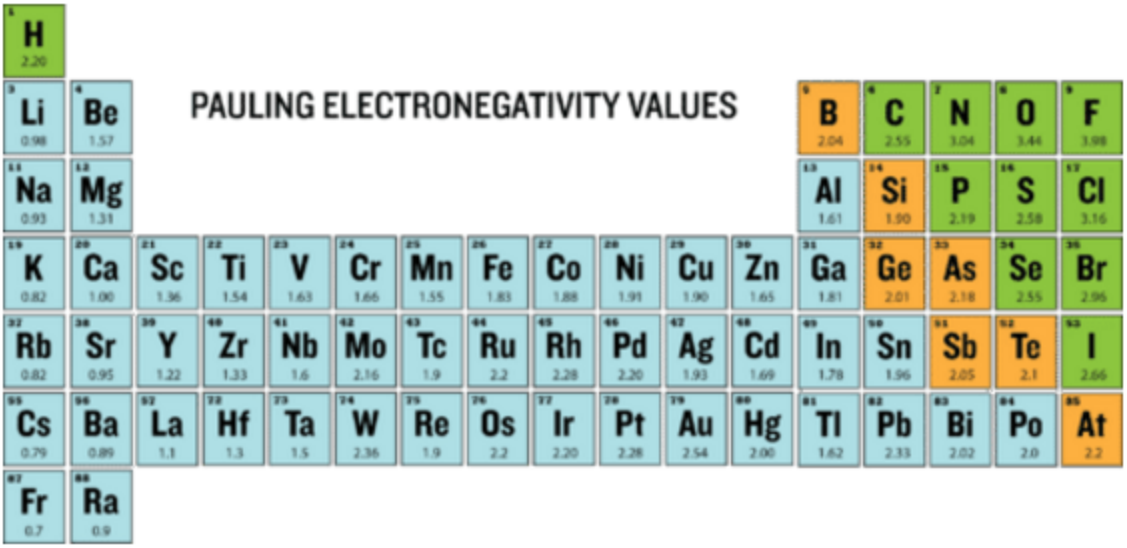


Figure 7.4.2 (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Naming anions is slightly different than naming cations. The ending of the element's name is dropped and replaced with the *-ide* suffix. For example, F^- is the fluoride ion, while O^{2-} is the oxide ion. As is the case with cations, the charge on the anion is indicated by a superscript following the symbol. Common anions are listed in the table below:

Table 7.4.1

Anion Name	Symbol and Charge
fluoride	F^-
chloride	Cl^-
bromide	Br^-
iodide	I^-
oxide	O^{2-}
sulfide	S^{2-}
nitride	N^{3-}

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Uses for Anions

Fluoride ion is widely used in water supplies to help prevent tooth decay. Chloride is an important component in ion balance in blood. Iodide ion is needed by the thyroid gland to make the hormone thyroxine.

Summary

- Anions are formed by the addition of one or more electrons to the outer shell of an atom.
- Group 17 elements add one electron to the outer shell, group 16 elements add two electrons, and group 15 elements add three electrons.
- Anions are named by dropping the ending of the element's name and adding *-ide*.

Review

1. What is an anion?
2. How are anions formed?
3. Why do anions form?
4. How are anions named?
5. List three examples of anions with names, charges, and chemical symbols.
6. List three ways anions are used.

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7.5: Transition Metal Ions



Figure 7.5.1 (Public Domain; the US Mint via Commons Wikimedia, Platinium eagle 101 [commons.wikimedia.org])

What kind of coin is this?

Most of us are familiar with the common coins: penny, nickel, dime, quarter. In some areas (such as Las Vegas), you might see large amounts of silver dollars (these get a little heavy in your pocket). But most of us have probably never seen a platinum eagle—an eagle-depicting coin that is held primarily by collectors. If you were to take a one-ounce platinum eagle into a store to buy one hundred dollars worth of items, the store owner will most likely not believe you when you claim that the coin is worth one hundred dollars. It would also be awkward and annoying if you lost one of these coins out of your pocket. Platinum is just one of several transition metals that is worth a lot of money (gold is another one).

The group 1 and 2 elements form cations through a simple process that involves the loss of one or more outer shell electrons. These electrons come from the s orbital and are removed very readily.

Transition Metal Ions

Most **transition metals** differ from the metals of Groups 1, 2, and 13 in that they are capable of forming more than one cation with different ionic charges. As an example, iron commonly forms two different ions. It can sometimes lose two electrons to form the Fe^{2+} ion, while at other times it loses three electrons to form the Fe^{3+} ion. Tin and lead, though members of the p block rather than the d **block**, also are capable of forming multiple ions.

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Figure 7.5.2: (Public Domain; Cepheus, modified by CK-12 Foundation via Commons Wikimedia, Periodic Table [commons.wikimedia.org])

Ionic formation for transition metals is complicated by the fact that these elements have unfilled inner d shells. Although the next higher s orbitals are actually at a lower energy level than the d level, these s electrons are the ones that are removed during ionization.

The table below lists the names and formulas of some of the common transition metal ions:

Common Transition Metal Ions

Table 7.5.1: Common Transition Metal Ions

1 +	2 +	3 +	4 +
copper (I), Cu^{+}	cadmium, Cd^{2+}	chromium (III), Cr^{3+}	lead (IV), Pb^{4+}
gold (I), Au^{+}	chromium (II), Cr^{2+}	cobalt (III), Co^{3+}	tin (IV), Sn^{4+}
mercury (I), $Hg_2{}^{2+}$	cobalt (II), Co^{2+}	gold (III), Au^{3+}	
silver, Ag^{+}	copper (II), Cu^{2+}	iron (III), Fe^{3+}	
	iron (II), Fe^{2+}		
	lead (II), Pb^{2+}		
	manganese (II), Mn^{2+}		
	mercury (II), Hg^{2+}		
	nickel (II), Ni^{2+}		
	platinum (II), Pt^{2+}		
	tin (II), Sn^{2+}		
	zinc, Zn^{2+}		

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Uses for Transition Metals

Because there are so many metals in this group, there are a wide variety of uses. Many of the metals are used in electronics, while others (such as gold and silver) are used in monetary systems. Iron is a versatile structural material. Cobalt, nickel, platinum, and other metals are employed as catalysts in a number of chemical reactions. Zinc is a significant component of batteries.

Summary

- Transition metals have unfilled inner d electron shells.
- Ions form primarily through loss of s electrons.
- Many transition metals can form more than one ion.
- Transition metals have a wide variety of applications.

Review

1. What is unique about the electron configurations of transition metals?
2. Which electrons of transition metal elements are most likely to be lost during ion formation?
3. How many ions can iron form?
4. Which transition metal forms only one ion?
5. List several uses for transition metals.

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7.6: Naming Binary Ionic Compounds



Figure 7.6.1 (Public Domain; Carl Gustav Calwer, Gustav Jäger, Emil Hochdanz via Wikipedia)

Do you know the proper name for this insect?

Proper naming (or *nomenclature*) is important for identification purposes. Medicine names must be precise so that the correct drug is given—one that will help the patient and not harm them. Biological classification of species requires accurate naming for proper categorization. Insects are categorized by genus and species, so they are uniquely identified. Names must be correct for the tracing of family trees in genealogical studies. Compounds used in chemical reactions must be correctly specified in order for the reaction to occur.

Binary Ionic Compounds

A **binary** ionic compound is a compound composed of a **monatomic** metal **cation** and a monatomic nonmetal **anion**.

Naming Binary Ionic Compounds

When examining the formula of a compound in order to name it, you must first decide what kind of compound it is. For a binary ionic compound, a metal will always be the first element in the formula, while a nonmetal will always be the second. The metal cation is named first, followed by the nonmetal anion. Subscripts in the formula do not affect the name. The table below shows three examples.

Table 7.6.1: Naming Binary Ionic Compounds

Formula	Name
KF	potassium fluoride
$\mathrm{Na_3N}$	sodium nitride
$\mathrm{Ca_3P_2}$	calcium phosphide

Notice that in each of the formulas above, the overall charge of the compound is zero. Potassium ion is $\mathrm{K^+}$, while fluoride ion is $\mathrm{F^-}$. Since the magnitude of the charges is equal, the formula contains one of each ion. This would also be the case for a compound such as MgS, in which the ions are $\mathrm{Mg^{2+}}$ and $\mathrm{S^{2-}}$. For sodium nitride, the sodium ion is $\mathrm{Na^+}$, while the nitride ion is $\mathrm{N^{3-}}$. In order to make a neutral compound, three of the $1+$ sodium ions are required in order to balance out the single $3-$ nitride ion. So the Na is given a subscript of 3. For calcium phosphide, the calcium ion is $\mathrm{Ca^{2+}}$, while the phosphide ion is $\mathrm{P^{3-}}$.

The least common multiple of 2 and 3 is 6. To make the compound neutral, three calcium ions have a total charge of $6+$, while two phosphide ions have a total charge of $6-$. The Ca is given a subscript of 3, while the P is given a subscript of 2.

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Summary

- A binary ionic compound is a compound composed of a monatomic metal cation and a monatomic nonmetal anion.
- The metal part of the compound is named as the element.
- The nonmetallic part of the compound is named by dropping the end of the element and adding *-ide*.
- For binary compounds, it is not necessary to indicate the number of ions in the compound.

Review

1. What is a binary compound?
2. Which of the following is not a binary compound?
   a. NaCl
   b. KH2PO4
   c. KBr
3. Name the following compounds:
   a. NaBr
   b. MgCl2
   c. LiI
   d. CaO
   e. CuBr2
   f. FeO

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7.7: Stock System Naming



Figure 7.7.1 (Public Domain; XAVeRY via Wikipedia)

How is an mp3 player designed?

In describing many technological items, it's not enough to simply say what brand or model we have. We talk about details such as how much horsepower is “under the hood” for a car or how fast the chip is for our computer. Even a simple device like an mp3 player has more than one size. We can get an 8 MB player, or a 16 MB player. Designation of the item often is incomplete without other information as to its capabilities. Descriptive nomenclature such as we use in chemistry provides such information for compounds.

Transition metals have more than one possibility for ion formation. In order to name these compounds correctly, we need to be able to indicate which ion is involved in any given compound.

Naming Compounds Using the Stock System

Naming compounds that involve transition metal cations necessitates use of the **Stock system**. Consider the binary ionic compound $\mathrm{FeCl_3}$. To simply name this compound "iron chloride" would be incomplete, because iron is capable of forming two ions with different charges. The name of any iron-containing compound must reflect which iron ion is in the compound. In this case, the subscript in the formula indicates that there are three chloride ions, each with a $1-$ charge. Therefore, the charge of the single iron ion must be $3+$. The correct name of $\mathrm{FeCl_3}$ is iron (III) chloride, with the cation charge written as the Roman numeral. Here are several other examples:

Table 7.7.1

Formula	Name
$\mathrm{Cu_2O}$	copper (I) oxide
CuO	copper (II) oxide
$\mathrm{SnO_2}$	tin (IV) oxide

The first two examples are both oxides of copper (shown in the figure below). The ratio of copper ions to oxide ions determines the name. Since the oxide ion is $\mathrm{O^{2-}}$, the charges of the copper ion must be $1+$ in the first formula and $2+$ in the second formula. In the third formula, there is one tin ion for every two oxide ions. This means that the tin must carry a $4+$ charge, making the name tin (IV) oxide.

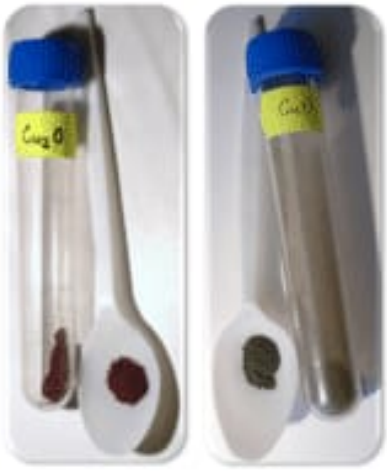


Figure 7.7.2: Copper (I) oxide, a red solid, and copper (II) oxide, a black solid, are different compounds because of the charge of the copper ion. (Public Domain; Ondřej Mangl via Copper(I) oxide: http://commons.wikimedia.org/wiki/File:Cu2O.png; Copper(II) oxide: http://commons.wikimedia.org/wiki/File:Oxid_m%25C4%259B%25C4%258Fnat%25C3%25BD.PNG)

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Summary

- The Stock system allows for the specification of transition metal ionic charge when naming ionic compounds.
- Roman numerals are used to indicate the amount of positive charge on the cation.

Review

1. What is the Stock system?
2. For which group of metal ions would we use the Stock system?
3. What does the Roman numeral stand for?
4. Assign a Roman numeral to each of the following cations:
   a. Sn^{4+}
   b. Fe^{3+}
   c. Co^{2+}
   d. Pb^{4+}

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7.8: Formulas for Binary Ionic Compounds

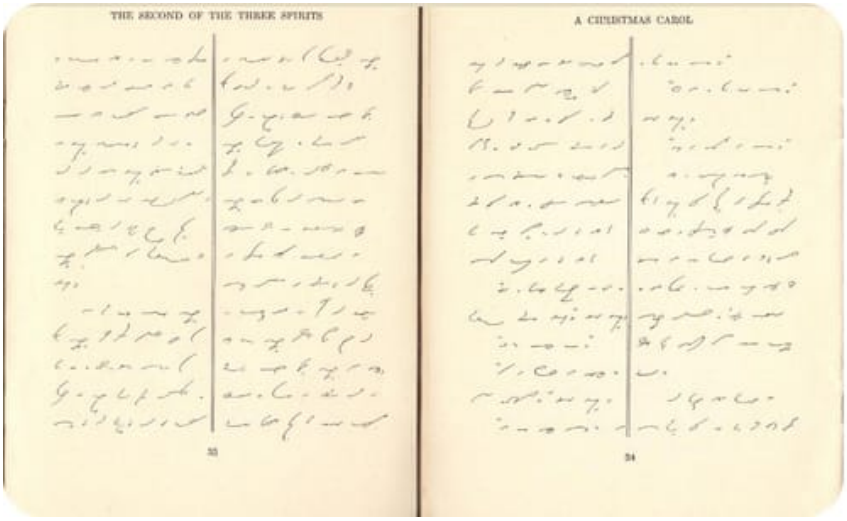


Figure 7.8.1 (Public Domain; Winifred Kenna via Commons Wikimedia, Gregg shorthand A Christmas Carol [commons.wikimedia.org])

How does shorthand work?

Shorthand was a very popular way of recording speech, especially in dictating letters and in court testimony. Instead of trying to write out all the words, the person taking the dictation would use a set of symbols that represented syllables or words. The pages above show a shorthand version of "A Christmas Carol" written by Charles Dickens. Unless you know shorthand, the passage is meaningless. But knowing shorthand allows you to read this classic story. Different professions also use a type of shorthand in communication to save time. Chemists use chemical symbols in combination to indicate specific compounds. There are two advantages to this approach:

1. The compound under discussion is clearly described so that there can be no confusion about its identity.
2. Chemical symbols represent a universal language that all chemists can understand, no matter what their native language is.

Writing Formulas for Binary Ionic Compounds

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. Consider the compound aluminum nitride. The ions are:

[redacted]

Since the ions have charges that are equal in magnitude, one of each will be the lowest ratio of ions in the formula. The formula for aluminum nitride is AlN.

The ions for the compound lithium oxide are:

Li^{+} [redacted]

In this case, two lithium ions are required to balance out the charge of one oxide ion. The formula of lithium oxide is $\mathrm{Li_2O}$.

An alternative way to write a correct formula for an ionic compound is to use the crisscross method. In this method, the numerical value of each of the ion charges is crossed over to become the subscript of the other ion. Signs of the charges are dropped. Shown below is the crisscross method for aluminum oxide.

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The red arrows indicate that the 3 from the $3+$ charge will cross over to become the subscript of the O. The 2 from the $2-$ charge will cross over to become the subscript of the Al. The formula for aluminum oxide is $\mathrm{Al_2O_3}$.

Be aware that ionic compounds are empirical formulas and so must be written as the lowest ratio of the ions. In the case of aluminum nitride, the crisscross method would yield a formula of $\mathrm{Al_3N_3}$, which is not correct. It must be reduced to AlN. Following the crisscross method to write the formula for lead (IV) oxide would involve the following steps:

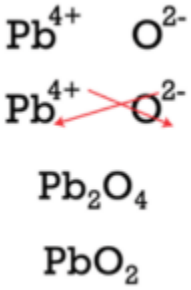


Figure 7.8.3 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The crisscross first yields $\mathrm{Pb_2O_4}$ for the formula, but that must be reduced to the lower ratio, and $\mathrm{PbO_2}$ is the correct formula.

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Summary

- Formulas for binary compounds begin with the metal followed by the nonmetal.
- Positive and negative charges must cancel each other out.
- Ionic compound formulas are written using the lowest ratio of ions.

Review

1. Write formulas for the binary ionic compounds formed between the following pairs of elements:
   a. cesium and fluorine
   b. calcium and sulfur
   c. aluminum and chlorine
   d. zinc and nitrogen
2. Write the formula and give the name for the compound formed by the following ions:
   a. Fe^{3+} and O^{2-}
   b. Ni^{2+} and S^{2-}
   c. Au^{+} and Cl^{-}
   d. Sn^{4+} and I^{-}
3. Give names for the following compounds:
   a. Ag_2S
   b. PdO
   c. $PtCl_4$
   d. V_2O_5

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7.9: Polyatomic Ions

Shakespeare wrote, "A rose by any other name would smell as sweet" as part of the romantic balcony scene between Romeo and Juliet in the famous play. The two families were bitter rivals, but Juliet meant this statement to mean that she loved Romeo no matter what his name was. Some names are simple—we know Romeo mainly as Romeo. Most of us have a first name, middle name (often not used), and last name. In some cultures, names will be much more complex. The full name of the famous 20th century artist Pablo Picasso is Pablo Diego José Francisco de Paula Juan Neponuceno María de los Remedios Cipriano de la Santísima Trinidad Martyr Patricio Clito Ruíz y Picasso.

Many materials exist as simply binary compounds composed of a metal cation and a nonmetal anion, with each ion consisting of only one type of atom. Other combinations of atoms also exist, either as larger ionic complexes or complete molecules. Some of the most useful materials we work with contain polyatomic ions.

Polyatomic Ions

A **polyatomic** ion is an ion composed of more than one atom. The ammonium ion consists of one nitrogen atom and four hydrogen atoms. Together, they comprise a single ion with a $1+$ charge and a formula of $\text{NH}_4{}^+$. The carbonate ion consists of one carbon atom and three oxygen atoms, and carries an overall charge of $2-$. The formula of the carbonate ion is $\text{CO}_3{}^{2-}$. The atoms of a polyatomic ion are tightly bonded together, and so the entire ion behaves as a single unit. The figures below show several models.

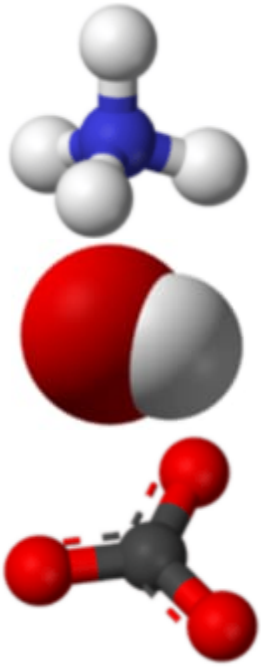


Figure 7.9.1**:** *(Top) The ammonium ion* $(\text{NH}_4{}^+)$ *is a nitrogen atom (blue) bonded to four hydrogen atoms (white). (Middle) The hydroxide ion* (OH^-) *is an oxygen atom (red) bonded to a hydrogen atom. (Bottom) The carbonate ion* $(\text{CO}_3{}^{2-})$ *is a carbon atom (black) bonded to three oxygen atoms.*

The table below lists a number of polyatomic ions by name and by formula. The heading for each column indicates the charge on the polyatomic ions in that group. Note that the vast majority of the ions listed are anions—there are very few polyatomic cations.

Table 7.9.1*: Common Polyatomic Ions*

$1-$	$2-$	$3-$	$1+$	$2+$
acetate, CH_3COO^-	carbonate, $\text{CO}_3{}^{2-}$	arsenate, $\text{AsO}_3{}^{3-}$	ammonium, $\text{NH}_4{}^+$	dimercury, $\text{Hg}_2{}^{2+}$

1 −	2 −	3 −	1 +	2 +
bromate, BrO_3^-	chromate, CrO_4^{2-}	phosphite, PO_3^{3-}		
chlorate, ClO_3^-	dichromate, $Cr_2O_7^{2-}$	phosphate, PO_4^{3-}		
chlorite, ClO_2^-	hydrogen phosphate, HPO_4^{2-}			
cyanide, CN^-	oxalate, $C_2O_4^{2-}$			
dihydrogen phosphate, $H_2PO_4^-$	peroxide, O_2^{2-}			
bicarbonate, HCO_3^-	silicate, SiO_3^{2-}			
hydrogen sulfate, HSO_4^-	sulfate, SO_4^{2-}			
hydrogen sulfide, HS^-	sulfite, SO_3^{2-}			
hydroxide, OH^-				
hypochlorite, ClO^-				
nitrate, NO_3^-				
nitrite, NO_2^-				
perchlorate, ClO_4^-				
permanganate, MnO_4^-				

The vast majority of polyatomic ions are anions, many of which end in *-ate* or *-ite*. Notice that in some cases, such as nitrate (NO_3^-) and nitrite (NO_2^-), there are multiple anions that consist of the same two elements. In these cases, the difference between the ions is in the number of oxygen atoms present, while the overall charge is the same. As a class, these are called oxoanions. When there are two oxoanions for a particular element, the one with the greater number of oxygen atoms gets the *-ate* suffix, while the one with the fewer number of oxygen atoms gets the *-ite* suffix. The four oxoanions of chlorine are as follows:

- ClO^-, hypochlorite
- ClO_2^-, chlorite
- ClO_3^-, chlorate
- ClO_4^-, perchlorate

In cases such as this, the ion with one more oxygen atom than the *-ate* anion is given a *per-* prefix. The ion with one fewer oxygen atom than the *-ite* anion is given a *hypo-* prefix.

Summary

- Polyatomic ions contain more than one type of atom in the ion.
- The majority of polyatomic ions are anions that are named ending in *-ate* or *-ite*.
- Some anions have multiple forms and are named accordingly.

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7.10: Names and Formulas of Ternary Ionic Compounds



Figure 7.10.1 (Public Domain; Jon Sullivan via Commons Wikimedia, Pepperoni Pizza [commons.wikimedia.org])

How many varieties of pizza can you think of?

Many people enjoy eating pizza. When someone orders a pizza (either by phone, online, or in a restaurant), they don't just order "pizza". Even the simplest "binary" pizza will have a topping and cheese—maybe pepperoni, maybe something else. However, many pizzas quickly become more complex. One person may want pepperoni and sausage, another may wish to order Canadian bacon and pineapple; then you have the folks that can't decide, so they order half-this and half-that. The combinations may be more complex, but the same basic ideas about pizza are valid.

Ternary Ionic Compounds

Not all ionic compounds are composed of only monatomic ions. A **ternary ionic compound** is an ionic compound composed of three elements. In a typical ternary ionic compound, there is still one type of cation and one type of anion involved. The cation, the anion, or both, are polyatomic ions.

Naming Ternary Ionic Compounds

The process of naming ternary ionic compounds is the same as naming binary ionic compounds. The cation is named first, followed by the anion. Some examples are shown in the table below:

Table 7.10.1: Examples of Ternary Ionic Compounds

Formula	Name
$\mathrm{NaNO_3}$	sodium nitrate
$\mathrm{NH_4Cl}$	ammonium chloride
$\mathrm{Fe(OH)_3}$	iron (III) hydroxide

When more than one polyatomic ion is present in a compound, the formula of the ion is placed in parentheses with a subscript outside of the parentheses that indicates how many of those ions are in the compound. In the last example above, there is one $\mathrm{Fe^{3+}}$ cation and three $\mathrm{OH^-}$ anions.

Writing Formulas for Ternary Ionic Compounds

Writing a formula for a ternary ionic compound 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. Use the crisscross method to ensure that the final formula is neutral. Calcium nitrate is composed of a calcium cation and a nitrate anion.

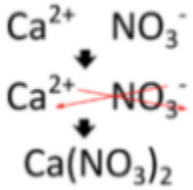


Figure 7.10.2 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The charge is balanced by the presence of two nitrate ions and one calcium ion. Parentheses are used around the nitrate ion because more than one of the polyatomic ion is needed. If only one polyatomic ion is in a formula, parentheses are not used. As an example, the formula for calcium carbonate is $\mathrm{CaCO_3}$. The carbonate ion carries a $2-$ charge and so exactly balances the $2+$ charge of the calcium ion.

There are two polyatomic ions that produce unusual formulas. The $\mathrm{Hg_2}^{2+}$ ion is called either the dimercury ion or, preferably, the mercury (I) ion. When bonded with an anion with a $1-$ charge, such as chloride, the formula is $\mathrm{Hg_2Cl_2}$. Because the cation consists of two Hg atoms bonded together, this formula is not reduced to HgCl. Likewise, the peroxide ion, $\mathrm{O_2}^{2-}$, is also a unit that must stay together in its formulas. For example, the formula for potassium peroxide is $\mathrm{K_2O_2}$.

Summary

- Ternary compounds are composed of three or more elements.
- Ternary compounds are named by stating the cation first, followed by the anion.
- Positive and negative charges must balance.

Review

1. What is a ternary compound?
2. What is the basic rule for naming ternary compounds?
3. Write the formulas for the following compounds:
   a. mercury(II) nitrate
   b. ammonium phosphate
   c. calcium silicate
   d. lead(II) chromate
4. Name the following compounds:
   a. $KClO_3$
   b. Rb_2SO_4
   c. $Cd(NO_3)_2$
   d. NaCN

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7.11: Naming Binary Molecular Compounds



Figure 7.11.1 (Public Domain; Anonymous via Wikipedia)

Why do so many relatives in royalty share the same name?

Some families name a son (usually the first born) after his father. So, it is somewhat common to meet a John Smith, Jr. named after John Smith the father. Certain families may name a grandson John Smith III. Countries with long histories of royalty take the naming even further. One line of kings named Henry goes up to Henry the Eighth (not the nicest guy in the world—he had six wives and two of them met untimely ends). The use of numbering for names adds clarity to a system—it's easily discernible which Henry is being spoken of.

Inorganic chemical compounds can be broadly classified into two groups: ionic compounds and molecular compounds. The structure of all ionic compounds is an extended three-dimensional array of alternating positive and negative ions. Since ionic compounds do not take the form of individual molecules, they are represented by empirical formulas. Now we will begin to examine the formulas and nomenclature of molecular compounds.

Molecular Compounds

Molecular compounds are inorganic compounds that take the form of discrete **molecules**. Examples include such familiar substances as water $(\mathrm{H_2O})$ and carbon dioxide $(\mathrm{CO_2})$. 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 **valence electrons** in such a way that a **bond** forms between pairs 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.

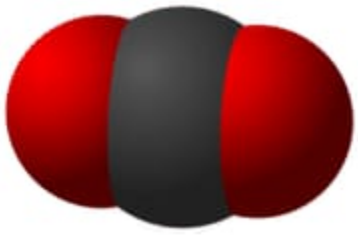


Figure 7.11.2: Carbon dioxide molecules consist of a central carbon atom bonded to two oxygen atoms. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

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 $\mathrm{H_2O}$. 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 $\mathrm{C_8H_{18}}$.



Figure 7.11.3: Nitrogen dioxide $(\mathrm{NO_2})$ is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines. (Public Domain; User:Greenhorn1/Wikimedia Commons via Wikipedia)

A binary molecular compound is a molecular compound that is composed of two elements. The elements that combine to form binary molecular compounds are both nonmetal atoms. This contrasts with ionic compounds, which are formed from a metal ion and a nonmetal ion. Therefore, binary molecular compounds are different because ionic charges cannot be used to name them or to write their formulas. Another difference is that two nonmetal atoms will frequently combine with one another in a variety of ratios. Consider the elements nitrogen and oxygen. They combine to make several compounds including NO, $\mathrm{NO_2}$, and $\mathrm{N_2O}$. They cannot all be called nitrogen oxide. How would someone know which one you were talking about? Each of the three compounds has very different properties and reactivities. A system to distinguish between compounds such as these is necessary.

Prefixes are used in the names of binary molecular compounds to identify the number of atoms of each element. The table below shows the prefixes up to ten.

Table 7.11.1: Numerical Prefixes

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

The rules for using the prefix system of nomenclature of binary compounds can be summarized as follows:

1. 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.
2. When naming, the appropriate prefix is used only if there are more than one atom of that element in the formula.
3. The second element is named after the first, but with the ending of the element's name changed to *-ide*. The appropriate prefix is always used for the second element.
4. 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 are tetroxide instead of tetraoxide.

Some examples of molecular compounds are listed in Table 7.11.2.

Table 7.11.2

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.

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Summary

- Molecular compounds are inorganic compounds that take the form of discrete molecules.
- The atoms of these compounds are held together by covalent bonds.
- Prefixes are used to indicate the number of atoms of an element that are in the compound.

Review

1. What are molecular compounds?
2. How do they differ from ionic compounds?
3. What holds the atoms of a molecule together?
4. Are the elements of molecular compounds metals or non-metals?

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5. What is the prefix for the number of atoms of an element when there are four atoms of that element in the compound?
6. What is the prefix for the number of atoms of an element when there are 7 atoms of that element in the compound?
7. Name the following compounds:
   a. ClF_3
   b. As_2O_5
   c. B_4H_{10}

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7.12: Naming Acids



Figure 7.12.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

How is gold tested?

A spot test for gold has been in use for decades. The sample is first treated with nitric acid. Other metals may react or dissolve in this acid, but gold will not. Then the sample is added to a mixture of nitric acid and hydrochloric acid. Gold will only dissolve in this mixture. The term "acid test" arose from the California gold rush in the late 1840s, when this combination was used to test for the presence of real gold. It has since come to mean "tested and approved" in a number of fields.

Acids

An **acid** can be defined in several ways. The most straightforward definition is that an acid is a molecular compound that contains one or more hydrogen atoms and produces hydrogen ions (H^+) when dissolved in water.



Figure 7.12.2: (A) Vinegar comes in a variety of types, but all contain acetic acid. (B) Citrus fruits like grapefruit contain citric and ascorbic acids. (CC BY 2.0; (A) Becky Cortino; (B) Flickr: isox4 via (A) http://www.flickr.com/photos/mediasavvy/8239231530; (B) http://www.flickr.com/photos/isox4/5167980026)

This is a different type of compound than the others we have seen so far. Acids are molecular, which means that in their pure state they are individual molecules and do not adopt the extended three-dimensional structures of ionic compounds like NaCl. However, when these molecules are dissolved in water, the chemical bond between the hydrogen atom and the rest of the molecule breaks, leaving a positively-charged hydrogen ion and an anion. This can be symbolized in a chemical equation:

$$HCl \rightarrow H^+ + Cl^-$$

Since acids produce H^+ cations upon dissolving in water, the H of an acid is written first in the formula of an inorganic acid. The remainder of the acid (other than the H) is the anion after the acid dissolves. Organic acids are also an important class of compounds, but will not be discussed here. A **binary acid** is an acid that consists of hydrogen and one other element. The most common binary acids contain a halogen. An **oxoacid** is an acid that consists of hydrogen, oxygen, and a third element. The third element is usually a nonmetal.

Naming Acids

Since all acids contain hydrogen, the name of an acid is based on the anion that goes with it. These anions can either be monatomic or polyatomic. The name of all monatomic ions ends in *-ide*. The majority of polyatomic ions end in either *-ate* or *-ite*, though there are a few exceptions such as the cyanide ion (CN^-). It is the suffix of the anion that determines how the acid is named, as displayed in the rules and table below.

Table 7.12.1: Naming System for Acids

Anion Suffix	Example	Name of Acid	Example
-ide	chloride (Cl^-)	hydro____ic acid	hydrochloric acid (HCl)
-ate	sulfate $(SO_4{}^{2-})$	_____ic acid	sulfuric acid (H_2SO_4)
-ite	nitrite $(NO_2{}^-)$	___ous acid	nitrous acid (HNO_2)

The three different suffixes that are possible for the anions lead to three rules:

1. When the anion ends in *-ide*, the acid name begins with the prefix *hydro-*. The root of the anion name goes in the blank (*chlor* for chloride), followed by the suffix *-ic*. HCl is hydrochloric acid because Cl^- is the chloride ion. HCN is hydrocyanic acid because CN^- is the cyanide ion.
2. When the anion ends in *-ate*, the name of the acid is the root of the anion followed by the suffix *-ic*. There is no prefix. H_2SO_4 is sulfuric acid (not sulfic) because $SO_4{}^{2-}$ is the sulfate ion.
3. When the anion ends in *-ite*, the name of the acid is the root of the anion followed by the suffix *-ous*. Again, there is no prefix. HNO_2 is nitrous acid because $NO_2{}^-$ is the nitrite ion.

Note how the root for a sulfur-containing oxoacid is *sulfur-* instead of just *sulf-*. The same is true for a phosphorus-containing oxoacid. The root is *phosphor-* instead of simply *phosph-*.

Many foods and beverages contain citric acid. Vinegar is a dilute solution of acetic acid. Car batteries contain sulfuric acid that helps in the release of electrons to create electricity.

Writing Formulas for Acids

Like other compounds that we have studied, acids are electrically neutral. Therefore, the charge of the anion part of the formula must be exactly balanced out by the H^+ ions. Since H^+ ions carry a single positive charge, the number of H^+ ions in the formula is equal to the quantity of negative charge on the anion. Two examples from the table above illustrate this point. The chloride ion carries a $1-$ charge, so only one H is needed in the formula of the acid (HCl). The sulfate ion carries a $2-$ charge, so two Hs are needed in the formula of the acid (H_2SO_4). Another way to write the correct formula is to utilize the crisscross method, shown below for sulfuric acid.

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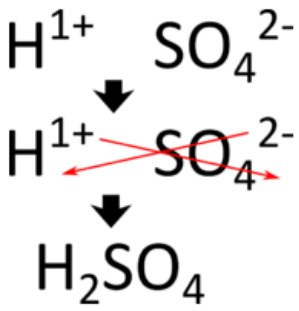


Figure 7.12.3: Crisscross approach to writing the formula for sulfuric acid. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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Summary

- Acids are molecular compounds that release hydrogen ions.
- A binary acid consists of hydrogen and one other element.
- Oxoacids contain hydrogen, oxygen, and one other element.
- The name of the acid is based on the anion attached to the hydrogen.
- Writing formulas for acids follows the same rules as for binary ionic compounds.

Review

1. Define “acid.”
2. What is a binary acid?
3. What is an oxoacid?
4. Name the following acids:
   a. H_2SO_4
   b. HCN
   c. HCl
   d. H_3PO_4
5. Write formulas for the following acids:
   a. hydrobromic acid
   b. perchloric acid
   c. nitrous acid

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7.13: Names and Formulas of Bases



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What different things are these workers doing to make soap?

Soap making has a long history. Until recently, soap was made using animal fats and lye from wood ashes. The lye served as a base to break down the fats and help form the soap. Needless to say, unless the soap was washed to remove the lye, it was very harsh on the skin. Many families would make their own soap by boiling the lye and fat in a large kettle over an open fire—a long and hot task.

Bases

A **base** can be simply defined as an ionic compound that produces **hydroxide ions** when dissolved in water. One of the most commonly used bases is sodium hydroxide, illustrated below.

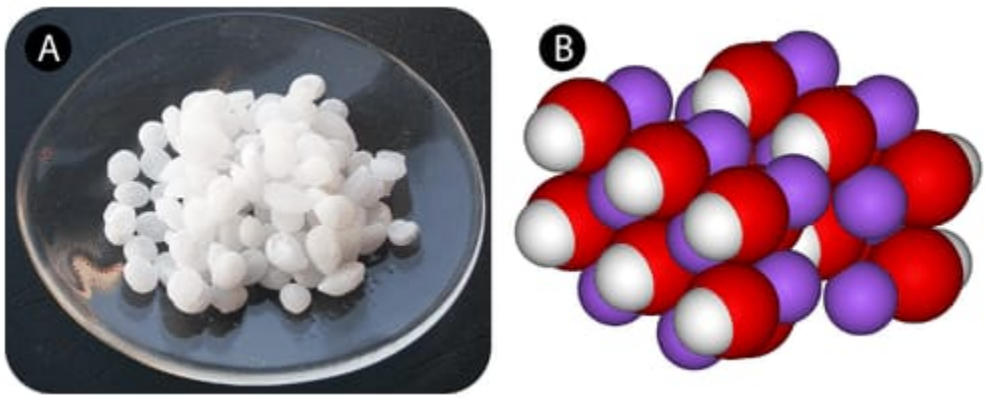


Figure 7.13.2: (A) Sodium hydroxide, a base, is a solid that is typically produced as small white pellets. (B) The structure of sodium hydroxide is an extended three-dimensional network. The purple spheres are the sodium ions (Na^+). The red and white spheres are oxygen and hydrogen atoms, respectively, which are bonded together to form hydroxide ions (OH^-). (Public Domain; (A) Martin Walker (Wikimedia: Walkerma); (B) Ben Mills (Wikimedia: Benjah-bmm27) via (A) http://commons.wikimedia.org/wiki/File:SodiumHydroxide.jpg; (B) http://commons.wikimedia.org/wiki/File:Sodium-hydroxide-crystal-3D-vdW.png)

Names and Formulas of Bases

There is no special system for naming bases. Since they all contain the OH^- anion, names of bases end in *hydroxide*. The cation is simply named first. Some examples of names and formulas for bases are shown in the table below.

Names and Formulas of Bases

Table 7.13.1

Formula	Name
$NaOH$	sodium hydroxide
$Ca(OH)_2$	calcium hydroxide
NH_4OH	ammonium hydroxide

Notice that because bases are ionic compounds, the number of hydroxides in the formula does not affect the name. The compound must be neutral, so the charges of the ions are balanced just as for other ionic compounds. The sodium ion (Na^+) requires one OH^- ion to balance the charge, so the formula is $NaOH$. Calcium ([redacted]) requires two OH^- ions to balance the charge, so the formula is $Ca(OH)_2$. Hydroxide ion is a polyatomic ion, and must be in parentheses when there is more than one in a formula.

Summary

- Bases are ionic compounds that produce hydroxide ions when dissolved in water.
- The cation is named first followed by "*hydroxide*."

Review

1. What is a base?
2. What is the charge on the hydroxide anion?
3. Name the following bases:
   a. LiOH
   b. $Mg(OH)_2$
   c. $Fe(OH)_3$
4. Write the formulas for the following bases:
   a. nickel (II) hydroxide
   b. aluminum hydroxide
   c. silver hydroxide

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8: IONIC BONDING

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8.3: Cation Formation

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8.5: Transition Metal Ion Formation

8.6: Ionic Bonding

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8.1: Electron Dot Diagrams

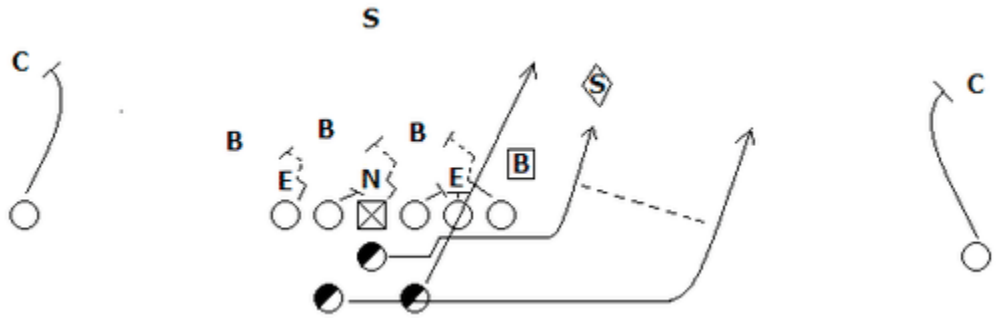


Figure 8.1.1 (Public Domain; User:Veatchw/Wikipedia via Commons Wikimedia, Veer vs 34 [commons.wikimedia.org])

How do we show electrons in atoms?

Diagrams contain a lot of useful information in a compact format. What does the diagram above tell us? The football play diagrammed above describes the lineup of each player on the team and describes how they will move when the ball is snapped. Diagrams of electrons give similar information about where certain electrons are. We can mark these electrons and indicate what happens to them when an element reacts.

Electron Dot Diagrams

Recall that the valence electrons of an atom are the electrons located in the highest occupied principal energy level. Valence electrons are primarily responsible for the chemical properties of elements. The number of valence electrons can be easily determined from the electron configuration. Several examples from the second period elements are shown in the table below.

Table 8.1.1:

Element	Electron Configuration	Number of Valence Electrons
lithium	$1s^2\ 2s^1$	1
beryllium	$1s^2\ 2s^2$	2
nitrogen	$1s^2\ 2s^2\ 2p^3$	5
neon	$1s^2\ 2s^2\ 2p^6$	8

In each case, valence electrons are those in the second principal energy level. As one proceeds left to right across a period, the number of valence electrons increases by one. In the s block, Group 1 elements have one valence electron, while Group 2 elements have two valence electrons. In the p block, the number of valence electrons is equal to the group number minus ten. Group 13 has three valence electrons, Group 14 has four, and so on...up through Group 18 with eight. The eight valence electrons, a full outer s and p sublevel, give the noble gases their special stability.

When examining chemical bonding, it is necessary to keep track of the valence electrons of each atom. **Electron dot diagrams** are diagrams in which the valence electrons of an atom are shown as dots distributed around the element's symbol. A beryllium atom, with two valence electrons, has the electron dot diagram below:

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Since electrons repel each other, the dots for a given atom are distributed evenly around the symbol before they are paired. The table below shows the electron dot diagrams for the entire second period.

Table 8.1.2: Electron Dot Diagrams for the Second Period Elements

Electron Dot Diagrams for the Second Period Elements

Group Number	Electron Dot Diagram
1	 Figure 8.1.3 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
2	 Figure 8.1.4 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
13	 Figure 8.1.5 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
14	 Figure 8.1.6 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
15	 Figure 8.1.7 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
16	 Figure 8.1.8 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
17	 Figure 8.1.9 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)
18	 Figure 8.1.10 (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)

Electron dot diagrams would be the same for each element in the representative element groups. Most transition elements have two valence electrons, though some that have unusual electron configurations have only one.

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Summary

- Electron dot diagrams show the valence electrons for an atom.
- The dot diagrams are the same for each element in the representative element groups.

Review

1. What are valence electrons primarily responsible for?
2. Calcium would have the same electron dot structure as which element pictured in the table?
3. What is the symbol for an element that would have the same electron dot structure as carbon?
4. Would you expect the group 18 elements to have the same electron dot diagram as neon?

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8.2: Octet Rule



Figure 8.2.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

How are electrons organized in atoms?

Graduations are exciting events in one's life. The processional, the ceremonies, even the speeches can be handled because this marks the end of a stage of life and the beginning of another one. The ceremony itself is challenging for those organizing it. There are just enough seats on the stage for the graduates. How do you avoid getting too many people in a row and not enough in the next row? Someone is stationed at the end of the row to count the students as they enter. Only so many are allowed to go into a row, and then you begin to fill the next row. Electrons in atoms behave the same way. There are rules that determine where electrons go in compounds.

Octet Rule

The noble gases are unreactive because of their electron configurations. The noble gas neon has the electron configuration of $1s^2\ 2s^2\ 2p^6$. It has a full outer shell and cannot incorporate any more electrons into the valence shell. The other noble gases have the same outer shell electron configuration, even though they have different numbers of inner-shell electrons.



Figure 8.2.2: Electron configuration of neon atom. (CC BY-NC-SA 3.0; Joy Sheng via CK-12 Foundation)

American chemist Gilbert Lewis (1875-1946) used this observation to explain the types of ions and molecules that are formed by other elements. He called his explanation the **octet rule**. The octet rule states that atoms tend to form compounds in ways that give them eight valence electrons, and thus the electron configuration of a noble gas. An exception to an octet of electrons is in the case of the first noble gas, helium, which only has two valence electrons. This primarily affects the element hydrogen, which forms stable compounds by achieving two valence electrons. Lithium, an alkali metal with three valence electrons, is also an exception to the octet rule. Lithium tends to lose one electron to take on the electron configuration of the nearest noble gas, helium, leaving it with two valence electrons.

There are two ways in which atoms can satisfy the octet rule. One way is by sharing their valence electrons with other atoms. The second way is by transferring valence electrons from one atom to another. Atoms of metals tend to lose all of their valence electrons, which leaves them with an octet from the next lowest principal energy level. Atoms of nonmetals tend to gain electrons in order to fill their outermost principal energy level with an octet.

Watch the following video to learn how to use the octet rule to predict the charge of an ion.

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Summary

- Atoms form compounds in ways that give them eight valence electrons.
- Metals tend to lose electrons to achieve this configuration.
- Nonmetals tend to gain electrons to achieve this configuration.

Review

1. What is the electron configuration of a noble gas?
2. Why is this configuration important?
3. What are the two ways in which atoms can satisfy the octet rule?
4. How do metals change to obey the octet rule?
5. How do nonmetals change to obey the octet rule?

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8.3: Cation Formation



Figure 8.3.1 (Public Domain; User:Bugman/Wikipedia via Wikipedia)

How are cations formed?

In many parts of the country, the water contains high concentrations of minerals that stain clothes, build up deposits on bathtubs and water heaters, and create problems with soap foaming properly. This problem is caused by "hard water". The water contains excessive amounts of cations such as iron and calcium. These ions create a lot of problems in the water. Ion exchange resins can remove these minerals and clean up the water.

Cation Formation

Cations are the positive ions formed by the loss of one or more electrons. The most commonly formed cations of the representative elements are those that involve the loss of all of the valence electrons. Consider the alkali metal sodium (Na) . It has one valence electron in the third principal energy level. Upon losing that electron, the sodium ion now has an octet of electrons from the second principal energy level. The equation below illustrates this process.

$$\begin{array}{ll} \mathrm{Na} & \rightarrow \quad \mathrm{Na^+ + e^-} \\ 1s^2\ 2s^2\ 2p^6\ 3s^1 & \quad\ 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

The electron configuration of the sodium ion is now the same as that of the noble gas neon. The term **isoelectronic** refers to an atom and an ion of a different atom (or two different ions) that have the same electron configuration. The sodium ion is isoelectronic with the neon atom. Consider a similar process with magnesium and with aluminum:

$$\begin{array}{ll} \mathrm{Mg} & \rightarrow \quad \blacksquare + 2\mathrm{e^-} \\ 1s^2\ 2s^2\ 2p^6\ 3s^2 & \quad\ 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

$$\begin{array}{ll} \mathrm{Al} & \rightarrow \quad \blacksquare + 3\mathrm{e^-} \\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1 & \quad\ 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

In this case, the magnesium atom loses its two valence electrons in order to achieve the same noble gas configuration. The aluminum atom loses its three valence electrons. The $\mathrm{Mg^{2+}}$ ion, the $\mathrm{Al^{3+}}$ ion, the $\mathrm{Na^+}$ ion, and the $ceNe$ atom are all isoelectronic. For representative elements under typical conditions, three electrons are the maximum number that will be lost.

We can also show the loss of valence electron(s) with an electron dot diagram:

$$\mathrm{Na\cdot \rightarrow Na^+ + e^-}$$
$$\mathrm{\cdot Mg\cdot \rightarrow} \blacksquare + 2\mathrm{e^-}$$

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Summary

- Cations form when an atom loses one or more electrons.
- The resulting cation has the electron configuration of the noble gas atom in the row above it in the periodic table.

Review

1. What is a cation?
2. How many valence electrons does the sodium atom have?
3. Which atom is the sodium ion isoelectronic with?
4. How many electrons does magnesium lose to form the magnesium ion?

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8.4: Anion Formation



Figure 8.4.1: A pressed glass open salt dish made in the early 1830s. (Public Domain; User:Daderot/Wikimedia Commons via Wikipedia)

How do you make chlorine safe to eat?

How do you transform a deadly gas into something you can sprinkle on your eggs and eat for breakfast? Chlorine in its free form is very dangerous if you breathe the fumes or come in contact with the gas. However, in reaction with sodium, sodium chloride forms as the sodium atom gives up an electron to chlorine, which accepts the electron to form the chloride anion.

Anions

Anions are the negative ions formed from the gain of one or more electrons. When nonmetal atoms gain electrons, they often do so until their outermost principal energy level achieves an octet. This process is illustrated below for the elements fluorine, oxygen, and nitrogen:

$$\begin{array}{lcl} \mathrm{F} + \mathrm{e}^- & \rightarrow & \mathrm{F}^- \\ 1s^2\ 2s^2\ 2p^5 & & 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

$$\begin{array}{lcl} \mathrm{O} + 2\mathrm{e}^- & \rightarrow & \blacksquare \\ 1s^2\ 2s^2\ 2p^4 & & 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

$$\begin{array}{lcl} \mathrm{N} + 3\mathrm{e}^- & \rightarrow & \blacksquare \\ 1s^2\ 2s^2\ 2p^3 & & 1s^2\ 2s^2\ 2p^6\ (\text{octet}) \end{array}$$

All of these anions are isoelectronic with each other and with neon. They are also isoelectronic with the three cations from the previous section. Under typical conditions, three electrons are the maximum that will be gained in the formation of anions.

Outer electron configurations are constant within a group, so this pattern of ion formation repeats itself for Periods 3, 4, and following (see below).

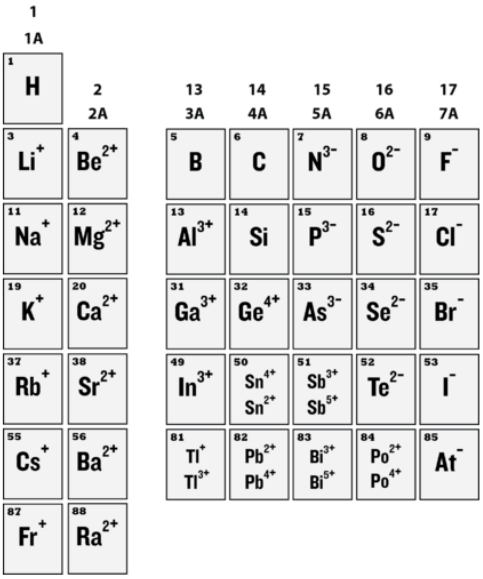


Figure 8.4.2: Ion charges. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

It is important not to misinterpret the concept of being isoelectronic. A sodium ion is very different from a neon atom, because the nuclei of the two contain different numbers of protons. One is an essential ion that is a part of table salt, while the other is an unreactive gas that is a very small part of the atmosphere. Likewise, sodium ions are very different than magnesium ions, fluoride ions, and all the other members of this isoelectronic series ([redacted], [redacted], F^-, Ne, Na^+, [redacted], [redacted]).

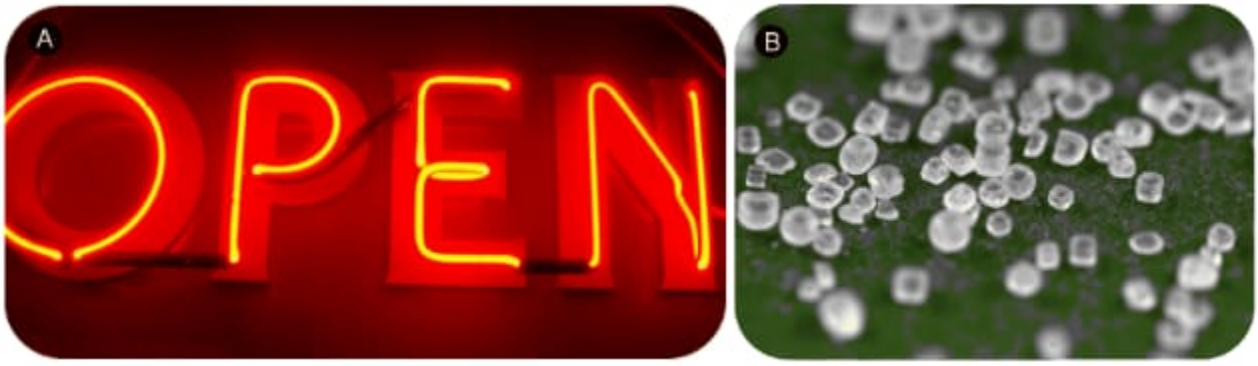


Figure 8.4.3: Neon gas (A) and sodium chloride crystals (B). Neon atoms and sodium ions are isoelectronic. Neon is a colorless and unreactive gas that glows a distinctive red-orange color in a gas discharge tube. Sodium ions are most commonly found in crystals of sodium chloride—ordinary table salt. (CC BY 2.0; (A) Andy Wright (Flickr: rightee); (B) Kevin Dooley;User:Bugman/Wikipedia via (A) http://www.flickr.com/photos/rightee/4356950; (B) http://www.flickr.com/photos/pagedooley/2769134850; http://commons.wikimedia.org/wiki/File:Ion_exchange_resin_beads.jpg)

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Summary

- Anions are negative ions formed by accepting electrons.
- The outermost principal energy level of an anion is usually an octet.

Review

1. What is an anion?
2. Write the electronic configurations for the chlorine atom and the chloride anion.
3. What does isoelectronic mean?

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8.5: Transition Metal Ion Formation

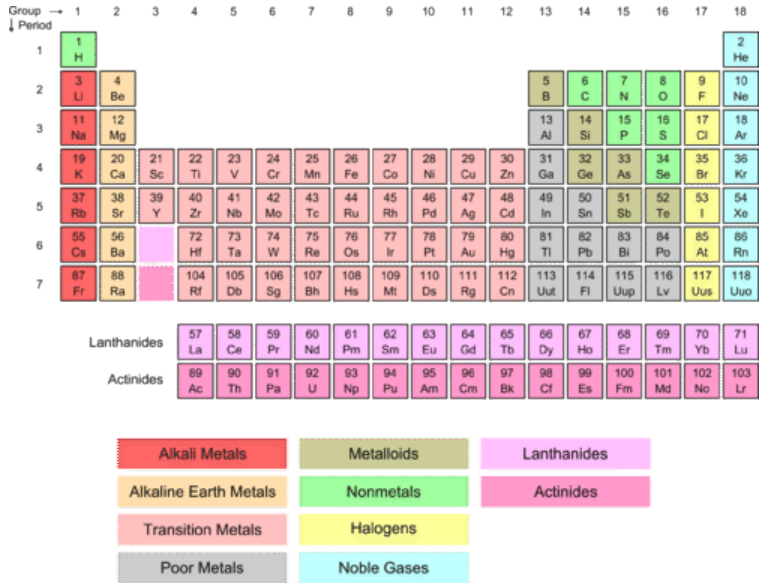


Figure 8.5.1 (Public Domain; User:Cepheus/Wikimedia Commons, modified by CK-12 Foundation via Wikipedia)

How do transition metals form ions?

The transition metals are an interesting and challenging group of elements. They have perplexing patterns of electron distribution that do not always follow the electron filling rules. Predicting how they will form ions is also not always obvious.

Transition Metal Ions

Transition metals belong to the d block, meaning that the d sublevel of electrons is in the process of being filled with up to ten electrons. Many transition metals cannot lose enough electrons to attain a noble-gas electron configuration. In addition, the majority of transition metals are capable of adopting ions with different charges. Iron, which forms either the Fe^{2+} or Fe^{3+} ions, loses electrons as shown below.

$$\begin{array}{lcl} \mathrm{Fe} & \rightarrow & \blacksquare + 2\mathrm{e}^- \\ [\mathrm{Ar}]\ 3d^6\ 4s^2 & & [\mathrm{Ar}]\ 3d^6 \end{array}$$

$$\begin{array}{lcl} \mathrm{Fe} & \rightarrow & \blacksquare + 3\mathrm{e}^- \\ [\mathrm{Ar}]\ 3d^6\ 4s^2 & & [\mathrm{Ar}]\ 3d^5 \end{array}$$

According to the Aufbau process, the electrons fill the $4s$ sublevel before beginning to fill the $3d$ sublevel. However, the outermost s electrons are always the first to be removed in the process of forming transition metal cations. Because most transition metals have two valence electrons, the charge of $2+$ is a very common one for their ions. This is the case for iron above. A half-filled d sublevel (d^5) is particularly stable, which is the result of an iron atom losing a third electron.

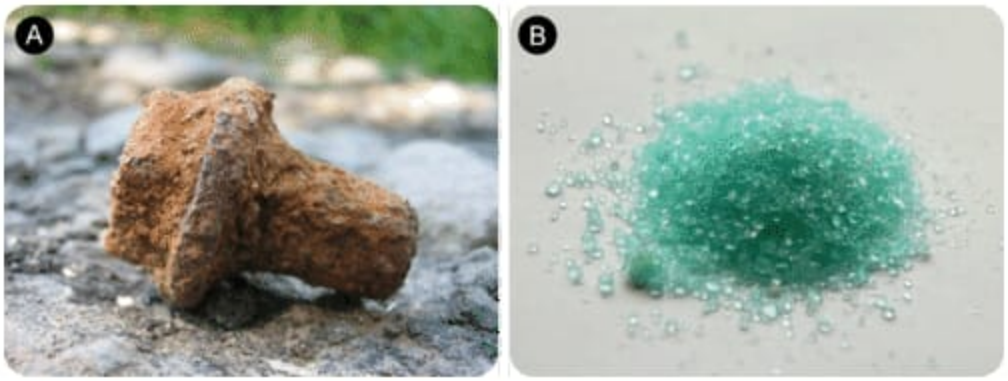


Figure 8.5.2: A. Rust is a complex combination of oxides of iron, among them iron (III) oxide, $\mathrm{Fe_2O_3}$. B. Iron (II) sulfate, $\mathrm{FeSO_4}$, has been known since ancient times as green vitriol and was used for centuries in the manufacture of inks. (CC by 2.0; Public Domain; (A) Paulnasca; (B) Ben Mills (Wikimedia: Benjah-bmm27) via (A) http://commons.wikimedia.org/wiki/File:Rust_screw.jpg; (B) http://commons.wikimedia.org/wiki/File:Iron%2528II%2529-sulfate-heptahydrate-sample.jpg)

Some transition metals that have relatively few d electrons may attain a noble-gas electron configuration. Scandium is an example.

$$\underset{[\mathrm{Ar}]\ 3d^1\ 4s^2}{\mathrm{Sc}} \rightarrow \underset{[\mathrm{Ar}]}{\blacksquare} + 3\mathrm{e}^-$$

Others may attain configurations with a full d sublevel, such as zinc and copper.

$$\underset{[\mathrm{Ar}]\ 3d^{10}\ 4s^2}{\mathrm{Zn}} \rightarrow \underset{[\mathrm{Ar}]\ 3d^{10}}{\blacksquare} + 2\mathrm{e}^-$$

$$\underset{[\mathrm{Ar}]\ 3d^{10}\ 4s^1}{\mathrm{Cu}} \rightarrow \underset{[\mathrm{Ar}]\ 3d^{10}}{\mathrm{Cu}^+} + \mathrm{e}^-$$

The resulting configuration above, with 18 electrons in the outermost principal energy level, is referred to as a **pseudo noble-gas electron configuration**. It gives particular stability to the Zn^{2+} and Cu^+ ions.

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Summary

- Transition metal ion formation is more complex than simple cation formation.
- Transition metal ions often involve rearrangements of both d and s electrons.

Review

1. What block do the transition metals fall in?
2. When writing the electron configuration for a transition metal, which sublevel (s, p, d, f) is filled first?
3. When forming transition metal ions, which sublevel (s, p, d, f) loses electrons first? What is the pseudo noble-gas electron configuration?

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8.6: Ionic Bonding



Figure 8.6.1 (Public Domain; Leon Brooks via Commons Wikimedia, Waneroo Beach [commons.wikimedia.org])

Does the sea really have salt in it?

We can get common table salt from several sources. It can be mined in the solid form in salt mines, or found as a solid in deposits. We can also get salt from the ocean, but it really does not exist as a salt when in solution. The sodium ions and chloride ions are dissolved, but not combined into a structure until all the water is removed.

Most of the rocks and minerals that make up the Earth's crust are composed of positive and negative ions held together by ionic bonding. An ionic compound is an electrically neutral compound consisting of positive and negative ions. You are very familiar with some ionic compounds, such as sodium chloride (NaCl). A sodium chloride crystal consists of equal numbers of positive sodium ions (Na^{+}) and negative chloride ions (Cl^{-}).

Ionic Bonds

Oppositely charged particles attract each other. This attractive force is often referred to as an **electrostatic force**. An **ionic bond** is the electrostatic force that holds ions together in an **ionic compound**. The strength of the ionic bond is directly dependent upon the quantity of the charges and inversely dependent on the distance between the charged particles. A cation with a $2+$ charge will make a stronger ionic bond than a cation with a $1+$ charge. A larger ion makes a weaker ionic bond because of the greater distance between its electrons and the nucleus of the oppositely charged ion.

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Electron Dot Diagrams

We will use sodium chloride as an example to demonstrate the nature of the ionic bond and how it forms. As you know, sodium is a metal and loses its one valence electron to become a cation. Chlorine is a nonmetal and gains one electron in becoming an anion. Both achieve a noble-gas electron configuration. However, electrons cannot be simply "lost" to nowhere in particular. A more accurate way to describe what is happening is that a single electron is transferred from the sodium atom to the chlorine atom, as shown below.

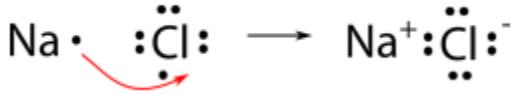


Figure 8.6.2 (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)

The ionic bond is the attraction of the Na^{+} ion for the Cl^{-} ion. It is conventional to show the cation without dots around the symbol to emphasize that the original energy level that contained the valence electron is now empty. The anion is now shown with a complete octet of electrons.

For a compound such as magnesium chloride, it is not quite as simple. Because magnesium has two valence electrons, it needs to lose both to achieve the noble-gas configuration. Therefore, two chlorine atoms will be needed.

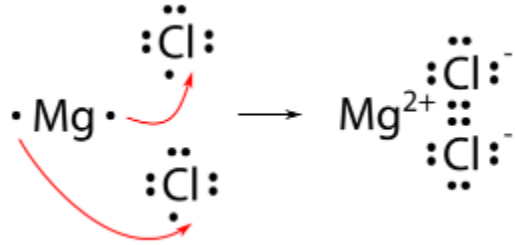


Figure 8.6.3 (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)

The final formula for magnesium chloride is MgCl_2.

Summary

- An ionic compound contains positive and negative ions.
- An ionic bond is electrostatic in nature.
- Electron dot diagrams can be used to illustrate electron movements and ion formation.

Review

1. What is an ionic compound?
2. What is an ionic bond?
3. Which cation (Na^+ or Ca^{2+}) would form a stronger ionic bond with Cl^-?

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8.7: Physical Properties of Ionic Compounds

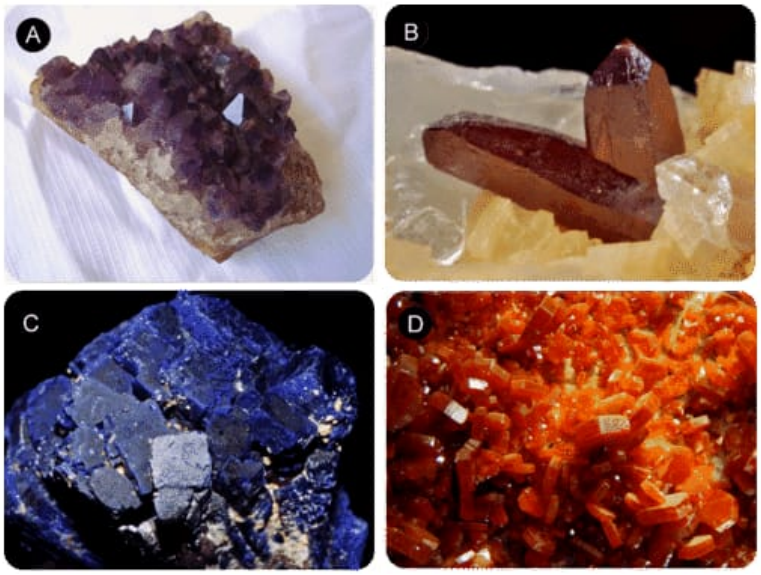


Figure 8.7.1: In nature, the ordered arrangement of ionic solids gives rise to beautiful crystals. (A) Amethyst: a form of quartz, $\mathrm{SiO_2}$, whose purple color comes from iron ions. (B) Cinnabar: the primary ore of mercury, is mercury (II) sulfide, HgS. (C) Azurite: a copper mineral, $\mathrm{Cu_3(CO_3)_2(OH)_2}$. (D) Vanadinite: the primary ore of vanadium, $\mathrm{Pb_3(VO_4)_3Cl}$. ((A; (A) Mauro Cateb; (B) Parent Géry; (C) Parent Géry; (D) User:vassil/Wikimedia Commons via (A) http://www.flickr.com/photos/mauroescritor/6544460363; (B) http://commons.wikimedia.org/wiki/File:Cinabre_macl%25C3%25A9_%2528Chine%2529_.jpg; (C) http://commons.wikimedia.org/wiki/File:Azurite_cristallis%25C3%25A9e_%2528Chine%2529_2_.jpg; (D) http://commons.wikimedia.org/wiki/File:Vanadinite_21207.jpg) CC BY 2.0; (B) Public Domain; (C) Public Domain; (D) Public Domain)

What produces colored crystals?

The figure above shows just a few examples of the color and brilliance of naturally occurring ionic crystals. The regular and orderly arrangement of ions in the crystal lattice is responsible for the various shapes of these crystals, while transition metal ions give rise to the colors.

Physical Properties of Ionic Compounds

Melting Points

Because of the many simultaneous attractions between cations and anions that occur, ionic crystal lattices are very strong. The process of melting an ionic compound requires the addition of large amounts of energy in order to break all of the ionic bonds in the crystal. For example, sodium chloride has a melting temperature of about $800^{\circ}\mathrm{C}$.

Shattering

Ionic compounds are generally hard, but **brittle.** Why? It takes a large amount of mechanical force, such as striking a crystal with a hammer, to force one layer of ions to shift relative to its neighbor. However, when that happens, it brings ions of the same charge next to one another (see below). The repulsive forces between like-charged ions cause the crystal to shatter. When an ionic crystal breaks, it tends to do so along smooth planes because of the regular arrangement of the ions.

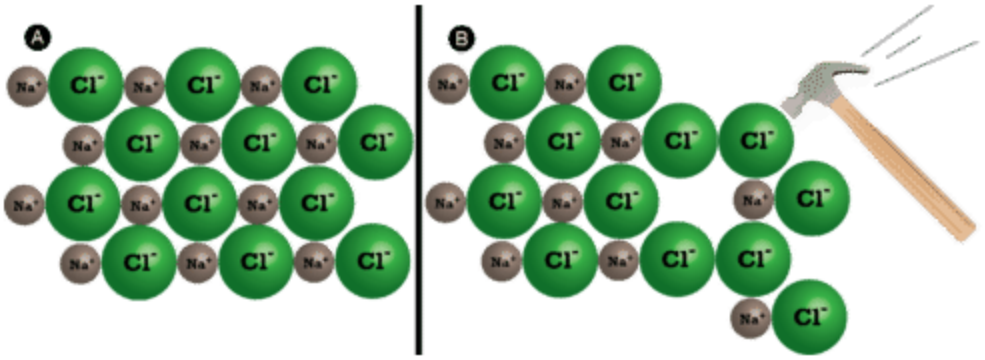


Figure 8.7.2: (A) The sodium chloride crystal is shown in two dimensions. (B) When struck by a hammer, the negatively-charged chloride ions are forced near one another and the repulsive force causes the crystal to shatter. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Conductivity

Another characteristic property of ionic compounds is their **electrical conductivity**. The figure below shows three experiments in which two electrodes that are connected to a light bulb are placed in beakers containing three different substances.

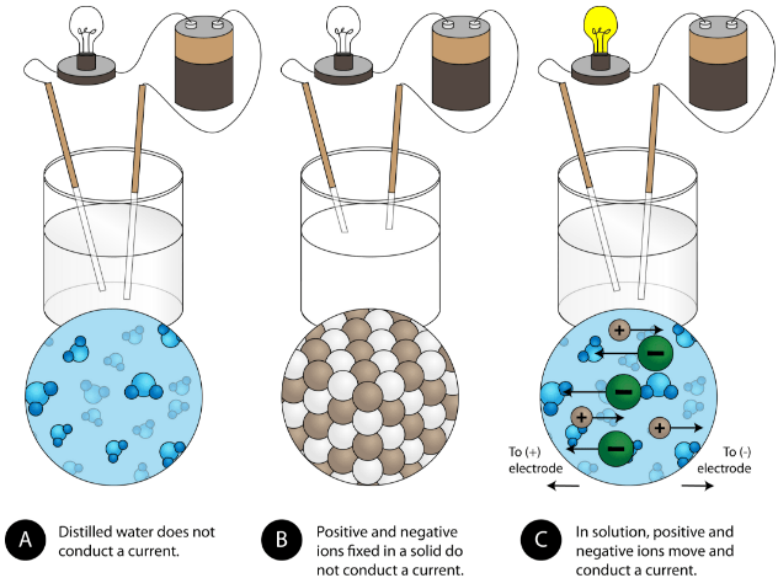


Figure 8.7.3: (A) Distilled water does not conduct electricity. (B) A solid ionic compound does not conduct electricity either. (C) A water solution of an ionic compound conducts electricity well. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

In the first beaker, distilled water does not conduct a current because water is a molecular compound. In the second beaker, solid sodium chloride also does not conduct a current. Despite being ionic and thus composed of charged particles, the solid crystal lattice does not allow the ions to move between the electrodes. Mobile charged particles are required for the circuit to be complete and the light bulb to light up. In the third beaker, the NaCl has been dissolved into the distilled water. Now the crystal lattice has been broken apart and the individual positive and negative ions can move. Cations move to one electrode, while anions move to the other, allowing electricity to flow (see figure below). Melting an ionic compound also frees the ions to conduct a current. Ionic compounds conduct an electric current when melted or dissolved in water.

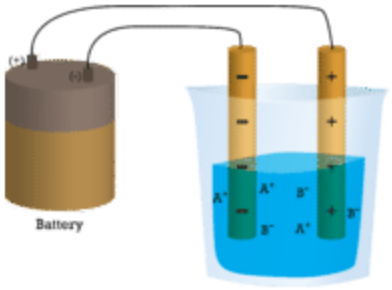


Figure 8.7.4: In an ionic solution, the A^+ ions migrate toward the negative electrode, while the B^- ions migrate toward the positive electrode. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

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Summary

- Ionic compounds have high melting points.
- Ionic compounds are hard and brittle.
- Solutions of ionic compounds and melted ionic compounds conduct electricity, but solid materials do not.

Review

1. Why are ionic compounds brittle?
2. Why are melting points high for ionic compounds?
3. What happens when an electric current is passed through a solution of an ionic compound?

Explore More

Watch the video below and answer the following questions:

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1. Do all ionic compounds form crystals?
2. Will melted ionic compounds conduct electricity?
3. What are the melting and boiling points of KI?

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8.8: Metallic Bonding



Figure 8.8.1: A nineteenth century copper plate. (Public Domain; John Pearson via Wikipedia)

Why do metals behave the way they do?

The image below is of a copper plate that was made in 1893. The utensil has a great deal of elaborate decoration, and the item is very useful. What would have happened if this plate was made of copper (I) chloride instead? Copper (I) chloride does contain copper, after all. However, the CuCl would end up as a powder when a metalworker pounded on it to shape it. Metals behave in unique ways. The bonding that occurs in a metal is responsible for its distinctive properties: luster, malleability, ductility, and excellent conductivity.

The Metallic Bond

Pure metals are crystalline solids, but unlike ionic compounds, every point in the crystal lattice is occupied by an identical atom. The electrons in the outer energy levels of a metal are mobile and capable of drifting from one metal atom to another. This means that the metal is more properly viewed as an array of positive ions surrounded by a sea of mobile valence electrons. Electrons which are capable of moving freely throughout the empty orbitals of the metallic crystal are called **delocalized electrons** (see below). A **metallic bond** is the attraction of the stationary metal cations to the surrounding mobile electrons.

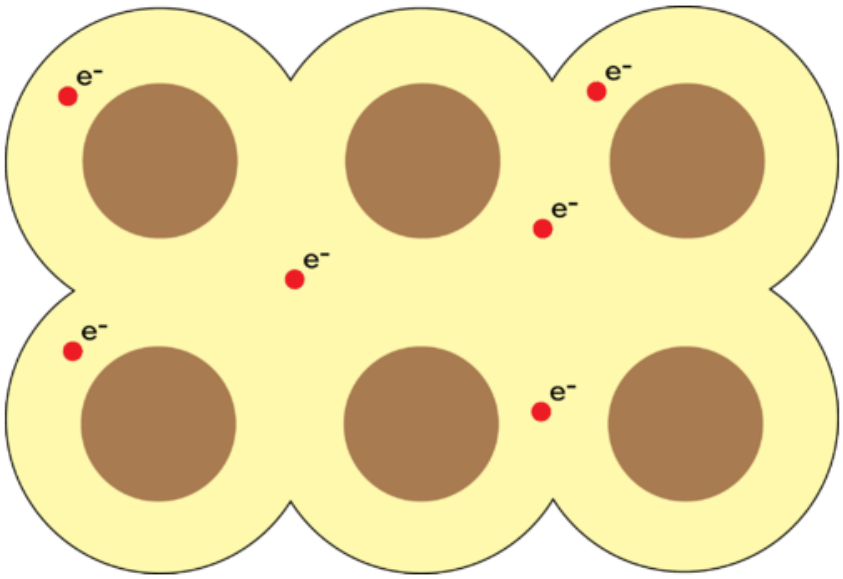


Figure 8.8.2: In a metal, the stationary metal cations are surrounded by a sea of mobile valence electrons that are not associated with any one cation. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

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Properties of Metals

The metallic bonding model explains the physical properties of metals. Metals conduct electricity and heat very well because of their free-flowing electrons. As electrons enter one end of a piece of metal, an equal number of electrons flow outward from the other end. When light is shone onto the surface of a metal, its electrons absorb small amounts of energy and become excited into one of its many empty orbitals. The electrons immediately fall back down to lower energy levels and emit light. This process is responsible for the high **luster** of metals.



Figure 8.8.3: The American Platinum Eagle is the official platinum bullion coin of the United States and was first minted in 1997. The luster of a metal is due to its metallic bonds. (Public Domain; the US Mint via Front: http://commons.wikimedia.org/wiki/File:2005_AEPlat_Proof_Obv.png; Back: http://commons.wikimedia.org/wiki/File:American_Platinum_Eagle_2008_Proof_Rev.jpg)

Recall that ionic compounds are very brittle. Application of a force results in like-charged ions in the crystal coming too close to one another, causing the crystal to shatter. When a force is applied to a metal, the free-flowing electrons can slip in between the stationary cations and prevent them from coming in contact. Imagine ball bearings that have been coated with oil sliding past one another. As a result, metals are very **malleable** and **ductile**. They can be hammered into shapes, rolled into thin sheets, or pulled into thin wires.

Summary

- The metallic bond is responsible for the properties of metals.
- Metals conduct electricity and heat well.
- Metals are ductile and malleable.
- Metals have luster.

Review

1. What is a delocalized electron?

2. Why do metals conduct electricity and heat well?
3. Why do metals have luster?

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9: COVALENT BONDING

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

9.1: Energy and Covalent Bond Formation

9.2: Lewis Electron-Dot Structures

9.3: Single Covalent Bonds

9.4: Multiple Covalent Bonds

9.5: Coordinate Covalent Bond

9.6: Covalent Bonding in Polyatomic Ions

9.7: Resonance

9.8: Exceptions to the Octet Rule

9.9: Bond Energy

9.10: VSEPR Theory

9.11: Molecular Shapes- No Lone Pairs on Central Atoms

9.12: Molecular Shapes - Lone Pair(s) on Central Atom

9.13: Periodic Trends- Electronegativity

9.14: Bond Polarity

9.15: Polar Molecules

9.16: Van der Waals Forces

9.17: Hydrogen Bonding

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9.1: Energy and Covalent Bond Formation



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Why do $BeCl_2$ and LiCl bond differently?

We have learned that halide salts of elements in group 1 are typically ionic compounds. We would expect LiCl to exist as Li^+ cations and Cl^- anions (and it does). However, if we move one column to the right, lithium's neighbor beryllium forms a different type of bond altogether. This bond consists of shared electrons between the Be and Cl atoms, not electrostatic attraction among ions.

Energy and Covalent Bond Formation

Molecular compounds are those that take the form of an individual molecule. Molecular compounds are generally comprised of two or more nonmetal atoms. Familiar examples include water $(\mathrm{H_2O})$, carbon dioxide $(\mathrm{CO_2})$, and ammonia $(\mathrm{NH_3})$. Recall that the molecular formula shows the number of each atom that occurs in a molecule of that compound. One molecule of water contains two hydrogen atoms and one oxygen atom. Hydrogen $(\mathrm{H_2})$ is an example of an element that exists naturally as a diatomic molecule. A **diatomic molecule** is a molecule containing two atoms.

Most atoms attain a lower potential energy when they are bonded to other atoms than when they are separated. Consider two isolated hydrogen atoms that are separated by a distance large enough to prevent any interaction between them. At this distance, the potential energy of the system is said to be equal to zero (see figure below).

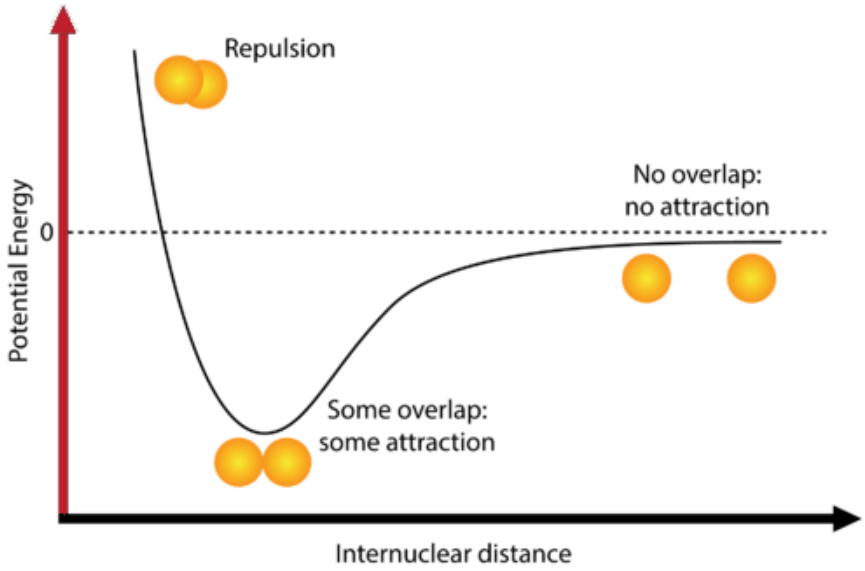


Figure 9.1.2: The graph shows how the potential energy of two hydrogen atoms changes as a function of their separation distance. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

As the atoms approach one another, their electron clouds gradually begin to overlap. Now there are several interactions which begin to occur. One is that the single electrons that each hydrogen atom possesses begin to repel each other. This repulsive force would tend to make the potential energy of the system increase. However, the electron of each atom begins to be attracted to the nucleus of the other atom. This attractive force tends to make the potential energy of the system decrease.

As the atoms first begin to interact, the attractive force is stronger than the repulsive force and so the potential energy of the system decreases, as seen in the diagram. Remember that the lower potential energy increases the stability of the system. As the two hydrogen atoms move closer and closer together, the potential energy is at its lowest possible point. If the hydrogen

atoms move any closer together, a third interaction begins to dominate, and that is the repulsive force between the two positively-charged nuclei. This repulsive force is very strong, as can be seen by the sharp rise in energy at the far left of the diagram.

The point at which the potential energy reached its minimum represents the ideal distance between hydrogen atoms for a stable chemical bond to occur. This type of chemical bond is called a covalent bond. A **covalent bond** is a bond in which two atoms share one or more pairs of electrons. The single electrons from each of the two hydrogen atoms are shared when the atoms come together to form a hydrogen molecule (H_2).

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Summary

- Covalent bonds are formed when atoms share electrons between them.

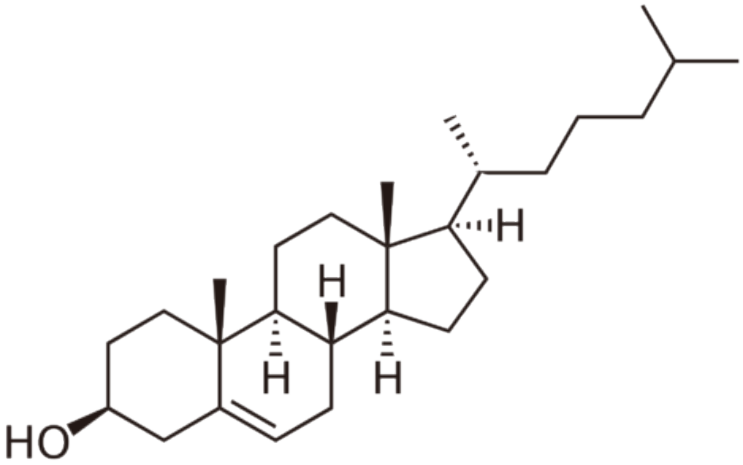
Review

1. What is a diatomic molecule?
2. What does lower potential energy do to a system?
3. What is the ideal distance for two atoms?

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9.2: Lewis Electron-Dot Structures



Figure

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What does cholesterol really look like?

We can write the structure of the cholesterol molecule a couple of different ways. The simplest approach is to simply write $C_{27}H_{46}O$. This "structure" is not very useful because it does not tell us how the carbons, hydrogens, and oxygen are connected to one another. The structure in the figure above is much more helpful—we see how the different atoms are connected together to form the molecule.

Lewis Electron-Dot Structures

In a previous chapter, you learned that the valence electrons of an atom can be shown in a simple way with an electron dot diagram. A hydrogen atom is shown as $H\cdot$ because of its one valence electron. The structures of molecules that are held together by covalent bonds can be diagrammed by **Lewis electron-dot structures**. The hydrogen molecule is shown in the figure below.

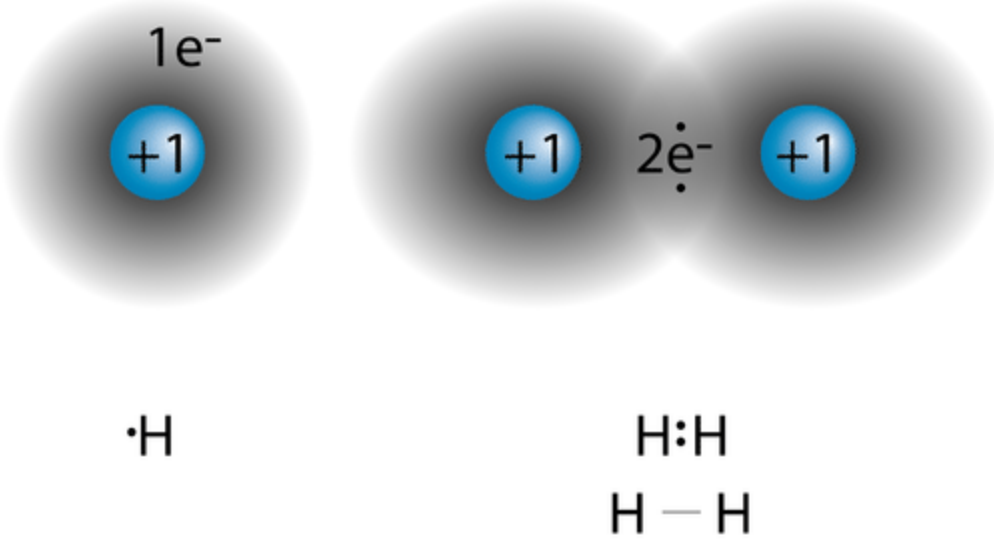


Figure $9.2.2$: On the left is a single hydrogen atom with one electron. On the right is an H_2 molecule showing the electron cloud overlap. (CC BY-NC 3.0; Jodi So via CK-12 Foundation)

The shared pair of electrons is shown as two dots in between the two H symbols $(H : H)$. This is called a **single covalent bond**, when two atoms are joined by the sharing of one pair of electrons. The single covalent bond can also be shown by a dash in between the two symbols $(H—H)$. A **structural formula** is a formula that shows the arrangement of atoms in a molecule and represents covalent bonds between atoms by dashes.

The Octet Rule and Covalent Bonds

When ions form, they conform to the **octet rule** by either losing or gaining electrons in order to achieve the electron configuration of the nearest noble gas. In a similar way, nonmetal atoms share electrons in the formation of a covalent bond in such a way that each of the atoms involved in the bond can attain a noble-gas electron configuration. The shared electrons are "counted" for each of the atoms involved in the sharing. For hydrogen (H_2), the shared pair of electrons means that each of the atoms is able to attain the electron configuration of helium, the noble gas with two electrons. For atoms other than hydrogen, the sharing of electrons will usually provide each of the atoms with eight valence electrons.

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Summary

- Lewis electron-dot structures show the bonding in covalent molecules.
- Covalent bonds between atoms can be indicated either with dots $(:)$ or a dash $(-)$.

Review

1. What is a single covalent bond?
2. How can covalently-bound atoms obey the octet rule?
3. Does the hydrogen molecule obey the octet rule?

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9.3: Single Covalent Bonds

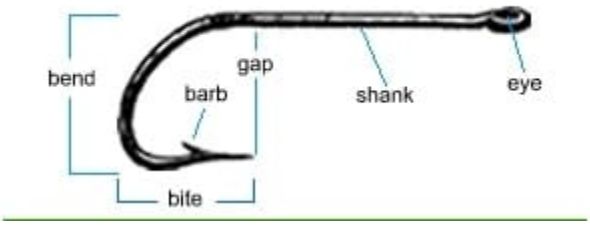


Figure 9.3.1 (Public Domain; the US Department of the Interior via Commons Wikimedia, Fishhook [commons.wikimedia.org])

What holds molecules together?

In one form or another, the idea of atoms connecting to form larger substances has been with us for a long time. The Greek philosopher Democritus (460-370 BCE) believed that atoms had hooks that allowed them to connect with one another. Today, we believe that atoms are held together by bonds formed when two atoms share a set of electrons—a much more complicated picture than the simple hooks that Democritus believed in.

Single Covalent Bonds

A covalent bond forms when two orbitals with one electron each overlap one another. For the hydrogen molecule, this can be shown as:

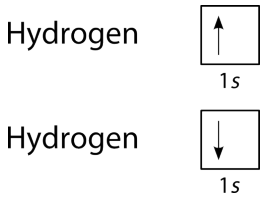


Figure 9.3.2 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Upon formation of the H_2 molecule, the shared electrons must have opposite spin, so they are shown with opposite spin in the atomic $1s$ orbital.

The halogens also form single covalent bonds in their diatomic molecules. An atom of any halogen, such as fluorine, has seven valence electrons. Its unpaired electron is located in the $2p$ orbital.

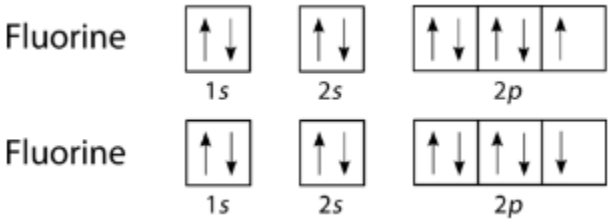


Figure 9.3.3 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The single electrons in the third $2p$ orbital combine to form the covalent bond:

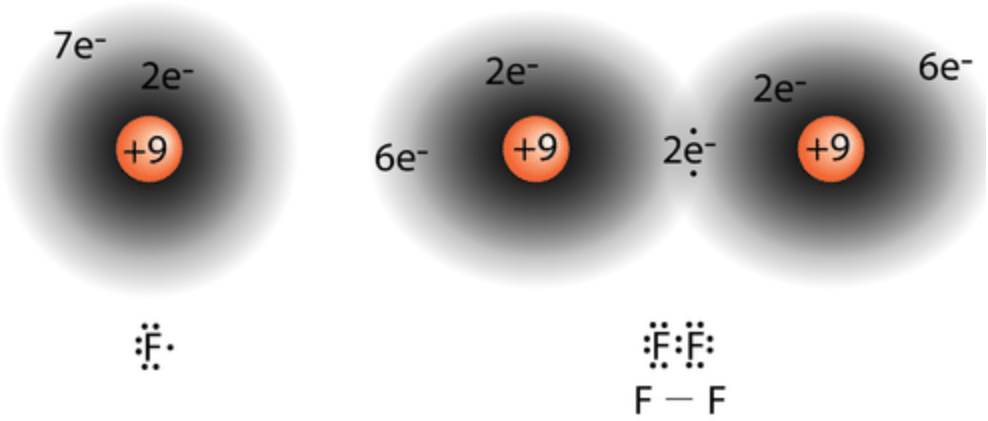


Figure 9.3.4: On the left is a fluorine atom with seven valence electrons. On the right is the F_2 molecule. (CC BY-NC 3.0; Jodi So via CK-12 Foundation)

The diatomic fluorine molecule (F_2) contains a single shared pair of electrons. Each F atom also has three pair of electrons that are not shared with the other atom. A **lone pair** is a pair of electrons in a Lewis electron-dot structure that is not shared between atoms. The oxygen atom in the water molecule shown below has two lone pair sets of electrons. Each F atom has three lone pairs. Combined with the two electrons in the covalent bond, each F atom follows the octet rule.

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Example 9.3.1

Draw the Lewis electron dot structure for water.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Molecular formula of water $= H_2O$
- 1 O atom $= 6$ valence electrons
- 2 H atoms $= 2 \times 1 = 2$ valence electrons
- Total number of valence electrons $= 8$

Use the periodic table to determine the number of valence electrons for each atom and the total number of valence electrons. Arrange the atoms and distribute the electrons so that each atom follows the octet rule. The oxygen atom will have 8 electrons, while the hydrogen atoms will each have 2.

Step 2: Solve.

Electron dot diagrams for each atom are:



Figure 9.3.5 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Each hydrogen atom with its single electron will form a covalent bond with the oxygen atom where it has a single electron. The resulting Lewis electron dot structure is:



Figure 9.3.6 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Step 3: Think about your result.

The oxygen atom follows the octet rule with two pairs of bonding electrons and two lone pairs. Each hydrogen atom follows the octet rule with one bonding pair of electrons.

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Summary

- Covalent bonds form when electrons in two atoms form overlapping orbitals.
- Lone pair electrons in an atom are not shared with any other atom.

Review

1. How does a covalent bond form?
2. What relationship do the spins of shared electrons have with each other?
3. Do lone pair electrons form covalent bonds?

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9.4: Multiple Covalent Bonds

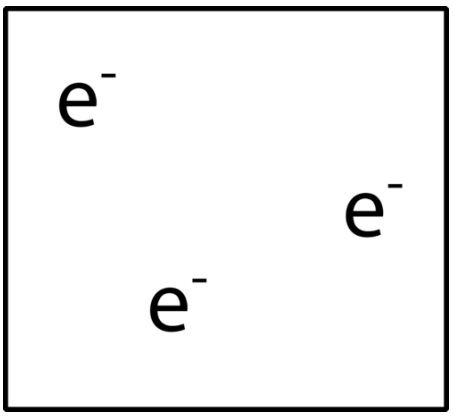


Figure 9.4.1 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

What do you do with your leftovers?

When working with covalent structures, it sometimes looks like there are leftover electrons. Maybe you have applied the rules you learned so far, but 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 $\mathrm{C_2H_4}$. The carbon atoms are bonded together, with each carbon also 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:

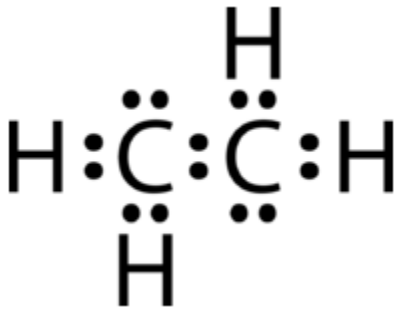


Figure 9.4.2: Incorrect dot structure of ethene. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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 the carbon atoms share two pairs, instead of only one pair, of electrons.

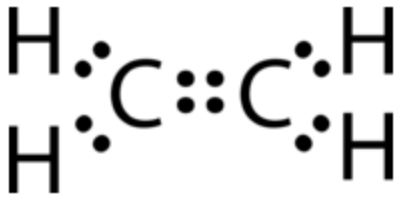


Figure 9.4.3: Correct dot structure for ethene. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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.

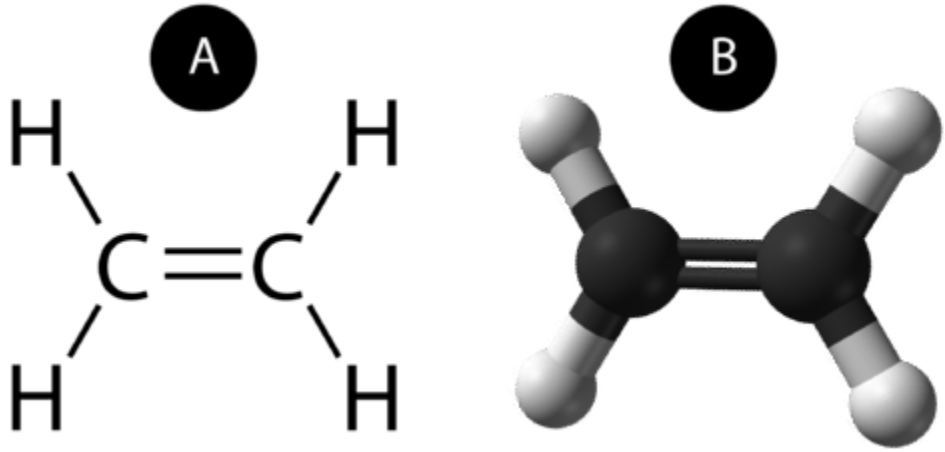


Figure 9.4.4: (A) The structural model for C_2H_4 consists of a double covalent bond between the two carbon atoms and single bonds to the hydrogen atoms. (B) Molecular model of C_2H_4. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via (A) http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-dimensions-2D-Vector.svg; (B) http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png)

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.

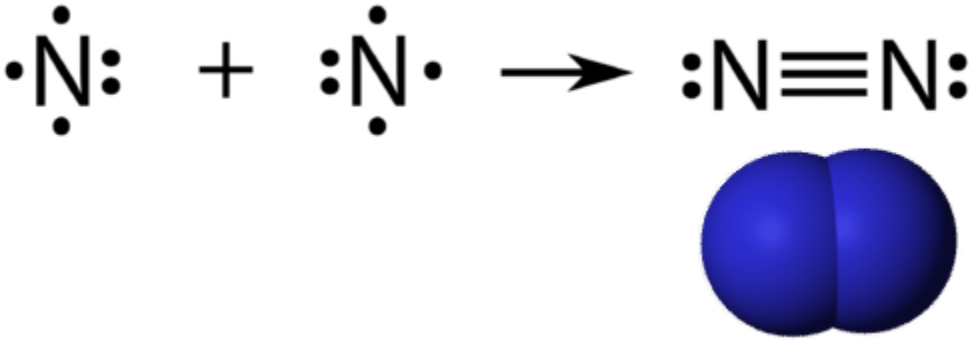


Figure 9.4.5: Triple bond in N_2. (CC BY-NC-SA 3.0; Molecule Public Domain; CK-12 Foundation - Joy Sheng, using 3D molecular structure by Ben Mills (Wikimedia: Benjah-bmm27) via Molecular structure: http://commons.wikimedia.org/wiki/File:Nitrogen-3D-vdW.png)

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

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Summary

- Lewis structures can be drawn for molecules that share multiple pairs of electrons.
- A double covalent bond is a covalent bond formed by atoms that share two pairs of electrons.
- A triple covalent bond is a covalent bond formed by atoms that share three pairs of electrons.

Review

1. Why is the first ethene Lewis structure incorrect?
2. What do the single electrons in nitrogen do to form a triple bond?
3. Draw the Lewis structure for ethyne C_2H_2.

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9.5: Coordinate Covalent Bond



Figure 9.5.1 (Public Domain; P. Periyannan (Wikimedia: TRYPPN) via Wikipedia)

Is sharing a good thing?

Remember when you were younger, and were told to share your favorite toy with your brother, sister, or friend? You probably didn't want to share, but did anyway. It likely turned out that you had more fun playing with the toy together than if you had kept it to yourself. Atoms often share electrons with other atoms that have nothing to contribute to the situation; the end result is a new structure.

Coordinate Covalent Bonds

Each of the covalent bonds that we have looked at so far has involved each of the atoms that are bonding contributing one of the electrons to the shared pair. There is an alternate type of covalent bond in which one of the atoms provides both of the electrons in a shared pair. Carbon monoxide, CO, is a toxic gas that is released as a byproduct during the burning of fossil fuels. The bonding between the C atom and the O atom can be thought of in the following procession:

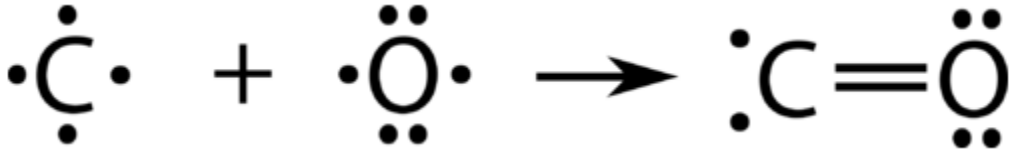


Figure 9.5.2: Formation of a CO double bond (incorrect structure). (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

At this point, a double bond has formed between the two atoms, with each atom providing one of the electrons to each bond. The oxygen atom now has a stable octet of electrons, but the carbon atom only has six electrons and is unstable. This situation is resolved if the oxygen atom contributes one of its lone pairs in order to make a third bond with the carbon atom.



Figure 9.5.3: Correct CO structure. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The carbon monoxide molecule is correctly represented by a triple covalent bond between the carbon and oxygen atoms. One of the bonds is a **coordinate covalent bond**, a covalent bond in which one of the atoms contributes both of the electrons in the shared pair.

Once formed, a coordinate covalent bond is the same as any other covalent bond. It is not as if the two conventional bonds in the CO molecule are stronger or different in any other way than the coordinate covalent bond.

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Summary

- Coordinate covalent bonds can form when one atom provides a lone pair of electrons to the bond.
- Coordinate covalent bonds are as strong as other covalent bonds.

Review

1. Where does the third covalent bond in the CO molecule come from?
2. Why is the incorrect structure for CO above wrong?
3. Are coordinate covalent bonds stronger or weaker than regular covalent bonds?

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9.6: Covalent Bonding in Polyatomic Ions



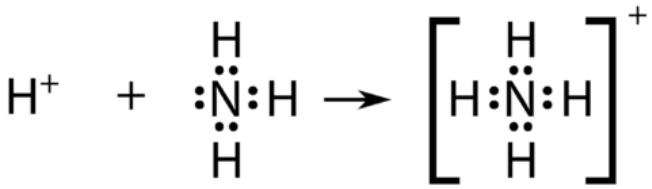
Figure 9.6.1 (CC by 2.0; Jan Seifert via Flickr)

How do we extend basic principles?

The United States Supreme Court has the unenviable task of determining what the law dictates. This responsibility can be a major challenge when there is no clear principle involved, or when they face a situation not previously encountered. Chemistry faces the same challenge in extending basic concepts to fit a new situation. Drawing of Lewis structures for polyatomic ions uses the same approach, but tweaks the process a little to fit a somewhat different set of circumstances.

Polyatomic Ions

Recall that a **polyatomic ion** is a group of atoms that are covalently bonded together, and which carry an overall electrical charge. The ammonium ion, $\mathrm{NH_4}^+$, is formed when a hydrogen ion (H^+) attaches to the lone pair of an ammonia $(\mathrm{NH_3})$ molecule in a coordinate covalent bond.



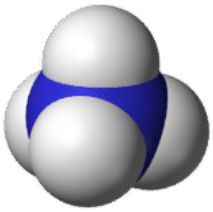


Figure 9.6.2: The ammonium ion. (CC BY-NC-SA 3.0; CK-12 Foundation - Joy Sheng, using 3D molecular structure by Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia; Molecule Public Domain)

When drawing the Lewis structure of a polyatomic ion, the charge of the ion is reflected in the number of total valence electrons in the structure. In the case of the ammonium ion:

$$1\ \mathrm{N\ atom}\ = 5\ \text{valence electrons}$$

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

subtract 1 electron for the $1+$charge of the ion

total of 8 valence electrons in the ion

It is customary to put the Lewis structure of a polyatomic ion into a large set of brackets, with the charge of the ion as a superscript outside of the brackets.

Example $9.6.1$

Draw the Lewis electron dot structure for the sulfate ion.

Solution

Step 1: List the known quantities and plan the problem.

Known

Molecular formula of sulfate ion: $\mathrm{SO_4}^{2-}$

$1\ \mathrm{S}$ atom $= 6$ valence electrons

$4\ \mathrm{O}$ atoms $= 4 \times 6 = 24$ valence electrons

add 2 electrons for the $2-$ charge of the ion

= total of 32 valence electrons

The less electronegative sulfur atom is the central atom in the structure. Place the oxygen atoms around the sulfur atom, each with a single covalent bond. Distribute lone pairs to each oxygen atom in order to satisfy the octet rule. Count the total number of atoms. If there are too many electrons in the structure, make multiple bonds between the S and O.

Step 2: Solve.

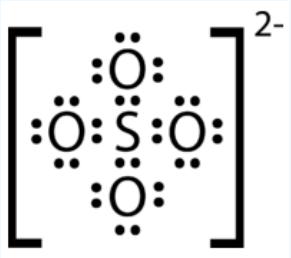


Figure $9.6.3$: The sulfate ion. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Step 3: Think about your result.

The Lewis structure for the sulfate ion consists of a central sulfur atom with four single bonds to oxygen atoms. This yields the expected total of 32 electrons. Since the sulfur atom started with six valence electrons, two of the $\mathrm{S-O}$ bonds are coordinate covalent.

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Summary

- Lewis structures for polyatomic ions follow the same rules as those for other covalent compounds.

Review

1. What are two characteristics of polyatomic ions?
2. Which atom becomes the central atom in the structure?
3. Where is the charge on an ion placed in a lewis dot diagram?

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9.7: Resonance



Figure 9.7.1 (Public Domain; User:Gouwenaar/Nl.Wikipedia via Wikipedia)

Is this a reflection or a doorway?

You may look at the picture below and think you are looking at the image of a room as reflected in a mirror (and you probably are). But the picture can be cropped in such a way as to give the impression of looking at the real room through a door. You would see the same thing and receive the same information, but it would be from a different perspective. There are molecules that can be represented in different ways, such that reality becomes a matter of interpretation.

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) molecule in the figure below. There are a total of 18 electrons in the structure, and so the following two structures are possible:

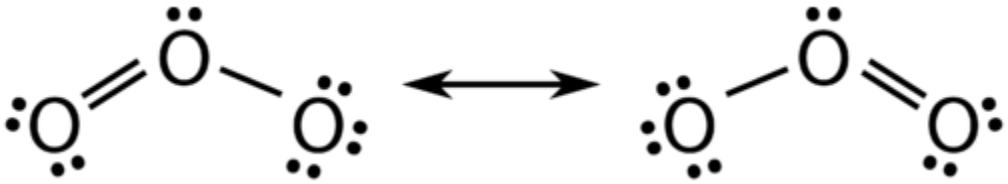


Figure 9.7.2: Resonance forms of ozone. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The structure on the left (see figure above) 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 Lewis structures to represent the covalent bonding in a molecule. One of the valid structures is referred to as a resonance structure. It is now understood that the true structure of a molecule which displays resonance is that of an average or a hybrid of all the resonance structures. In the case of the O_3 molecule, each of the covalent bonds between O atoms are 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 (see figure below).



Figure 9.7.3: "Half-bond" model of ozone molecule. (CC BY-NC 3.0; Molecule Public Domain; CK-12 Foundation - Joy Sheng, using 3D molecular structure by Ben Mills (Wikimedia: Benjah-bmm27) via Molecular structure: http://commons.wikimedia.org/wiki/File:Ozone-CRC-MW-3D-balls.png)

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Many polyatomic ions also display resonance. In some cases, the true structure may be an average of three valid resonance structures, as in the case of the nitrate ion, NO_3^- (see figure below).

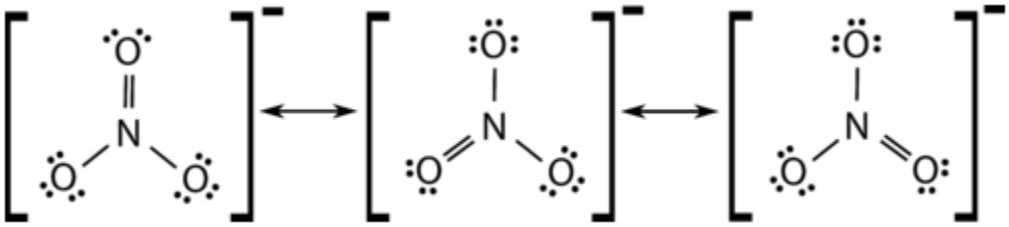


Figure 9.7.4: Resonance structure of nitrate anion. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The bond lengths between the central N atom and each O atom are identical and the bonds can be approximated as being equal to one and one-third bonds.

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Summary

- Resonance structures are averages of different Lewis structure possibilities.
- Bond lengths are intermediate between covalent bonds and covalent double bonds.

Review

1. How many electrons total are in the ozone structure?

2. What is changed in the two resonance structures of ozone?
3. How can we think of the covalent bonds in ozone?

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9.8: Exceptions to the Octet Rule



Figure 9.8.1 (Public Domain; the US Internal Revenue Service via Wikipedia)

Are rules always followed?

Every spring, millions of Americans file their income tax forms. The different rules determine how much tax a person pays. There are also exceptions to the rules. You pay less tax if you are married and/or have children. There are certain limits on how much money you can make before paying taxes. The rule is that you pay taxes, but there are also exceptions based on your personal situation. The bonding rules for molecules are generally applicable, but there are some exceptions allowed.

Exceptions to the Octet Rule

As the saying goes, all rules are made to be broken. When it comes to the octet rule, this is true. Exceptions to the octet rule fall into one of three categories: (1) an **incomplete octet**, (2) **odd-electron molecules**, and (3) an **expanded octet**.

Incomplete Octet

In some compounds, the number of electrons surrounding the central atom in a stable molecule is fewer than eight. Beryllium is an alkaline earth metal and so may be expected to form ionic bonds. However, its very small size and somewhat higher ionization energy compared to other metals actually leads beryllium to form primarily molecular compounds. Since beryllium only has two valence electrons, it does not typically attain an octet through sharing of electrons. The Lewis structure of gaseous beryllium hydride $(\mathrm{BeH_2})$ consists of two single covalent bonds between Be and H (see figure below).

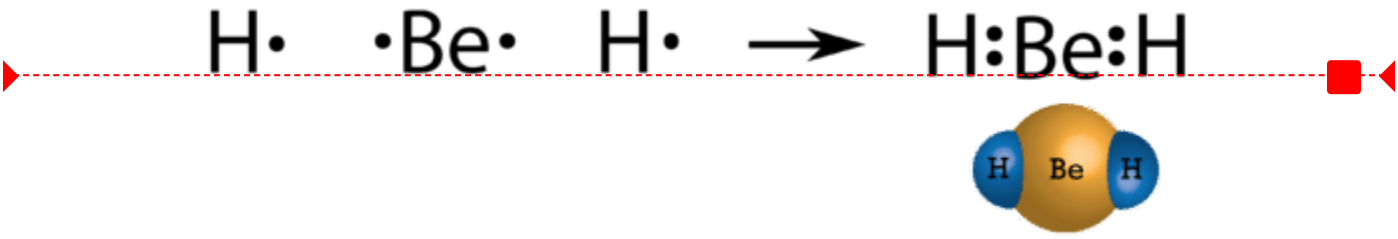


Figure 9.8.2: Beryllium hydride. (CC BY-NC 3.0; Christopher Auyeung, Joy Sheng via CK-12 Foundation)

Boron and aluminum, with three valence electrons, also tend to form covalent compounds with an incomplete octet. The central boron atom in boron trichloride (BCl_3) has six valence electrons, as shown in the figure below.

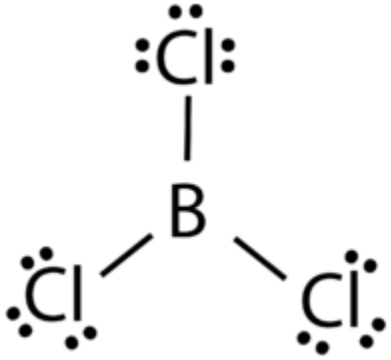


Figure $9.8.3$: Boron trichloride. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Odd-Electron Molecules

There are a number of molecules whose total number of valence electrons is an odd number. It is not possible for all of the atoms in such a molecule to satisfy the octet rule. An example is nitrogen dioxide (NO_2). Each oxygen atom contributes six valence electrons and the nitrogen atom contributes five for a total of seventeen. The Lewis structure for NO_2 appears in the figure below.

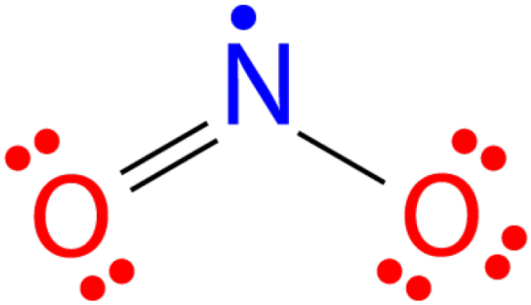


Figure $9.8.4$: The Lewis structure for nitrogen dioxide, an odd electron molecule. (CC BY-NC-SA; Todd Helmenstine via about.com)

Expanded Octets

Atoms of the second period cannot have more than eight valence electrons around the central atom. However, atoms of the third period and beyond are capable of exceeding the octet rule by having more than eight electrons around the central atom. Starting with the third period, the d sublevel becomes available, so it is possible to use these orbitals in bonding, resulting in an expanded octet.

Phosphorus and sulfur are two elements that react with halogen elements and make stable compounds with expanded octets. In phosphorus pentachloride, the central phosphorus atom makes five single bonds to chlorine atoms and, as a result, has ten electrons surrounding it (see figure below). In sulfur hexafluoride, the central sulfur atom has twelve electrons from its six bonds to fluorine atoms (see figure below).

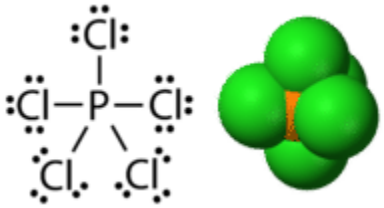


Figure $9.8.5$: Phosphorus pentachloride. Left image: Lewis structure. Right image: molecular model. (CC BY-NC 3.0; Molecule Public Domain; CK-12 Foundation - Joy Sheng, using 3D molecular structure by Ben Mills (Wikimedia: Benjah-bmm27) via Molecular structure: http://commons.wikimedia.org/wiki/File:Phosphorus-pentachloride-3D-vdW.png)

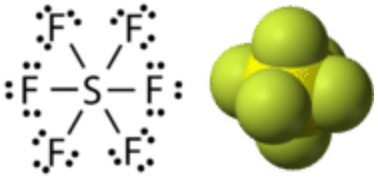


Figure 9.8.6: Sulfur hexafluoride. Left image: Lewis structure. Right image: molecular model. (CC BY-NC 3.0; Molecule Public Domain; CK-12 Foundation - Joy Sheng, using 3D molecular structure by Ben Mills (Wikimedia: Benjah-bmm27) via Molecular structure: http://commons.wikimedia.org/wiki/File:Sulfur-hexafluoride-3D-vdW.png)

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Summary

- Exceptions exist to the rules for covalent bonding
- These exceptions apply to atoms whose electrons will not accommodate the normal octet rule.

Review

1. What is an incomplete octet?
2. What is an odd-electron molecule?
3. Why are there extra electrons in the expanded octet?

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9.9: Bond Energy

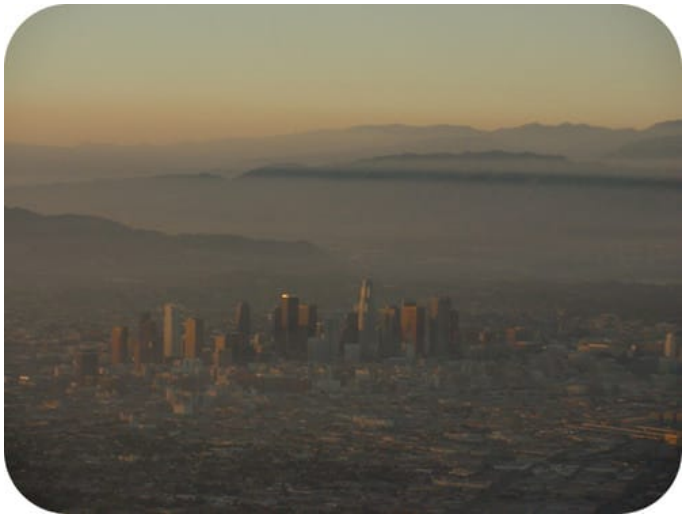


Figure 9.9.1 (CC by 2.0; Warner Hocker (Flickr: TravelingOtter) via Flickr)

How does smog form?

We know that nitrogen makes up about 79% of the gases in the air, and that this gas is chemically very inert. However, we also know that a major contributor to the production of smog is a mixture of nitrogen compounds referred to as NO_x. Due to the high combustion temperatures in modern car engines (brought about by the need for better fuel efficiency), we make more NO_x that will react with other materials to create smog. So, our inert nitrogen gas can be converted to other compounds if enough energy is available to break the triple bonds in the N_2 molecule.

Bond Energy

The formation of a chemical bond results in a decrease in potential energy. Consequently, breaking a chemical bond requires an input of energy. **Bond energy** is the energy required to break a covalent bond between two atoms. A high bond energy means that a bond is strong and the molecule that contains that bond is likely to be stable and less reactive. More reactive compounds will contain bonds that have generally lower bond energies. Some bond energies are listed in the table below.

Table 9.9.1: Bond Energies

Bond	Bond Energy $(\mathrm{kJ/mol})$
$\mathrm{H{-}H}$	436
$\mathrm{C{-}H}$	414
$\mathrm{C{-}C}$	347
$\mathrm{C{=}C}$	620
$\mathrm{C \equiv C}$	812
$\mathrm{F{-}F}$	157
$\mathrm{Cl{-}Cl}$	243
$\mathrm{Br{-}Br}$	193
$\mathrm{I{-}I}$	151
$\mathrm{N \equiv N}$	941

The halogen elements all exist naturally as diatomic molecules (F_2, Cl_2, Br_2, and I_2). They are generally very reactive and thus have relatively low bond energies.

As can be seen by a comparison of the bond energies for the various carbon-carbon bonds, double bonds are substantially stronger than single bonds. Likewise, triple bonds are even stronger. The triple bond that exists between the nitrogen atoms in nitrogen gas (N_2) makes it very unreactive. All plants and animals require the element nitrogen, but it cannot be obtained from the direct absorption of nitrogen gas from the atmosphere because of its strong, unreactive triple bond. Instead, bacteria convert the nitrogen to a more usable form such as ammonium and nitrate ions, which are then absorbed by plants from the soil. Animals only obtain nitrogen by eating those plants.

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Summary

- Bond energy is an indication of the amount of energy needed for a chemical reaction.
- The higher the bond energy, the less reactive the bond is.

Review

1. What does a high bond energy mean?
2. What kind of bonds do more reactive compounds have?
3. Which will react more readily: a C-H bond or a Cl-Cl bond?

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9.10: VSEPR Theory

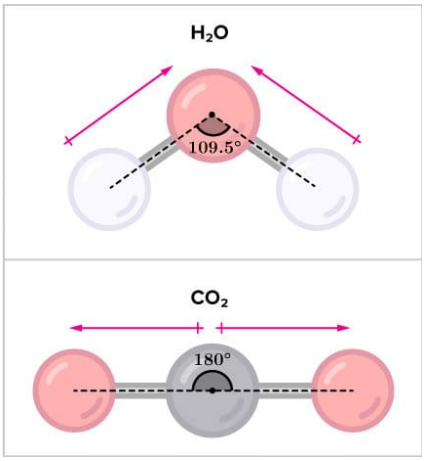


Figure 9.10.1: Copy and Paste Caption here. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Why is the water molecule bent like that?

The characteristic bent shape of the water molecule shown above was a puzzling discovery for scientists at first. The shape allows the molecule to be polar, increasing its boiling point and making it possible for life on earth to exist as we know it. But what makes it bend? The structure is almost the same as carbon dioxide which is known to be a gas at room temperature, why not water too?

Putting atoms together to form compounds can be done on paper or in the lab. However, when the shape of the molecule made in the lab is different from the shape of the molecule drawn on paper, then we need to rethink our ideas and find better explanations.

VSEPR Theory

In 1956, British scientists R.J. Gillespie and R.S. Nyholm recognized that the current model for explaining bond angles did not work well. The theory at that time relied on hybrid orbitals to explain all aspects of bonding. The problem was that this theory gave an incorrect prediction of bond angles for many compounds. They developed a new approach based on earlier work by other scientists that incorporated a consideration of electron pairs in predicting three-dimensional structure.

The **valence shell** is the outermost electron-occupied shell of an atom. The valence shell holds the electrons that are involved in bonding and are the electrons shown in a Lewis structure. The acronym VSEPR stands for the **valence-shell electron pair repulsion** model. The model states that electron pairs will repel each other such that the shape of the molecule will adjust, so that the valence electron-pairs stay as far apart from one another as possible. Molecules can be systematically classified according to the number of bonding pairs of electrons; as well as the number of nonbonding, or lone pairs, around the central atom. For the purposes of the VSEPR model, a double or triple bond is no different in terms of repulsion than a single bond.

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Summary

- VSEPR theory allows more accurate predictions of molecular shape.

Review

1. Who did the major work in developing the VSEPR theory?
2. Where are the electrons that are involved in bonding?
3. What is the basic idea behind the VSEPR theory?

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9.11: Molecular Shapes- No Lone Pairs on Central Atoms

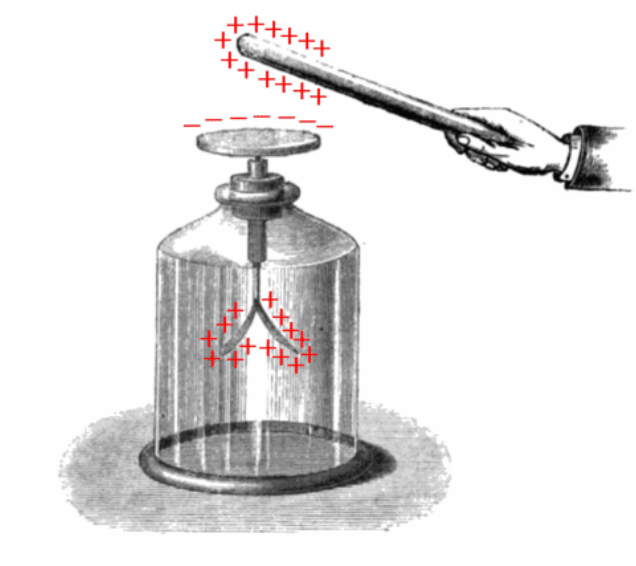


Figure 9.11.1 (Public Domain; Sylvanus P. Thompson via Wikipedia)

How does an electroscope work?

An electroscope is a device used to study charge. When a positively charged object (the rod) nears the upper post, electrons flow to the top of the jar, leaving the two gold leaves positively charged. The leaves repel each other since both hold positive, like charges. The VSEPR theory says that electron pairs, also a set of like charges, will repel each other such that the shape of the molecule will adjust, so that the valence electron pairs stay as far apart from each other as possible.

Central Atom with No Lone Pairs

In order to easily understand the types of molecules possible, we will use a simple system to identify the parts of any molecule:

- A = **central atom** in a molecule
- B = atoms surrounding the central atom

Subscripts after the B will denote the number of B atoms that are bonded to the central A atom. For example, AB_4 is a molecule with a central atom surrounded by four covalently bonded atoms. Again, it does not matter if those bonds are single, double, or triple bonds.

AB_2: Beryllium Hydride (BeH_2)

Beryllium hydride consists of a central beryllium atom with two single bonds to hydrogen atoms. Recall that it violates the octet rule.

$$H—Be—H$$

According to the requirement that electron pairs maximize their distance from one another, the two bonding pairs in the BeH_2 molecules will arrange themselves on directly opposite sides of the central Be atom. The resulting geometry is a linear molecule, shown in the figure below in a "ball and stick" model.

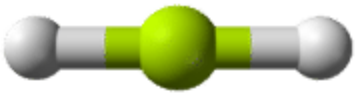


Figure 9.11.2: Beryllium hydride model. (Public Domain; Ben Mills (User:Benjah-bmm27/Wikimedia Commons) via Wikipedia)

The bond angle from $\mathrm{H{-}Be{-}H}$ is 180° because of its linear geometry.

Carbon dioxide is another example of a molecule which falls under the AB_2 category. Its Lewis structure consists of double bonds between the central carbon and the oxygen atoms (see figure below).

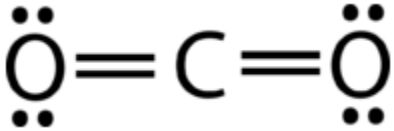


Figure 9.11.3: Carbon dioxide bonding. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The repulsion between the two groups of four electrons (two pairs) is no different than the repulsion of the two groups of two electrons (one pair) in the $\mathrm{BeH_2}$ molecule. Carbon dioxide is also linear (see figure below).

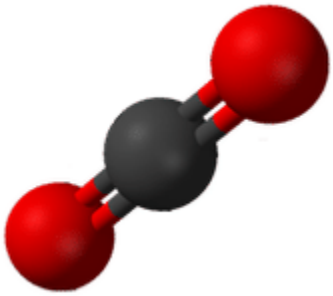


Figure 9.11.4: Carbon dioxide. (Public Domain; User:Benji9072/Wikimedia Commons via Wikipedia)

AB_3: Boron Trifluoride $(\mathrm{BF_3})$

Boron trifluoride consists of a central boron atom with three single bonds to fluorine atoms (see figure below). The boron atom also has an incomplete octet.

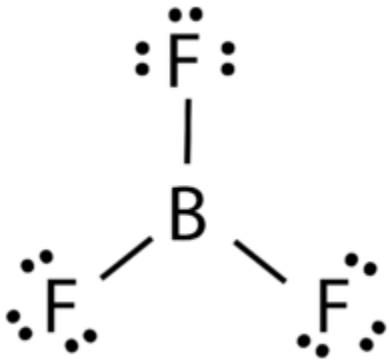


Figure 9.11.5: Boron trifluoride bonding. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The geometry of the $\mathrm{BF_3}$ molecule is called trigonal planar (figure 9.11.3). The fluorine atoms are positioned at the vertices of an equilateral triangle. The $\mathrm{F{-}B{-}F}$ angle is 120° and all four atoms lie in the same plane.

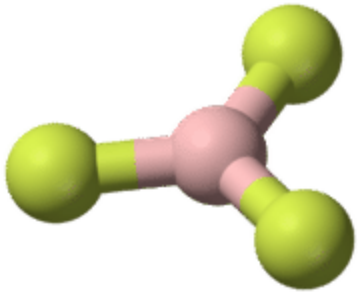


Figure 9.11.6: Boron trifluoride model. (Public Domain; User:Benji90072/Wikimedia Commons via Wikipedia)

AB_4: Methane (CH_4)

Methane is an organic compound that is the primary component of natural gas. Its structure consists of a central carbon atom with four single bonds to hydrogen atoms (see figure below). In order to maximize their distance from one another, the four groups of bonding electrons do not lie in the same plane. Instead, each of the hydrogen atoms lie at the corners of a geometrical shape called a tetrahedron. The carbon atom is at the center of the tetrahedron. Each face of a tetrahedron is an equilateral triangle.

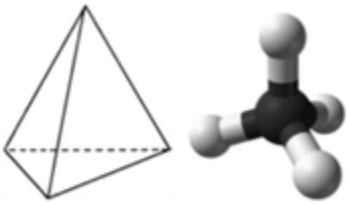


Figure 9.11.7: Tetrahedral structure of methane. (Public Domain; (Left) Pearson Scott Foresman; (Right) Ben Mills (Wikimedia: Benjah-bmm27) via (Left) http://commons.wikimedia.org/wiki/File:Tetrahedron_(PSF).png; (Right) http://commons.wikimedia.org/wiki/File:Methane-CRC-MW-3D-balls.png)

The molecular geometry of the methane molecule is tetrahedral (see figure 9.11.5). The $H-C-H$ bond angles are 109.5^{o}, which is larger than the 90° that they would be if the molecule was planar. When drawing a structural formula for a molecule such as methane, it is advantageous to be able to indicate the three-dimensional character of its shape. The structural formula below is called a perspective drawing. The dotted line bond is to be visualized as receding into the page, while the solid triangle bond is to be visualized as coming out of the page.

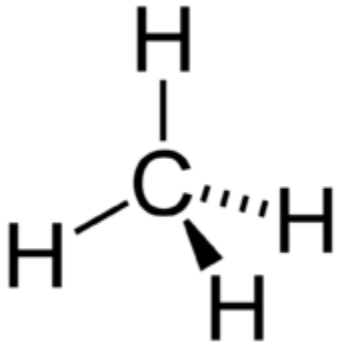


Figure 9.11.8: Methane perspective model. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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Summary

- Electron pairs repel each other and influence bond angles and molecular shape.

Review

1. What are the bond angles in carbon dioxide?
2. What molecule has bond angles of 109.5°?
3. What is the geometry of the BF_3 molecule?

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9.12: Molecular Shapes - Lone Pair(s) on Central Atom



Figure 9.12.1 (Used under license from Shutterstock.com; Image copyright Jandrie Lombard, 2014 via http://www.shutterstock.com)

How can all these clothes fit into such a small space?

When we travel, we often take a lot more stuff than we need. Trying to fit it all into a suitcase can be a real challenge. We may have to repack or just squeeze it all in. Atoms often have to rearrange where the electrons are in order to create a more stable structure.

Central Atom with One or More Lone Pairs

The molecular geometries of molecules change when the central atom has one or more lone pairs of electrons. The total number of electron pairs, both bonding pairs and lone pairs, leads to what is called the **electron domain geometry**. When one or more of the bonding pairs of electrons is replaced with a lone pair, the molecular geometry (actual shape) of the molecule is altered. In keeping with the A and B symbols established in the previous section, we will use E to represent a lone pair on the central atom (A). A subscript will be used when there is more than one lone pair. Lone pairs on the surrounding atoms (B) do not affect the geometry.

AB_3E: Ammonia, NH_3

The ammonia molecule contains three single bonds and one lone pair on the central nitrogen atom (see figure below).

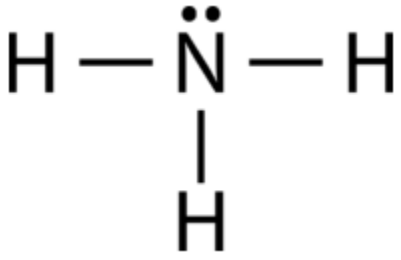


Figure 9.12.2: Lone pair electrons in ammonia. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The domain geometry for a molecule with four electron pairs is tetrahedral, as was seen with CH_4. In the ammonia molecule, one of the electron pairs is a lone pair rather than a bonding pair. The molecular geometry of NH_3 is called trigonal pyramidal (see figure below).

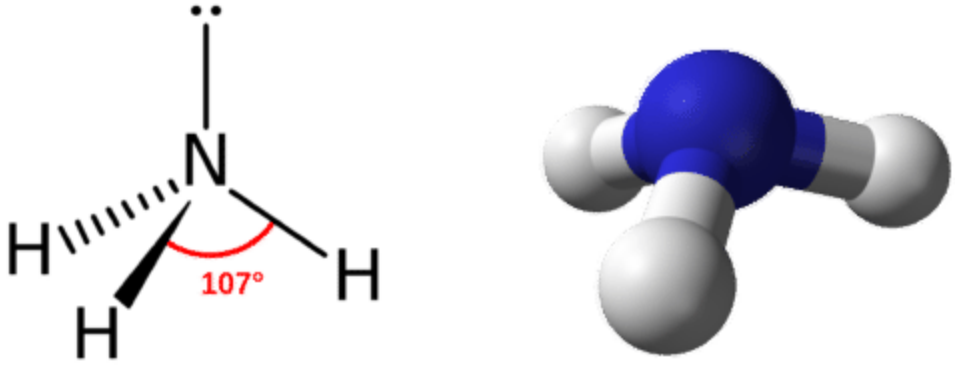


Figure 9.12.3: Ammonia molecule. (Public Domain; (Left) User: Booyabazooka/ Wikimedia Commons; (Right) Ben Mills (Wikimedia: Benjah-bmm27) via (Left) http://commons.wikimedia.org/wiki/File:Ammonia_lone_electron_pair.svg; (Right) http://commons.wikimedia.org/wiki/File:Ammonia-3D-balls-A.png)

Recall that the bond angle in the tetrahedral $\mathrm{CH_4}$ molecule is 109.5°. Again, the replacement of one of the bonded electron pairs with a lone pair compresses the angle slightly. The $\mathrm{H{-}N{-}H}$ angle is approximately 107°.

$\mathrm{AB_2E_2}$: Water, $\mathrm{H_2O}$

A water molecule consists of two bonding pairs and two lone pairs (see figure below).



Figure 9.12.4: Lone pair electrons on water. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

As for methane and ammonia, the domain geometry for a molecule with four electron pairs is tetrahedral. In the water molecule, two of the electron pairs are lone pairs rather than bonding pairs. The molecular geometry of the water molecule is bent. The $\mathrm{H{-}O{-}H}$ bond angle is 104.5°, which is smaller than the bond angle in $\mathrm{NH_3}$ (see figure below).

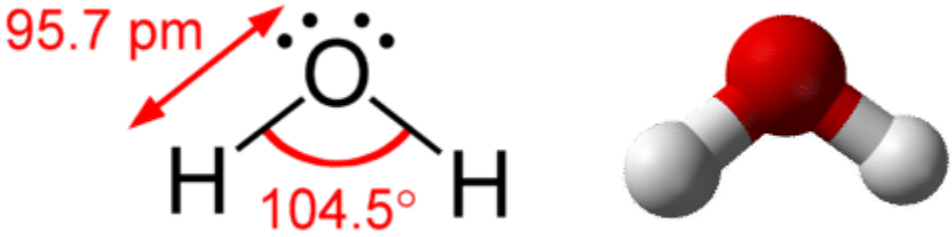


Figure 9.12.5: Water molecule. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via (Left) http://commons.wikimedia.org/wiki/File:Water-dimensions-from-Greenwood%2526Earnshaw-2D.png; (Right) http://commons.wikimedia.org/wiki/File:Water-3D-balls-A.png; http://commons.wikimedia.org/wiki/File:Zeepzieden.jpg)

$\mathrm{AB_4E}$: Sulfur Tetrafluoride, $\mathrm{SF_4}$

The Lewis structure for $\mathrm{SF_4}$ contains four single bonds and a lone pair on the sulfur atom (see figure below).

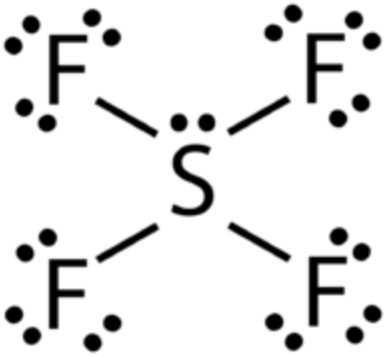


Figure 9.12.6: Lone pair electrons in SF_4. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

The sulfur atom has five electron groups around it, which corresponds to the trigonal bipyramidal domain geometry, as in PCl_5 (see figure below). Recall that the trigonal bipyramidal geometry has three equatorial atoms and two axial atoms attached to the central atom. Because of the greater repulsion of a lone pair, it is one of the equatorial atoms that are replaced by a lone pair. The geometry of the molecule is called a distorted tetrahedron, or seesaw.

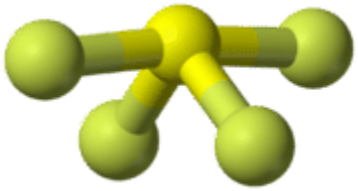


Figure 9.12.7: Ball and stick model for SF_4. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

Table 9.12.1: Geometries in Which the Central Atom Has One or More Lone Pairs

Total Number of Electron Pairs	Number of Bonding Pairs	Number of Lone Pairs	Electron Domain Geometry	Molecular Geometry	Examples
3	2	1	Trigonal Planar	Bent	O_3
4	3	1	Tetrahedral	Trigonal Pyramidal	NH_3
4	2	2	Tetrahedral	Bent	H_2O
5	4	1	Trigonal Bipyramidal	Distorted Tetrahedron (Seesaw)	SF_4
5	3	2	Trigonal Bipyramidal	T-shaped	ClF_3
5	2	3	Trigonal Bipyramidal	Linear	I_3^-
6	5	1	Octahedral	Square Pyramidal	BrF_5
6	4	2	Octahedral	Square Planar	XeF_4

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Summary

- The presence of lone pair electrons influences the three-dimensional shape of the molecule.

Review

1. Why does water have a bent geometry?
2. Why is ammonia not a planar molecule?
3. How would we write the configuration for xenon tetrafluoride using the ABE system?

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9.13: Periodic Trends- Electronegativity



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Is it easy or hard for you to make new friends?

Have you ever noticed how some people attract others to them? Whether it be their personality, attractiveness, or athletic skills—something pulls people toward them; while others have a smaller group of friends and acquaintances. Atoms do the same thing. One atom may pull electrons strongly to it, while a second type of atom has much less "pulling power".

Electronegativity

Valence electrons of both atoms are always involved when those two atoms come together to form a chemical bond. Chemical bonds are the basis for how elements combine with one another to form compounds. When these chemical bonds form, atoms of some elements have a greater ability to attract the valence electrons involved in the bond than other elements.

Electronegativity is a measure of the ability of an atom to attract the electrons when the atom is part of a compound. Electronegativity differs from electron affinity because electron affinity is the actual energy released when an atom gains an electron. Electronegativity is not measured in energy units, but instead a relative scale. All elements are compared to one another, with the most electronegative element, fluorine, being assigned an electronegativity value of 3.98. Fluorine attracts electrons better than any other element. The table below shows the electronegativity values for the elements.

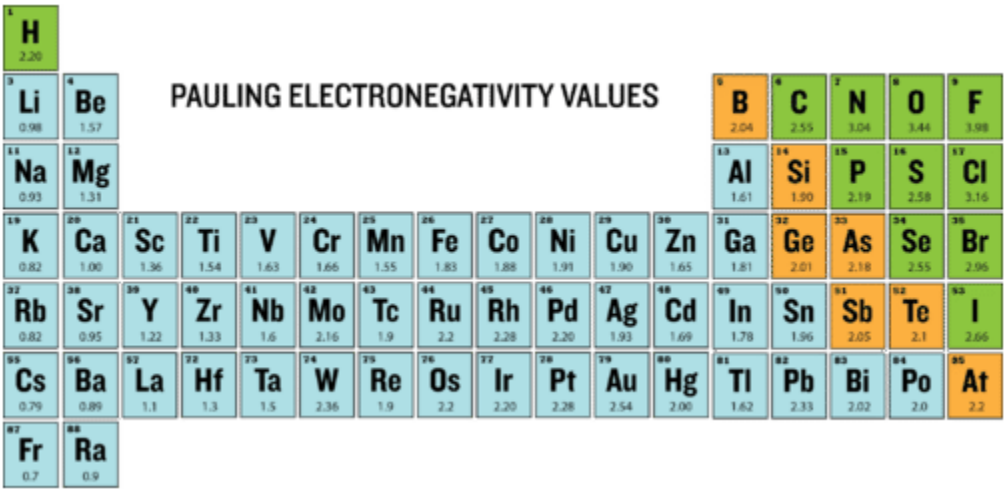


Figure 9.13.2: The electronegativity scale was developed by Nobel Prize winning American chemist Linus Pauling. The largest electronegativity (3.98) is assigned to fluorine and all other electronegativity measurements are on a relative scale. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Since metals have few valence electrons, they tend to increase their stability by losing electrons to become cations. Consequently, the electronegativities of metals are generally low. Nonmetals have more valence electrons and increase their stability by gaining electrons to become anions. The electronegativities of nonmetals are generally high.

Trends

Electronegativities generally increase from left to right across a period. This is due to an increase in nuclear charge. Alkali metals have the lowest electronegativities, while halogens have the highest. Because most noble gases do not form compounds, they do not have electronegativities. Note that there is little variation among the transition metals. Electronegativities generally decrease from top to bottom within a group, due to the larger atomic size.

Of the main group elements, fluorine has the highest electronegativity (EN $= 4.0$) and cesium the lowest (EN $= 0.79$). This indicates that fluorine has a high tendency to gain electrons from other elements with lower electronegativities. We can use these values to predict what happens when certain elements combine.

When the difference between atom electronegativities is greater than ~1.7, then a complete exchange of electrons occurs. Typically this exchange is between a metal and a nonmetal. For instance, sodium and chlorine will typically combine to form a new compound and each ion becomes isoelectronic with its nearest noble gas. When we compare the EN values, we see that the electronegativity for Na is 0.93 and the value for Cl is 3.2. The absolute difference between ENs is $|0.93 - 3.2| = 2.27$. This value is greater than 1.7, and therefore indicates the occurence of a complete electron exchange.

Summary

- Electronegativity is a measure of the ability of an atom to attract the electrons when the atom is part of a compound.
- Electronegativity values generally increase from left to right across the periodic table.
- Electronegativities generally decrease from the top to bottom of a group.
- The highest electronegativity value is for fluorine.

Review

1. Define “electronegativity.”
2. How does electronegativity differ from electron affinity?
3. Why are the electronegativity values of metals generally low?
4. Describe the trend in electronegativities across the periodic table.
5. Describe the trends in electronegativities in a group of the periodic table.

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9.14: Bond Polarity



Figure 9.14.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

What makes people share?

Have you ever spent time with someone you really didn't like? You had nothing in common with them and did not want to have anything to do with them. On the other hand, there are people that you enjoy being with. You have a lot in common and like to share with them. Atoms work the same way. If there are strong differences in their attraction of electrons, one atom gets the electrons and the other atom loses them. If they are similar, they share electrons to form a covalent bond.

Bond Polarity

Electronegativity is defined as the ability of an atom to attract electrons when the atoms are in a compound. Electronegativities of elements are shown in the periodic table below.

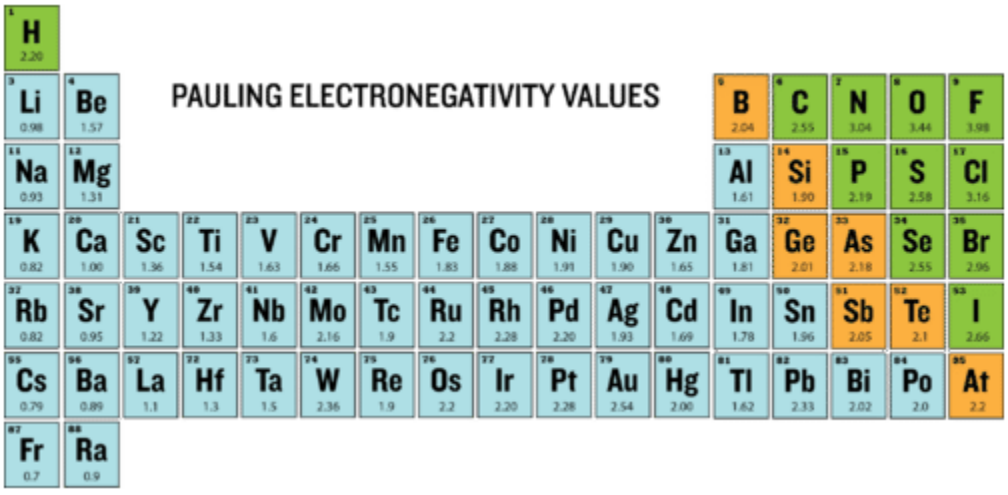


Figure 9.14.2: Electronegativities of elements. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The degree to which a given bond is ionic or covalent is determined by calculating the difference in electronegativity between the two atoms involved in the bond.

As an example, consider the bond that occurs between an atom of potassium and an atom of fluorine. Using the table, the difference in electronegativity is equal to $4.0 - 0.8 = 3.2$. Since the difference in electronegativity is relatively large, the bond between the two atoms is ionic. Since the fluorine atom has a much larger attraction for electrons than the potassium atom does, the valence electron from the potassium atom is completely transferred to the fluorine atom. The diagram below shows how difference in electronegativity relates to the ionic or covalent character of a chemical bond.

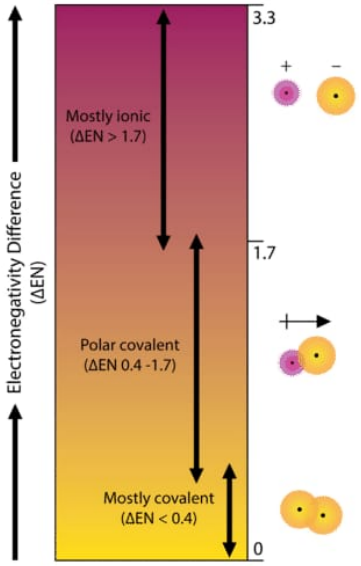


Figure $9.14.3$: Bond type is predicted on the difference in electronegativity of the two elements involved in the bond. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 1.7 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.

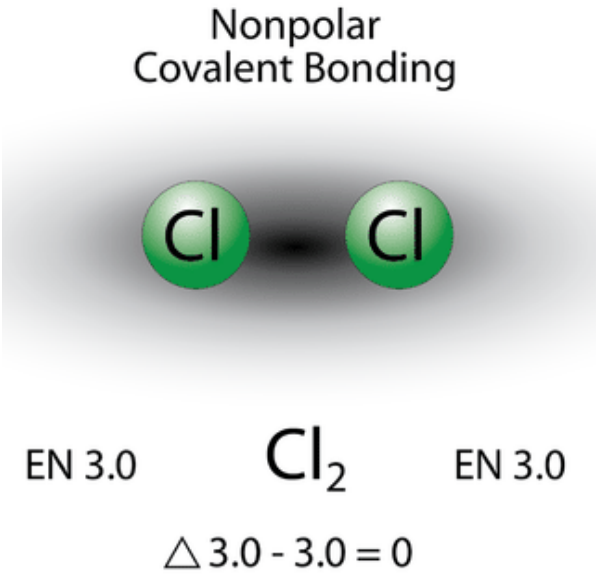


Figure $9.14.4$: A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal. (CC BY-NC 3.0; Jodi So via CK-12 Foundation)

The two chlorine atoms share the pair of electrons in the single covalent bond equally, and the electron density surrounding the

Cl_2 molecule is symmetrical. Note that molecules in which the electronegativity difference is very small (<0.4) are also considered nonpolar covalent. An example would be a bond between chlorine and bromine ($\Delta EN = 3.0 - 2.8 = 0.2$).

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.4 and 1.7 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.

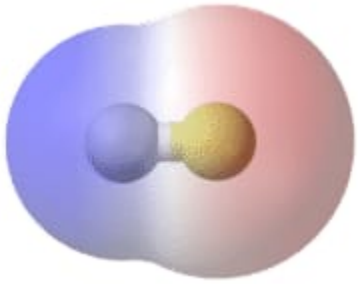


Figure 9.14.5: In the polar covalent bond of HF, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Commons Wikimedia, Hydrogen fluoride elpot transparent 3D balls [commons.wikimedia.org])

An easy way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta (δ).

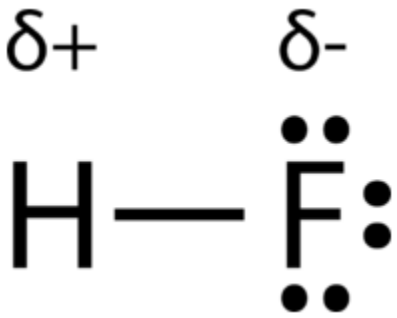


Figure 9.14.6: Use of δ to indicate partial charge. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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 crossed arrow can also be used to indicate the direction of greater electron density.

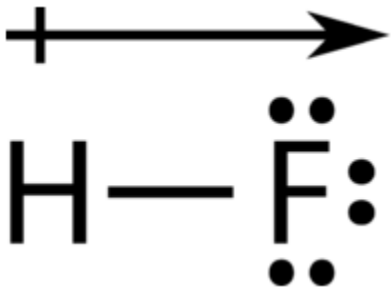


Figure 9.14.7: Use of crossed arrow to indicate polarity. (Credit: Joy Sheng Source: CK-12 Foundation; License: CC BY-NC 3.0)

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Simulation

If it wasn't for our understanding of bond polarity, we'd have a really tough time cleaning our clothing. Use this simulation to understand how the unique polarity of soap molecules helps to clean our clothing!

Summary

- The electronegativity of an atom determines how strongly it attracts electrons to itself.
- The polarity of a bond is affected by the electronegativity values of the two atoms involved in that bond.

Review

1. If two atoms bonded together have an electronegativity difference of 1.9, what is the bond type?
2. What would be the bond type for BH_2?
3. Your friend tells you that the LiF bond is covalent. Are they correct? Why or why not.

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9.15: Polar Molecules

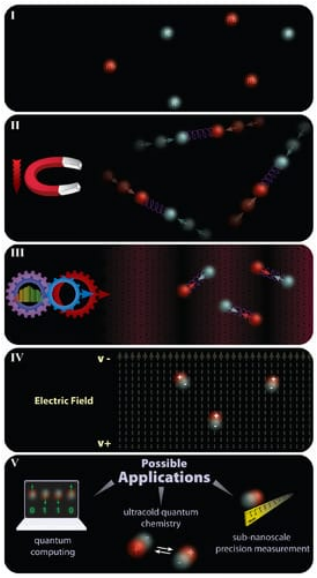


Figure $9.15.1$ (Public Domain; G. Kuebler/JILA and the NIST via Patapsco Image Gallery [patapsco.nist.gov])

How cold are ultracold polar molecules?

Ultracold polar molecules have been created by research scientists. These "molecules" consist of Rb and K atoms excited by lasers to form a type of $\mathrm{Rb}{-}\mathrm{K}$ compound where the Rb has a positive charge and the K has a negative charge. The material is formed at temperatures extremely close to absolute zero. The researchers believe these techniques will help them make new reactions and new materials.

Polar Molecules

A **polar molecule** is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative. 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 $9.15.2$: A dipole is any molecule with a positive end and a negative end, resulting from unequal distribution of electron density throughout the molecule. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Polar molecules orient themselves in the presence of an electric field with the positive ends of the molecule being attracted to the negative plate, while the negative ends of the molecules are attracted to the positive plate (see figure below).

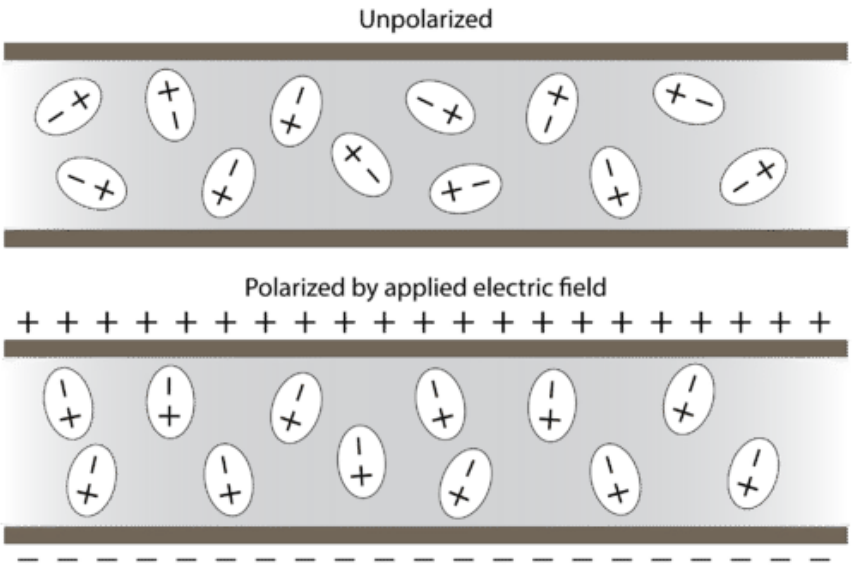


Figure 9.15.3: Polar molecules are randomly oriented in the absence of an applied electric field (top). In an electric field, the molecules orient themselves because of the attraction of opposite charges (bottom). (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

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 (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 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.

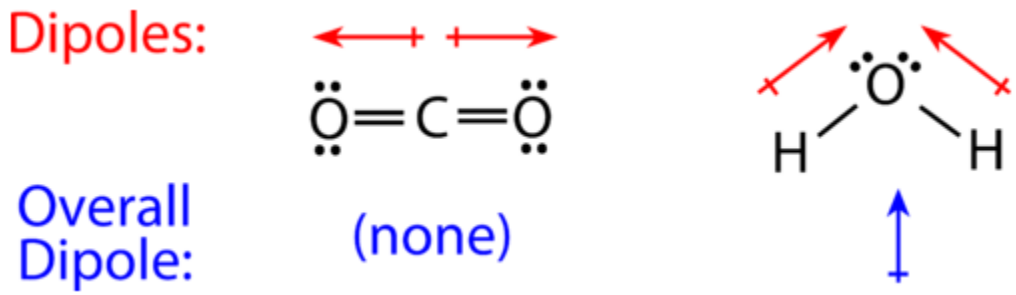


Figure 9.15.4: The molecular geometry of a molecule affects its polarity. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Some other molecules are shown in the figure below. Notice that a tetrahedral molecule such as 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 (BF_3) may be nonpolar if all three peripheral atoms are the same, but a trigonal pyramidal molecule (NH_3) is polar.

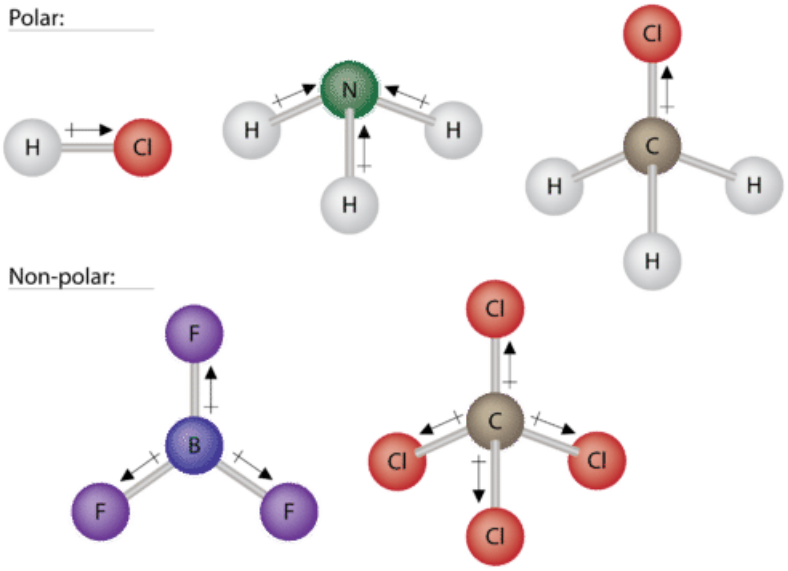


Figure 9.15.5: Some examples of polar and nonpolar molecules based on molecular geometry. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

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Simulations

Why does water form droplets? How does soap get our clothes clean? Hint: both actions have to do with polarity. Try these simulations to learn more:

How are stains removed from your clothing?

Why do liquids form droplets?

Summary

- Polar molecules result from differences in the electronegativity of the atoms in the molecule.
- Dipoles that are directly opposite one another cancel each other out.

Review

1. Why is a polar molecule called a dipole?
2. The molecular geometry of NH_3 is trigonal pyramidal. What about the molecular geometry of NH_3 makes it a polar molecule?
3. What happens to polar molecules when an electric field is applied?

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9.16: Van der Waals Forces



Figure 9.16.1 (Public Domain; User:MartinSaunders/Wikimedia Commons via Wikipedia)

How to Keep it Cool?

Magnetic resonance imaging (MRI) devices use liquid nitrogen to cool the superconducting magnets. Nitrogen is a gas at room temperature and liquefies at $-195.8^{\circ}\mathrm{C}$. Its neighbor on the periodic table (oxygen) boils at $-182.95^{\circ}\mathrm{C}$. The interactions between nitrogen molecules (N_2) are weaker, so the boiling point is lower. Interactions between nonpolar molecules depend on the degree of electron fluctuation within the molecule.

Van der Waals Forces

The first type of intermolecular forces that we consider are van der Waals forces, after Dutch chemist Johannes van der Waals (1837-1923). **Van der Waals forces** are the weakest intermolecular force, and consist of dipole-dipole forces and dispersion forces.

Dipole-Dipole Forces

Dipole-dipole forces are the attractive forces that occur between polar molecules. A molecule of hydrogen chloride has a partially positive hydrogen atom and a partially negative chlorine atom. In a collection of many hydrogen chloride molecules, the molecules will align themselves so that the oppositely charged regions of neighboring molecules are near each other.

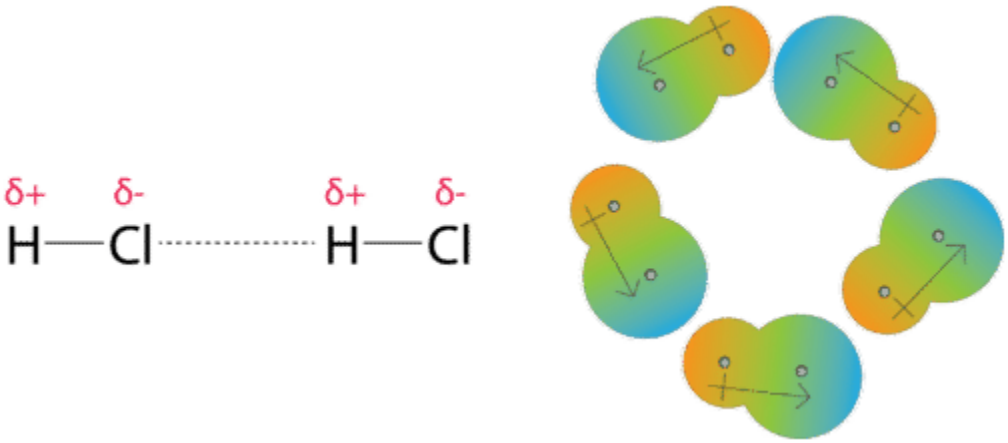


Figure 9.16.2: Dipole-dipole forces are a result of the attraction of the positive end of one dipole to the negative end of a neighboring dipole. (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Dipole-dipole forces are similar in nature to ionic bonds, but much weaker.

London Dispersion Forces

Dispersion forces are also considered a type of van der Waals force and are the weakest of all intermolecular forces. They are often called London dispersion forces after Fritz London (1900-1954), who first proposed their existence in 1930. **London dispersion forces** are the intermolecular forces that occur between atoms, and between nonpolar molecules as a result of the motion of electrons.

The electron cloud of a helium atom contains two electrons, which can normally be expected to be equally distributed spatially around the nucleus. However, at any given moment the electron distribution may be uneven, resulting in an instantaneous dipole. This weak and temporary dipole subsequently influences neighboring helium atoms through electrostatic attraction and repulsion. It induces a dipole on nearby helium atoms (see figure below).

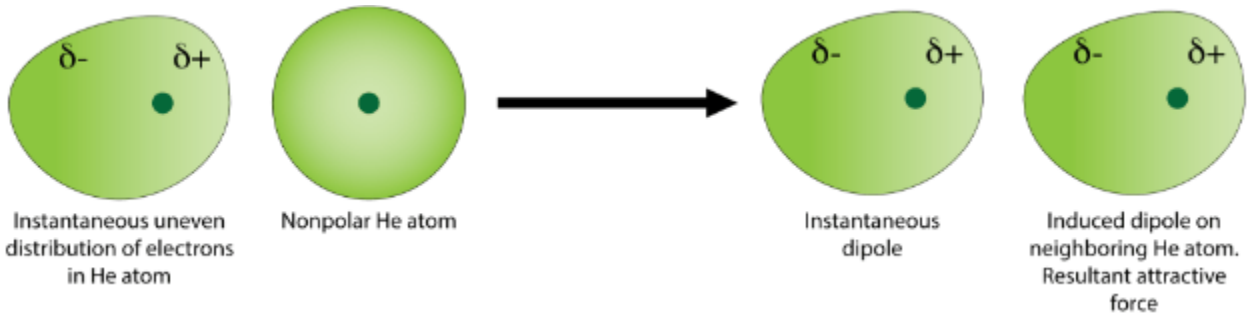


Figure 9.16.3: A short-lived or instantaneous dipole in a helium atom. (CC BY-NC 3.0; Zachary Wilson via CK-12 Foundation)

The instantaneous and induced dipoles are weakly attracted to one another. The strength of dispersion forces increases as the number of electrons in the atoms or nonpolar molecules increases.

The halogen group consists of four elements that all take the form of nonpolar diatomic molecules. The table below shows a comparison of the melting and boiling points for each.

Melting and Boiling Points of Halogens

Table 9.16.1: Melting and Boiling Points of Halogens

Molecule	Total Number of Electrons	Melting Point (°C)	Boiling Point (°C)	Physical State at Room Temperature
F_2	18	-220	-188	gas
Cl_2	34	-102	-34	gas
Br_2	70	-7	59	liquid
I_2	106	114	184	solid

The dispersion forces are strongest for iodine molecules because they have the greatest number of electrons. The relatively stronger forces result in melting and boiling points that are the highest of the halogen group. These forces are strong enough to hold iodine molecules close together in the solid state at room temperature. The dispersion forces are progressively weaker for bromine, chloride, and fluorine; this is illustrated in their steadily lower melting and boiling points. Bromine is a liquid at room temperature, while chlorine and fluorine are gases whose molecules are much further apart from one another. Intermolecular forces are nearly nonexistent in the gas state, and so the dispersion forces in chlorine and fluorine only become measurable as the temperature decreases and they condense into the liquid state.

Interactive Content

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Simulation

Why does water form droplets? Hint: it has to do with the interactions between water molecules. Try out this simulation to learn more.

Summary

- Van der Waals forces are weak interactions between molecules that involve dipoles.
- Polar molecules have permanent dipole-dipole interactions.
- Nonpolar molecules can interact by way of London dispersion forces.

Review

1. What attractive forces develop between polar molecules?
2. What creates London dispersion forces?
3. Are London dispersion forces permanent or temporary?
4. Are the dispersion forces for Cl_2 stronger or weaker than the ones for Br_2?

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9.17: Hydrogen Bonding

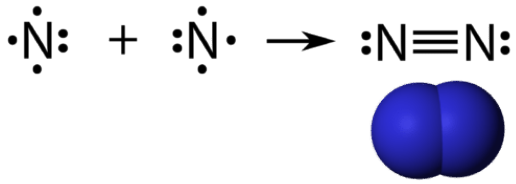


Figure 9.17.1 (CC BY-NC-SA 3.0; Molecule Public Domain; CK-12 Foundation - Joy Sheng, using 3D molecular structure by Ben Mills (Wikimedia: Benjah-bmm27) via Molecular structure: http://commons.wikimedia.org/wiki/File:Nitrogen-3D-vdW.png)

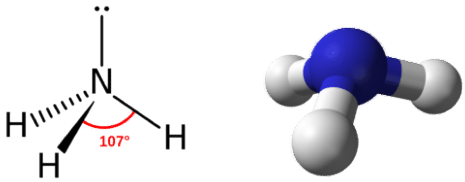


Figure 9.17.2 (Public Domain; (Left) User:Booyabazooka/Wikimedia Commons; (Right) Ben Mills (Wikimedia: Benjah-bmm27) via (Left) http://commons.wikimedia.org/wiki/File:Ammonia_lone_electron_pair.svg; (Right) http://commons.wikimedia.org/wiki/File:Ammonia-3D-balls-A.png)

What's the difference between these two molecules?

A rough rule of thumb is that higher molecular-weight materials have higher boiling points than their lower molecular-weight counterparts. More energy is needed to move the larger molecule from the liquid state to the vapor state. However, ammonia has a boiling point of $-33.34^{\circ}\mathrm{C}$ and a molecular weight of 17, while nitrogen (molecular weight 28) has a boiling point of $-195.8^{\circ}\mathrm{C}$. The lighter ammonia molecule must have other factors that influence its physical properties.

Hydrogen Bonding

The attractive force between water molecules is a dipole interaction. The hydrogen atoms are bound to the highly electronegative oxygen atom (which also possesses two lone pair sets of electrons, making for a very polar bond). The partially positive hydrogen atom of one molecule is then attracted to the oxygen atom of a nearby water molecule (see figure below).

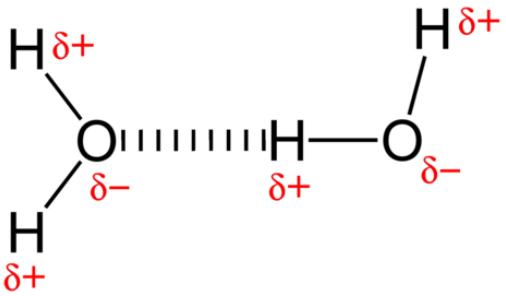


Figure 9.17.3: A hydrogen bond in water occurs between the hydrogen atom of one water molecule and the lone pair of electrons on an oxygen atom of a neighboring water molecule. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

A **hydrogen bond** is an intermolecular attractive force in which a hydrogen atom that is covalently bonded to a small, highly electronegative atom is attracted to a lone pair of electrons on an atom in a neighboring molecule. Hydrogen bonds are very strong compared to other dipole interactions. The strength of a typical hydrogen bond is about 5% of that of a covalent bond.

Hydrogen bonding occurs only in molecules where hydrogen is covalently bonded to one of three elements: fluorine, oxygen, or nitrogen. These three elements are so electronegative that they withdraw the majority of the electron density in the covalent bond with hydrogen, leaving the H atom very electron-deficient. The H atom nearly acts as a bare proton, leaving it very attracted to

lone pair electrons on a nearby atom.

The hydrogen bonding that occurs in water leads to some unusual, but very important, properties. Most molecular compounds that have a mass similar to water are gases at room temperature. Because of the strong hydrogen bonds, water molecules are able to stay condensed in the liquid state. The figure below shows how the bent shape, and two hydrogen atoms per molecule, allows each water molecule to be able to hydrogen bond to two other molecules.

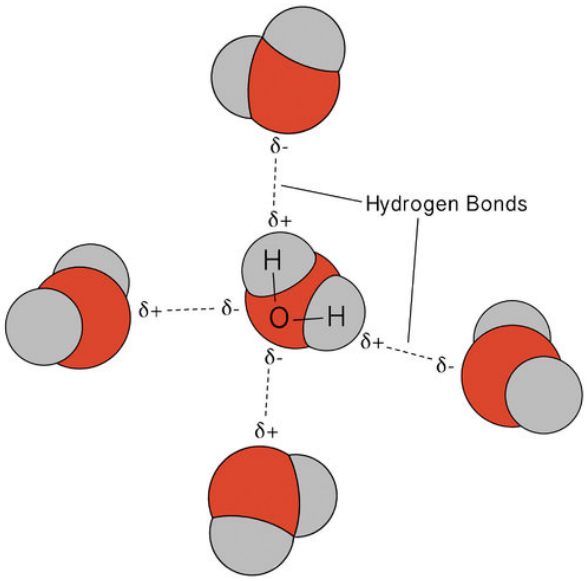


Figure 9.17.4: Multiple hydrogen bonds occur simultaneously in water because of its bent shape and the presence of two hydrogen atoms per molecule. (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

In the liquid state, the hydrogen bonds of water can break and reform as the molecules flow from one place to another. When water is cooled, the molecules begin to slow down. Eventually, when water is frozen to ice, the hydrogen bonds become permanent and form a very specific network (see figure below).

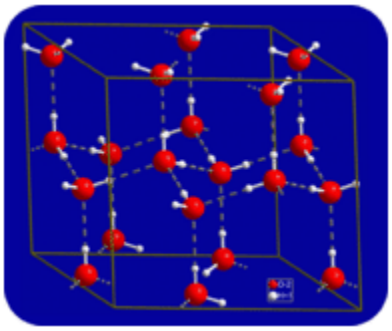


Figure 9.17.5: When water freezes to ice, the hydrogen bonding network becomes permanent. Each oxygen atom has an approximately tetrahedral geometry—two covalent bonds and two hydrogen bonds. (Public Domain; User:Materialscientist/Wikimedia Commons via Wikipedia)

The bent shape of the molecules leads to gaps in the hydrogen bonding network of ice. Ice has a very unusual property—its solid state is less dense than its liquid state. Ice floats in water. Virtually all other substances are denser in the solid state than in the liquid state. Hydrogen bonds play a very important biological role in the physical structures of proteins and nucleic acids.

Interactive Content

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Simulation

Why does water form droplets? Hint: it has to do with the attraction between water molecules. Learn more (and see how many drops of water you can get a penny can hold) in this simulation.

Summary

- Hydrogen bonds form when a H attached to a N, O, or F atom interacts with another N, O, or F atom.

Review

1. How strong is a hydrogen bond?
2. What happens when H is covalently bonded to N, O, or F?
3. How does the shape of the water molecule affect the properties of ice?

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10: CHEMICAL REACTIONS

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10.1: Word Equations

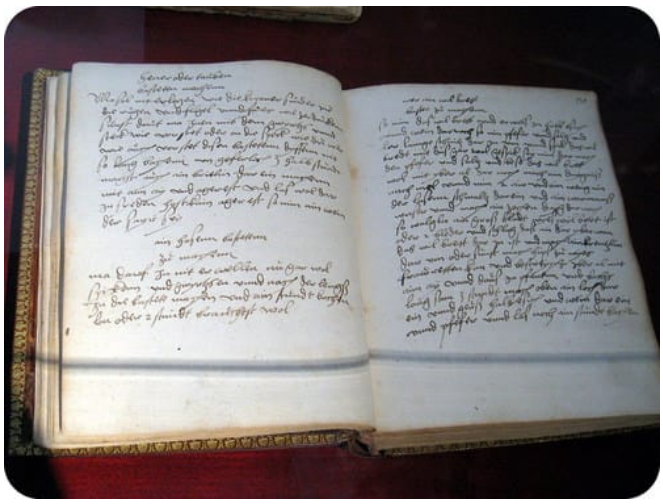


Figure 10.1.1 (Public Domain; User:Daderot/Wikimedia Commons via Wikipedia)

What's for Dinner?

Various ways of recording recipes have developed over the centuries. The cookbook shown above was written by a woman who probably collected all her own recipes. Later, printed cookbooks became available (even guys had no excuse for not being able to cook). Today we can find recipes on a number of internet sites and can quickly search for information on how to cook anything we want. Reading a recipe sometimes requires we understand a few codes and symbols (what's the difference between a tsp and a Tsp?), but the information on what we start with and what we end up with is there.

Writing Chemical Equations

Chemical reactions are occurring all around us. Plants use sunlight to drive their photosynthetic process and produce energy. Cars and other vehicles burn gasoline in order to power their engines. Batteries use electrochemical reactions to produce energy and power many everyday devices. Many chemical reactions are going on within the human body as well, particularly during the digestion of food.

In math class, you have written and solved many mathematical equations. Chemists keep track of chemical reactions by writing equations as well. In any chemical reaction, one or more substances—called **reactants**—are converted into one or more new substances—called **products**. The general form of the equation for such a process looks like this:

$$\text{Reactants} \rightarrow \text{Products}$$

Unlike a math equation, a chemical equation does not use an equal sign. Instead, the arrow is called a yield sign and so the equation is described as "reactants yield products".

Word Equations

You can describe a chemical reaction by writing a **word equation**. When silver metal is exposed to sulfur, it reacts to form silver sulfide. Silver sulfide is commonly known as tarnish and turns the surface of silver objects dark and streaky black (see figure below). The sulfur that contributes to tarnish can come from traces of sulfur in the air, or from foods such as eggs. The word equation for the process is:

$$\text{Silver} + \text{sulfur} \rightarrow \text{Silver sulfide}$$

The silver and the sulfur are the reactants in the equation, while the silver sulfide is the product.



Figure 10.1.2**:** You can see dark spots of tarnished (silver sulfide) forming on this ring as it reacts with sulfur compounds in the air. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Another common chemical reaction is the burning of methane gas. Methane is the major component of natural gas and is commonly burned on a gas stove or in a Bunsen burner (figure below). Burning is a chemical reaction in which some type of fuel is reacted with oxygen gas. The products of the reaction in the burning of methane, as well as other fuels, are carbon dioxide and water. The word equation for this reaction is:

$$\text{Methane} + \text{oxygen} \rightarrow \text{carbon dioxide} + \text{water}$$



Figure 10.1.3: A Bunsen burner is commonly used to heat substances in a chemistry lab. Methane is reacted with oxygen to form carbon dioxide and water. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Word equations can be very useful, but do have one major drawback—they cannot be used for any quantitative work. A word equation does not tell how many moles of each material are needed, or how many moles of product are formed.

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Summary

- Word equations are used to describe the conversion of reactants to products.

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Review

1. Write the generic form of a chemical reaction.
2. What are reactants?
3. What are products?

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10.2: Chemical Equations



Figure 10.2.1 (Public Domain; jons2 at pdphoto.org via Wikipedia)

How do you make shrimp gumbo?

Shrimp gumbo is one of many dishes that are part of the Cajun culture in Louisiana. It's a spicy dish that needs careful control of all the ingredients so that it has a "kick", but is not overwhelming. Recipes convey not only what the preparation entails, but also describe how much of each ingredient is needed, and the details of how to cook the meal. Similarly, we need this type of information in order to carry out chemical reactions successfully and safely.

Chemical Equations

Word equations are time-consuming to write and do not prove to be convenient for many of the things that chemists need to do with equations. A **chemical equation** is a representation of a chemical reaction that displays the reactants and products with chemical formulas. The chemical equation for the reaction of methane with oxygen is shown:

$$\mathrm{CH_4 + O_2 \rightarrow CO_2 + H_2O}$$

The equation above, called a **skeleton equation**, is an equation that shows only the formulas of the reactants and products with nothing to indicate the relative amounts. The first step in writing an accurate chemical equation is to write the skeleton equation, making sure that the formulas of all substances involved are written correctly. All reactants are written to the left of the yield arrow, separated from one another by a plus sign. Likewise, products are written to the right of the yield arrow, also separated with a plus sign.

It is often important to know the physical states of the reactants and products taking part in a reaction. To do this, put the appropriate symbol in parentheses after each formula: (s) for solid, (l) for liquid, (g) for gas, and (aq) for an aqueous (water-based) solution. The previous reaction becomes:

$$\mathrm{CH_4\,(g) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(l)}$$

The table below shows a listing of symbols used in chemical equations. Some, such as the double arrow which represents equilibrium, and the use of a catalyst in a reaction, will be detailed in other concepts.

Table 10.2.1: Symbols Used in Chemical Equations

Symbols	Description
$+$	Used to separate multiple reactants or products.
$\rightarrow$	Yield sign; separates reactants from products.
$\rightleftharpoons$	Replaces the yield sign for reversible reactions that reach equilibrium.
(s)	Reactant or product in the solid state.
(l)	Reactant or product in the liquid state.
(g)	Reactant or product in the gaseous state.
(aq)	Reactant or product in an aqueous solution (dissolved in water).
$\overset{\text{Pt}}{\rightarrow}$	Formula written above the arrow is used as a catalyst in the reaction.
$\overset{\Delta}{\rightarrow}$	Triangle indicates that the reaction is being heated.

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Summary

- A chemical equation is representation a chemical reaction using chemical formulas reactants and products, as well as symbols.
- Symbols used in chemical equations are described and explained.

Review

1. What does a skeleton equation tell you?
2. Why would you want to know the physical state of materials in a chemical reaction?
3. In a chemical reaction what does symbol → mean?
4. If I see Δ over the arrow, what does this indicate?

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10.3: Balancing Equations



Figure 10.3.1 (Public Domain; Joseph Allen via Wikipedia)

Any leftovers?

When you cook a meal, quite often there are leftovers because you prepared more than people would eat at one sitting. Sometimes when you repair a piece of equipment, you end up with what are called "pocket parts"—small pieces you put in your pocket because you're not sure where they belong. Chemistry tries to avoid leftovers and pocket parts. In normal chemical processes, we cannot create or destroy matter (law of conservation of mass). If we start out with ten carbon atoms, we need to end up with ten carbon atoms. John Dalton's atomic theory said that chemical reactions basically involve the rearrangement of atoms. Chemical equations need to follow these principles in order to be correct.

Balancing Chemical Equations

A **balanced equation** is a chemical equation in which mass is conserved and there are equal numbers of atoms of each element on both sides of the equation. We can write a chemical equation for the reaction of carbon with hydrogen gas to form methane $(\mathrm{CH_4})$:

$$\underset{2\text{ C atoms}}{\mathrm{C}\,(s)} \quad + \quad \underset{2\text{ H atoms}}{\mathrm{H_2}\,(g)} \quad \rightarrow \quad \underset{1\text{ C atom},\,4\text{ H atoms}}{\mathrm{CH_4}\,(g)}$$

In order to write a correct equation, you must first write the correct skeleton equation with the correct chemical formulas. Recall that hydrogen is a diatomic molecule and so is written as $\mathrm{H_2}$.

When we count the number of atoms of both elements, shown under the equation, we see that the equation is not balanced. There are only 2 atoms of hydrogen on the reactant side of the equation, while there are 4 atoms of hydrogen on the product side. We can balance the above equation by adding a coefficient of 2 in front of the formula for hydrogen.

$$\mathrm{C}\,(s) + 2\mathrm{H_2}\,(g) \rightarrow \mathrm{CH_4}\,(g)$$

A **coefficient** is a small whole number placed in front of a formula in an equation in order to balance it. The 2 in front of the $\mathrm{H_2}$ means that there are a total of $2 \times 2 = 4$ atoms of hydrogen as reactants. Visually, the reaction looks like the figure below.

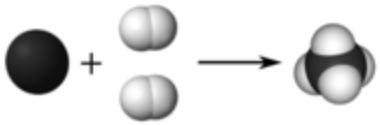


Figure 10.3.2: Reaction between carbon and hydrogen to form methane. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27 via Wikipedia and http://commons.wikimedia.org/wiki/File:Methane-3D-space-filling.png)

In the balanced equation, there is one atom of carbon and four atoms of hydrogen on both sides of the arrow. Below are guidelines for writing and balancing chemical equations.

1. Determine the correct chemical formulas for each reactant and product.
2. Write the skeleton equation.
3. Count the number of atoms of each element that appears as a reactant and as a product. If a polyatomic ion is unchanged on both sides of the equation, count it as a unit.
4. Balance each element one at a time by placing coefficients in front of the formulas.
   a. It is best to begin by balancing elements that only appear in one chemical formula on each side of the equation.
   b. No coefficient is written for a 1.
   c. NEVER change the subscripts in a chemical formula—you can only balance equations by using coefficients.
5. Check each atom or polyatomic ion to be sure that they are equal on both sides of the equation.
6. Make sure that all coefficients are in the lowest possible ratio. If necessary, reduce to the lowest ratio.

Example 10.3.1: Balancing Chemical Equations

Aqueous solutions of lead (II) nitrate and sodium chloride are mixed. The products of the reaction are an aqueous solution of sodium nitrate and a solid precipitate of lead (II) chloride. Write the balanced chemical equation for this reaction.

Solution

Step 1: Plan the problem.

Follow the steps for writing and balancing a chemical equation listed in the text.

Step 2: Solve.

Write the skeleton equation with the correct formulas.

$$\mathrm{Pb(NO_3)_2}\,(aq) + \mathrm{NaCl}\,(aq) \rightarrow \mathrm{NaNO_3}\,(aq) + \mathrm{PbCl_2}\,(s)$$

Count the number of each atom or polyatomic ion on both sides of the equation.

Reactants	Products
1 Pb atom	1 Pb atom
2 $\mathrm{NO_3}^-$ ions	1 $\mathrm{NO_3}^-$ ions
1 Na atom	1 Na atom
1 Cl atom	2 Cl atoms

The nitrate ions and the chlorine atoms are unbalanced. Start by placing a 2 in front of the NaCl. This increases the reactant counts to 2 Na atoms and 2 Cl atoms. Then place a 2 in front of the $\mathrm{NaNO_3}$. The result is:

$$\mathrm{Pb(NO_3)_2}\,(aq) + 2\mathrm{NaCl}\,(aq) \rightarrow 2\mathrm{NaNO_3}\,(aq) + \mathrm{PbCl_2}\,(s)$$

The new count for each atom and polyatomic ion becomes:

Reactants	Products
1 Pb atom	1 Pb atom
2 $\mathrm{NO_3}^-$ ions	2 $\mathrm{NO_3}^-$ ions
2 Na atom	2 Na atom
2 Cl atom	2 Cl atoms

Step 3: Think about your result.

The equation is now balanced since there are equal numbers of atoms of each element on both sides of the equation.

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Simulation

Practice balancing chemical equations with this simulation:

Can you balance a chemical equation?

Summary

- The process of balancing chemical equations is described.

Review

1. What is the law of conservation of mass?
2. How did Dalton describe the process of a chemical reaction?
3. Why don't we change the subscripts in order to balance an equation?

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10.4: Combination Reactions



Figure 10.4.1 (Public Domain; Rim: User:Relaxatiallc/Wikimedia Commons; Wheel: Christopher Ziemnowicz via Rim: http://commons.wikimedia.org/wiki/File:Ats_amgpenta.jpg; Wheel: http://commons.wikimedia.org/wiki/File:1975_AACA_AMC_Pacer_X_red-white_wheel.jpg)

How useful is a wheel rim?

A wheel rim is not very useful by itself. Driving on the rim can damage it and make for a very rough ride. When the rim is combined with a tire, the product can be put on a car and used for a safe and comfortable ride. The two separate items combine to make something that improves the car ride.

Combination Reactions

A **combination reaction** is a reaction in which two or more substances combine to form a single new substance. Combination reactions can also be called synthesis reactions. The general form of a combination reaction is:

$$\mathrm{A} + \mathrm{B} \rightarrow \mathrm{AB}$$

One combination reaction is two elements combining to form a compound. Solid sodium metal reacts with chlorine gas to produce solid sodium chloride.

$$2\mathrm{Na}\,(s) + \mathrm{Cl_2}\,(g) \rightarrow 2\mathrm{NaCl}\,(s)$$

In order to write and balance the equation correctly, it is important to remember the seven elements that exist in nature as diatomic molecules: $\mathrm{H_2}$, $\mathrm{N_2}$, $\mathrm{O_2}$, $\mathrm{F_2}$, $\mathrm{Cl_2}$, $\mathrm{Br_2}$, and $\mathrm{I_2}$.

One sort of combination reaction that occurs frequently is the reaction of an element with oxygen to form an oxide. Metals and nonmetals both react readily with oxygen under most conditions. Magnesium reacts rapidly and dramatically when ignited, combining with oxygen from the air to produce a fine powder of magnesium oxide.

$$2\mathrm{Mg}\,(s) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{MgO}\,(s)$$

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Sulfur reacts with oxygen to form sulfur dioxide.

$$\mathrm{S}\,(s) + \mathrm{O_2}\,(g) \rightarrow \mathrm{SO_2}\,(g)$$

When nonmetals react with one another, the product is a molecular compound. Often, the nonmetal reactants can combine in different ratios and produce different products. Sulfur can also combine with oxygen to form sulfur trioxide.

$$2\mathrm{S}\,(s) + 3\mathrm{O_2}\,(g) \rightarrow 2\mathrm{SO_3}\,(g)$$

Transition metals are capable of adopting multiple positive charges within their ionic compounds. Therefore, most transition metals are capable of forming different products in a combination reaction. Iron reacts with oxygen to form both iron (II) oxide and iron (III) oxide:

$$2\mathrm{Fe}\,(s) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{FeO}\,(s)$$

$$4\mathrm{Fe}\,(s) + 3\mathrm{O_2}\,(g) \rightarrow 2\mathrm{Fe_2O_3}\,(s)$$

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Example 10.4.1**: Combination Reactions**

Potassium is a very reactive alkali metal that must be stored under oil in order to prevent it from reacting with air. Write the balanced chemical equation for the combination reaction of potassium with oxygen.

Solution

Step 1: Plan the problem.

Make sure formulas of all reactants and products are correct before balancing the equation. Oxygen gas is a diatomic molecule. Potassium oxide is an ionic compound and so its formula is constructed by the crisscross method. Potassium as an ion becomes $\mathrm{K^+}$, while the oxide ion is $\mathrm{O^{2-}}$.

Step 2: Solve.

The skeleton (unbalanced) equation:

$$\mathrm{K}\,(s) + \mathrm{O_2}\,(g) \rightarrow \mathrm{K_2O}\,(s)$$

The equation is then easily balanced with coefficients.

$$4\mathrm{K}\,(s) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{K_2O}\,(s)$$

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Step 3: Think about your result.

Formulas are correct and the resulting combination reaction is balanced.

Exercise 10.4.1

Can you build a balanced, combination reaction for the formation of aluminum oxide (Al_2O_3)?

Answer available at the end of this section.

Combination reactions can also take place when an element reacts with a compound to form a new compound composed of a larger number of atoms. Carbon monoxide reacts with oxygen to form carbon dioxide, according to the equation:

$$2\mathrm{CO}\,(g) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{CO_2}\,(g)$$

Two compounds may also react to form a more complex compound. A very common example is the reaction of an oxide with water. Calcium oxide reacts readily with water to produce an aqueous solution of calcium hydroxide.

$$\mathrm{CaO}\,(s) + \mathrm{H_2O}\,(l) \rightarrow \mathrm{Ca(OH)_2}\,(aq)$$

Sulfur trioxide gas reacts with water to form sulfuric acid. This is an unfortunately common reaction that occurs in the atmosphere in some places where oxides of sulfur are present as pollutants. The acid formed in the reaction falls to the ground as acid rain.

$$\mathrm{SO_3}\,(g) + \mathrm{H_2O}\,(l) \rightarrow \mathrm{H_2SO_4}\,(aq)$$

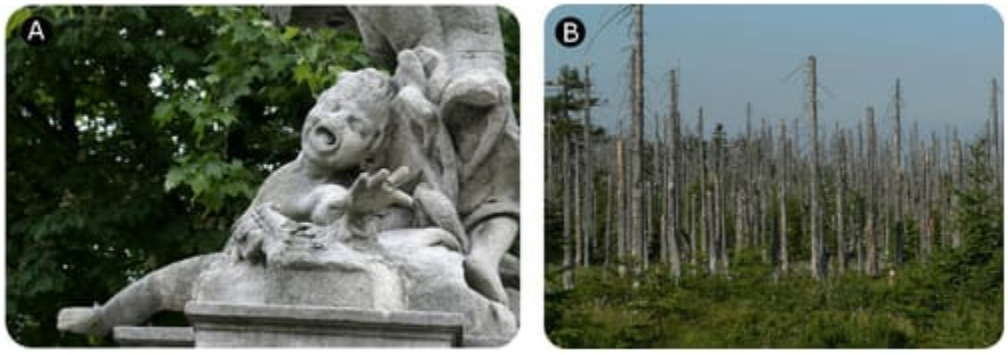


Figure 10.4.2: Acid rain has severe consequences on both natural and manmade objects. Acid rain degrades marble statues like the one on the left (A). The trees in the forest on the right (B) have been killed by acid rain. ((A; (A) Nipik; (B) Nino Barbieri via (A) http://commons.wikimedia.org/wiki/File:Acid_rain_woods1.JPG; (B) http://commons.wikimedia.org/wiki/File:Pollution_-_Damaged_by_acid_rain.jpg) Public Domain; (B) CC by 2.5)

Summary

- Combination reactions occur when two or more substances combine to form a new substance.

Review

1. What are combination reactions?
2. Write the product of the following reaction: $Mg + H_2O \rightarrow$
3. Is $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ a combination reaction? Explain your answer.

Answers

Exercise 10.4.1

$4Al + 3O_2 \rightarrow 2Al_2O_3$

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10.5: Decomposition Reactions

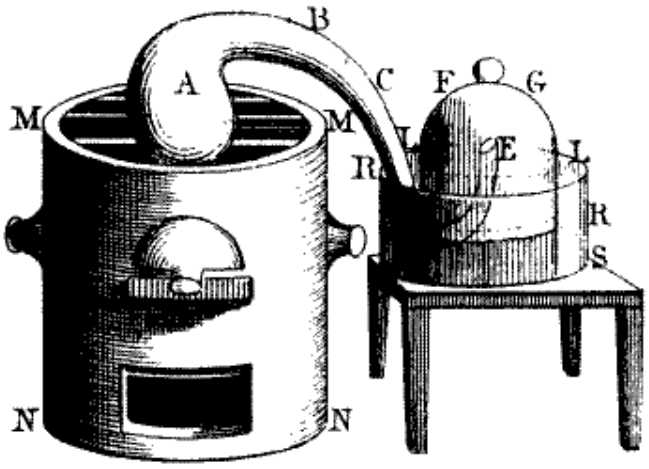


Figure 10.5.1 (Public Domain; Madame Lavoisier, modified by User: Cdang/Wikimedia Commons via Wikipedia)

How does a decomposition reaction work?

Antoine Lavoisier is widely known as the “father of modern chemistry.” He was one of the first to study chemical reactions in detail. Lavoisier reacted mercury with oxygen to form mercuric oxide as part of his studies on the composition of the atmosphere. He was then able to show that the decomposition of mercuric oxide produced mercury and oxygen. The diagram above shows the apparatus used by Lavoisier to study the formation and decomposition of mercuric oxide.

Decomposition Reactions

A **decomposition reaction** is a reaction in which a compound breaks down into two or more simpler substances. The general form of a decomposition reaction is:

$$\mathrm{AB} \rightarrow \mathrm{A} + \mathrm{B}$$

Most decomposition reactions require an input of energy in the form of heat, light, or electricity.

Binary compounds are compounds composed of just two elements. The simplest kind of decomposition reaction is when a binary compound decomposes into its elements. Mercury (II) oxide, a red solid, decomposes when heated to produce mercury and oxygen gas.

$$2\mathrm{HgO}\,(s) \rightarrow 2\mathrm{Hg}\,(l) + \mathrm{O_2}\,(g)$$

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Figure 10.5.2: *Mercury (II) oxide is a red solid. When it is heated, it decomposes into mercury metal and oxygen gas.* (Public Domain; Ben Mills (User:Benjah-bmm27/Wikimedia Commons) via Wikipedia)

A reaction is also considered to be a decomposition reaction even when one or more of the products are still compounds. A metal carbonate decomposes into a metal oxide and carbon dioxide gas. For example, calcium carbonate decomposes into calcium oxide and carbon dioxide:

$$\mathrm{CaCO_3}\,(s) \rightarrow \mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g)$$

Metal hydroxides decompose on heating to yield metal oxides and water. Sodium hydroxide decomposes to produce sodium oxide and water:

$$2\mathrm{NaOH}\,(s) \rightarrow \mathrm{Na_2O}\,(s) + \mathrm{H_2O}\,(g)$$

Some unstable acids decompose to produce nonmetal oxides and water. Carbonic acid decomposes easily at room temperature into carbon dioxide and water:

$$\mathrm{H_2CO_3}\,(aq) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

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Exercise 10.5.1

Can you build a balanced Chemical reaction for the decomposition of Potassium chlorate ($KClO_3$)?

Answer available at the end of this section.

Example 10.5.1: Decomposition Reactions

When an electric current is passed through pure water, it decomposes into its elements. Write a balanced equation for the decomposition of water.

Solution

Step 1: Plan the problem.

Water is a binary compound composed of hydrogen and oxygen. The hydrogen and oxygen gases produced in the reaction are both diatomic molecules.

Step 2: Solve.

The skeleton (unbalanced) equation:

$$\mathrm{H_2O}\,(l) \overset{\text{elec}}{\rightarrow} \mathrm{H_2}\,(g) + \mathrm{O_2}\,(g)$$

Note the abbreviation "elec" above the arrow to indicate the passage of an electric current to initiate the reaction. Balance the equation.

$$2\mathrm{H_2O}\,(l) \overset{\text{elec}}{\rightarrow} 2\mathrm{H_2}\,(g) + \mathrm{O_2}\,(g)$$

Step 3: Think about your result.

The products are elements and the equation is balanced.

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Simulations

How does an airbag explode at just the right time in the event of a car accident? Learn more in this simulation:

How does a chemical reaction inflate and airbag

Summary

- A definition of decomposition reaction and example reactions are given.

Review

1. What is a decomposition reaction?
2. What is usually needed for a decomposition reaction to take place?
3. Are elements always the product of a decomposition reaction?

Answers

Exercise 10.5.1

$2KClO_3 \rightarrow 2KCl + 3O_2$

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10.6: Combustion Reactions



Figure 10.6.1 (CC BY 2.0; Nina Hale via Wikipedia)

How do you cook the perfect marshmallow?

Roasting marshmallows over an open fire is a favorite past-time for campers, outdoor cook-outs, and just gathering around a fire in the back yard. The trick is to get the marshmallow a nice golden brown without catching it on fire. Too often we are not successful and we see the marshmallow burning on the stick – a combustion reaction taking place right in front of us.

Combustion Reactions

A **combustion reaction** is a reaction in which a substance reacts with oxygen gas, releasing energy in the form of light and heat. Combustion reactions must involve $\mathrm{O_2}$ as one reactant. The combustion of hydrogen gas produces water vapor:

$$\mathrm{2H_2}\,(g) + \mathrm{O_2}\,(g) \rightarrow \mathrm{2H_2O}\,(g)$$

Notice that this reaction also qualifies as a combination reaction.

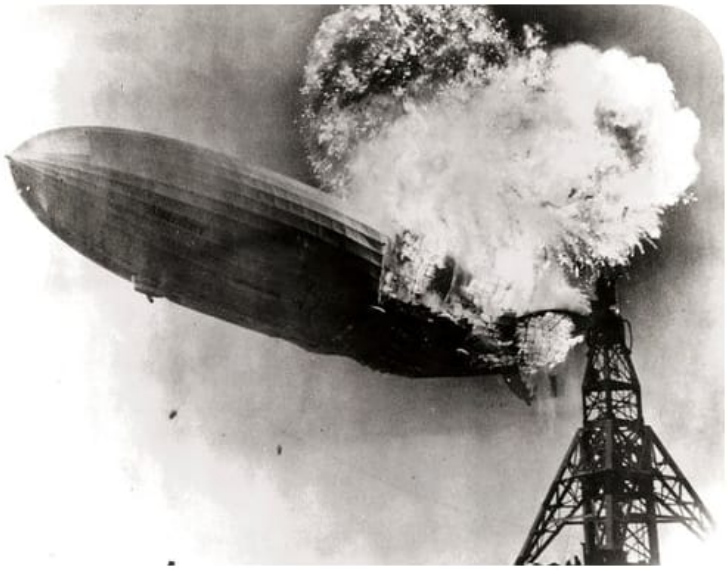


Figure 10.6.2: Explosion of the Hindenberg. (Public Domain; Gus Pasquerella/US Navy via Wikipedia)

The *Hindenberg* was a hydrogen-filled airship that suffered an accident upon its attempted landing in New Jersey in 1937. The hydrogen immediately combusted in a huge fireball, destroying the airship and killing 36 people. The chemical reaction was a simple one: hydrogen combining with oxygen to produce water.

Many combustion reactions occur with a hydrocarbon, a compound made up solely of carbon and hydrogen. The products of the combustion of hydrocarbons are carbon dioxide and water. Many hydrocarbons are used as fuel because their combustion releases very large amounts of heat energy. Propane $(\mathrm{C_3H_8})$ is a gaseous hydrocarbon that is commonly used as the fuel source in gas grills.

$$\mathrm{C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(g)}$$

Example 10.6.1: Combustion Reactions

Ethanol can be used as a fuel source in an alcohol lamp. The formula for ethanol is $\mathrm{C_2H_5OH}$. Write the balanced equation for the combustion of ethanol.

Solution

Step 1: Plan the problem.

Ethanol and oxygen are the reactants. As with a hydrocarbon, the products of the combustion of an alcohol are carbon dioxide and water.

Step 2: Solve.

Write the skeleton equation:

$$\mathrm{C_2H_5OH\,(l) + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(g)}$$

Balance the equation.

$$\mathrm{C_2H_5OH\,(l) + 3O_2\,(g) \rightarrow 2CO_2\,(g) + 3H_2O\,(g)}$$

Step 3: Think about your result.

Combustion reactions must have oxygen as a reactant. Note that the water produced is in the gas state, rather than the liquid state, because of the high temperatures that accompany a combustion reaction.

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Summary

- Combustion reaction is defined and examples are given.

Review

1. What is needed for a combustion reaction to take place?
2. What is formed in any combustion reaction?
3. Mercury reacts with oxygen to form mercuric oxide. Is this a combustion reaction?
4. What are the products of any combustion reaction involving a hydrocarbon?

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10.7: Single Replacement Reactions



Figure 10.7.1 (Pixabay License; wisteriax via Pixabay)

Why is the silver dark?

The platter and pitcher shown above provides an example of tarnish, a chemical reaction caued when silver metal reacts with hydrogen sulfide gas produced by some industrial processes or as a result of decaying animal or plant materials:

$$2\mathrm{Ag} + \mathrm{H_2S} \rightarrow \mathrm{Ag_2S} + \mathrm{H_2}$$

The tarnish can be removed using a number of polishes, but the process also removes a small amount of silver along with the tarnish.

Single-Replacement Reactions

A **single-replacement reaction** is a reaction in which one element replaces a similar element in a compound. The general form of a single-replacement (also called single-displacement) reaction is:

$$\mathrm{A} + \mathrm{BC} \rightarrow \mathrm{AC} + \mathrm{B}$$

In this general reaction, element A is a metal and replaces element B (also a metal) in the compound. When the element that is doing the replacing is a nonmetal, it must replace another nonmetal in a compound, and the general equation becomes:

$$\mathrm{Y} + \mathrm{XZ} \rightarrow \mathrm{XY} + \mathrm{Z}$$

(Where Y is a nonmetal and replaces the nonmetal Z in the compound with X.)

Metal Replacement

Magnesium is a more reactive metal than copper. When a strip of magnesium metal is placed in an aqueous solution of copper (II) nitrate, it replaces the copper. The products of the reaction are aqueous magnesium nitrate and solid copper metal.

$$\mathrm{Mg}\,(s) + \mathrm{Cu(NO_3)_2}\,(aq) \rightarrow \mathrm{Mg(NO_3)_2}\,(aq) + \mathrm{Cu}\,(s)$$

This subcategory of single-replacement reactions is called a metal replacement reaction because it is a metal that is being replaced (copper).

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Hydrogen Replacement

Many metals react easily with acids and when they do so, one of the products of the reaction is hydrogen gas. Zinc reacts with hydrochloric acid to produce aqueous zinc chloride and hydrogen (figure below).

$$\mathrm{Zn}\,(s) + 2\mathrm{HCl}\,(aq) \rightarrow \mathrm{ZnCl_2}\,(aq) + \mathrm{H_2}\,(g)$$

In a hydrogen replacement reaction, the hydrogen in the acid is replaced by an active metal.



Figure 10.7.2: Zinc metal reacts with hydrochloric acid to give off hydrogen gas in a single-displacement reaction. (Public Domain; User: Chemicalinterest/Wikimedia Commons via Wikipedia)

Some metals are so reactive that they are capable of replacing the hydrogen in water. The products of such a reaction are the metal hydroxide and hydrogen gas. All group 1 metals undergo this type of reaction. Sodium reacts vigorously with water to produce aqueous sodium hydroxide and hydrogen (see figure below).

$$2\mathrm{Na}\,(s) + 2\mathrm{H_2O}\,(l) \rightarrow 2\mathrm{NaOH}\,(aq) + \mathrm{H_2}\,(g)$$



Figure 10.7.3: Sodium metal reacts vigorously with water, giving off hydrogen gas. A large piece of sodium will often generate so much heat that the hydrogen will ignite. (Public Domain; User:Ajhalls/Wikimedia Commons via Wikipedia)

Halogen Replacement

The element chlorine reacts with an aqueous solution of sodium bromide to produce aqueous sodium chloride and elemental bromine:

$$\mathrm{Cl_2}\,(g) + 2\mathrm{NaBr}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{Br_2}\,(l)$$

The reactivity of the halogen group (group 17) decreases from top to bottom within the group. Fluorine is the most reactive halogen, while iodine is the least. Since chlorine is above bromine, it is more reactive than bromine and can replace it in a halogen replacement reaction.

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Summary

- The activity series describes the relative reactivities of metals and halogens.

Review

1. What is a metal replacement reaction?
2. Will a non-metal replace a metal?
3. What is the most reactive halogen?
4. What products will I get if I add potassium metal to water?

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10.8: Double Replacement Reactions



Figure 10.8.1 (Public Domain; F. S. Church, published in Harper's Weekly, January 17, 1874, p. 61 via Wikipedia)

Want to trade?

The practice of barter (trading one thing for another) has been in existence for a very long time. In the illustration above, items like chickens were bartered for newspapers. You have something I want, and I have something you want. So we trade and we each have something new. Some chemical reactions are like that. Compounds swap parts and you have new materials.

Double-Replacement Reactions

A **double-replacement reaction** is a reaction in which the positive and negative ions of two ionic compounds exchange places to form two new compounds. The general form of a double-replacement (also called double-displacement) reaction is:

$$\mathrm{AB} + \mathrm{CD} \rightarrow \mathrm{AD} + \mathrm{CB}$$

In this reaction, A and C are positively-charged cations, while B and D are negatively-charged anions. Double-replacement reactions generally occur between substances in aqueous solution. In order for a reaction to occur, one of the products is usually a solid precipitate, a gas, or a molecular compound such as water.

Formation of a Precipitate

A precipitate forms in a double-replacement reaction when the cations from one of the reactants combine with the anions from the other reactant to form an insoluble ionic compound. When aqueous solutions of potassium iodide and lead (II) nitrate are mixed, the following reaction occurs:

$$2\mathrm{KI}\,(aq) + \mathrm{Pb(NO_3)_2}\,(aq) \rightarrow 2\mathrm{KNO_3}\,(aq) + \mathrm{PbI_2}\,(s)$$

There are very strong attractive forces that occur between $\mathrm{Pb^{2+}}$ and $\mathrm{I^-}$ ions and the result is a brilliant yellow precipitate (see figure below). The other product of the reaction, potassium nitrate, remains soluble.



Figure 10.8.2: Formation of lead iodide precipitate. (CC-BY 2.0; Paige Powers - "Lead Iodide" via Flickr)

Formation of a Gas

Some double-replacement reactions produce a gaseous product which then bubbles out of the solution and escapes into the air. When solutions of sodium sulfide and hydrochloric acid are mixed, the products of the reaction are aqueous sodium chloride and hydrogen sulfide gas.

$$\mathrm{Na_2S}\,(aq) + 2\mathrm{HCl}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{H_2S}\,(g)$$

Formation of a Molecular Compound

Another kind of double-replacement reaction is one that produces a molecular compound as one of its products. Many examples in this category are reactions that produce water. When aqueous hydrochloric acid is reacted with aqueous sodium hydroxide, the products are aqueous sodium chloride and water:

$$\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$$

Example 10.8.1: Double-Replacement Reactions

Write a complete and balanced chemical equation for the following double-replacement reactions. One product is indicated as a guide.

a. $\mathrm{NaCN}\,(aq) + \mathrm{HBr}\,(aq) \rightarrow$ (hydrogen cyanide gas is formed)
b. $(\mathrm{NH_4})_2\mathrm{SO_4}\,(aq) + \mathrm{Ba(NO_3)_2}\,(aq) \rightarrow$ (a precipitate of barium sulfate forms)

Solution

Step 1: Plan the problem.

In A, the production of a gas drives the reaction. In B, the production of a precipitate drives the reaction. In both cases, use the ionic charges of both reactants to construct the correct formulas of the products.

Step 2: Solve.

A. The cations of both reactants are $+1$ charged ions, while the anions are -1 charged ions. After exchanging partners, the balanced equation is:

$$\mathrm{NaCN}\,(aq) + \mathrm{HBr}\,(aq) \rightarrow \mathrm{NaBr}\,(aq) + \mathrm{HCN}\,(g)$$

B. Ammonium ion and nitrate ion are $1+$ and $1-$ respectively, while barium and sulfate are $2+$ and $2-$. This must be taken into account when exchanging partners and writing the new formulas. Then, the equation is balanced.

$$(\mathrm{NH_4})_2\mathrm{SO_4}\,(aq) + \mathrm{Ba(NO_3)_2}\,(aq) \rightarrow 2\mathrm{NH_4NO_3}\,(aq) + \mathrm{BaSO_4}\,(s)$$

Step 3: Think about your result.

Both are double replacement reactions. All formulas are correct and the equations are balanced.

Occasionally, a reaction will produce both a gas and a molecular compound. The reaction of a sodium carbonate solution with

hydrochloric acid produces aqueous sodium chloride, carbon dioxide gas, and water.

$$\mathrm{Na_2CO_3}\,(aq) + 2\mathrm{HCl}\,(aq) \rightarrow 2\ \text{DEL}\ \mathrm{NaCl}\,(aq) + \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

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Summary

- The double-replacement reaction generally takes the form of AB + CD → AD + CB where A and C are positively-charged cations, while B and D are negatively-charged anions.
- In a double replacement reactions, typically one of the products is a precipitate, a gas, or a molecular compound.

Review

1. In a double-replacement reaction, what type of compounds are usually the reactants? A double-replacement reaction occurs between sodium sulfide and hydrogen chloride. Write the rest of the chemical equation and balance it.
2. Na_2S + HCl →
3. In double-replacement reaction, one of three possible types of products usually form. What are the three types?

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10.9: Molecular and Ionic Equations

One of the unfortunate byproducts of industrialized society is acid rain. Sulfur dioxide from burning coal and nitrogen oxides from vehicle emissions both form acids. When these acids react with limestone (calcium carbonate), reactions occur that dissolve the limestone and release water and carbon dioxide. Over a period of time, serious damage is caused to the limestone structure.

Molecular and Ionic Equations

When ionic compounds are dissolved into water, the polar water molecules break apart the solid crystal lattice, resulting in the hydrated ions being evenly distributed through the water. This process is called dissociation and is the reason that all ionic compounds are strong electrolytes. When two different ionic compounds that have been dissolved in water are mixed, a chemical reaction may occur between certain pairs of the hydrated ions. Consider the double-replacement reaction that occurs when a solution of sodium chloride is mixed with a solution of silver nitrate:

$$\mathrm{NaCl}\,(aq) + \mathrm{AgNO_3}\,(aq) \rightarrow \mathrm{NaNO_3}\,(aq) + \mathrm{AgCl}\,(s)$$

The driving force behind this reaction is the formation of the silver chloride precipitate (see figure below).

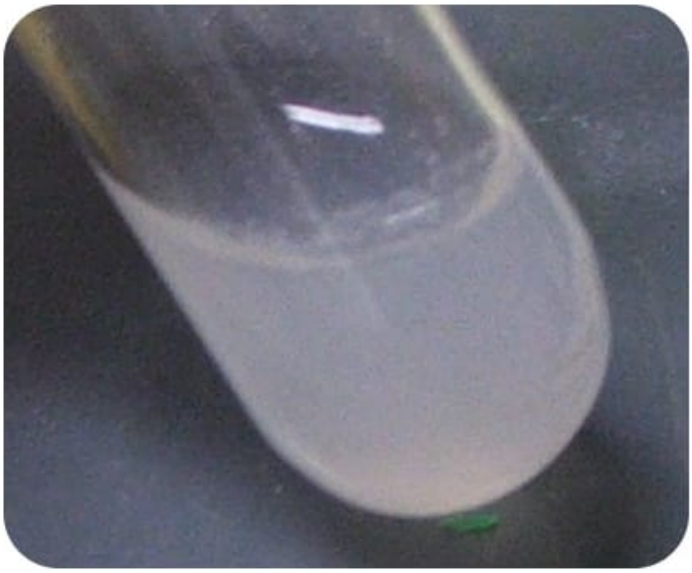


Figure 10.9.1: The white silver chloride precipitate instantly forms when a solution of silver nitrate is added to a solution of sodium chloride.

This is called a molecular equation. A **molecular equation** is an equation in which the formulas of the compounds are written as though all substances exist as molecules. However, there is a better way to show what is happening in this reaction. All of the aqueous compounds should be written as ions, because they are present in the water as separated ions—a result of their dissociation.

$$\mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq) + \mathrm{Ag^+}\,(aq) + \mathrm{NO_3{}^-}\,(aq) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{NO_3{}^-}\,(aq) + \mathrm{AgCl}\,(s)$$

This equation is called an **ionic equation**, an equation in which dissolved ionic compounds are shown as free ions.

Some other double-replacement reactions do not produce a precipitate as one of the products. The production of a gas and/or a molecular compound such as water may also drive the reaction. For example, consider the reaction of a solution of sodium carbonate with a solution of hydrochloric acid (HCl). The products of the reaction are aqueous sodium chloride, carbon dioxide, and water. The balanced molecular equation is:

$$\mathrm{Na_2CO_3}\,(aq) + 2\mathrm{HCl}\,(aq) \rightarrow 2\mathrm{NaCl}\,(aq) + \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

The ionic equation is:

$$2\mathrm{Na}^{+}\left(aq\right)+\mathrm{CO_3}^{2-}\left(aq\right)+2\mathrm{H}^{+}\left(aq\right)+2\mathrm{Cl}^{-}\left(aq\right)\rightarrow 2\mathrm{Na}^{+}\left(aq\right)+2\mathrm{Cl}^{-}\left(aq\right)+\mathrm{CO_2}\left(g\right)+\mathrm{H_2O}\left(l\right)$$

A single-replacement reaction is one in which an element replaces another element in a compound. An element is in either the solid, liquid, or gas state and is not an ion. The example below shows the reaction of solid magnesium metal with aqueous silver nitrate to form aqueous magnesium nitrate and silver metal.

Balanced molecular equation:

$$\mathrm{Mg}\left(s\right)+2\mathrm{AgNO_3}\left(aq\right)\rightarrow\mathrm{Mg(NO_3)_2}\left(aq\right)+2\mathrm{Ag}\left(s\right)$$

Ionic equation:

$$\mathrm{Mg}\left(s\right)+2\mathrm{Ag}^{+}\left(aq\right)+2\mathrm{NO_3}^{-}\left(aq\right)\rightarrow\mathrm{Mg}^{2+}\left(aq\right)+2\mathrm{NO_3}^{-}\left(aq\right)+\mathrm{Ag}\left(s\right)$$

This type of single-replacement reaction is called a metal replacement. Other common categories of single-replacement reactions are: hydrogen replacement and halogen replacement.

Summary

- A molecular equation is an equation in which the formulas of the compounds are written as though all substances exist as molecules.
- An ionic equation is one in which dissolved ionic compounds are shown as free ions.
- Examples of molecular and ionic equations are shown.

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10.10: Net Ionic Equations

At sports events around the world, a small number of athletes fiercely compete on fields and in stadiums. They get tired, dirty, and sometimes hurt as they try to win the game. Surrounding them are thousands of spectators watching and cheering. Would the game be different without the spectators? Definitely! Spectators provide encouragement to the team and generate enthusiasm. Although the spectators are not playing the game, they are certainly a part of the process.

Net Ionic Equations

We can write a molecular equation for the formation of silver chloride precipitate:

$$\mathrm{NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl}$$

The corresponding ionic equation is:

$$\mathrm{Na^+\,(aq) + Cl^-\,(aq) + Ag^+\,(aq) + NO_3{}^-\,(aq) \rightarrow Na^+\,(aq) + NO_3{}^-\,(aq) + AgCl\,(s)}$$

If you look carefully at the ionic equation, you will notice that the sodium ion and the nitrate ion appear unchanged on both sides of the equation. When the two solutions are mixed, neither the $\mathrm{Na^+}$ nor the $\mathrm{NO_3{}^-}$ ions participate in the reaction. They can be eliminated from the reaction.

$$\mathrm{\cancel{Na^+\,(aq)} + Cl^-\,(aq) + Ag^+\,(aq) + \cancel{NO_3{}^-\,(aq)} \rightarrow \cancel{Na^+\,(aq)} + \cancel{NO_3{}^-\,(aq)} + AgCl\,(s)}$$

A **spectator ion** is an ion that does not take part in the chemical reaction and is found in solution both before and after the reaction. In the above reaction, the sodium ion and the nitrate ion are both spectator ions. The equation can now be written without the spectator ions:

$$\mathrm{Ag^+\,(aq) + Cl^-\,(aq) \rightarrow AgCl\,(s)}$$

The **net ionic equation** is the chemical equation that shows only those elements, compounds, and ions that are directly involved in the chemical reaction. Notice that in writing the net ionic equation, the positively-charged silver cation was written first on the reactant side, followed by the negatively-charged chloride anion. This is somewhat customary because that is the order in which the ions must be written in the silver chloride product. However, it is not absolutely necessary to order the reactants in this way.

Net ionic equations must be balanced by both mass and charge. Balancing by mass means ensuring that there are equal masses of each element on the product and reactant sides. Balancing by charge means making sure that the overall charge is the same on both sides of the equation. In the above equation, the overall charge is zero, or neutral, on both sides of the equation. As a general rule, if you balance the molecular equation properly, the net ionic equation will end up being balanced by both mass and charge.

Example 10.10.1

When aqueous solutions of copper (II) chloride and potassium phosphate are mixed, a precipitate of copper (II) phosphate is formed. Write a balanced net ionic equation for this reaction.

Solution

Step 1: Plan the problem.

Write and balance the molecular equation first, making sure that all formulas are correct. Then write the ionic equation, showing all aqueous substances as ions. Carry through any coefficients. Finally, eliminate spectator ions and write the net ionic equation.

Step 2: Solve.

Molecular equation:

$$3\mathrm{CuCl_2}\,(aq) + 2\mathrm{K_3PO_4}\,(aq) \rightarrow 6\mathrm{KCl}\,(aq) + \mathrm{Cu_3(PO_4)_2}\,(s)$$

Ionic equation:

$$3[\blacksquare]\,(aq) + 6\mathrm{Cl^-}\,(aq) + 6\mathrm{K^+}\,(aq) + 2{\mathrm{PO_4}}^{3-}\,(aq) \rightarrow 6\mathrm{K^+}\,(aq) + 6\mathrm{Cl^-}\,(aq) + \mathrm{Cu_3(PO_4)_2}\,(s)$$

Notice that the balance of the equation is carried through when writing the dissociated ions. For example, there are six chloride ions on the reactant side because the coefficient of 3 is multiplied by the subscript of 2 on the copper (II) chloride formula. The spectator ions, $\mathrm{K^+}$ and $\mathrm{Cl^-}$, can be eliminated.

Net ionic equation:

$$3[\blacksquare]\,(aq) + 2{\mathrm{PO_4}}^{3-}\,(aq) \rightarrow \mathrm{Cu_3(PO_4)_2}\,(s)$$

Step 3: Think about your result.

For a precipitation reaction, the net ionic equation always shows the two ions that come together to form the precipitate. The equation is balanced by mass and charge.

Summary

- A spectator ion is an ion that does not take part in the chemical reaction and is found in solution both before and after the reaction.
- The net ionic equation is the chemical equation that shows only those elements, compounds, and ions that are directly involved in the chemical reaction.
- An example of writing a net ionic equation is outlined.

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10.11: Predicting Precipitates Using Solubility Rules

Predicting the weather is tricky business. A thorough examination of a large amount of data is needed to make the daily forecast. Wind patterns, historical data, barometric pressure—these and many other data are fed into computers that then use a set of rules to predict what will happen based on past history.

Solubility Rules

Some combinations of aqueous reactants result in the formation of a solid precipitate as a product. However, some combinations will not produce such a product. If solutions of sodium nitrate and ammonium chloride are mixed, no reaction occurs. One could write a molecular equation showing a double-replacement reaction, but both products, sodium chloride and ammonium nitrate, are soluble and would remain in the solution as ions. Every ion is a spectator ion and there is no net ionic equation at all.

It is useful to be able to predict when a precipitate will occur in a reaction. To do so, you can use a set of guidelines called the **solubility rules** (shown in Table 10.11.1).

Table 10.11.1: Solubility Rules for Ionic Compounds in Water

Soluble	Compounds containing the alkali metal ions (Li^+, Na^+, K^+, Rb^+, Cs^+) and ammonium ion (NH_4^+).
Soluble	Compounds containing the nitrate ion (NO_3^-), acetate ion (CH_3COO^-), chlorate ion (ClO_3^-), and bicarbonate ion (HCO_3^-).
Mostly Soluble	Compounds containing the chloride ion (Cl^-), bromide ion (Br^-), and iodide ion (I^-). Exceptions are those of silver (Ag^+), mercury (I) (Hg_2^{2+}), and lead (II) ([redacted]).
Mostly Soluble	Compounds containing the sulfate ion (SO_4^{2-}). Exceptions are those of silver (Ag^+), calcium ([redacted]), strontium ([redacted]), barium ([redacted]), mercury (I) (Hg_2^{2+}), and lead (II) ([redacted]).
Mostly Insoluble	Compounds containing the carbonate ion (CO_3^{2-}), phosphate ion (PO_4^{3-}), chromate ion (CrO_4^{2-}), sulfide ion ([redacted]), and silicate ion (SiO_3^{2-}). Exceptions are those of the alkali metals and ammonium.
Mostly Insoluble	Compounds containing the hydroxide ion (OH^-). Exceptions are those of the alkali metals and the barium ion ([redacted]).

For practice using the solubility rules, predict if a precipitate will form when solutions of cesium bromide and lead (II) nitrate are mixed.

$$Cs^+(aq) + Br^-(aq) + [redacted](aq) + 2NO_3^-(aq) \rightarrow ?$$

The potential precipitates from a double-replacement reaction are cesium nitrate and lead (II) bromide. According to the solubility rules table, cesium nitrate is soluble because all compounds containing the nitrate ion, as well as all compounds containing the alkali metal ions, are soluble. Most compounds containing the bromide ion are soluble, but lead (II) is an exception. Therefore, the cesium and nitrate ions are spectator ions and the lead (II) bromide is a precipitate. The balanced net ionic reaction is:

$$[redacted](aq) + 2Br^-(aq) \rightarrow PbBr_2(s)$$

Summary

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10.12: Redox Reactions and Ionic Compounds

Nitric acid has many uses in the manufacture of fertilizers and explosives. Most nitric acid is manufactured from ammonia using a three-step process. The ammonia is oxidized to HNO_3 through the formation of several nitrogen oxides, finally resulting in the acid.

Redox Reactions and Ionic Compounds

In the course of a chemical reaction between a metal and a nonmetal, electrons are transferred from the metal atoms to the nonmetal atoms. For example, when zinc metal is mixed with sulfur and heated, the compound zinc sulfide is produced. Two valence electrons from each zinc atom are transferred to each sulfur atom.

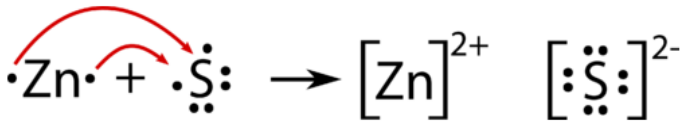


Figure 10.12.1: Reaction between zinc and sulfur.

Since the zinc is losing electrons in the reaction, it is being oxidized. The sulfur is gaining electrons and is thus being reduced. An **oxidation-reduction reaction** is a reaction that involves the full or partial transfer of electrons from one reactant to another. **Oxidation** is the full or partial loss of electrons or the gain of oxygen. **Reduction** is the full or partial gain of electrons or the loss of oxygen. A **redox reaction** is another term for an oxidation-reduction reaction.

Each of these processes can be shown in a separate equation called a half-reaction. A **half-reaction** is an equation that shows either the oxidation or the reduction reaction that occurs during a redox reaction.

$$\text{Oxidation: } Zn \rightarrow \blacksquare + 2e^-$$
$$\text{Reduction: } S + 2e^- \rightarrow \blacksquare$$

It is important to remember that the two half-reactions occur simultaneously. The resulting ions that are formed are then attracted to one another in an ionic bond.

Another example of an oxidation-reduction reaction involving electron transfer is the well-known combination of metallic sodium and chlorine gas to form sodium chloride:

$$2Na + Cl_2 \rightarrow 2NaCl$$

The half reactions are as follows:

$$\text{Oxidation: } Na \rightarrow Na^+ + e^-$$
$$\text{Reduction: } Cl + e^- \rightarrow Cl^-$$

We will concern ourselves with the balancing of these equations at another time.

Summary

- An oxidation-reduction reaction, or redox reaction, is a reaction that involves the full or partial transfer of electrons from one reactant to another.
- Oxidation is the full or partial loss of electrons or the gain of oxygen.
- Reduction is the full or partial gain of electrons or the loss of oxygen.
- A half-reaction is an equation that shows either the oxidation or the reduction reaction that occurs during a redox reaction.
- Examples of oxidation-reduction reactions are shown.

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10.13: Identifying Reaction Types

The reaction of copper wire with nitric acid produces a colorful mix of products that include copper (II) nitrate, nitrogen dioxide, and water. Copper salts are blue in solution, reflecting the rather unique arrangements of electrons in the d orbital as the copper ionizes from metallic copper.

Identifying Reaction Types

A redox reaction must involve a change in oxidation number for two of the elements involved in the reaction. The oxidized element increases in oxidation number, while the reduced element decreases in oxidation number.

Single-replacement reactions are redox reactions because two different elements appear as free elements (oxidation number of zero) on one side of the equation and as part of a compound on the other side. Therefore, oxidation numbers must change.

$$\mathrm{Zn} + 2\mathrm{HCl} \rightarrow \mathrm{ZnCl_2} + \mathrm{H_2}$$

Zn is oxidized from $\mathrm{Zn^0}$ to $\mathrm{Zn^{2+}}$ and the H is reduced from $\mathrm{H^+}$ to $\mathrm{H^0}$.

Combustion reactions are redox reactions because elemental oxygen $(\mathrm{O_2})$ acts as the oxidizing agent and is itself reduced.

$$\mathrm{CH_4} + 2\mathrm{O_2} \rightarrow \mathrm{CO_2} + 2\mathrm{H_2O}$$

Most combination and decomposition reactions are redox reactions, since elements are usually transformed into compounds and vice-versa. The thermite reaction involves ferric oxide and metallic aluminum:

$$\mathrm{Fe_2O_3} + 2\mathrm{Al} \rightarrow \mathrm{Al_2O_3} + 2\mathrm{Fe}$$



Figure 10.13.1: *Thermite grenade demonstration.*

We see that the iron is reduced and the aluminum oxidized during the course of the reaction.

So what types of reactions are not redox reactions? Double-replacement reactions, such as the one below, are not redox reactions because ions are simply recombined without any transfer of electrons.

$$\overset{+1\ +6}{\mathrm{Na_2\,S}}\blacksquare\mathrm{O_4}\,(aq) + \overset{+2\ +5\ -2}{\mathrm{Ba(NO_3)}}\,(aq) \rightarrow 2\overset{+1\ +5\ -2}{\mathrm{NaNO_3}}\,(aq) + \overset{+2\ +6\ -2}{\mathrm{BaSO_4}}\,(s)$$

Note that the oxidation numbers for each element remain unchanged in the reaction.

Acid-base reactions involve a transfer of a hydrogen ion instead of an electron. Acid-base reactions, like the one below, are also not redox reactions.

$$\overset{+1-1}{\mathrm{H\,F}}\,(aq) + \overset{-3\,+1}{\mathrm{NH_3}}\,(aq) \rightarrow \overset{-3\ \ +1}{\mathrm{NH_4}^+}\,(aq) + \overset{-1}{\mathrm{F}^-}$$

Again, the transfer of an H^+ ion leaves the oxidation numbers unaffected. In summary, redox reactions can always be recognized by a change in oxidation number of two of the atoms in the reaction. Any reaction in which no oxidation numbers change is not a redox reaction.

Summary

- A redox reaction must involve a change in oxidation number for two of the elements involved in the reaction.
- The oxidized element increases in oxidation number, while the reduced element decreases in oxidation number.
- Single-replacement reactions and combustion reactions **are** redox reactions.
- **Most** combination and decomposition reactions are redox reactions.
- Double-replacement reactions and acid-base reactions **are not** redox reactions.

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11: THE MOLE

The mole is the unit of measurement in the International System of Units (SI) for amount of substance. It is defined as the amount of a chemical substance that contains as many elementary entities (e.g., atoms, molecules, ions, electrons, or photons). This number is expressed by the Avogadro constant, which has a value of $6.022140857 \times 10^{23} \ mol^{-1}$. The mole is one of the base units of the SI, and has the unit symbol mol.

11.1: Avogadro's Number

11.2: Conversions Between Moles and Atoms

11.3: Molar Mass

11.4: Conversions Between Moles and Mass

11.5: Conversions Between Mass and Number of Particles

11.6: Percent Composition

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11.1: Avogadro's Number



Figure 11.1.1 (Public Domain; User:OSX/Wikimedia Commons via Wikipedia)

Is there an easier way to load this truck?

When the weather is nice, many people begin to work on their yards and homes. For many projects, sand is needed as a foundation for a walk or to add to other materials. You could order up twenty million grains of sand and have people really stare at you. You could order by the pound, but that takes a lot of time weighing out. The best bet is to order by the yard, meaning a cubic yard. The loader can easily scoop up what you need and put it directly in your truck.

Avogadro's Number

It certainly is easy to count objects such as bananas, or something as large as elephants (as long as you stay out of their way). However, counting grains of sugar from a sugar canister would take a long, long time. Atoms and molecules are extremely small—far, far smaller than grains of sugar. Counting atoms or molecules is not only unwise, it is absolutely impossible. One drop of water contains about 10^{22} molecules of water. If you counted 10 molecules every second for 50 years, without stopping, you would have counted only 1.6×10^{10} molecules. Put another way, at that counting rate, it would take you over 30 trillion years to count the water molecules in one tiny drop.

Chemists of the past needed a name that could stand for a very large number of items. Amadeo Avogadro (1776-1856), an Italian scientist, provided such a number. He is responsible for the counting unit of measure called the mole. A **mole** (mol) is the amount of a substance that contains 6.02×10^{23} representative particles of that substance. The mole is the SI unit for amount of a substance. Just like the dozen and the gross, it is a name that stands for a number. There are therefore 6.02×10^{23} water molecules in a mole of water molecules. There also would be 6.02×10^{23} bananas in a mole of bananas, if such a huge number of bananas ever existed.



Figure 11.1.2: Italian scientist Amadeo Avogadro, whose work led to the concept of the mole as a counting unit in chemistry. (Public Domain; C. Sentier via Wikipedia)

The number 6.02×10^{23} is called **Avogadro's number**, the number of representative particles in a mole. It is an experimentally determined number. A **representative particle** is the smallest unit in which a substance naturally exists. For the majority of elements, the representative particle is the atom. Iron, carbon, and helium consist of iron atoms, carbon atoms, and helium atoms, respectively. Seven elements exist in nature as diatomic molecules and they are H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. The representative particle for these elements is the molecule. Likewise, all molecular compounds such as H_2O and CO_2 exist as molecules and so the molecule is their representative particle. For ionic compounds such as $NaCl$ and $Ca(NO_3)_2$, the representative particle is the formula unit. A mole of any substance contains Avogadro's number $\left(6.02 \times 10^{23}\right)$ of representative particles.



Figure 11.1.3: The animal mole is very different than the counting unit of the mole. Chemists nonetheless have adopted the mole as their unofficial mascot. National Mole Day is a celebration of chemistry that occurs on October 23rd (10/23) of each year. ((left; Left: Michael David Hill, 2005 (Mikiwikipikidikipedia); Right: chrisbb@prodigy.net;(left) Michael David Hill, 2005 (Mikiwikipikidikipedia); (right) chrisbb@prodigy.net via Left: http://commons.wikimedia.org/wiki/File:Close-up_of_mole.jpg; Right: http://www.flickr.com/photos/chrisbrenschmidt/436990097;(left) http://commons.wikimedia.org/wiki/File:Close-up_of_mole.jpg; (right) http://www.flickr.com/photos/chrisbrenschmidt/436990097) CC-BY 2.5; (right) CC-BY 2.0)

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Summary

- A mole of any substance contains Avogadro's number $\left(6.02 \times 10^{23}\right)$ of representative particles.

Review

1. What is the SI unit for amount of a substance?
2. What is the representative particle for an element?
3. The formula unit is the representative particle for what?

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11.2: Conversions Between Moles and Atoms



Figure 11.2.1 (Public Domain; Diane A. Reid/National Cancer Institute via Wikipedia)

Big numbers or little numbers?

Do you hate to type subscripts and superscripts? Even with a good word-processing program, having to click on an icon to get a superscript and then remembering to click off after you type the number can be a real hassle. If we did not know about moles and just knew about numbers of atoms or molecules (those big numbers that require lots of superscripts), life would be much more complicated and we would make many more typing errors.

Conversions Between Moles and Atoms

Conversions Between Moles and Number of Particles

Using our unit conversion techniques, we can use the mole label to convert back and forth between the number of particles and moles.

Example 11.2.1: Converting Number of Particles to Moles

The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?

Solution

Step 1: List the known quantities and plan the problem.

Known

- number of C atoms $= 4.72 \times 10^{24}$
- 1 mole $= 6.02 \times 10^{23}$ atoms

Unknown

- 4.72 x 10^{24} = ? mol C

One conversion factor will allow us to convert from the number of C atoms to moles of C atoms.

Step 2: Calculate.

$$4.72 \times 10^{24} \text{ atoms C} \times \frac{1 \text{ mol C}}{6.02 \times 10^{23} \text{ atoms C}} = 7.84 \text{ mol C}$$

Step 3: Think about your result.

The given number of carbon atoms was greater than Avogadro's number, so the number of moles of C atoms is greater than 1 mole. Since Avogadro's number is a measured quantity with three significant figures, the result of the calculation is rounded to three significant figures.

Suppose that you want to know how many hydrogen atoms are in a mole of water molecules. First, you need to know the chemical formula for water, which is $\mathrm{H_2O}$. There are two atoms of hydrogen in each molecule of water. How many atoms of hydrogen are in two water molecules? There are $2 \times 2 = 4$ hydrogen atoms. How about in a dozen? In that case, a dozen is 12; so $12 \times 2 = 24$ hydrogen atoms in a dozen water molecules. To get the answers (4 and 24), you multiply the given number of molecules by two atoms of hydrogen per molecule. So, to find the number of hydrogen atoms in a mole of water molecules, the problem can be solved using conversion factors:

$$1 \text{ mol } \mathrm{H_2O} \times \frac{6.02 \times 10^{23} \text{ molecules } \mathrm{H_2O}}{1 \text{ mol } \mathrm{H_2O}} \times \frac{2 \text{ atoms H}}{1 \text{ molecule } \mathrm{H_2O}} = 1.20 \times 10^{24} \text{ atoms H}$$

The first conversion factor converts from moles of particles to the number of particles. The second conversion factor reflects the number of atoms contained within each molecule.

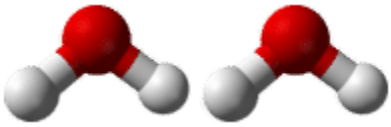


Figure 11.2.2: Two water molecules contain 4 hydrogen atoms and 2 oxygen atoms. A mole of water molecules contains 2 moles of hydrogen atoms and 1 mole of oxygen atoms. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

Example 11.2.2: Atoms, Molecules, and Moles

Sulfuric acid has the chemical formula $\mathrm{H_2SO_4}$. A certain quantity of sulfuric acid contains 4.89×10^{25} atoms of oxygen. How many moles of sulfuric acid is the sample?

Solution:

Step 1: List the known quantities and plan the problem.

Known

- $4.89 \times 10^{25} = \mathrm{O}$ atoms
- 1 mole $= 6.02 \times 10^{23}$ molecules $\mathrm{H_2SO_4}$

Unknown

- mol of $\mathrm{H_2SO_4}$ molecules

Two conversion factors will be used. First, convert atoms of oxygen to molecules of sulfuric acid. Then, convert molecules of sulfuric acid to moles of sulfuric acid.

Step 2: Calculate.

$$4.89 \times 10^{25} \text{ atoms O} \times \frac{1 \text{ molecule } \mathrm{H_2SO_4}}{4 \text{ atoms O}} \times \frac{1 \text{ mol } \mathrm{H_2SO_4}}{6.02 \times 10^{23} \text{ molecules } \mathrm{H_2SO_4}} = 20.3 \text{ mol } \mathrm{H_2SO_4}$$

Step 3: Think about your result.

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The original number of oxygen atoms was about 80 times larger than Avogadro's number. Since each sulfuric acid molecule contains 4 oxygen atoms, there are about 20 moles of sulfuric acid molecules.

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Summary

- Methods are described for conversions between moles, atoms, and molecules.

Review

1. What conversion factor would we need to convert moles of helium to atoms of helium?
2. I want to convert atoms to moles. My friend tells me to multiply the number of atoms by 6.02×10^{23} atoms/mole. Is my friend correct?
3. Why do you need to know the formula for a molecule in order to calculate the number of moles of one of the atoms?
4. How many atoms of fluorine are in 5.6×10^{22} molecules of MgF_2?

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11.3: Molar Mass



Figure 11.3.1 (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

When creating a solution, how do I know how much of each substance to put in?

I want to make a solution that contains 1.8 moles of potassium dichromate. You do not have a balance calibrated in moles, but you do have one calibrated in grams. If you know the relationship between moles and the number of grams in a mole, you can use your balance to measure out the needed amount of material.

Molar Mass

Molar mass is defined as the mass of one mole of representative particles of a substance. By looking at a periodic table, we can conclude that the molar mass of lithium is $6.94\ \text{g}$, the molar mass of zinc is $65.38\ \text{g}$, and the molar mass of gold is $196.97\ \text{g}$. Each of these quantities contains 6.02×10^{23} atoms of that particular element. The units for molar mass are grams per mole, or g/mol.

Molar Masses of Compounds

The molecular formula of the compound carbon dioxide is CO_2. One molecule of carbon dioxide consists of 1 atom of carbon and 2 atoms of oxygen. We can calculate the mass of one molecule of carbon dioxide by adding together the masses of 1 atom of carbon and 2 atoms of oxygen:

$$12.01\ \text{amu} + 2\,(16.00\ \text{amu}) = 44.01\ \text{amu}$$

The **molecular mass** of a compound is the mass of one molecule of that compound. The molecular mass of carbon dioxide is $44.01\ \text{amu}$.

The molar mass of any compound is the mass in grams of one mole of that compound. One mole of carbon dioxide molecules has a mass of $44.01\ \text{g}$, while one mole of sodium sulfide formula units has a mass of $78.04\ \text{g}$. The molar masses are $44.01\ \text{g/mol}$ and $78.04\ \text{g/mol}$ respectively. In both cases, that is the mass of 6.02×10^{23} representative particles. The representative particle of CO_2 is the molecule, while for Na_2S it is the formula unit.

Example 11.3.1: Molar Mass of a Compound

Calcium nitrate, $\text{Ca(NO}_3)_2$, is used as a component in fertilizer. Determine the molar mass of calcium nitrate.

Solution

Step 1: List the known and unknown quantities and plan the problem.

Known

- Formula $= \mathrm{Ca(NO_3)_2}$
- Molar mass $\mathrm{Ca} = 40.08\ \mathrm{g/mol}$
- Molar mass $\mathrm{N} = 14.01\ \mathrm{g/mol}$
- Molar mass $\mathrm{O} = 16.00\ \mathrm{g/mol}$

Unknown

- molar mass $Ca(NO_3)_2$

First we need to analyze the formula. Since the Ca lacks a subscript, there is one Ca atom per formula unit. The 2 outside the parentheses means that there are two nitrate ions per formula unit and each nitrate ion consists of one nitrogen atom and three oxygen atoms per formula unit. Thus, $1\ \mathrm{mol}$ of calcium nitrate contains $1\ \mathrm{mol}$ of Ca atoms, $2\ \mathrm{mol}$ of N atoms, and $6\ \mathrm{mol}$ of O atoms.

Step 2: Calculate

Use the molar masses of each atom together with the number of atoms in the formula and add together.

$$1\ \mathrm{mol\ Ca} \times \frac{40.08\ \mathrm{g\ Ca}}{1\ \mathrm{mol\ Ca}} = 40.08\ \mathrm{g\ Ca}$$

$$2\ \mathrm{mol\ N} \times \frac{14.01\ \mathrm{g\ N}}{1\ \mathrm{mol\ N}} = 28.02\ \mathrm{g\ N}$$

$$6\ \mathrm{mol\ O} \times \frac{16.00\ \mathrm{g\ O}}{1\ \mathrm{mol\ O}} = 96.00\ \mathrm{g\ O}$$

$$\text{Molar mass of } \mathrm{Ca(NO_3)_2} = 40.08\ \mathrm{g} + 28.02\ \mathrm{g} + 96.00\ \mathrm{g} = 164.10\ \mathrm{g/mol}$$

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Summary

- Calculations are described for the determination of molar mass of an atom or a compound.

Review

1. What is the molar mass of Pb?
2. Where do you find the molar mass of an element?
3. How many moles of Cl are in one mole of the $CaCl_2$?
4. How many moles of H are in one mole of the compound $(NH_4)_3PO_4$?
5. Calculate the molar mass of $CaCl_2$.

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11.4: Conversions Between Moles and Mass



Figure 11.4.1 (Public Domain; User:Estormiz/Wikimedia Commons via Wikipedia)

How can we get more product?

Chemical manufacturing plants are always seeking to improve their processes. One way that improvement comes about is through measuring the amount of material produced in a reaction. By knowing how much is made, the scientists and engineers can try different ways of getting more product at less cost.

Conversions Between Moles and Mass

The molar mass of any substance is the mass in grams of one mole of representative particles of that substance. The representative particles can be atoms, molecules, or formula units of ionic compounds. This relationship is frequently used in the laboratory. Suppose that for a certain experiment, you need 3.00 moles of calcium chloride $(\mathrm{CaCl_2})$. Since calcium chloride is a solid, it would be convenient to use a balance to measure the mass that is needed. The molar mass of $\mathrm{CaCl_2}$ is $110.98\ \mathrm{g/mol}$. The conversion factor that can be used is then based on the equality that $1\ \mathrm{mol} = 110.98\ \mathrm{g}\ \mathrm{CaCl_2}$. Dimensional analysis will allow you to calculate the mass of $\mathrm{CaCl_2}$ that you should measure.

$$3.00\ \mathrm{mol\ CaCl_2} \times \frac{110.98\ \mathrm{g\ CaCl_2}}{1\ \mathrm{mol\ CaCl_2}} = 333\ \mathrm{g\ CaCl_2}$$

When you measure the mass of $333\ \mathrm{g}$ of $\mathrm{CaCl_2}$, you are measuring 3.00 moles of $\mathrm{CaCl_2}$.



Figure 11.4.2: Calcium chloride is used as a drying agent and as a road deicer. (Public Domain; Martin Walker (User:Walkerma/Wikimedia Commons) via Wikipedia)

Example 11.4.1: Converting Moles to Mass

Chromium metal is used for decorative electroplating of car bumpers and other surfaces. Find the mass of 0.560 moles of chromium.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Molar mass of $\mathrm{Cr} = 52.00\ \mathrm{g/mol}$
- $0.560\ \mathrm{mol\ Cr}$

Unknown

- $0.560\ \mathrm{mol\ Cr}$ = ? g

One conversion factor will allow us to convert from the moles of Cr to mass.

Step 2: Calculate.

$$0.560\ \mathrm{mol\ Cr} \times \frac{52.00\ \mathrm{g\ Cr}}{1\ \mathrm{mol\ Cr}} = 29.1\ \mathrm{g\ Cr}$$

Step 3: Think about your result.

Since the desired amount was slightly more than one half of a mole, the mass should be slightly more than one half of the molar mass. The answer has three significant figures because of the $0.560\ \mathrm{mol}$.

A similar conversion factor utilizing molar mass can be used to convert from the mass of a substance to moles. In a laboratory situation, you may perform a reaction and produce a certain amount of a product which can be massed. It will often then be necessary to determine the number of moles of the product that was formed. The next problem illustrates this situation.

Example 11.4.2: Converting Mass to Moles

A certain reaction produces $2.81\ \mathrm{g}$ of copper (II) hydroxide, $\mathrm{Cu(OH)_2}$. Determine the number of moles produced in the reaction.

Step 1: List the known quantities and plan the problem.

Known

- mass = 2.81g

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Unknown

- mol $Cu(OH)_2$

One conversion factor will allow us to convert from mass to moles.

Step 2: Calculate.

First, it is necessary to calculate the molar mass of $\mathrm{Cu(OH)_2}$ from the molar masses of Cu, O, and H. The molar mass is $97.57\ \mathrm{g/mol}$.

$$2.81\ \mathrm{g\ Cu(OH)_2} \times \frac{1\ \mathrm{mol\ Cu(OH)_2}}{97.57\ \mathrm{g\ Cu(OH)_2}} = 0.0288\ \mathrm{mol\ Cu(OH)_2}$$

Step 3: Think about your result.

The relatively small mass of product formed results in a small number of moles.

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Simulation

Practice converting mass to moles in everyday objects in this simulation:

How many atoms are in a gold ring, a grain of sand, or the human body?

Summary

- Calculations involving conversions between moles of a material and the mass of that material are described.

Review

1. You have 19.7 grams of a material and wonder how many moles were formed. Your friend tells you to multiply the mass by grams/mole. Is your friend correct?
2. How many grams of MgO are in 3.500 moles?
3. How many moles of H_2O are in 15.2 grams of pure ice?

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11.5: Conversions Between Mass and Number of Particles

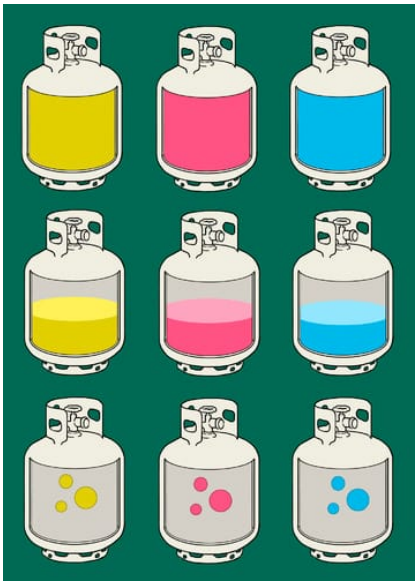


Figure 11.5.1 (CC BY-NC 3.0; Laura Guerin via CK-12 Foundation)

How much gas is there?

Avogadro was interested in studying gases. He theorized that equal volumes of gases under the same conditions contained the same number of particles. Other researchers studied how many gas particles were in specific volumes of gases. Eventually, scientists were able to develop the relationship between number of particles and mass, using the idea of moles.

Conversions Between Mass and Number of Particles

In "Conversions Between Moles and Mass", you learned how to convert back and forth between moles and the number of representative particles. Now you have seen how to convert back and forth between moles and the mass of a substance in grams. We can combine the two types of problems into one. Mass and number of particles are both related to grams. In order to convert from mass to number of particles or vice-versa, a conversion to moles is required.



Figure 11.5.2: Conversion from number of particles to mass, or from mass to number of particles requires two steps. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Example 11.5.1: Converting Mass to Particles

How many molecules is $20.0\ \text{g}$ of chlorine gas, Cl_2?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Molar mass $\mathrm{Cl_2} = 70.90\ \mathrm{g/mol}$
- $20.0\ \mathrm{g}\ \mathrm{Cl_2}$

Unknown

- number of molecules of Cl_2

Use two conversion factors. The first converts grams of $\mathrm{Cl_2}$ to moles. The second converts moles of $\mathrm{Cl_2}$ to the number of molecules.

Step 2: Calculate.

$$20.0\ \mathrm{g\ Cl_2} \times \frac{1\ \mathrm{mol\ Cl_2}}{70.90\ \mathrm{g\ Cl_2}} \times \frac{6.02 \times 10^{23}\ \mathrm{molecules\ Cl_2}}{1\ \mathrm{mol\ Cl_2}} = 1.70 \times 10^{23}\ \mathrm{molecules\ Cl_2}$$

The problem is done using two consecutive conversion factors. There is no need to explicitly calculate the moles of $\mathrm{Cl_2}$.

Step 3: Think about your result.

Since the given mass is less than half of the molar mass of chlorine, the resulting number of molecules is less than half of Avogadro's number.

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Summary

- Calculations are illustrated for conversions between mass and number of particles.

Review

1. Why can't we convert directly from number of particles to grams?
2. How many atoms of chlorine are present in 1.70×10^{23} molecules Cl_2?
3. How many molecules of BH_3 are in 14.32 grams BH_3?

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11.6: Percent Composition



Figure 11.6.1 (CC BY-NC 3.0; Joy Sheng via CK-12 Foundation)

Is there anything healthy in this jar?

Packaged foods that you eat typically have nutritional information provided on the label. The label on a jar of peanut butter (shown above) reveals that one serving size is considered to be $32\ \text{g}$. The label also gives the masses of various types of compounds that are present in each serving. One serving contains $7\ \text{g}$ of protein, $15\ \text{g}$ of fat, and $3\ \text{g}$ of sugar. By calculating the fraction of protein, fat, or sugar in one serving size of peanut butter and converting to percent values, we can determine the composition of peanut butter on a percent by mass basis.

Percent Composition

Chemists often need to know what elements are present in a compound and in what percentage. The **percent composition** is the percent by mass of each element in a compound. It is calculated in a similar way to what was just indicated for the peanut butter.

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

Percent Composition from Mass Data

The sample problem below shows the calculation of the percent composition of a compound based on mass data.

Example 11.6.1: Percent Composition from Mass

A certain newly synthesized compound is known to contain the elements zinc and oxygen. When a $20.00\ \text{g}$ sample of the compound is decomposed, $16.07\ \text{g}$ of zinc remains. Determine the percent composition of the compound.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass of compound $= 20.00\ \text{g}$
- Mass of $\text{Zn} = 16.07\ \text{g}$

Unknown

- Percent $\mathrm{Zn} = ?\%$
- Percent $\mathrm{O} = ?\%$

Subtract to find the mass of oxygen in the compound. Divide each element's mass by the mass of the compound to find the percent by mass.

Step 2: Calculate.

$$\text{Mass of oxygen} = 20.00\ \text{g} - 16.07\ \text{g} = 3.93\ \text{g O}$$

$$\%\ \mathrm{Zn} = \frac{16.07\ \text{g Zn}}{20.00\ \text{g}} \times 100\% = 80.35\%\ \mathrm{Zn}$$

$$\%\ \mathrm{O} = \frac{3.93\ \text{g O}}{20.00\ \text{g}} \times 100\% = 19.65\%\ \mathrm{O}$$

Step 3: Think about your result.

The calculations make sense because the sum of the two percentages adds up to 100%. By mass, the compound is mostly zinc.

Percent Composition from a Chemical Formula

The percent composition of a compound can also be determined from the formula of the compound. The subscripts in the formula are first used to calculate the mass of each element in one mole of the compound. That is divided by the molar mass of the compound and multiplied by 100%.

$$\%\ \text{by mass} = \frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100\%$$

The percent composition of a given compound is always the same as long as the compound is pure.

Example $11.6.2$: Percent Composition from Chemical Formula

Dichlorineheptoxide $(\mathrm{Cl_2O_7})$ is a highly reactive compound used in some organic synthesis reactions. Calculate the percent composition of dichlorineheptoxide.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass of Cl in $1\ \text{mol}\ \mathrm{Cl_2O_7} = 70.90\ \text{g}$
- Mass of O in $1\ \text{mol}\ \mathrm{Cl_2O_7} = 112.00\ \text{g}$
- Molar mass of $\mathrm{Cl_2O_7} = 182.90\ \text{g/mol}$

Unknown

- Percent $\mathrm{Cl} = ?\%$
- Percent $\mathrm{O} = ?\%$

Calculate the percent by mass of each element by dividing the mass of that element in 1 mole of the compound by the molar mass of the compound and multiplying by 100%.

Step 2: Calculate.

$$\%\text{Cl} = \frac{70.90\text{ g Cl}}{182.90\text{ g}} \times 100\% = 38.76\%\text{ Cl}$$

$$\%\text{ O} = \frac{112.00\text{ g O}}{182.90\text{ g}} \times 100\% = 61.24\%\text{ O}$$

Step 3: Think about your result.

The percentages add up to 100%.

Percent composition can also be used to determine the mass of a certain element that is contained in any mass of a compound. In the previous sample problem, it was found that the percent composition of dichlorine heptoxide is $38.76\%\ \text{Cl}$ and $61.24\%\ \text{O}$. Suppose that you need to know the masses of chlorine and oxygen present in a $12.50\ \text{g}$ sample of dichlorine heptoxide. You can set up a conversion factor based on the percent by mass of each element:

$$12.50\text{ g } \text{Cl}_2\text{O}_7 \times \frac{38.76\text{ g Cl}}{100\text{ g } \text{Cl}_2\text{O}_7} = 4.845\text{ g Cl}$$

$$12.50\text{ g } \text{Cl}_2\text{O}_7 \times \frac{61.24\text{ g O}}{100\text{ g } \text{Cl}_2\text{O}_7} = 7.655\text{ g O}$$

The sum of the two masses is $12.50\ \text{g}$, the mass of the sample size.

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Science Friday: Stained Glass Conservation

Stained glass from the Middle Ages is often hundreds of years old. Unfortunately, many of these relics are in need of cleaning and maintenance. In this video by Science Friday, conservator Mary Higgins discusses the methods used to protect the stained glass.

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Summary

- Processes are described for calculating the percent composition of a material based on mass or on chemical composition.

Review

1. What is the formula for calculating percent composition?
2. What information do you need to calculate percent composition by mass?
3. What do subscripts in a chemical formula tell you?

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11.7: Percent of Water in a Hydrate



Figure 11.7.1 (Public Domain; Doug Smith (Wikimedia: Xsmith) via Wikipedia)

Why does the color change?

If you look at a typical bottle of copper sulfate, it will be a bluish-green. If someone were to tell you that copper sulfate is white, you likely would not believe them. However, it turns out that you are both right; it just depends on the copper sulfate. Your blue-green copper sulfate has several water molecules attached to it, while your friend's copper sulfate is anhydrous (no water attached). Why the difference? The water molecules interact with some of the d electrons in the copper ion and produce the color. When the water is removed, the electron configuration changes and the color disappears.

Percent of Water in a Hydrate

Many ionic compounds naturally contain water as part of the crystal lattice structure. A **hydrate** is a compound that has one or more water molecules bound to each formula unit. Ionic compounds that contain a transition metal are often highly colored. Interestingly, it is common for the hydrated form of a compound to be of a different color than the **anhydrous** form, which has no water in its structure. A hydrate can usually be converted to the anhydrous compound by heating. For example, the anhydrous compound cobalt (II) chloride is blue, while the hydrate is a distinctive magenta color.

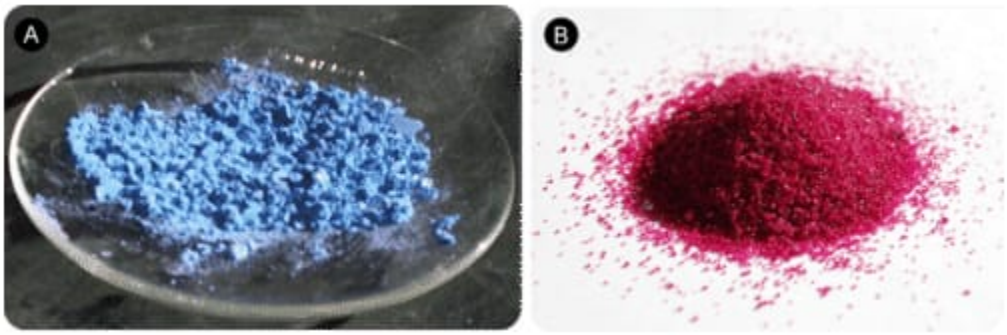


Figure 11.7.2: On the left is anhydrous cobalt (II) chloride, $\mathrm{CoCl_2}$. On the right is the hydrated form of the compound called cobalt (II) chloride hexahydrate, $\mathrm{CoCl_2 \cdot 6H_2O}$. (Public Domain; (A) Martin Walker (Wikimedia: Walkerma); (B) Ben Mills (Wikimedia: Benjah-bmm27) via (A) http://commons.wikimedia.org/wiki/File:Cobalt%2528II%2529_chloride.jpg; (B) http://commons.wikimedia.org/wiki/File:Cobalt%2528II%2529-chloride-hexahydrate-sample.jpg)

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The hydrated form of cobalt (II) chloride contains six water molecules in each formula unit. The name of the compound is cobalt (II) chloride hexahydrate and its formula is $\mathrm{CoCl_2 \cdot 6H_2O}$. The formula for water is set apart at the end of the formula with a dot, followed by a coefficient that represents the number of water molecules per formula unit.

It is useful to know the percent of water contained within a hydrate. The sample problem below demonstrates the procedure.

Example 11.7.1: Percent of Water in a Hydrate

Find the percent water in cobalt (II) chloride hexahydrate, $\mathrm{CoCl_2 \cdot 6H_2O}$.

Solution

Step 1: List the known quantities and plan the problem.

The mass of water in the hydrate is the coefficient (6) multiplied by the molar mass of $\mathrm{H_2O}$. The molar mass of the hydrate is the molar mass of the $\mathrm{CoCl_2}$ plus the mass of water.

Known

- Mass of $\mathrm{H_2O}$ in $1\ \mathrm{mol}$ hydrate $= 108.12\ \mathrm{g}$
- Molar mass of hydrate $= 237.95\ \mathrm{g/mol}$

Unknown

- percent H_2O = ?%

Calculate the percent by mass of water by dividing the mass of $\mathrm{H_2O}$ in 1 mole of the hydrate by the molar mass of the hydrate and multiplying by 100%.

Step 2: Calculate.

$$\%\ \mathrm{H_2O} = \frac{108.12\ \mathrm{g\ H_2O}}{237.95\ \mathrm{g}} \times 100\% = 45.44\%\ \mathrm{H_2O}$$

Step 3: Think about your result.

Nearly half of the mass of the hydrate is composed of water molecules within the crystal.

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Summary

- The process of calculating the percent water in a hydrate is described.

Review

1. What is a hydrate?
2. How can you convert a hydrate to an anhydrous compound?
3. What does hexahydrate mean?

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11.8: Determining Empirical Formulas



Figure 11.8.1 (Public Domain; Harriet Moore via Wikipedia)

What is occuring in this picture?

In the early days of chemistry, there were few tools available for the detailed study of compounds. Much of the information regarding the composition of compounds came from the elemental analysis of inorganic materials. The "new" field of organic chemistry (the study of carbon compounds) faced the challenge of not being able to characterize a compound completely. The relative amounts of elements could be determined, but many of these materials had carbon, hydrogen, oxygen, and possibly nitrogen in simple ratios. Chemists had no way to determine the exact amounts of these atoms that were contained in specific molecules.

Determining Empirical Formulas

An **empirical formula** is one that shows the lowest whole-number ratio of the elements in a compound. Because the structure of ionic compounds is an extended three-dimensional network of positive and negative ions, all formulas of ionic compounds are empirical. However, we can also consider the empirical formula of a molecular compound. Ethene is a small hydrocarbon compound with the formula $\mathrm{C_2H_4}$ (see figure below). While $\mathrm{C_2H_4}$ is its molecular formula and represents its true molecular structure, it has an empirical formula of $\mathrm{CH_2}$. The simplest ratio of carbon to hydrogen in ethene is 1:2. There are two ways to view that ratio. Considering one molecule of ethene, the ratio is 1 carbon atom for every 2 atoms of hydrogen. Considering one mole of ethene, the ratio is 1 mole of carbon for every 2 moles of hydrogen. So, the subscripts in a formula represent the mole ratio of the elements in that formula.

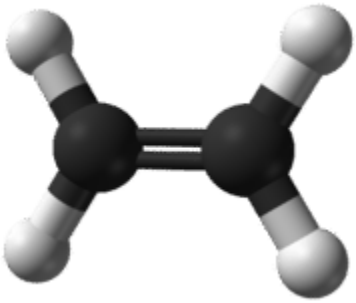


Figure 11.8.2: Ball-and-stick model of ethene, $\mathrm{C_2H_4}$. (Public Domain; Ben Mills (Wikimedia: Benjah-bmm27) via Wikipedia)

In a procedure called **elemental analysis**, an unknown compound can be analyzed in the laboratory in order to determine the percentages of each element contained within it. These percentages can be transformed into the mole ratio of the elements,

which leads to the empirical formula. The steps to be taken are outlined below.

1. Assume a $100\ \text{g}$ sample of the compound, so that the given percentages can be directly converted into grams.
2. Use each element's molar mass to convert the grams of each element to moles.
3. In order to find a whole-number ratio, divide the moles of each element by whichever of the moles from step 2 is the smallest.
4. If all the moles at this point are whole numbers (or very close), the empirical formula can be written with the moles as the subscript of each element.
5. In some cases, one or more of the moles calculated in step 3 will not be whole numbers. Multiply each of the moles by the smallest whole number that will convert each into a whole number. Write the empirical formula.

Example $11.8.1$**: Determining the Empirical Formula of a Compound**

A compound of iron and oxygen is analyzed and found to contain 69.94% iron and 30.06% oxygen. Find the empirical formula of the compound.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\%$ of $\text{Fe} = 69.94\%$
- $\%$ of $\text{O} = 30.06\%$

Unknown

- Empirial formula = $\text{Fe}_?\text{O}_?$

Steps to follow are outlined in the text.

Step 2: Calculate.

1. Assume a $100\ \text{g}$ sample.

$$69.94\ \text{g Fe}$$

$$30.06\ \text{g O}$$

2. Convert to moles.

$$69.94\ \text{g Fe} \times \frac{1\ \text{mol Fe}}{55.85\ \text{g Fe}} = 1.252\ \text{mol Fe}$$

$$30.06\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \text{g O}} = 1.879\ \text{mol O}$$

3. Divide both moles by the smallest of the results.

$$\frac{1.252\ \text{mol Fe}}{1.252} = 1\ \text{mol Fe} \quad \frac{1.879\ \text{mol O}}{1.252} = 1.501\ \text{molO}$$

4/5. Since the moles of O is still not a whole number, both moles can be multiplied by 2, while rounding to a whole number.

$$1\ \text{mol Fe} \times 2 = 2\ \text{mol Fe} \quad 1.501\ \text{mol O} \times 2 = 3\ \text{mol O}$$

The empirical formula of the compound is Fe_2O_3.

Step 3: Think about your result.

The subscripts are whole numbers and represent the mole ratio of the elements in the compound. The compound is the ionic compound iron (III) oxide.

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Summary

- A process is described for the calculation of the empirical formula for a compound based on the percent composition of that compound.

Review

1. What is an empirical formula?
2. What does an empirical formula tell you?
3. What does it not tell you?

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11.9: Determining Molecular Formulas

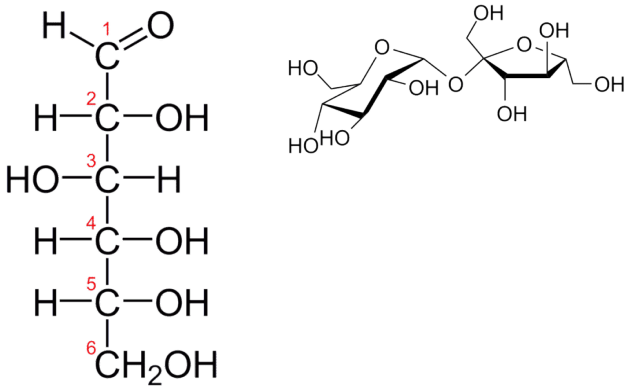


Figure 11.9.1 (Left: Credit: Ben Mills (Wikimedia: Benjah-bmm27) Source: http://commons.wikimedia.org/wiki/File:D-glucose-chain-2D-Fischer.png; http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png License: Public Domain) (Right: Credit: User:glycoform/ Wikimedia Commons; Ben Mills (Wikimedia: Benjah-bmm27); Source: http://commons.wikimedia.org/wiki/File:Sucrose_3Dprojection.png; http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png; License: Public Domain)

How can you determine the differences between these two molecules?

Above we see two carbohydrates: glucose and sucrose. Sucrose is almost exactly twice the size of glucose, although their empirical formulas are very similar. Some people could distinguish them on the basis of taste, but it's not a good idea to go around tasting chemicals. The best way is to determine the molecular weights—this approach allows you to easily tell which compound is which.

Molecular Formulas

Molecular formulas give the kind and number of atoms of each element present in a molecular compound. In many cases, the molecular formula is the same as the empirical formula. The molecular formula of methane is $\mathrm{CH_4}$ and because it contains only one carbon atom, that is also its empirical formula. Sometimes, however, the molecular formula is a simple whole-number multiple of the empirical formula. Acetic acid is an organic acid that is the main component of vinegar. Its molecular formula is $\mathrm{C_2H_4O_2}$. Glucose is a simple sugar that cells use as a primary source of energy. Its molecular formula is $\mathrm{C_6H_{12}O_6}$. The structures of both molecules are shown in the figure below. They are very different compounds, yet both have the same empirical formula of $\mathrm{CH_2O}$.

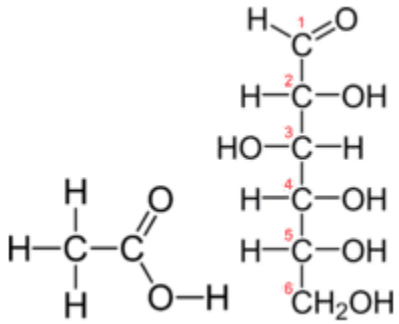


Figure 11.9.2: Acetic acid (left) has a molecular formula of $\mathrm{C_2H_4O_2}$, while glucose (right) has a molecular formula of $\mathrm{C_6H_{12}O_6}$. Both have the empirical formula $\mathrm{CH_2O}$. (Public Domain; (left) Ben Mills (Wikimedia: Benjah-bmm27); (right) Ben Mills (Wikimedia: Benjah-bmm27), User:Yikrazuul/Wikimedia Commons; Ben Mills (Wikimedia: Benjah-bmm27) via Acetic acid: http://commons.wikimedia.org/wiki/File:Acetic-acid-2D-flat.png; Glucose: http://commons.wikimedia.org/wiki/File:D-glucose-chain-2D-Fischer.png; http://commons.wikimedia.org/wiki/File:Ethylene-CRC-MW-3D-balls.png)

Empirical formulas can be determined from the percent composition of a compound. In order to determine its molecular formula, it is necessary to know the molar mass of the compound. Chemists use an instrument called a mass spectrometer to determine

the molar mass of compounds. In order to go from the empirical formula to the molecular formula, follow these steps:

1. Calculate the **empirical formula mass (EFM)**, which is simply the molar mass represented by the empirical formula.
2. Divide the molar mass of the compound by the empirical formula mass. The result should be a whole number or very close to a whole number.
3. Multiply all the subscripts in the empirical formula by the whole number found in step 2. The result is the molecular formula.

Example $11.9.1$**: Determining the Molecular Formula of a Compound**

The empirical formula of a compound of boron and hydrogen is $\mathrm{BH_3}$. Its molar mass is $27.7\ \mathrm{g/mol}$. Determine the molecular formula of the compound.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Empirical formula $= \mathrm{BH_3}$
- Molar mass $= 27.7\ \mathrm{g/mol}$

Unknown

- molecular formula = ?

Steps to follow are outlined in the text.

Step 2: Calculate.

$$\text{Empirical formula mass (EFM)} = 13.84\,\mathrm{g/mol}$$

$$\frac{\text{molar mass}}{\text{EFM}} = \frac{17.7}{13.84} = 2$$

$$\mathrm{BH_3} \times 2 = \mathrm{B_2H_6}$$

The molecular formula of the compound is $\mathrm{B_2H_6}$.

Step 3: Think about your result.

The molar mass of the molecular formula matches the molar mass of the compound.

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Summary

- A procedure is described for the calculation of the exact molecular formula of a compound.

Review

1. What is the difference between an empirical formula and a molecular formula?
2. In addition to the elemental analysis, what do you need to know to calculate the molecular formula?
3. What does the empirical formula mass tell you?

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11.10: Avogadro's Hypothesis and Molar Volume



Figure 11.10.1 (Public Domain; User:Mark.murphy/Wikimedia Commons via Wikipedia)

How do scuba divers know if they will run out of gas?

Knowing how much gas is available for a dive is crucial to a diver's survival. The tank on the diver's back is equipped with gauges to indicate how much gas is present and what the pressure is. A basic knowledge of gas behavior allows the diver to assess how long they can stay underwater without developing problems.

Avogadro's Hypothesis and Molar Volume

Volume is a third way to measure the amount of matter, after item count and mass. With liquids and solids, volume varies greatly depending on the density of the substance. This is because solid and liquid particles are packed close together with very little space in between the particles. However, gases are largely composed of empty spaces between the actual gas particles (see figure below).

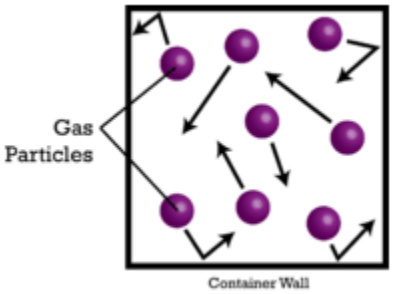


Figure 11.10.2: Gas particles are very small compared to the large amounts of empty space between them. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

In 1811, Amadeo Avogadro explained that the volumes of all gases can be easily determined. **Avogadro's hypothesis** states that equal volumes of all gases at the same temperature and pressure contain equal numbers of particles. Since the total volume that a gas occupies is made up primarily of the empty space between the particles, the actual size of the particles themselves is nearly negligible. A given volume of a gas with small light particles, such as hydrogen (H_2), contains the same number of particles as the same volume of a heavy gas with large particles, such as sulfur hexafluoride, SF_6.

Gases are compressible, meaning that when put under high pressure, the particles are forced closer to one another. This

decreases the amount of empty space and reduces the volume of the gas. Gas volume is also affected by temperature. When a gas is heated, its molecules move faster and the gas expands. Because of the variation in gas volume due to pressure and temperature changes, the comparison of gas volumes must be done at standard temperature and pressure. **Standard temperature and pressure (STP)** is defined as $0^{\circ}\mathrm{C}\ (273.15\ \mathrm{K})$ and $1\ \mathrm{atm}$ pressure. The **molar volume** of a gas is the volume of one mole of a gas at STP. At STP, one mole (6.02×10^{23} representative particles) of any gas occupies a volume of $22.4\ \mathrm{L}$ (figure below).

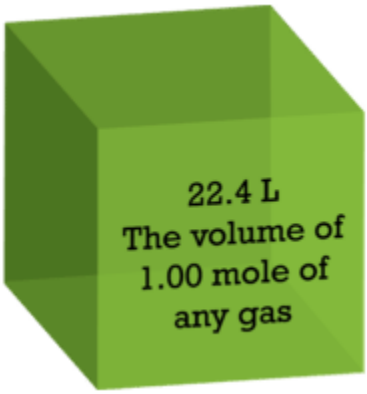
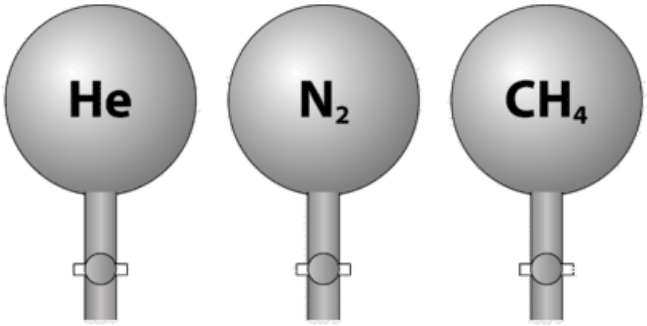


Figure $11.10.3$: A mole of any gas occupies $22.4\ \mathrm{L}$ at standard temperature and pressure ($0^{\circ}\mathrm{C}$ and $1\ \mathrm{atm}$). (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The figure below illustrates how molar volume can be seen when comparing different gases. Samples of helium (He), nitrogen $(\mathrm{N_2})$, and methane $(\mathrm{CH_4})$ are at STP. Each contains 1 mole or 6.02×10^{23} particles. However, the mass of each gas is different and corresponds to the molar mass of that gas: $4.00\ \mathrm{g/mol}$ for He, $28.0\ \mathrm{g/mol}$ for $\mathrm{N_2}$, and $16.0\ \mathrm{g/mol}$ for $\mathrm{CH_4}$.



Volume	22.4 L	22.4 L	22.4 L
Pressure	1 atm	1 atm	1 atm
Temperature	0°C	0°C	0°C
Mass of gas	4.00 g	28.0 g	16.0 g
Number of gas molecules	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

Figure $11.10.4$: Avogadro's hypothesis states that equal volumes of any gas at the same temperature and pressure contain the same number of particles. At standard temperature and pressure, 1 mole of any gas occupies $22.4\ \mathrm{L}$. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

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Summary

- Equal volumes of gases at the same conditions contain the same number of particles.
- Standard temperature and pressure is abbreviated (STP).
- Standard temperature is 0°C (273.15 K) and standard pressure is 1 atm.
- At STP, one mole of any gas occupies a volume of 22.4 L

Review

- A container is filled with gas, what do we know about the space actually taken up by a gas?
- Why do we need to do all our comparisons at the same temperature and pressure?
- At standard temperature and pressure, 1 mole of gas is always equal to how many liters?

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12: STOICHIOMETRY

12.1: Everyday Stoichiometry

12.2: Mole Ratios

12.3: Mass-Mole and Mole-Mass Stoichiometry

12.4: Mass-Mass Stoichiometry

12.5: Limiting Reactant

12.6: Determining the Limiting Reactant

12.7: Theoretical Yield and Percent Yield

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12.1: Everyday Stoichiometry



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How much equipment do you need for an experiment?

You are in charge of setting out the lab equipment for a chemistry experiment. If you have twenty students in the lab (and they will be working in teams of two) and the experiment calls for three beakers and two test tubes, how much glassware do you need to set out? Figuring this out involves a type of balanced equation and the sort of calculations that you would do for a chemical reaction.

Everyday Stoichiometry

You have learned about chemical equations and the techniques used in order to balance them. Chemists use balanced equations to allow them to manipulate chemical reactions in a quantitative manner. Before we look at a chemical reaction, let's consider an equation for the ideal ham sandwich.



Figure 12.1.2: The ideal ham sandwich. (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Our ham sandwich is composed of 2 slices of ham (H), a slice of cheese (C), a slice of tomato (T), 5 pickles (P), and 2 slices of bread (B). The equation for our sandwich is:

$$2\mathrm{H} + \mathrm{C} + \mathrm{T} + 5\mathrm{P} + 2\mathrm{B} \rightarrow \mathrm{H_2CTP_5B_2}$$

Now let us suppose that you are having some friends over and need to make five ham sandwiches. How much of each sandwich ingredient do you need? You take the number of each ingredient required for one sandwich (its coefficient in the above equation) and multiply by five. Using ham and cheese as examples, and using a conversion factor, you can calculate:

$$5\mathrm{H_2CTP_5B_2} \times \frac{2\ \mathrm{H}}{1\mathrm{H_2CTP_5B_2}} = 10\ \mathrm{H}$$

$$5\mathrm{H_2CTP_5B_2} \times \frac{1\mathrm{C}}{1\mathrm{H_2CTP_5B_2}} = 5\ \mathrm{C}$$

The conversion factors contain the coefficient of each specific ingredient as the numerator and the formula of one sandwich as the denominator. The result is what you would expect. In order to make five ham sandwiches, you need 10 slices of ham and 5 slices of cheese.

This type of calculation demonstrates the use of stoichiometry. **Stoichiometry** is the calculation of the amount of substances in a chemical reaction from the balanced equation. The sample problem below is another stoichiometry problem involving ingredients of the ideal ham sandwich.

Example 12.1.1**: Ham Sandwich Stoichiometry**

Kim looks in the refrigerator and finds that she has 8 slices of ham. In order to make as many sandwiches as possible, how many pickles does she need? Use the equation above.

Solution

Step 1: List the known quantities and plan the problem.

- Have 8 ham slices (H)
- $2\ \mathrm{H} = 5\ \mathrm{P}$ (conversion factor)

Unknown

- How many pickles (P) needed?

The coefficients for the two reactants (ingredients) are used to make a conversion factor between ham slices and pickles.

Step 2: Solve.

$$8\ \mathrm{H} \times \frac{5\ \mathrm{P}}{2\ \mathrm{H}} = 20\ \mathrm{P}$$

Since 5 pickles combine with 2 ham slices in each sandwich, 20 pickles are needed to fully combine with 8 ham slices.

Step 3: Think about your result.

The 8 ham slices will make 4 ham sandwiches. With 5 pickles per sandwich, the 20 pickles are used in the 4 sandwiches.

Summary

- An example of everyday stoichiometry is given.

Review

1. A smoothie contains 1 banana (B), 4 strawberries (St), 1 container of yogurt (Y), and 3 ice cubes (Ic). Write a balanced

equation to describe the relationship.
2. Write a conversion factor to show the relationship between the number of ice cubes and the number of smoothies produced.
3. How many strawberries would you need to make 12 smoothies?

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12.2: Mole Ratios



Figure 12.2.1 (Public Domain; User:Ilemme/Wikipedia via Commons Wikimedia, File:Col 03 scup porch chairs.JPG [commons.wikimedia.org])

What does this porch need?

You want to add some sections to the porch seen above. Before you go to the hardware store to buy lumber, you need to determine the unit composition (the material between two large uprights). You count how many posts, how many boards, how many rails – then you decide how many sections you want to add before you calculate the amount of building material needed for your porch expansion.

Mole Ratios

Stoichiometry problems can be characterized by two things: (1) the information *given* in the problem, and (2) the information that is to be solved for, referred to as the *unknown*. The given and the unknown may both be reactants, both be products, or one may be a reactant while the other is a product. The amounts of the substances can be expressed in moles. However, in a laboratory situation, it is common to determine the amount of a substance by finding its mass in grams. The amount of a gaseous substance may be expressed by its volume. In this concept, we will focus on the type of problem where both the given and the unknown quantities are expressed in moles.

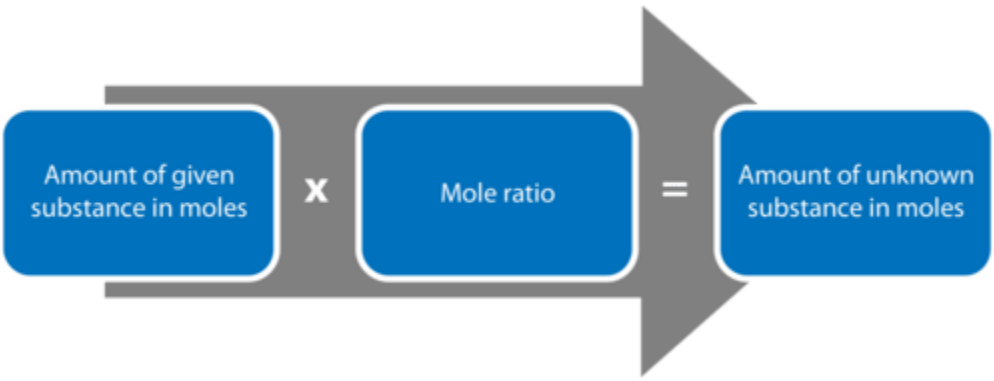


Figure 12.2.2: Mole ratio relationship. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

Chemical equations express the amounts of reactants and products in a reaction. The coefficients of a balanced equation can represent either the number of molecules or the number of moles of each substance. The production of ammonia (NH_3) from

nitrogen and hydrogen gases is an important industrial reaction called the Haber process, after German chemist Fritz Haber.

$$\mathrm{N_2}\,(g) + 3\mathrm{H_2}\,(g) \rightarrow 2\mathrm{NH_3}\,(g)$$

The balanced equation can be analyzed in several ways, as shown in the figure below.

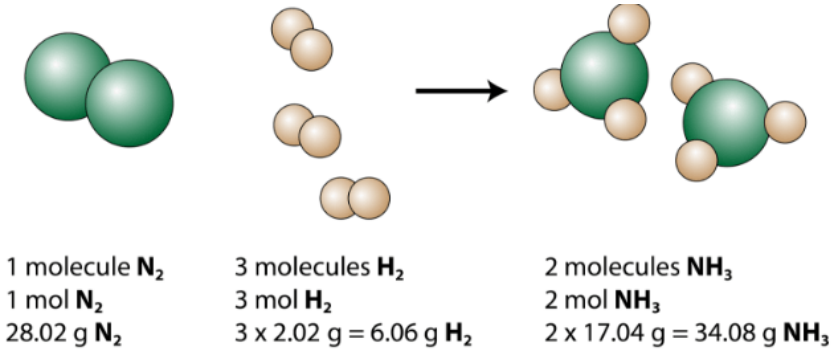


Figure 12.2.3: This representation of the production of ammonia from nitrogen and hydrogen shows several ways to interpret the quantitative information of a chemical reaction. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

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. $1\ \mathrm{mol}$ of nitrogen has a mass of $28.02\ \mathrm{g}$, while $3\ \mathrm{mol}$ of hydrogen has a mass of $6.06\ \mathrm{g}$, and $2\ \mathrm{mol}$ of ammonia has a mass of $34.08\ \mathrm{g}$.

$$28.02\ \mathrm{g}\ \mathrm{N_2} + 6.06\ \mathrm{g}\ \mathrm{H_2} \rightarrow 34.08\ \mathrm{g}\ \mathrm{NH_3}$$

Mass and the number of atoms must be conserved in any chemical reaction. The number of molecules is not necessarily conserved.



Figure 12.2.4: An apparatus for running the Haber process. (Public Domain; User: JGvBerkel/Wikimedia Commons via Commons Wikimedia, Haber Ammonia [commons.wikimedia.org])

A **mole ratio** is a conversion factor that relates the amounts in moles of any two substances in a chemical reaction. The numbers in a conversion factor come from the coefficients of the balanced chemical equation. The following six mole ratios can be written for the ammonia forming reaction above.

$$\frac{1\ \text{mol}\ \text{N}_2}{3\ \text{mol}\ \text{H}_2} \quad or \quad \frac{3\ \text{mol}\ \text{H}_2}{1\ \text{mol}\ \text{N}_2}$$
$$\frac{1\ \text{mol}\ \text{N}_2}{2\ \text{mol}\ \text{NH}_3} \quad or \quad \frac{2\ \text{mol}\ \text{NH}_3}{1\ \text{mol}\ \text{N}_2}$$
$$\frac{3\ \text{mol}\ \text{H}_2}{2\ \text{mol}\ \text{NH}_3} \quad or \quad \frac{2\ \text{mol}\ \text{NH}_3}{3\ \text{mol}\ \text{H}_2}$$

In a mole ratio problem, the given substance, expressed in moles, is written first. The appropriate conversion factor is chosen in order to convert from moles of the given substance to moles of the unknown.

Example 12.2.1**: Mole Ratio**

How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?

Solution

Step 1: List the known quantities and plan the problem.

Known

- given: H_2 = 4.20 mol

Unknown

- mol of NH_3

The conversion is from $\text{mol}\ \text{H}_2$ to $\text{mol}\ \text{NH}_3$. The problem states that there is an excess of nitrogen, so we do not need to be concerned with any mole ratio involving N_2. Choose the conversion factor that has the NH_3 in the numerator and the H_2 in the denominator.

Step 2: Solve.

$$4.20\ \text{mol}\ \text{H}_2 \times \frac{2\ \text{mol}\ \text{NH}_3}{3\ \text{mol}\ \text{H}_2} = 2.80\ \text{mol}\ \text{NH}_3$$

The reaction of $4.20\ \text{mol}$ of hydrogen with excess nitrogen produces $2.80\ \text{mol}$ of ammonia.

Step 3: Think about your result.

The result corresponds to the 3:2 ratio of hydrogen to ammonia from the balanced equation.

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Summary

- Mole ratios allow comparison of the amounts of any two materials in a balanced equation.
- Calculations can be made to predict how much product can be obtained from a given number of moles of reactant.

Review

- If a reactant is in excess, why do we not worry about the mole ratios involving that reactant?
- What is the mole ratio of H to N in the ammonia molecule?
- The formula for ethanol is CH_3CH_2OH. What is the mole ratio of H to C in this molecule?

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12.3: Mass-Mole and Mole-Mass Stoichiometry

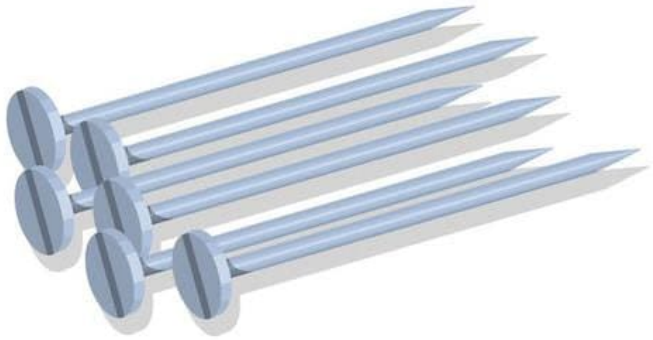


Figure 12.3.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Need nails?

When you are doing a large construction project, you have a good idea of how many nails you will need (lots!). When you go to the hardware store, you don't want to sit there and count out several hundred nails. You can buy nails by weight, so you determine how many nails are in a pound, calculate how many pounds you need, and you're on your way to begin building.

While the mole ratio is ever-present in all stoichiometry calculations, amounts of substances in the laboratory are most often measured by mass. Therefore, we need to use mole-mass calculations in combination with mole ratios to solve several different types of mass-based stoichiometry problems.

Mass to Moles Problems

In this type of problem, the mass of one substance is given, usually in grams. From this, you are to determine the amount in moles of another substance that will either react with or be produced from the given substance.

$$\text{mass of given} \rightarrow \text{moles of given} \rightarrow \text{moles of unknown}$$

The mass of the given substance is converted into moles by use of the molar mass of that substance from the periodic table. Then, the moles of the given substance are converted into moles of the unknown by using the mole ratio from the balanced chemical equation.

Example 12.3.1: Mass-Mole Stoichiometry

Tin metal reacts with hydrogen fluoride to produce tin (II) fluoride and hydrogen gas, according to the following balanced equation.

$$\mathrm{Sn}\,(s) + 2\mathrm{HF}\,(g) \rightarrow \mathrm{SnF_2}\,(s) + \mathrm{H_2}\,(g)$$

How many moles of hydrogen fluoride are required to react completely with $75.0\ \mathrm{g}$ of tin?

Solution:

Step 1: List the known quantities and plan the problem.

Known

- Given: $75.0\ \mathrm{g\ Sn}$
- Molar mass of $\mathrm{Sn} = 118.69\ \mathrm{g/mol}$
- $1\ \mathrm{mol\ Sn} = 2\ \mathrm{mol\ HF}$ (mole ratio)

Unknown

- mol HF

Use the molar mass of Sn to convert the grams of Sn to moles. Then use the mole ratio to convert from $\mathrm{mol\ Sn}$ to $\mathrm{mol\ HF}$. This will be done in a single two-step calculation.

$\mathrm{g\ Sn \rightarrow mol\ Sm \rightarrow mol\ HF}$

Step 2: Solve.

$$75.0\ \mathrm{g\ Sn} \times \frac{1\ \mathrm{mol\ Sn}}{118.69\ \mathrm{g\ Sn}} \times \frac{2\ \mathrm{mol\ HF}}{1\ \mathrm{mol\ Sn}} = 1.26\ \mathrm{mol\ HF}$$

Step 3: Think about your result.

The mass of tin is less than one mole, but the 1:2 ratio means that more than one mole of HF is required for the reaction. The answer has three significant figures because the given mass has three significant figures.

Moles to Mass Problems

In this type of problem, the amount of one substance is given in moles. From this, you are to determine the mass of another substance that will either react with or be produced from the given substance.

$$\mathrm{moles\ of\ given \rightarrow moles\ of\ unknown \rightarrow mass\ of\ unknown}$$

The moles of the given substance are first converted into moles of the unknown by using the mole ratio from the balanced chemical equation. Then, the moles of the unknown are converted into mass in grams by use of the molar mass of that substance from the periodic table.

Example $12.3.2$**: Mole-Mass Stoichiometry**

Hydrogen sulfide gas burns in oxygen to produce sulfur dioxide and water vapor:

$$2\mathrm{H_2S}\,(g) + 3\mathrm{O_2}\,(g) \rightarrow 2\mathrm{SO_2}\,(g) + 2\mathrm{H_2O}\,(g)$$

What mass of oxygen gas is consumed in a reaction that produces $4.60\ \mathrm{mol\ SO_2}$?

Solution:

Step 1: List the known quantities and plan the problem.

Known

- Given: $4.60\ \mathrm{mol\ SO_2}$
- $2\ \mathrm{mol\ SO_2} = 3\ \mathrm{mol\ O_2}$ (mole ratio)
- Molar mass of $\mathrm{O_2} = 32.00\ \mathrm{g/mol}$

Unknown

- mass O_2 = ? g

Use the mole ratio to convert from $\mathrm{mol\ SO_2}$ to $\mathrm{mol\ O_2}$. Then convert $\mathrm{mol\ O_2}$ to grams. This will be done in a single two-step calculation.

$\mathrm{mol\ SO_2 \rightarrow mol\ O_2 \rightarrow g\ O_2}$

Step 2: Solve.

$$4.60 \text{ mol } SO_2 \times \frac{3 \text{ mol } O_2}{2 \text{ mol } SO_2} \times \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 221 \text{ g } O_2$$

Step 3: Think about your result.

According to the mole ratio, $6.90 \text{ mol } O_2$ is produced with a mass of 221 g. The answer has three significant figures because the given number of moles has three significant figures.

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Summary

- Calculations involving conversions of mass to moles and moles to mass are described.

Review

1. In the first problem, what would happen if you multiply grams Sn by 118.69 grams/mole Sn?
2. Why is a balanced equation needed?
3. Does the physical form of the material matter for these calculations?

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12.4: Mass-Mass Stoichiometry



Figure 12.4.1 (CC by 2.0; Pauline Mak via Flickr)

How many walnuts are needed to equal 250 grams?

I want to send 250 grams of shelled walnuts to a friend (don't ask why- just go with the question). How many walnuts in shells do I need to buy? To figure this out, I need to know how much the shell of a walnut weighs (about 40% of the total weight of the unshelled walnut). I can then calculate the mass of walnuts that will yield 250 grams of shelled walnuts, and then determine how many walnuts I need to buy.

Mass to Mass Problems

Mass-mass calculations are the most practical of all mass-based stoichiometry problems. Moles cannot be measured directly, while the mass of any substance can generally be easily measured in the lab. This type of problem is three steps, and is a combination of the two previous types.

$$\text{mass of given} \rightarrow \text{moles of given} \rightarrow \text{moles of unknown} \rightarrow \text{mass of unknown}$$

The mass of the given substance is converted into moles by use of the molar mass of that substance from the periodic table. Then, the moles of the given substance are converted into moles of the unknown by using the mole ratio from the balanced chemical equation. Finally, the moles of the unknown are converted to mass by use of its molar mass.

Example 12.4.1: Mass-Mass Stoichiometry

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, $45.7\ \mathrm{g}$ of ammonium nitrate is decomposed. Find the mass of each of the products formed.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Given: $45.7\ \mathrm{g}\ \mathrm{NH_4NO_3}$
- $1\ \mathrm{mol}\ \mathrm{NH_4NO_3} = 1\ \mathrm{mol}\ \mathrm{N_2O} = 2\ \mathrm{mol}\ \mathrm{H_2O}$

- Molar mass of $\mathrm{NH_4NO_3} = 80.06\ \mathrm{g/mol}$
- Molar mass of $\mathrm{N_2O} = 44.02\ \mathrm{g/mol}$
- Molar mass of $\mathrm{H_2O} = 18.02\ \mathrm{g/mol}$

Unknown

- Mass $\mathrm{N_2O} = ?\ \mathrm{g}$
- Mass $\mathrm{H_2O} = ?\ \mathrm{g}$

Perform two separate three-step mass-mass calculations, as shown below.

$$\mathrm{g\ NH_4NO_3 \rightarrow mol\ NH_4NO_3 \rightarrow mol\ N_2O \rightarrow g\ N_2O}$$

$$\mathrm{g\ NH_4NO_3 \rightarrow mol\ NH_4NO_3 \rightarrow mol\ H_2O \rightarrow g\ H_2O}$$

Step 2: Solve.

$$45.7\ \mathrm{g\ NH_4NO_3} \times \frac{1\ \mathrm{mol\ NH_4NO_3}}{80.06\ \mathrm{g\ NH_4NO_3}} \times \frac{1\ \mathrm{mol\ N_2O}}{1\ \mathrm{mol\ NH_4NO_3}} \times \frac{44.02\ \mathrm{g\ N_2O}}{1\ \mathrm{mol\ N_2O}} = 25.1\ \mathrm{g\ N_2O}$$

$$45.7\ \mathrm{g\ NH_4NO_3} \times \frac{1\ \mathrm{mol\ NH_4NO_3}}{80.06\ \mathrm{g\ NH_4NO_3}} \times \frac{2\ \mathrm{H_2O}}{1\ \mathrm{mol\ NH_4NO_3}} \times \frac{18.02\ \mathrm{g\ H_2O}}{1\ \mathrm{mol\ H_2O}} = 20.6\ \mathrm{g\ H_2O}$$

Step 3: 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.

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Summary

- Mass-mass calculations involve converting the mass of a reactant to moles of reactant, then using mole ratios to determine moles of product which can then be converted to mass of product.

Review

1. If matter is neither created nor destroyed, why can't we just go directly from grams of reactant to grams of product?
2. Why is it important to get the subscripts correct in the formulas?
3. Why do the coefficients need to be correct?

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12.5: Limiting Reactant

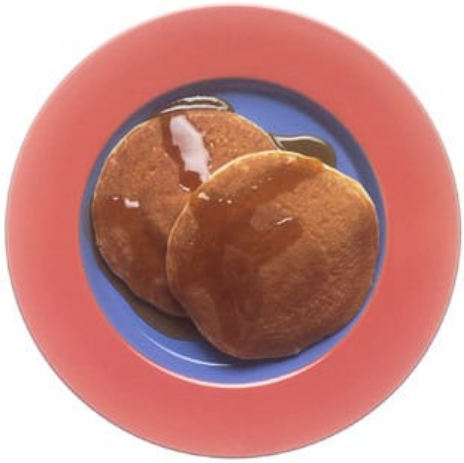


Figure 12.5.1 (Public Domain; Renee Comet and the National Cancer Institute; CK-12 Foundation - Christopher Auyeung via Commons Wikimedia, Pancakes [commons.wikimedia.org])

Don't you hate running out of cooking ingredients?

Cooking is a great example of everyday chemistry. In order to correctly follow a recipe, a cook needs to make sure that he has plenty of all the necessary ingredients in order to make his dish. Let us suppose that you are deciding to make some pancakes for a large group of people. The recipe on the box indicates that the following ingredients are needed for each batch of pancakes:

1 cup of pancake mix

$\frac{3}{4}$ cup milk

1 egg

1 tablespoon vegetable oil

Now you check the pantry and the refrigerator and see that you have the following ingredients available:

2 boxes of pancake mix (8 cups)

Half a gallon of milk (4 cups)

2 eggs

Full bottle of vegetable oil (about 3 cups)

The question that you must ask is: how many batches of pancakes can I make? The answer is two. Even though you have enough pancake mix, milk, and oil to make many more batches of pancakes, you are *limited* by the fact that you only have two eggs. As soon as you have made two batches of pancakes, you will be out of eggs and your "reaction" will be complete.

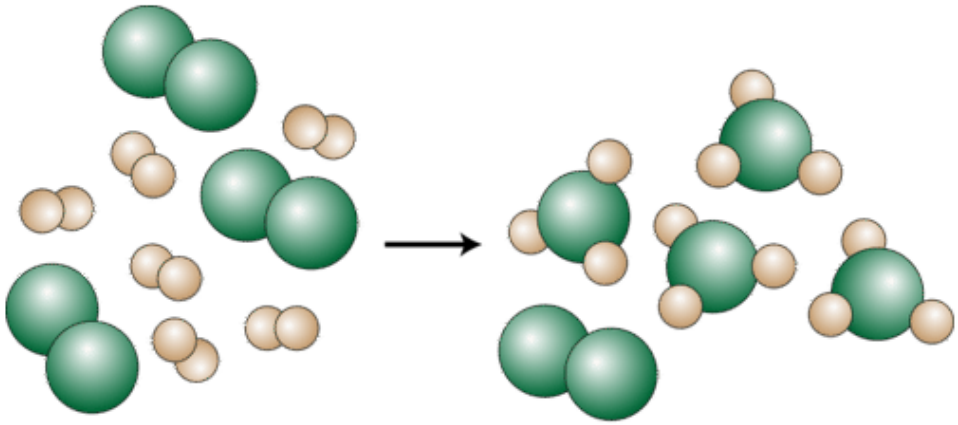
Limiting Reactant

For a chemist, the balanced chemical equation is the recipe that must be followed. As you have seen earlier, the Haber process is a reaction in which nitrogen gas is combined with hydrogen gas to form ammonia. The balanced equation is shown below.

$$\mathrm{N_2}\,(g) + 3\mathrm{H_2}\,(g) \rightarrow 2\mathrm{NH_3}\,(g)$$

We know that the coefficients of the balanced equation indicate the mole ratio that is required for this reaction to occur. One mole of $\mathrm{N_2}$ will react with three moles of $\mathrm{H_2}$ to form two moles of $\mathrm{NH_3}$.

Now let us suppose that a chemist were to react three moles of $\mathrm{N_2}$ with six moles of $\mathrm{H_2}$ (see figure below).



	Reactants		Products
Before reaction	3 mol N_2	6 mol H_2	0 mol NH_3
After reaction	1 mol N_2	0 mol H_2	4 mol NH_3

Figure 12.5.2: Reaction in presence of limiting reagent. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

So what happened in this reaction? The chemist started with 3 moles of $\mathrm{N_2}$. You may think of this as being 3 times as much as the "recipe" (the balanced equation) requires since the coefficient for the $\mathrm{N_2}$ is a 1. However, the 6 moles of $\mathrm{H_2}$ that the chemist started with is only two times as much as the "recipe" requires, since the coefficient for the $\mathrm{H_2}$ is a 3 and $3 \times 2 = 6$. So, after the reaction is complete, the hydrogen gas will be completely used up; while there will be 1 mole of nitrogen gas left over. Finally, the reaction will produce 4 moles of $\mathrm{NH_3}$ because that is also two times as much as shown in the balanced equation. The overall reaction that occurred in words:

$$2 \text{ mol } \mathrm{N_2} + 6 \text{ mol } \mathrm{H_2} \rightarrow 4 \text{ mol } \mathrm{NH_3}$$

All the amounts are doubled from the original balanced equation.

The **limiting reactant** (or **limiting reagent**) is the reactant that determines the amount of product that can be formed in a chemical reaction. The reaction proceeds until the limiting reactant is completely used up. In our example above, the $\mathrm{H_2}$ is the limiting reactant. The **excess reactant** (or **excess reagent**) is the reactant that is initially present in a greater amount than will eventually be reacted. In other words, there is always excess reactant left over after the reaction is complete. In the above example, the $\mathrm{N_2}$ is the excess reactant.

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Summary

- The amount of limiting reactant determines how much product will be formed in a chemical reaction.

Review

1. In the Haber reaction illustrated above, how do we know that hydrogen is the limiting reactant?
2. What if hydrogen were left over?
3. Which material would be limiting if no hydrogen or nitrogen were left over?

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12.6: Determining the Limiting Reactant



Figure 12.6.1 (Public Domain; Renee Comet and the National Cancer Institute via Wikipedia)

Who's coming for dinner?

You have ten people that show up for a dinner party. One of the guest brings twenty brownies for dessert. The decision about serving dessert is easy: two brownies are placed on every plate. If someone wants more brownies, they will have to wait until they go to the store. There are only enough brownies for everyone to have two.

Determining the Limiting Reactant

In the real world, amounts of reactants and products are typically measured by mass or by volume. It is first necessary to convert the given quantities of each reactant to moles in order to identify the limiting reactant.

Example 12.6.1: Determining the Limiting Reactant

Silver metal reacts with sulfur to form silver sulfide according to the following balanced equation:

$$2\mathrm{Ag}\,(s) + \mathrm{S}\,(s) \rightarrow \mathrm{Ag_2S}\,(s)$$

What is the limiting reactant when $50.0\ \mathrm{g\ Ag}$ is reacted with $10.0\ \mathrm{g\ S}$?

Solution:

Step 1: List the known quantities and plan the problem.

Known

- Given: $50.0\ \mathrm{g\ Ag}$
- Given: $10.0\ \mathrm{g\ S}$

Unknown

- limiting reactant

Use the atomic masses of Ag and S to determine the number of moles of each present. Then, use the balanced equation to calculate the number of moles of sulfur that would be needed to react with the number of moles of silver present. Compare this result to the actual number of moles of sulfur present.

Step 2: Solve.

First, calculate the number of moles of Ag and S present:

$$50.0\ \mathrm{g\ Ag} \times \frac{1\ \mathrm{mol\ Ag}}{107.87\ \mathrm{g\ Ag}} = 0.464\ \mathrm{mol\ Ag}$$

$$10.0\ \mathrm{g\ S} \times \frac{1\ \mathrm{mol\ S}}{32.07\ \mathrm{g\ S}} = 0.312\ \mathrm{mol\ S}$$

Second, find the moles of S that would be required to react with all of the given Ag:

$$0.464\ \mathrm{mol\ Ag} \times \frac{1\ \mathrm{mol\ S}}{2\ \mathrm{mol\ Ag}} = 0.232\ \mathrm{mol\ S\ (required)}$$

The amount of S actually present is 0.312 moles. The amount of S that is required to fully react with all of the Ag is 0.232 moles. Since there is more sulfur present than what is required to react, the sulfur is the excess reactant. Therefore, silver is the limiting reactant.

Step 3: Think about your result.

The balanced equation indicates that the necessary mole ratio of Ag to S is 2:1. Since there were not twice as many moles of Ag present in the original amounts, that makes silver the limiting reactant.

There is a very important point to consider about the preceding problem. Even though the mass of silver present in the reaction $(50.0\ \mathrm{g})$ was greater than the mass of sulfur $(10.0\ \mathrm{g})$, silver was the limiting reactant. This is because chemists must always convert to molar quantities and consider the mole ratio from the balanced chemical equation.

There is one other component to be determined in a limiting reactant problem—the quantity of the excess reactant that will be left over after the reaction is complete. We will return to the sample problem, above, to answer this question below.

Example 12.6.2: Determining the Amount of Excess Reactant Left Over

What is the mass of excess reactant remaining when $50.0\ \mathrm{g\ Ag}$ reacts with $10.0\ \mathrm{g\ S}$?

$$2\mathrm{Ag}\,(s) + \mathrm{S}\,(s) \rightarrow \mathrm{Ag_2S}\,(s)$$

Solution:

Step 1: List the known quantities and plan the problem.

Known

- Excess reactant $= 0.312\ \mathrm{mol\ S}$ (from Example 12.6.1)
- Amount of excess reactant needed $= 0.232\ \mathrm{mol\ S}$ (from Example 12.6.1)

Unknown

- Mass of excess reactant remaining after the reaction = ? g

Subtract the amount (in moles) of the excess reactant that will react from the amount that is originally present. Convert moles to grams.

Step 2: Solve.

$$0.312\ \mathrm{mol\ S} - 0.232\ \mathrm{mol\ S} = 0.080\ \mathrm{mol\ S\ (remaining\ after\ reaction)}$$

$$0.080\ \mathrm{mol\ S} \times \frac{32.07\ \mathrm{g\ S}}{1\ \mathrm{mol\ S}} = 2.57\ \mathrm{g\ S}$$

There are $2.57\ \mathrm{g}$ of sulfur remaining when the reaction is complete.

Step 3: Think about your result.

There were $10.0\ \text{g}$ of sulfur present before the reaction began. If $2.57\ \text{g}$ of sulfur remain after the reaction, then $7.43\ \text{g S}$ reacted.

$$7.43\ \text{g S} \times \frac{1\ \text{mol S}}{32.07\ \text{g S}} = 0.232\ \text{mol S}$$

This is the amount of sulfur that reacted. The problem is internally consistent.

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Summary

- Determining the limiting reactant requires that all mass quantities first be converted to moles to evaluate the equation.

Review

1. Why do all mass values need to be converted to moles before determining the limiting reactant?
2. Silver metal reacts with sulfur to form silver sulfide according to the following balanced equation:

$$2\text{Ag}\,(s) + \text{S}\,(s) \rightarrow \text{Ag}_2\text{S}\,(s)$$

   a. If 0.700 moles Ag is reacted with 10.0 g S, is sulfur or aluminum the limiting reactant?
   b. How many grams of Ag_2S will be produced?

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12.7: Theoretical Yield and Percent Yield



Figure 12.7.1 (Public Domain; Linda Bartlett and the National Cancer Institute via Wikipedia)

Can we save some money?

The world of pharmaceutical production is an expensive one. Many drugs have several steps in their synthesis and use costly chemicals. A great deal of research takes place to develop better ways to make drugs faster and more efficiently. Studying how much of a compound is produced in any given reaction is an important part of cost control.

Percent Yield

Chemical reactions in the real world do not always go exactly as planned on paper. In the course of an experiment, many things will contribute to the formation of less product than would be predicted. Besides spills and other experimental errors, there are often losses due to an incomplete reaction, undesirable side reactions, etc. Chemists need a measurement that indicates how successful a reaction has been. This measurement is called the percent yield.

To compute the percent yield, it is first necessary to determine how much of the product should be formed based on stoichiometry. This is called the **theoretical yield**, the maximum amount of product that could be formed from the given amounts of reactants. The **actual yield** is the amount of product that is actually formed when the reaction is carried out in the laboratory. The **percent yield** is the ratio of the actual yield to the theoretical yield, expressed as a percentage:

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Percent yield is very important in the manufacture of products. Much time and money is spent improving the percent yield for chemical production. When complex chemicals are synthesized by many different reactions, one step with a low percent yield can quickly cause a large waste of reactants and unnecessary expense.

Typically, percent yields are understandably less than 100% because of the reasons previously indicated. However, percent yields greater than 100% are possible if the measured product of the reaction contains impurities that cause its mass to be greater than it actually would be if the product was pure. When a chemist synthesizes a desired chemical, he or she is always careful to purify the products of the reaction.

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Example 12.7.1: Calculating the Theoretical Yield and the Percent Yield

Potassium chlorate decomposes upon slight heating in the presence of a catalyst, according to the reaction below.

$$2\mathrm{KClO_3}\,(s) \rightarrow 2\mathrm{KCl}\,(s) + 3\mathrm{O_2}\,(g)$$

In a certain experiment, $40.0\ \mathrm{g\ KClO_3}$ is heated until it completely decomposes. What is the theoretical yield of oxygen gas? The experiment is performed, the oxygen gas is collected, and its mass is found to be $14.9\ \mathrm{g}$. What is the percent yield for the reaction?

Solution

First, we will calculate the theoretical yield based on the stoichiometry.

Step 1: List the known quantities and plan the problem.

Known

- Given: Mass of $\mathrm{KClO_3} = 40.0\ \mathrm{g}$
- Molar mass $\mathrm{KClO_3} = 122.55\ \mathrm{g/mol}$
- Molar mass $\mathrm{O_2} = 32.00\ \mathrm{g/mol}$

Unknown

- theoretical yield O_2 = ? g

Apply stoichiometry to convert from the mass of a reactant to the mass of a product:

$$\mathrm{g\ KClO_3} \rightarrow \mathrm{mol\ KClO_3} \rightarrow \mathrm{mol\ O_2} \rightarrow \mathrm{g\ O_2}$$

Step 2: Solve.

$$40.0\ \mathrm{g\ KClO_3} \times \frac{1\ \mathrm{mol\ KClO_3}}{122.55\ \mathrm{g\ KClO_3}} \times \frac{3\ \mathrm{mol\ O_2}}{2\ \mathrm{mol\ KClO_3}} \times \frac{32.00\ \mathrm{g\ O_2}}{1\ \mathrm{mol\ O_2}} = 15.7\ \mathrm{g\ O_2}$$

The theoretical yield of $\mathrm{O_2}$ is $15.7\ \mathrm{g}$.

Step 3: Think about your result.

The mass of oxygen gas must be less than the $40.0\ \mathrm{g}$ of potassium chlorate that was decomposed.

Now we will use the actual yield and the theoretical yield to calculate the percent yield.

Step 1: List the known quantities and plan the problem.

Known

- Actual yield $= 14.9\ \mathrm{g}$
- Theoretical yield $= 15.7\ \mathrm{g}$

Unknown

- Percent yield = ? %

$$\mathrm{Percent\ Yield} = \frac{\mathrm{Actual\ Yield}}{\mathrm{Theoretical\ Yield}} \times 100\%$$

Use the percent yield equation above.

Step 2: Solve.

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$$\text{Percent Yield} = \frac{14.9\text{ g}}{15.7\text{ g}} \times 100\% = 94.9\%$$

Step 3: Think about your result.

Since the actual yield is slightly less than the theoretical yield, the percent yield is just under 100%.

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Summary

- Theoretical yield is calculated based on the stoichiometry of the chemical equation.
- The actual yield is experimentally determined.
- The percent yield is determined by calculating the ratio of actual yield/theoretical yield.

Review

1. What do we need in order to calculate theoretical yield?
2. If I spill some of the product before I weigh it, how will that affect the actual yield?
3. How will spilling some of the product affect the percent yield?
4. I make a product and weigh it before it is dry. How will that affect the actual yield?

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13: THE BEHAVIOR OF GASES

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13.1: Compressibility

When we pack to go on vacation, there seems to always be "one more" thing that we need to get in our suitcase. Maybe it's another bathing suit, pair of shoes, book—whatever the item, we need to get it in. Fortunately, we can usually squeeze things together somehow. Perhaps there is a little space between folds of clothing, or we can rearrange the shoes; somehow we can get that last item in and close the suitcase.

Compressibility

Scuba diving is a form of underwater diving in which a diver carries his own breathing gas, usually in the form of a tank of compressed air. The pressure in most commonly used scuba tanks ranges from 200 to 300 atmospheres. Gases are unlike other states of matter in that a gas expands to fill the shape and volume of its container. For this reason, gases can also be compressed so that a relatively large amount of gas can be forced into a small container. If the air in a typical scuba tank were transferred to a container at the standard pressure of $1\ \mathrm{atm}$, the volume of that container would need to be about 2500 liters.

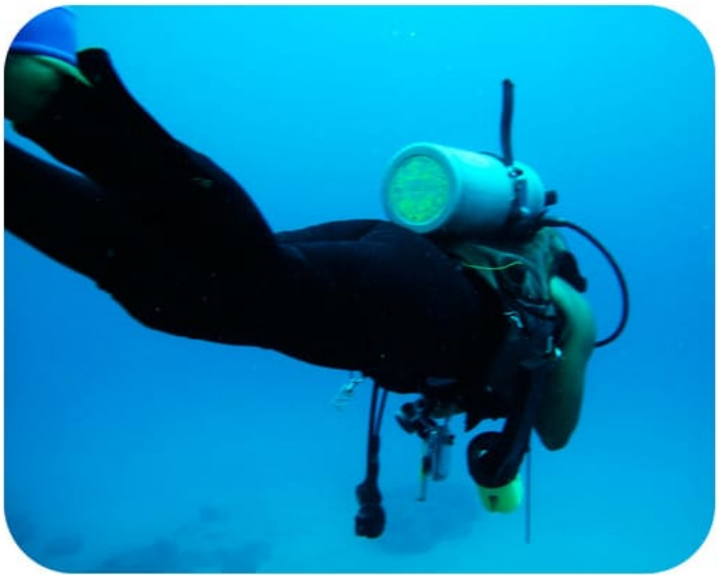


Figure 13.1.1: Scuba diver. (CC BY-NC; CK-12)

Compressibility is the measure of how much a given volume of matter decreases when placed under pressure. If we put pressure on a solid or a liquid, there is essentially no change in volume. The atoms, ions, or molecules that make up the solid or liquid are very close together. There is no space between the individual particles, so they cannot pack together.

The kinetic-molecular theory explains why gases are more compressible than either liquids or solids. Gases are compressible because most of the volume of a gas is composed of the large amounts of empty space between the gas particles. At room temperature and standard pressure, the average distance between gas molecules is about ten times the diameter of the molecules themselves. When a gas is compressed, as when the scuba tank is being filled, the gas particles are forced closer together.

Compressed gases are used in many situations. In hospitals, oxygen is often used for patients who have damaged lungs to help them breathe better. If a patient is having a major operation, the anesthesia that is administered will frequently be a compressed gas. Welding requires very hot flames produced by compressed acetylene and oxygen mixtures. Many summer barbeque grills are fueled by compressed propane.



Figure 13.1.2: *Oxygen tank.* (CC BY-NC; CK-12)

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13.2: Factors Affecting Gas Pressure

The pressure of the air in a basketball has to be adjusted so that the ball bounces to the correct height. Before a game, officials check the ball by dropping it from shoulder height and seeing how far back up it bounces. What would an official do if the ball did not bounce up high enough, or if it bounced too high?

The pressure inside a container is dependent on the amount of gas inside the container. If a basketball does not bounce high enough, an official could remedy the situation by using a hand pump and adding more air to the ball. Conversely, if it bounces too high, the official could let some air out of the ball.

Factors Affecting Gas Pressure

Recall from the kinetic-molecular theory that gas particles move randomly and in straight lines until they elastically collide with either other gas particles, or with one of the walls of the container. It is these collisions with the walls of the container that define the pressure of the gas. Four variables are used to describe the condition of a gas: pressure (P), volume (V), temperature (T), and the amount of the gas as measured by the number of moles (\left(n \right)\). We will examine separately how the volume, temperature, and amount of gas each affect the pressure of an enclosed gas sample.

Amount of Gas

The figure below shows what happens when air is added to a **rigid container**. A rigid container is one that is incapable of expanding or contracting. A steel canister is an example of a rigid container.

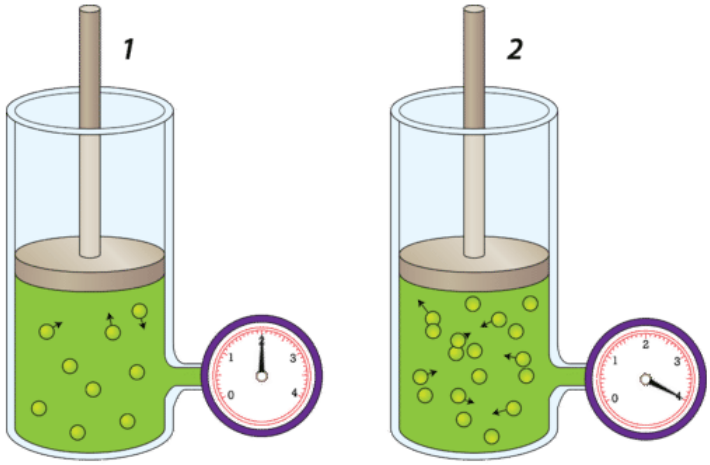


Figure 13.2.1: Increase in pressure with increase in number of gas particles. (CC BY-NC; CK-12)

The canister on the left contains a gas at a certain pressure. The attached air pump is then used to double the amount of gas in the canister. Since the canister cannot expand, the increased number of air molecules will strike the inside walls of the canister twice as frequently as they did before. The result is that the pressure inside the canister doubles. As you might imagine, if more and more air is continually added to a rigid container, it may eventually burst. Reducing the number of molecules in a rigid container has the opposite effect, and the pressure decreases.

Volume

Pressure is also affected by the volume of the container. If the volume of a container is decreased, the gas molecules have less space in which to move around. As a result, they will strike the walls of the container more often, and the pressure increases.

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The figure below shows a cylinder of gas whose volume is controlled by an adjustable piston. On the left, the piston is pulled mostly out and the gauge reads a certain pressure. On the right, the piston has been pushed so that the volume of the enclosed portion of the container where the gas is located has been cut in half. The pressure of the gas doubles. Increasing the volume of the container would have the opposite effect, and the pressure of the gas would decrease.

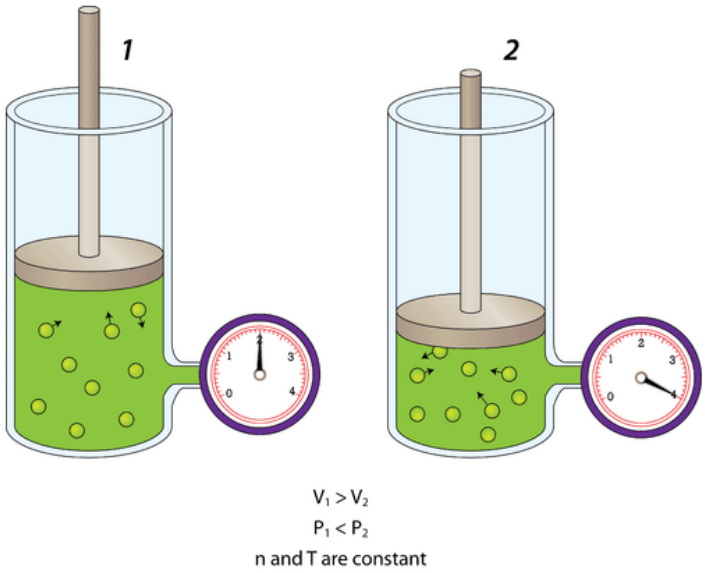


Figure 13.2.2: Decrease in gas volume produced increase in gas pressure. (CC BY-NC; CK-12)

Temperature

It would be very inadvisable to place a can of soup over a campfire without venting the can. As the can heats up, it may explode. The kinetic-molecular theory explains why. The air inside the rigid can of soup is given more kinetic energy by the heat coming from the campfire. The kinetic energy causes the air molecules to move faster; they impact the container walls more frequently and with more force. The increase in pressure inside may eventually exceed the strength of the can and it will explode. An additional factor is that the soup may begin boiling, which will then aid even more gas and more pressure to the inside of the can.

Shown in the figure below is a cylinder of gas (left) that is at room temperature $(300\ \mathrm{K})$. On the right, the cylinder has been heated until the Kelvin temperature has doubled to $600\ \mathrm{K}$. The kinetic energy of the gas molecules increases, so collisions with the walls of the container are now more forceful than they were before. As a result, the pressure of the gas doubles. Decreasing the temperature would have the opposite effect, and the pressure of an enclosed gas would decrease.

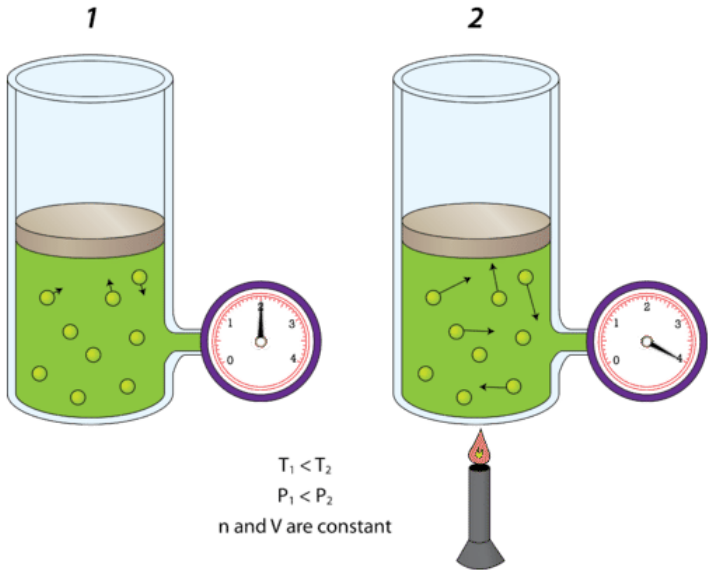


Figure 13.2.3: Increase in temperature produces increase in pressure. (CC BY-NC; CK-12)

Summary

- An increase in the number of gas molecules, while container volume stays constant, increases pressure.
- A decrease in container volume increases gas pressure.
- An increase in temperature of a gas in a rigid container increases the pressure.

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13.3: Boyle's Law

Each day, hundreds of weather balloons are launched. Made of a synthetic rubber and carrying a box of instruments, each helium-filled balloon rises up into the sky. As a balloon gains altitude, the atmospheric pressure becomes less and the balloon expands. At some point the balloon bursts due to the expansion; the instruments drop (aided by a parachute) to be retrieved and studied for information about the weather.

Boyle's Law

Robert Boyle (1627-1691), an English chemist, is widely considered to be one of the founders of the modern experimental science of chemistry. He discovered that doubling the pressure of an enclosed sample of gas, while keeping its temperature constant, caused the volume of the gas to be reduced by half. **Boyle's law** states that the volume of a given mass of gas varies inversely with the pressure when the temperature is kept constant. An inverse relationship is described in this way. As one variable increases in value, the other variable decreases.

Physically, what is happening? The gas molecules are moving and are a certain distance apart from one another. An increase in pressure pushes the molecules closer together, reducing the volume. If the pressure is decreased, the gases are free to move about in a larger volume.



Figure 13.3.1: Robert Boyle. (CC BY-NC; CK-12)

Mathematically, Boyle's law can be expressed by the equation:

$$P \times V = k$$

The k is a constant for a given sample of gas and depends only on the mass of the gas and the temperature. The table below shows pressure and volume data for a set amount of gas at a constant temperature. The third column represents the value of the constant (k) for this data and is always equal to the pressure multiplied by the volume. As one of the variables changes, the other changes in such a way that the product of $P \times V$ always remains the same. In this particular case, that constant is $500 \text{ atm} \cdot \text{mL}$.

Table 13.3.1: Pressure-Volume Data

Pressure (atm)	**Volume** (mL)	$P \times V = k \ (\text{atm} \cdot \text{mL})$
0.5	1000	500
0.625	800	500
1.0	500	500
2.0	250	500
5.0	100	500
8.0	62.5	500

Pressure (atm)	Volume (mL)	$P \times V = k\ (\text{atm} \cdot \text{mL})$
10.0	50	500

A graph of the data in the table further illustrates the inverse relationship nature of Boyle's Law (see figure below). Volume is plotted on the x-axis, with the corresponding pressure on the y-axis.

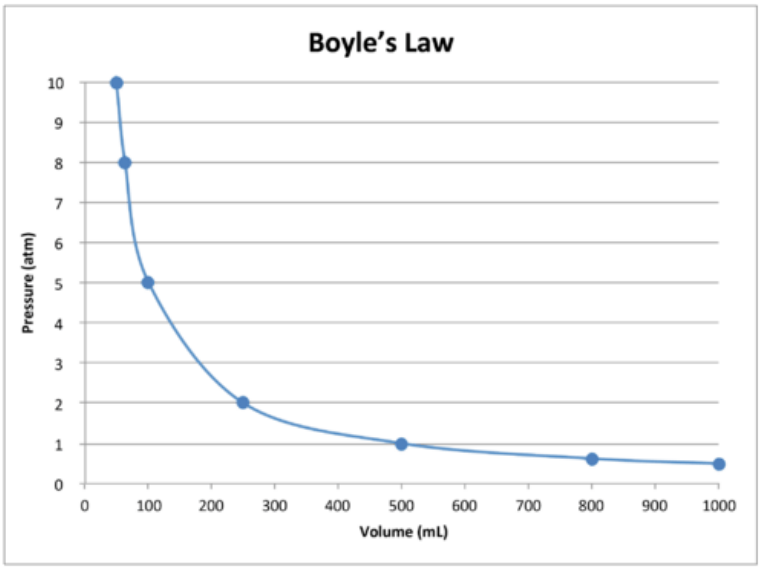


Figure 13.3.2: The pressure of a gas decreases as the volume increases, making Boyle's law an inverse relationship. (CC BY-NC; CK-12)

Boyle's Law can be used to compare changing conditions for a gas. We use P_1 and V_1 to stand for the initial pressure and initial volume of a gas. After a change has been made, P_2 and V_2 stand for the final pressure and volume. The mathematical relationship of Boyle's Law becomes:

$$P_1 \times V_1 = P_2 \times V_2$$

This equation can be used to calculate any one of the four quantities if the other three are known.

Example 13.3.1

A sample of oxygen gas has a volume of $425\ \text{mL}$ when the pressure is equal to $387\ \text{kPa}$. The gas is allowed to expand into a $1.75\ \text{L}$ container. Calculate the new pressure of the gas.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $P_1 = 387\ \text{kPa}$
- $V_1 = 425\ \text{mL}$
- $V_2 = 1.75\ \text{L} = 1750\ \text{mL}$

Unknown

Use Boyle's Law to solve for the unknown pressure (P_2). It is important that the two volumes (V_1 and V_2) are expressed in the same units, so V_2 has been converted to mL.

Step 2: Solve.

First, rearrange the equation algebraically to solve for P_2.

$$P_2 = \frac{P_1 \times V_1}{V_2}$$

Now substitute the known quantities into the equation and solve.

$$P_2 = \frac{387\ \text{kPa} \times 425\ \text{mL}}{1750\ \text{mL}} = 94.0\ \text{kPa}$$

Step 3: Think about your result.

The volume has increased to slightly over 4 times its original value and so the pressure is decreased by about one fourth. The pressure is in kPa and the value has three significant figures. Note that any pressure or volume units can be used as long as they are consistent throughout the problem.

Summary

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13.4: Charles's Law

Freshly-baked bread is light and fluffy as a result of the action of yeast on sugar. The yeast converts the sugar to carbon dioxide, which at high temperatures causes the dough to expand. The end result is an enjoyable treat, especially when covered with melted butter.

Charles's Law

French physicist Jacques Charles (1746-1823) studied the effect of temperature on the volume of a gas at constant pressure. **Charles's Law** states that the volume of a given mass of gas varies directly with the absolute temperature of the gas when pressure is kept constant. The absolute temperature is temperature measured with the Kelvin scale. The Kelvin scale must be used because zero on the Kelvin scale corresponds to a complete stoppage of molecular motion.

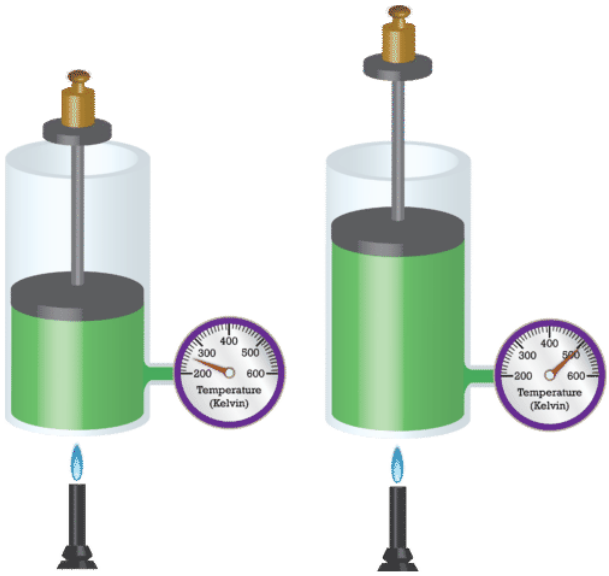


Figure 13.4.1: As a container of confined gas is heated, its molecules increase in kinetic energy and push the movable piston outward, resulting in an increase in volume. (CC BY-NC; CK-12)

Mathematically, the direct relationship of Charles's Law can be represented by the following equation:

$$\frac{V}{T} = k$$

As with Boyle's Law, k is constant only for a given gas sample. The table below shows temperature and volume data for a set amount of gas at a constant pressure. The third column is the constant for this particular data set and is always equal to the volume divided by the Kelvin temperature.

Table 13.4.1: Temperature-Volume Data

Temperature (K)	**Volume** (mL)	$\frac{V}{T} = k\ \left(\frac{\mathrm{mL}}{\mathrm{K}}\right)$
50	20	0.40
100	40	0.40
150	60	0.40
200	80	0.40
300	120	0.40
500	200	0.40

Temperature (K)	Volume (mL)	$\frac{V}{T} = k\ (\frac{\mathrm{mL}}{\mathrm{K}})$
1000	400	0.40

When this data is graphed, the result is a straight line, indicative of a direct relationship, shown in the figure below.

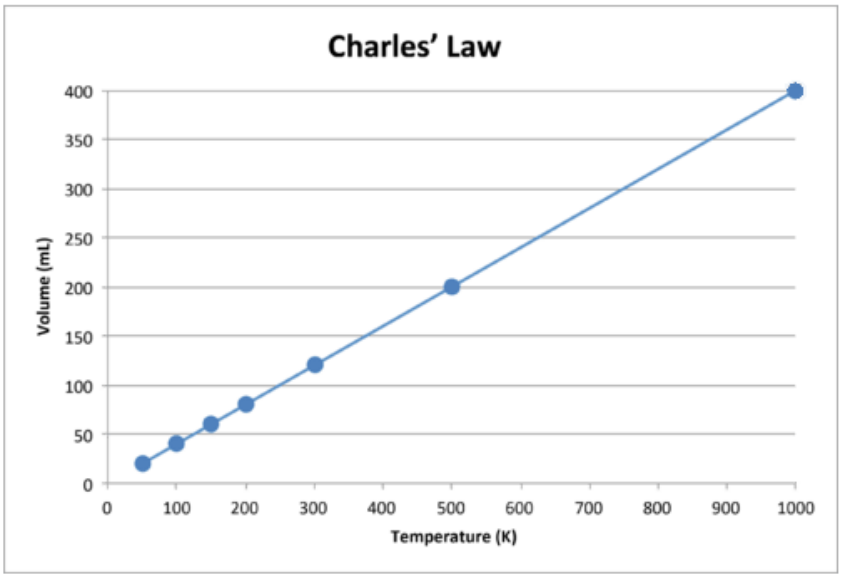


Figure $13.4.2$***:*** *The volume of a gas increases as the Kelvin temperature increases.*

Notice that the line goes exactly toward the origin, meaning that as the absolute temperature of the gas approaches zero, its volume approaches zero. However, when a gas is brought to extremely cold temperatures, its molecules would eventually condense into the liquid state before reaching absolute zero. The temperature at which this change into the liquid state occurs varies for different gases.

Charles's Law can also be used to compare changing conditions for a gas. Now we use V_1 and T_1 to stand for the initial volume and temperature of a gas, while V_2 and T_2 stand for the final volume and temperature. The mathematical relationship of Charles's Law becomes:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

This equation can be used to calculate any one of the four quantities if the other three are known. The direct relationship will only hold if the temperatures are expressed in Kelvin. Temperatures in Celsius will not work. Recall the relationship that $\mathrm{K} = {}^{\circ}\mathrm{C} + 273$.

Example $13.4.1$

A balloon is filled to a volume of $2.20\ \mathrm{L}$ at a temperature of $22^{\circ}\mathrm{C}$. The balloon is then heated to a temperature of $71^{\circ}\mathrm{C}$. Find the new volume of the balloon.

Step 1: List the known quantities and plan the problem.

Known

- $V_1 = 2.20\ \mathrm{L}$
- $T_1 = 22^{\circ}\mathrm{C} = 295\ \mathrm{K}$
- $T_2 = 71^{\circ}\mathrm{C} = 344\ \mathrm{K}$

Unknown

Use Charles's Law to solve for the unknown volume (V_2). The temperatures have first been converted to Kelvin.

Step 2: Solve.

First, rearrange the equation algebraically to solve for V_2.

$$V_2 = \frac{V_1 \times T_2}{T_1}$$

Now substitute the known quantities into the equation and solve.

$$V_2 = \frac{2.20\ \text{L} \times 344\ \text{K}}{295\ \text{K}} = 2.57\ \text{L}$$

Step 3: Think about your result.

The volume increases as the temperature increases. The result has three significant figures.

Summary

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13.5: Gay-Lussac's Law

Propane tanks are widely used with barbeque grills. However, it's not fun to find out half-way through grilling that you've run out of gas. You can buy gauges that measure the pressure inside the tank to see how much is left. The gauge measures pressure and will register a higher pressure on a hot day than it will on a cold day. So, you need to take the air temperature into account when you decide whether or not to refill the tank before your next cook-out.

Gay-Lussac's Law

When the temperature of a sample of gas in a rigid container is increased, the pressure of the gas increases as well. The increase in kinetic energy results in the molecules of gas striking the walls of the container with more force, resulting in a greater pressure. The French chemist Joseph Gay-Lussac (1778-1850) discovered the relationship between the pressure of a gas and its absolute temperature. **Gay-Lussac's Law** states that the pressure of a given mass of gas varies directly with the absolute temperature of the gas, when the volume is kept constant. Gay-Lussac's Law is very similar to Charles's Law, with the only difference being the type of container. Whereas the container in a Charles's Law experiment is flexible, it is rigid in a Gay-Lussac's Law experiment.



Figure 13.5.1: *Joseph Gay-Lussac.* (CC BY-NC; CK-12)

The mathematical expressions for Gay-Lussac's Law are likewise similar to those of Charles's Law:

$$\frac{P}{T} \quad \text{and} \quad \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

A graph of pressure vs. temperature also illustrates a direct relationship. As a gas is cooled at constant volume, its pressure continually decreases until the gas condenses to a liquid.

Example 13.5.1

The gas in an aerosol can is under a pressure of 3.00 atm at a temperature of 25°C. It is dangerous to dispose of an aerosol can by incineration. What would the pressure in the aerosol can be at a temperature of 845°C?

Solution

Known

- $P_1 = 3.00 \text{ atm}$
- $T_1 = 25^{\circ}\text{C} = 298 \text{ K}$
- $T_2 = 845^{\circ}\text{C} = 1118 \text{ K}$

Unknown

Use Gay-Lussac's Law to solve for the unknown pressure (P_2). The temperatures have first been converted to Kelvin.

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Step 2: Solve.

First, rearrange the equation algebraically to solve for P_2.

$$P_2 = \frac{P_1 \times T_2}{T_1}$$

Now substitute the known quantities into the equation and solve.

$$P_2 = \frac{3.00\ \text{atm} \times 1118\ \text{K}}{298\ \text{K}} = 11.3\ \text{atm}$$

Step 3: Think about your result.

The pressure increases dramatically due to large increase in temperature.

Summary

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13.6: Combined Gas Law

The modern refrigerator takes advantage of the gas laws to remove heat from a system. Compressed gas in the coils is allowed to expand. This expansion lowers the temperature of the gas and transfers heat energy from the material in the refrigerator to the gas. As the gas is pumped through the coils, the pressure on the gas compresses it and raises the gas temperature. This heat is then dissipated through the coils into the outside air. As the compressed gas is pumped through the system again, the process repeats itself.

Combined Gas Law

To this point, we have examined the relationships between any two of the variables of P, V, and T, while the third variable is held constant. However, situations do arise where all three variables change. The **combined gas law** expresses the relationship between the pressure, volume, and absolute temperature of a fixed amount of gas. For a combined gas law problem, only the amount of gas is held constant.

$$\frac{P \times V}{T} = k \text{ and } \frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$$

Example 13.6.1

$2.00\ \text{L}$ of a gas at 35°C and $0.833\ \text{atm}$ is brought to standard temperature and pressure (STP). What will be the new gas volume?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $P_1 = 0.833\ \text{atm}$
- $V_1 = 2.00\ \text{L}$
- $T_1 = 35^\circ\text{C} = 308\ \text{K}$
- $P_2 = 1.00\ \text{atm}$
- $T_2 = 0^\circ\text{C} = 273\ \text{K}$

Unknown

Use the combined gas law to solve for the unknown volume (V_2). STP is $273\ \text{K}$ and $1\ \text{atm}$. The temperatures have been converted to Kelvin.

Step 2: Solve.

First, rearrange the equation algebraically to solve for V_2.

$$V_2 = \frac{P_1 \times V_1 \times T_2}{P_2 \times T_1}$$

Now substitute the known quantities into the equation and solve.

$$V_2 = \frac{0.833\ \text{atm} \times 2.00\ \text{L} \times 273\ \text{K}}{1.00\ \text{atm} \times 308\ \text{K}} = 1.48\ \text{L}$$

Step 3: Think about your result.

Both the increase in pressure and the decrease in temperature cause the volume of the gas sample to decrease. Since both changes are relatively small, the volume does not decrease dramatically.

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It may seem challenging to remember all the different gas laws introduced so far. Fortunately, Boyle's, Charles's, and Gay-Lussac's laws can all be easily derived from the combined gas law. For example, consider a situation where a change occurs in the volume and pressure of a gas while the temperature is being held constant. In that case, it can be said that $T_1 = T_2$. Look at the combined gas law and cancel the T variable out from both sides of the equation. What is left over is Boyle's Law: $P_1 \times V_1 = P_2 \times V_2$. Likewise, if the pressure is constant, then $P_1 = P_2$ and cancelling P out of the equation leaves Charles's Law. If the volume is constant, then $V_1 = V_2$ and cancelling V out of the equation leaves Gay-Lussac's Law.

Summary

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

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13.7: Avogadro's Law

A flat tire is not very useful. It does not cushion the rim of the wheel and creates a very uncomfortable ride. When air is added to the tire, the pressure increases as more molecules of gas are forced into the rigid tire. The amount of air that should be put into a tire depends on the pressure rating for that tire. Too little pressure and the tire will not hold its shape. Too much pressure and the tire could burst.

Avogadro's Law

You have learned about Avogadro's hypothesis: equal volumes of any gas at the same temperature and pressure contain the same number of molecules. It follows that the volume of a gas is directly proportional to the number of moles of gas present in the sample. **Avogadro's Law** states that the volume of a gas is directly proportional to the number of moles of gas, when the temperature and pressure are held constant. The mathematical expression of Avogadro's Law is:

$$V = k \times n \quad \text{and} \quad \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

(Where n is the number of moles of gas and k is a constant). Avogadro's Law is in evidence whenever you blow up a balloon. The volume of the balloon increases as you add moles of gas to the balloon by blowing it up.

If the container holding the gas is rigid rather than flexible, pressure can be substituted for volume in Avogadro's Law. Adding gas to a rigid container makes the pressure increase.

Example 13.7.1

A balloon has been filled to a volume of $1.90\ \text{L}$ with $0.0920\ \text{mol}$ of helium gas. If $0.0210\ \text{mol}$ of additional helium is added to the balloon while the temperature and pressure are held constant, what is the new volume of the balloon?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $V_1 = 1.90\ \text{L}$
- $n_1 = 0.0920\ \text{mol}$
- $n_2 = 0.0920 + 0.0210 = 0.1130\ \text{mol}$

Unknown

Note that the final number of moles has to be calculated by adding the original number of moles to the moles of added helium. Use Avogadro's Law to solve for the final volume.

Step 2: Solve.

First, rearrange the equation algebraically to solve for V_2.

$$V_2 = \frac{V_1 \times n_2}{n_1}$$

Now substitute the known quantities into the equation and solve.

$$V_2 = \frac{1.90\ \text{L} \times 0.1130\ \text{mol}}{0.0920\ \text{mol}} = 2.33\ \text{L}$$

Step 3: Think about your result.

Since a relatively small amount of additional helium was added to the balloon, its volume increases slightly.

Summary

- Avogadro's Law states that the volume of a gas is directly proportional to the number of moles of gas, when the temperature and pressure are held constant.
- Calculations are shown for relationships between volume and number of moles of a gas.

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13.8: Ideal Gas Law

There are a number of chemical reactions that require ammonia. In order to carry out a reaction efficiently, we need to know how much ammonia we have for stoichiometric purposes. Using gas laws, we can determine the number of moles present in the tank if we know the volume, temperature, and pressure of the system.

Ideal Gas Law

The combined gas law shows that the pressure of a gas is inversely proportional to volume and directly proportional to temperature. Avogadro's Law shows that volume or pressure is directly proportional to the number of moles of gas. Putting these laws together gives us the following equation:

$$\frac{P_1 \times V_1}{T_1 \times n_1} = \frac{P_2 \times V_2}{T_2 \times n_2}$$

As with the other gas laws, we can also say that $\frac{(P \times V)}{(T \times n)}$ is equal to a constant. The constant can be evaluated provided that the gas being described is considered to be ideal.

The **ideal gas law** is a single equation which relates the pressure, volume, temperature, and number of moles of an ideal gas. If we substitute in the variable R for the constant, the equation becomes:

$$\frac{P \times V}{T \times n} = R$$

The ideal gas law is conveniently rearranged to look this way, with the multiplication signs omitted:

$$PV = nRT$$

The variable R in the equation is called the **ideal gas constant**.

Evaluating the Ideal Gas Constant

The value of R, the ideal gas constant, depends on the units chosen for pressure, temperature, and volume in the ideal gas equation. It is necessary to use Kelvin for the temperature and it is conventional to use the SI unit of liters for the volume. However, pressure is commonly measured in one of three units: kPa, atm, or $\mathrm{mm\ Hg}$. Therefore, R can have three different values.

We will demonstrate how R is calculated when the pressure is measured in kPa. Recall that the volume of $1.00\ \mathrm{mol}$ of any gas at STP is measured to be $22.414\ \mathrm{L}$. We can substitute $101.325\ \mathrm{kPa}$ for pressure, $22.414\ \mathrm{L}$ for volume, and $273.15\ \mathrm{K}$ for temperature into the ideal gas equation and solve for R.

$$R = \frac{PV}{nT} = \frac{101.325\ \mathrm{kPa} \times 22.414\ \mathrm{L}}{1.000\ \mathrm{mol} \times 273.15\ \mathrm{K}} = 8.314\ \mathrm{kPa \cdot L/K \cdot mol}$$

This is the value of R that is to be used in the ideal gas equation when the pressure is given in kPa. The table below shows a summary of this and the other possible values of R. It is important to choose the correct value of R to use for a given problem.

Table 13.8.1*: Values of the Ideal Gas Constant*

Unit of P	Unit of V	Unit of n	Unit of T	Value and Unit of R
kPa	L	mol	K	$8.314\ \mathrm{J/K \cdot mol}$
atm	L	mol	K	$0.08206\ \mathrm{L \cdot atm/K \cdot mol}$
$\mathrm{mm\ Hg}$	L	mol	K	$62.36\ \mathrm{L \cdot mm\ Hg/K \cdot mol}$

Notice that the unit for R when the pressure is in kPa has been changed to $\mathrm{J/K \cdot mol}$. A kilopascal multiplied by a liter is equal to the SI unit for energy, a joule (J).

Example 13.8.1

What volume is occupied by $3.760\ \text{g}$ of oxygen gas at a pressure of $88.4\ \text{kPa}$ and a temperature of 19°C? Assume the oxygen is an ideal gas.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $P = 88.4\ \text{kPa}$
- $T = 19^{\circ}\text{C} = 292\ \text{K}$
- Mass $\text{O}_2 = 3.760\ \text{g}$
- $\text{O}_2 = 32.00\ \text{g/mol}$
- $R = 8.314\ \text{J/K} \cdot \text{mol}$

Unknown

In order to use the ideal gas law, the number of moles of $\text{O}_2\ (n)$ must be found from the given mass and the molar mass. Then, use $PV = nRT$ to solve for the volume of oxygen.

Step 2: Solve.

$$3.760\ \text{g} \times \frac{1\ \text{mol O}_2}{32.00\ \text{g O}_2} = 0.1175\ \text{mol O}_2$$

Rearrange the ideal gas law and solve for V.

$$V = \frac{nRT}{P} = \frac{0.1175\ \text{mol} \times 8.314\ \text{J/K} \cdot \text{mol} \times 292\ \text{K}}{88.4\ \text{kPa}} = 3.23\ \text{L O}_2$$

Step 3: Think about your result.

The number of moles of oxygen is far less than one mole, so the volume should be fairly small compared to molar volume $(22.4\ \text{L/mol})$ since the pressure and temperature are reasonably close to standard. The result has three significant figures because of the values for T and P. Since a joule $(\text{J}) = \text{kPa} \cdot \text{L}$, the units cancel out correctly, leaving a volume in liters.

Summary

- The ideal gas constant is calculated.
- An example of calculations using the ideal gas law is shown.

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13.9: Calculating the Molar Mass of a Gas

Helium has long been used in balloons and blimps. Since it is much less dense than air, it will float above the ground. Small balloons filled with helium are often affordable and available at stores, but large ones are much more expensive (and require a lot more helium).

Calculating Molar Mass and Density of a Gas

A chemical reaction, which produces a gas, is performed. The produced gas is then collected and its mass and volume are determined. The molar mass and volume are determined. The molar mass of the unknown gas can be found using the ideal gas law, provided the temperature and pressure of the gas are also known.

Example 13.9.1

A certain reaction occurs, producing an oxide of nitrogen as a gas. The gas has a mass of $1.211\ \text{g}$ and occupies a volume of $677\ \text{mL}$. The temperature in the laboratory is 23°C and the air pressure is $0.987\ \text{atm}$. Calculate the molar mass of the gas and deduce its formula. Assume the gas is ideal.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass $= 1.211\ \text{g}$
- $V = 677\ \text{mL} = 0.677\ \text{L}$
- $T = 23^\circ\text{C} = 296\ \text{K}$
- $P = 0.987\ \text{atm}$
- $R = 0.08206\ \text{L} \cdot \text{atm/K} \cdot \text{mol}$

Unknown

- $n = ?\ \text{mol}$
- Molar Mass $= ?\ \text{g/mol}$

First the ideal gas law will be used to solve for the moles of unknown gas (n). Then the mass of the gas divided by the moles will give the molar mass.

Step 2: Solve.

$$n = \frac{PV}{RT} = \frac{0.987\ \text{atm} \times 0.677\ \text{L}}{0.08206\ \text{L} \cdot \text{atm/K} \cdot \text{mol} \times 296\ \text{K}} = 0.0275\ \text{mol}$$

Now divide g by mol to get the molar mass.

$$\text{molar mass} = \frac{1.211\ \text{g}}{0.0275\ \text{mol}} = 44.0\ \text{g/mol}$$

Since N has a molar mass of $14\ \text{g/mol}$ and O has a molar mass of $16\ \text{g/mol}$, the formula N_2O would produce the correct molar mass.

Step 3: Think about your result.

The R value that corresponds to a pressure in atm was chosen for this problem. The calculated molar mass gives a reasonable formula for dinitrogen monoxide.

Calculating Density of a Gas

The ideal gas law can be used to find the density of a gas at conditions that are not standard. For example, we will determine the density of ammonia gas $(\mathrm{NH_3})$ at $0.913\ \mathrm{atm}$ and $20^\circ\mathrm{C}$, assuming the ammonia is ideal. First, the molar mass of ammonia is calculated to be $17.04\ \mathrm{g/mol}$. Next, assume exactly $1\ \mathrm{mol}$ of ammonia $(n = 1)$ and calculate the volume that such an amount would occupy at the given temperature and pressure.

$$V = \frac{nRT}{P} = \frac{1.00\ \mathrm{mol} \times 0.08206\ \mathrm{L \cdot atm/K \cdot mol} \times 293\ \mathrm{K}}{0.913\ \mathrm{atm}} = 26.3\ \mathrm{L}$$

Now the density can be calculated by dividing the mass of one mole of ammonia by the volume above.

$$\mathrm{Density} = \frac{17.04\ \mathrm{g}}{26.3\ \mathrm{L}} = 0.648\ \mathrm{g/L}$$

As a point of comparison, this density is slightly less than the density of ammonia at STP, which is equal to $\frac{(170.4\ \mathrm{g/mol})}{(22.4\ \mathrm{L/mol})} = 0.761\ \mathrm{g/L}$. It makes sense that the density should be lower compared to that at STP since both the increase in temperature (from $0^\circ\mathrm{C}$ to $20^\circ\mathrm{C}$) and the decrease in pressure (from $1\ \mathrm{atm}$ to $0.913\ \mathrm{atm}$) would cause the $\mathrm{NH_3}$ molecules to spread out a bit further from one another.

Summary

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13.10: Gas Stoichiometry

The Haber cycle reaction of gaseous nitrogen and hydrogen to form ammonia is a critical step in the production of fertilizer from ammonia. It is important to have an excess of the starting materials so that a maximum yield of ammonia can be achieved. By knowing how much ammonia is needed for the manufacture of a batch of fertilizer, the proper amounts of nitrogen and hydrogen gases can be incorporated into the process.

Gas Stoichiometry

You have learned how to use molar volume to solve stoichiometry problems for chemical reactions involving one or more gases at STP. Now, we can use the ideal gas law to expand our treatment of chemical reactions to solve stoichiometry problems for reactions that occur at any temperature and pressure.

Example 13.10.1

What volume of carbon dioxide is produced by the combustion of $25.21\ \text{g}$ of ethanol $(\text{C}_2\text{H}_5\text{OH})$ at 54°C and $728\ \text{mm Hg}$?

Solution

Before using the ideal gas law, it is necessary to write and balance the chemical equation. Recall that in most combustion reactions, the given substance reacts with O_2 to form CO_2 and H_2O. Here is the balanced equation for the combustion of ethanol:

$$\text{C}_2\text{H}_5\text{OH}\,(l) + 3\text{O}_2\,(g) \rightarrow 2\text{CO}_2\,(g) + 3\text{H}_2\text{O}\,(l)$$

Step 1: List the known quantities and plan the problem.

Known

- Mass $\text{C}_2\text{H}_5\text{OH} = 25.21\ \text{g}$
- Molar mass $\text{C}_2\text{H}_5\text{OH} = 46.08\ \text{g/mol}$
- $P = 728\ \text{mm Hg}$
- $T = 54^\circ\text{C} = 327\ \text{K}$

Unknown

The number of moles of carbon dioxide gas is first calculated by stoichiometry. Then, the ideal gas law is used to calculate the volume of CO_2 produced.

Step 2: Solve.

$$25.21\ \text{g}\ \text{C}_2\text{H}_5\text{OH} \times \frac{1\ \text{mol}\ \text{C}_2\text{H}_5\text{OH}}{46.08\ \text{g}\ \text{C}_2\text{H}_5\text{OH}} \times \frac{2\ \text{mol}\ \text{CO}_2}{1\ \text{mol}\ \text{C}_2\text{H}_5\text{OH}} = 1.094\ \text{mol}\ \text{CO}_2$$

The moles of carbon dioxide (n) is now substituted into $PV = nRT$ to solve for the volume.

$$V = \frac{nRT}{P} = \frac{1.094\ \text{mol} \times 62.36\ \text{L} \cdot \text{mm Hg/K} \cdot \text{mol} \times 327\ \text{K}}{728\ \text{mm Hg}} = 30.6\ \text{L}$$

Step 3: Think about your result.

The mass of ethanol is slightly more than one half mole, meaning that the mole ratio results in slightly more than one mole of carbon dioxide being produced. Because of the elevated temperature and reduced pressure compared to STP, the resulting volume is larger than $22.4\ \text{L}$.

Summary

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13.11: Real and Ideal Gases

The behavior of a molecule depends a lot on its structure. Two compounds with the same number of atoms can act very differently. Ethanol $(\mathrm{C_2H_5OH})$ is a clear liquid that has a boiling point of about $79^{\circ}\mathrm{C}$. Dimethylether $(\mathrm{CH_3OCH_3})$ has the same number of carbons, hydrogens, and oxygens, but boils at a much lower temperature $(-25^{\circ}\mathrm{C})$. The difference lies in the amount of intermolecular interaction (strong H-bonds for ethanol, weak van der Waals force for the ether).

Real and Ideal Gases

An ideal gas is one that follows the gas laws at all conditions of temperature and pressure. To do so, the gas needs to completely abide by the kinetic-molecular theory. The gas particles need to occupy zero volume and they need to exhibit no attractive forces whatsoever toward each other. Since neither of those conditions can be true, there is no such thing as an ideal gas. A **real gas** is a gas that does not behave according to the assumptions of the kinetic-molecular theory. Fortunately, at the conditions of temperature and pressure that are normally encountered in a laboratory, real gases tend to behave very much like ideal gases.

Under what conditions then, do gases behave least ideally? When a gas is put under high pressure, its molecules are forced closer together as the empty space between the particles is diminished. A decrease in the empty space means that the assumption that the volume of the particles themselves is negligible is less valid. When a gas is cooled, the decrease in kinetic energy of the particles causes them to slow down. If the particles are moving at slower speeds, the attractive forces between them are more prominent. Another way to view it is that continued cooling of the gas will eventually turn it into a liquid and a liquid is certainly not an ideal gas anymore (see liquid nitrogen in the figure below). In summary, a real gas deviates most from an ideal gas at low temperatures and high pressures. Gases are most ideal at high temperature and low pressure.



Figure $13.11.1$: Nitrogen gas that has been cooled to $77\ \mathrm{K}$ has turned to a liquid and must be stored in a vacuum insulated container to prevent it from rapidly vaporizing. (CC BY-NC; CK-12)

The figure below shows a graph of $\frac{PV}{RT}$ plotted against pressure for $1\ \mathrm{mol}$ of a gas at three different temperatures—$200\ \mathrm{K}$, $500\ \mathrm{K}$, and 1000 \: \text{K}\). An ideal gas would have a value of 1 for that ratio at all temperatures and pressures, and the graph would simply be a horizontal line. As can be seen, deviations from an ideal gas occur. As the pressure begins to rise, the attractive forces cause the volume of the gas to be less than expected and the value of $\frac{PV}{RT}$ drops under 1. Continued pressure increase results in the volume of the particles to become significant and the value of $\frac{PV}{RT}$ rises to greater than 1. Notice that the magnitude of the deviations from ideality is greatest for the gas at $200\ \mathrm{K}$ and least for the gas at $1000\ \mathrm{K}$.

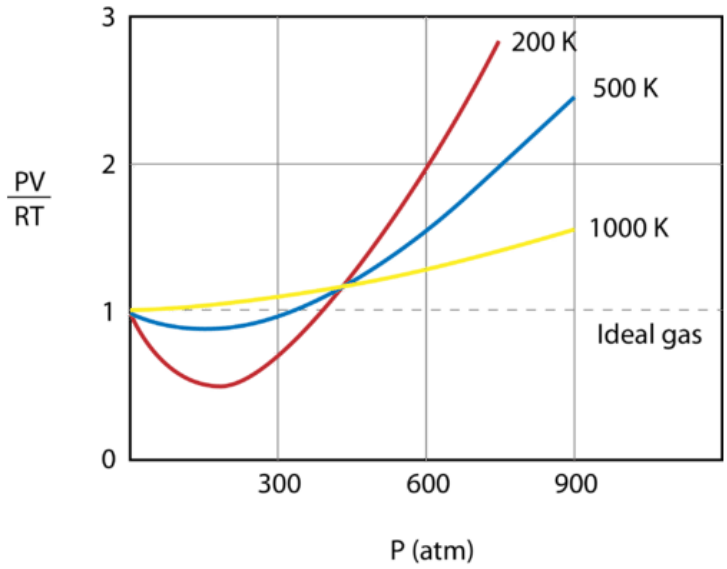


Figure 13.11.2: Real gases deviate from ideal gases at high pressures and low temperatures. (CC BY-NC; CK-12)

The ideality of a gas also depends on the strength and type of intermolecular attractive forces that exist between the particles. Gases whose attractive forces are weak are more ideal than those with strong attractive forces. At the same temperature and pressure, neon is more ideal than water vapor because neon's atoms are only attracted by weak dispersion forces, while water vapor's molecules are attracted by relatively strong hydrogen bonds. Helium is a more ideal gas than neon because its smaller number of electrons means that helium's dispersion forces are even weaker than those of neon.

Summary

- A real gas is a gas that does not behave according to the assumptions of the kinetic-molecular theory.
- The properties of real gases and their deviations from ideality are described.

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13.12: Mole Fraction

Sulfur dioxide is a by-product of many processes, both natural and human-made. Massive amounts of this gas are released during volcanic eruptions. Humans produce sulfur dioxide by burning coal. When in the atmosphere, the gas has a cooling effect by reflecting sunlight away from the earth. However, sulfur dioxide is also a component of smog and acid rain, both of which are harmful to the environment. Many efforts have been made to reduce $\mathrm{SO_2}$ levels to lower acid rain production. However, $\mathrm{SO_2}$ reduction efforts have an unforeseen complication: as we lower the concentration of this gas in the atmosphere, we lower its ability to cool, and consequently have global warming concerns.

Mole Fraction

One way to express relative amounts of substances in a mixture is with the mole fraction. **Mole fraction** X is the ratio of moles of one substance in a mixture to the total number of moles of all substances. For a mixture of two substances, A and B, the mole fractions of each would be written as follows:

$$X_A = \frac{\text{mol A}}{\text{mol A} + \text{mol B}} \quad \text{and} \quad X_B = \frac{\text{mol B}}{\text{mol A} + \text{mol B}}$$

If a mixture consists of $0.50\ \text{mol A}$ and $1.00\ \text{mol B}$, then the mole fraction of A would be $X_A = \frac{0.5}{1.5} + 0.33$. Similarly, the mole fraction of B would be $X_B = \frac{1.0}{1.5} = 0.67$. Mole fraction is a useful quantity for analyzing gas mixtures in conjunction with Dalton's law of partial pressures. Consider the following situation... A 20.0 liter vessel contains $1.0\ \text{mol}$ of hydrogen gas at a pressure of $600\ \text{mm Hg}$. Another 20.0 liter vessels contains $3.0\ \text{mol}$ of helium at a pressure of $1800\ \text{mm Hg}$. These two gases are mixed together in an identical 20.0 liter vessel. Because each will exert its own pressure according to Dalton's law, we can express the partial pressures as follows:

$$P_{H_2} = X_{H_2} \times P_{\text{Total}} \quad \text{and} \quad P_{He} = X_{He} \times P_{\text{Total}}$$

The partial pressure of a gas in a mixture is equal to its mole fraction multiplied by the total pressure. For our mixture of hydrogen and helium:

$$X_{H_2} = \frac{1.0\ \text{mol}}{1.0\ \text{mol} + 3.0\ \text{mol}} = 0.25 \quad \text{and} \quad X_{He} = \frac{3.0\ \text{mol}}{1.0\ \text{mol} + 3.0\ \text{mol}} = 0.75$$

The total pressure according to Dalton's law is $600\ \text{mm Hg} + 1800\ \text{mm Hg} = 2400\ \text{mm Hg}$. So, each partial pressure will be:

$$P_{H_2} = 0.25 \times 2400\ \text{mm Hg} = 600\ \text{mm Hg}$$

$$P_{He} = 0.75 \times 2400\ \text{mm Hg} = 1800\ \text{mm Hg}$$

The partial pressures of each gas in the mixture do not change, since they were mixed into the same size vessel and the temperature was not changed.

Example 13.12.1

A flask contains a mixture of 1.24 moles of hydrogen gas and 2.91 moles of oxygen gas. If the total pressure is $104\ \text{kPa}$, what is the partial pressure of each gas?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $1.24\ \text{mol}\ \mathrm{H_2}$
- $2.91\ \text{mol}\ \mathrm{O_2}$
- $P_{\text{Total}} = 104\ \text{kPa}$

Unknown

- $P_{H_2} = ?\ \text{kPa}$
- $P_{O_2} = ?\ \text{kPa}$

First, the mole fraction of each gas can be determined. Then, the partial pressure can be calculated by multiplying the mole fraction by the total pressure.

Step 2: Solve.

$$X_{H_2} = \frac{1.24\ \text{mol}}{1.24\ \text{mol} + 2.91\ \text{mol}} = 0.299 \qquad X_{O_2} = \frac{2.91\ \text{mol}}{1.24\ \text{mol} + 2.91\ \text{mol}} = 0.701$$
$$P_{H_2} = 0.299 \times 104\ \text{kPa} = 31.1\ \text{kPa} \qquad P_{O_2} = 0.701 \times 104\ \text{kPa} = 72.9\ \text{kPa}$$

Step 3: Think about your result.

The hydrogen is slightly less than one third of the mixture, so it exerts slightly less than one third of the total pressure.

Summary

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13.13: Gas Collection by Water Displacement

Imagine that you need to do a lab experiment where hydrogen gas is generated. In order to calculate the yield of gas, you have to know the pressure inside the tube where the gas is collected. But how can you get a barometer in there? Very simple—you don't. All you need is the atmospheric pressure in the room. As the gas pushes out the water, it is pushing against the atmosphere, so the pressure inside is equal to the pressure outside.

Gas Collection by Water Displacement

Gases that are produced in laboratory experiments are often collected by a technique called **water displacement** (see figure below). A bottle is filled with water and placed upside-down in a pan of water. The reaction flask is fitted with rubber tubing which is then fed under the bottle of water. As the gas is produced in the reaction flask, it exits through the rubber tubing and displaces the water in the bottle. When the bottle is full of the gas, it can be sealed with a lid.



Figure 13.13.1: A gas produced in a chemical reaction can be collected by water displacement. (CC BY-NC; CK-12)

Because the gas is collected over water, it is not pure, but is mixed with vapor from the evaporation of the water. Dalton's law can be used to calculate the amount of the desired gas by subtracting the contribution of the water vapor:

$$P_{\text{Total}} = P_g + P_{H_2O} \quad P_g \text{ is the pressure of the desired gas}$$
$$P_g = P_{Total} - P_{H_2O}$$

In order to solve a problem, it is necessary to know the vapor pressure of water at the temperature of the reaction (see table below). The sample problem illustrates the use of Dalton's law when a gas is collected over water.

Vapor Pressure of Water (mm Hg) *at Selected Temperatures* $(^\circ\text{C})$

Table 13.13.1: Vapor Pressure of Water (mm Hg) at Selected Temperatures $(^\circ\text{C})$

0	5	10	15	20	25	30	35	40	45	50	55	60
4.58	6.54	9.21	12.79	17.54	23.76	31.82	42.18	55.32	71.88	92.51	118.04	149.38

Example 13.13.1

A certain experiment generates $2.58\ \text{L}$ of hydrogen gas, which is collected over water. The temperature is 20°C and the atmospheric pressure is $98.60\ \text{kPa}$. Find the volume that the dry hydrogen would occupy at STP.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $V_{\text{Total}} = 2.58\ \text{L}$
- $T = 20^\circ\text{C} = 293\ \text{K}$

- $P_{\text{Total}} = 98.60 \text{ kPa} = 739.7 \text{ mm Hg}$

Unknown

The atmospheric pressure is converted from kPa to mm Hg in order to match units with the table. The sum of the pressures of the hydrogen and the water vapor is equal to the atmospheric pressure. The pressure of the hydrogen is found by subtraction. Then, the volume of the gas at STP can be calculated by using the combined gas law.

Step 2: Solve.

$$P_{H_2} = P_{\text{Total}} - P_{H_2O} = 739.7 \text{ mm Hg} - 17.54 \text{ mm Hg} = 722.2 \text{ mm Hg}$$

Now the combined gas law is used, solving for V_2, the volume of hydrogen at STP.

$$V_2 = \frac{P_1 \times V_1 \times T_2}{P_2 \times T_1} = \frac{722.2 \text{ mm Hg} \times 2.58 \text{ L} \times 273 \text{ K}}{760 \text{ mm Hg} \times 293 \text{ K}} = 2.28 \text{ L H}_2$$

Step 3: Think about your result.

If the hydrogen gas were to be collected at STP and without the presence of the water vapor, its volume would be 2.28 L. This is less than the actual collected volume because some of that is water vapor. The conversion using STP is useful for stoichiometry purposes.

Summary

- Gases that are produced in laboratory experiments are often collected by a technique called water displacement.
- The vapor pressure due to water in a sample can be corrected for in order to get the true value of the pressure of the gas.

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13.14: Dalton's Law of Partial Pressures

The atmosphere of Venus is markedly different from that of Earth. The gases in the Venusian atmosphere are 96.5% carbon dioxide and 3% nitrogen. The atmospheric pressure on Venus is roughly 92 times that of Earth, so the amount of nitrogen on Venus contributes pressure well over $2700\ \text{mm Hg}$. And there is no oxygen present, so humans could not breathe there. Not that anyone would want to go to Venus—the surface temperature is usually over 460°C.

Dalton's Law of Partial Pressures

Gas pressure results from collisions between gas particles and the inside walls of their container. If more gas is added to a rigid container, the gas pressure increases. The identities of the two gases do not matter. John Dalton, the English chemist who proposed the atomic theory, also studied mixtures of gases. He found that each gas in a mixture exerts a pressure independently of every other gas in the mixture. For example, our atmosphere is composed of about 78% nitrogen and 21% oxygen, with smaller amounts of several other gases making up the rest. Since nitrogen makes up 78% of the gas particles in a given sample of air, it exerts 78% of the pressure. If the overall atmospheric pressure is $1.00\ \text{atm}$, then the pressure of just the nitrogen in the air is $0.78\ \text{atm}$. The pressure of the oxygen in the air is $0.21\ \text{atm}$.

The **partial pressure** of a gas is the contribution that gas makes to the total pressure when the gas is part of a mixture. The partial pressure of nitrogen is represented by P_{N_2}. **Dalton's law of partial pressures** states that the total pressure of a mixture of gases is equal to the sum of all of the partial pressures of the component gases. Dalton's law can be expressed with the following equation:

$$P_{\text{total}} = P_1 + P_2 + P_3 + \cdots$$

The figure below shows two gases that are in separate, equal-sized containers at the same temperature and pressure. Each exerts a different pressure, P_1 and P_2, reflective of the number of particles in the container. On the right, the two gases are combined into the same container, with no volume change. The total pressure of the gas mixture is equal to the sum of the individual pressures. If $P_1 = 300\ \text{mm Hg}$ and $P_2 = 500\ \text{mm Hg}$, then $P_{\text{total}} = 800\ \text{mm Hg}$.

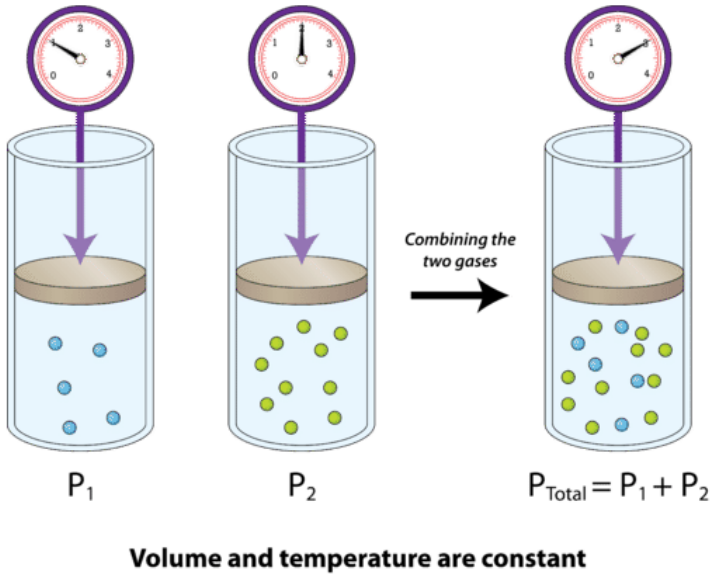


Figure 13.14.1: Dalton's law says that the pressure of a gas mixture is equal to the partial pressures of the combining gases. (CC BY-NC; CK-12)

Summary

- The partial pressure of a gas is the contribution that gas makes to the total pressure when the gas is part of a mixture.
- The total pressure in a system is equal to the sums of the partial pressures of the gases present, according to Dalton's Law

of Partial Pressures.

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13.15: Diffusion and Effusion and Graham's Law

We usually cannot see gases, so we need methods to detect their movements indirectly. The relative rates of diffusion of ammonia to hydrogen chloride can be observed in a simple experiment. Cotton balls are soaked with solutions of ammonia and hydrogen chloride (hydrochloric acid) and attached to two different rubber stoppers. These are simultaneously plugged into either end of a long glass tube. The vapors of each travel down the tube at different rates. Where the vapors meet, they react to form ammonium chloride $(\mathrm{NH_4Cl})$, a white solid that appears in the glass tube as a ring.

Graham's Law

When a person opens a bottle of perfume in one corner of a large room, it doesn't take very long for the scent to spread throughout the entire room. Molecules of the perfume evaporate and the vapor spreads out to fill the entire space. **Diffusion** is the tendency of molecules to move from an area of high concentration to an area of low concentration until the concentration is uniform. While gases diffuse rather quickly, liquids diffuse much more slowly. Solids essentially do not diffuse.

A related process to diffusion is effusion. **Effusion** is the process of a confined gas escaping through a tiny hole in its container. Effusion can be observed by the fact that a helium-filled balloon will stop floating and sink to the floor after a day or so. This is because the helium gas effuses through tiny pores in the balloon. Both diffusion and effusion are related to the speed at which various gas molecules move. Gases that have a lower molar mass effuse and diffuse at a faster rate than gases that have a higher molar mass.

Scottish chemist Thomas Graham (1805-1869) studied the rates of effusion and diffusion of gases. **Graham's law** states that the rate of effusion or diffusion of a gas is inversely proportional to the square root of the molar mass of the gas. Graham's law can be understood by comparing two gases (A and B) at the same temperature, meaning the gases have the same kinetic energy. The kinetic energy of a moving object is given by the equation $KE = \frac{1}{2}mv^2$ where m is mass and v is velocity. Setting the kinetic energies of the two gases equal to one another gives:

$$\frac{1}{2}m_A v_A^2 = \frac{1}{2}m_B v_B^2$$

The equation can be rearranged to solve for the ratio of the velocity of gas A to the velocity of gas B $\left(\frac{v_A}{v_B}\right)$.

$$\frac{v_A^2}{v_B^2} = \frac{m_B}{m_A} \quad \text{which becomes} \quad \frac{v_A}{v_B} = \sqrt{\frac{m_B}{m_A}}$$

For the purposes of comparing the rates of effusion or diffusion of two gases at the same temperature, the molar masses of each gas can be used in the equation for m.

Example 13.15.1

Calculate the ratio of diffusion rates of ammonia gas $(\mathrm{NH_3})$ to hydrogen chloride (HCl) at the same temperature and pressure.

Solution:

Step 1: List the known quantities and plan the problem.

Known

- Molar mass $\mathrm{NH_3} = 17.04\ \mathrm{g/mol}$
- Molar mass $\mathrm{HCl} = 36.46\ \mathrm{g/mol}$

Unknown

Substitute the molar masses of the gases into Graham's law and solve for the ratio.

Step 2: Solve.

$$\frac{v_{NH_3}}{v_{HCl}} = \sqrt{\frac{36.46\ \text{g/mol}}{17.04\ \text{g/mol}}} = 1.46$$

The rate of diffusion of ammonia is 1.46 times faster than the rate of diffusion of hydrogen chloride.

Step 3: Think about your result.

Since ammonia has a smaller molar mass than hydrogen chloride, the velocity of its molecules is greater and the velocity ratio is larger than 1.

Summary

- Diffusion is the tendency of molecules to move from an area of high concentration to an area of low concentration until the concentration is uniform.
- Effusion is the process of a confined gas escaping through a tiny hole in its container.
- Graham's law states that the rate of effusion or diffusion of a gas is inversely proportional to the square root of the molar mass of the gas.

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13.16: Conversions Between Moles and Gas Volume

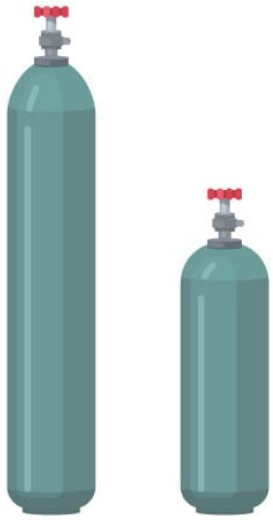


Figure 13.16.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

How can you tell how much gas is in these containers?

Small gas tanks are often used to supply gases for chemistry reactions. A gas gauge will give some information about how much is in the tank, but quantitative estimates are needed so that the reaction will be able to proceed to completion. Knowing how to calculate the necessary parameters for gases is very helpful to avoid running out earlier than desired.

Conversions Between Moles and Gas Volume

Molar volume at STP can be used to convert from moles to gas volume and from gas volume to moles. The equality of $1\ \text{mol} = 22.4\ \text{L}$ is the basis for the conversion factor.

Example 13.16.1: Converting Gas Volume to Moles

Many metals react with acids to produce hydrogen gas. A certain reaction produces $86.5\ \text{L}$ of hydrogen gas at STP. How many moles of hydrogen were produced?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $86.5\ \text{L}\ \text{H}_2$
- $1\ \text{mol} = 22.4\ \text{L}$

Unknown

- moles of H_2

Apply a conversion factor to convert from liters to moles.

Step 2: Calculate.

$$86.5\ \text{L}\ \text{H}_2 \times \frac{1\ \text{mol}\ \text{H}_2}{22.4\ \text{L}\ \text{H}_2} = 3.86\ \text{mol}\ \text{H}_2$$

Step 3: Think about your result.

The volume of gas produced is nearly four times larger than the molar volume. The fact that the gas is hydrogen plays no role in the calculation.

Example 13.16.2**: Convertig Moles to Gas Volume**

What volume does $4.96\ \text{mol}$ of O_2 occupy at STP?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $4.96\ \text{mol}\ \text{O}_2$
- $1\ \text{mol} = 22.4\ \text{L}$

Unknown

- volume of O_2

Step 2: Calculate.

$$4.96\ \text{mol} \times 22.4\ \text{L/mol} = 111.1\ \text{L}$$

Step 3: Think about your result.

The volume seems correct given the number of moles.

Example 13.16.3**: Converting Volume to Mass**

If we know the volume of a gas sample at STP, we can determine how much mass is present. Assume we have $867\ \text{L}$ of N_2 at STP. What is the mass of the nitrogen gas?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $867\ \text{L}\ \text{N}_2$
- $1\ \text{mol} = 22.4\ \text{L}$
- Molar mass $\text{N}_2 = 28.02\ \text{g/mol}$

Unknown

- mass of N_2

Step 2: Calculate.

We start by determining the number of moles of gas present. We know that 22.4 liters of a gas at STP equals one mole, so:

$$867\ \text{L} \times \frac{1\ \text{mol}}{22.4\ \text{L}} = 38.7\ \text{mol}$$

We also know the molecular weight of $\text{N}_2\ (28.0\ \text{g/mol})$, so we can then calculate the weight of nitrogen gas in 867 liters:

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$$38.7 \text{ mol} \times \frac{28 \text{ g}}{1 \text{ mol}} = 1083.6 \text{ g } N_2$$

Step 3: Think about your result.

In a multi-step problem, be sure that the units check out.

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Summary

- Conversions between moles and volume of a gas are shown.

Review

1. In the problems above, why was the gas always at standard temperature and pressure?
2. A container contains 45.2 L of N_2 gas at STP. How many moles of N_2 gas are in the container?
3. If the gas in the previous problem was CH_4 gas at STP instead of N_2 gas, then how many moles of CH_4 gas would there be?

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13.17: Gas Density



Figure 13.17.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

Why does carbon dioxide sink in air?

When we run a reaction to produce a gas, we expect it to rise into the air. Many students have done experiments where gases such as hydrogen are formed. The gas can be trapped in a test tube held upside-down over the reaction. Carbon dioxide, on the other hand, sinks when it is released. Carbon dioxide has a density greater than air, so it will not rise like the hydrogen gas.

Gas Density

As you know, density is defined as the mass per unit volume of a substance. Since gases all occupy the same volume on a per mole basis, the density of a particular gas is dependent on its molar mass. A gas with a small molar mass will have a lower density than a gas with a large molar mass. Gas densities are typically reported in $\mathrm{g/L}$. Gas density can be calculated from molar mass and molar volume.



Figure 13.17.2: Balloons filled with helium gas float in air because the density of helium is less than the density of air. (Public Domain; Photographer: Warren Denning, courtesy of the Pioneer Balloon Company via Wikipedia)

Example 13.17.1: Gas Density

What is the density of nitrogen gas at STP?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\mathrm{N_2} = 28.02\ \mathrm{g/mol}$
- $1\ \mathrm{mol} = 22.4\ \mathrm{L}$

Unknown

- density = ? g/L

Molar mass divided by molar volume yields the gas density at STP.

Step 2: Calculate.

$$\frac{28.02\ \mathrm{g}}{1\ \mathrm{mol}} \times \frac{1\ \mathrm{mol}}{22.4\ \mathrm{L}} = 1.25\ \mathrm{g/L}$$

When set up with a conversion factor, the mol unit cancels, leaving $\mathrm{g/L}$ as the unit in the result.

Step 3: Think about your result.

The molar mass of nitrogen is slightly larger than molar volume, so the density is slightly greater than $1\ \mathrm{g/L}$.

Alternatively, the molar mass of a gas can be determined if the density of the gas at STP is known.

Example 13.17.2: Molar Mass from Gas Density

What is the molar mass of a gas whose density is $0.761\ \mathrm{g/L}$ at STP?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\mathrm{N_2} = 28.02\ \mathrm{g/mol}$
- $1\ \mathrm{mol} = 22.4\ \mathrm{L}$

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Unknown

- molar mass = ? g/mol

Molar mass is equal to density multiplied by molar volume.

Step 2: Calculate.

$$\frac{0.761\ \text{g}}{1\ \text{L}} \times \frac{22.4\ \text{L}}{1\ \text{mol}} = 17.0\ \text{g/mol}$$

Step 3: Think about your result.

Because the density of the gas is less than $1\ \text{g/L}$, the molar mass is less than 22.4.

Summary

- Calculations are described showing conversions between molar mass and density for gases.

Review

1. How is density calculated?
2. How is molar mass calculated?
3. What would be the volume of 3.5 moles of a gas?

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13.18: Volume-Volume Stoichiometry



Figure 13.18.1 (Public Domain; User:Fluffball70/Wikipedia via Wikipedia)

How much propane is left in the tank?

As the weather gets warmer, more and more people want to cook out on the back deck or back yard. Many folks still use charcoal for grilling because of the added flavor. But increasing numbers of back yard cooks like to use a propane grill. The gas burns clean, the grill is ready to go as soon as the flame is lit – but how do you know how much propane is left in the tank? You can buy gauges at hardware stores that measure gas pressure and tell you how much is left in the tank.

Volume-Volume Stoichiometry

Avogadro's hypothesis states that equal volumes of all gases at the same temperature and pressure contain the same number of gas particles. Further, one mole of any gas at standard temperature and pressure (0°C and $1\ \text{atm}$) occupies a volume of $22.4\ \text{L}$. These characteristics make stoichiometry problems involving gases at STP very straightforward. Consider the reaction of nitrogen and oxygen gases to form nitrogen dioxide:

$$\begin{array}{ccccc} \text{N}_2\,(g) & + & 2\text{O}_2\,(g) & \rightarrow & 2\text{NO}_2\,(g) \\ 1\text{ molecule} & & 2\text{ molecules} & & 2\text{ molecules} \\ 1\text{ mol} & & 2\text{ mol} & & 2\text{ mol} \\ 1\text{ volume} & & 2\text{ volumes} & & 2\text{ volumes} \end{array}$$

Because of Avogadro's work, we know that the mole ratios between substances in a gas-phase reaction are also volume ratios. The six possible volume ratios for the above equation are:

$$\frac{1\text{ volume N}_2}{2\text{ volumes O}_2}\text{ or }\frac{2\text{ volumes O}_2}{1\text{ volume N}_2}$$
$$\frac{1\text{ volume N}_2}{2\text{ volumes NO}_2}\text{ or }\frac{2\text{ volumes NO}_2}{1\text{ volume N}_2}$$
$$\frac{2\text{ volumes O}_2}{2\text{ volumes NO}_2}\text{ or }\frac{2\text{ volumes NO}_2}{2\text{ volumes O}_2}$$

The volume ratios above can easily be used when the volume of one gas in a reaction is known, and you need to determine the volume of another gas that will either react with or be produced from the first gas. The pressure and temperature conditions of both gases need to be the same.

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Example 13.18.1: Volume-Volume Stoichiometry

The combustion of propane gas produces carbon dioxide and water vapor.

$$\mathrm{C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(g)}$$

What volume of oxygen is required to completely combust $0.650\ \mathrm{L}$ of propane? What volume of carbon dioxide is produced in the reaction?

Solution:

Step 1: List the known quantities and plan the problem.

Known

- Given: $0.650\ \mathrm{L\ C_3H_8}$
- 1 volume $\mathrm{C_3H_8} = 5$ volumes $\mathrm{O_2}$
- 1 volume $\mathrm{C_3H_8} = 3$ volumes $\mathrm{CO_2}$

Unknown

- Volume $\mathrm{O_2} = ?\ \mathrm{L}$
- Volume $\mathrm{CO_2} = ?\ \mathrm{L}$

Two separate calculations can be done using the volume ratios.

Step 2: Solve.

$$0.650\ \mathrm{L\ C_3H_8} \times \frac{5\ \mathrm{L\ O_2}}{1\ \mathrm{L\ C_3H_8}} = 3.25\ \mathrm{L\ O_2}$$

$$0.650\ \mathrm{L\ C_3H_8} \times \frac{3\ \mathrm{L\ CO_2}}{1\ \mathrm{L\ C_3H_8}} = 1.95\ \mathrm{L\ CO_2}$$

Step 3: Think about your result.

Because the coefficients of the $\mathrm{O_2}$ and the $\mathrm{CO_2}$ are larger than that of the $\mathrm{C_3H_8}$, the volumes for those two gases are greater. Note that total volume is not necessarily conserved in a reaction because moles are not necessarily conserved. In this reaction, 6 total volumes of reactants become 7 total volumes of products.

Summary

- Calculations of volume-volume ratios are based on Avogadro's hypothesis.
- Pressures and temperatures of the gases involved need to be the same.

Review

1. What is Avogadro's hypothesis?
2. How much volume is occupied by one mole of a gas at STP?
3. In the sample problem above, assume we combust 1.3 L of propane. How much CO_2 will be produced?

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13.19: Mass-Volume and Volume-Mass Stoichiometry



Figure 13.19.1 (CK-12 Curriculum Materials license; CK-12 Foundation via CK-12 Foundation)

How much azide is needed to fill an air bag?

Cars and many other vehicles have air bags in them. In case of a collision, a reaction is triggered so that the rapid decomposition of sodium azide produces nitrogen gas, filling the air bag. If too little sodium azide is used, the air bag will not fill completely and will not protect the person in the vehicle. Too much sodium azide could cause the formation of more gas that the bag can safely handle. If the bag breaks from the excess gas pressure, all protection is lost.

Mass to Volume and Volume to Mass Problems

Chemical reactions frequently involve both solid substances whose masses can be measured, as well as gases, for which volume measurements are more appropriate. Stoichiometry problems of this type are called either mass-volume or volume-mass problems.

$$\text{mass of given} \rightarrow \text{moles of given} \rightarrow \text{moles of unknown} \rightarrow \text{volume of unknown}$$

$$\text{volume of given} \rightarrow \text{moles of given} \rightarrow \text{moles of unknown} \rightarrow \text{mass of unknown}$$

Because both types of problems involve a conversion from either moles of gas to volume or vice-versa, we can use the molar volume of $22.4\ \text{L/mol}$, provided that the conditions for the reaction are at STP.

Example 13.19.1: Mass-Volume Stoichiometry

Aluminum metal reacts rapidly with aqueous sulfuric acid to produce aqueous aluminum sulfate and hydrogen gas:

$$2\ \text{Al}\,(s) + 3\text{H}_2\text{SO}_4\,(aq) \rightarrow \text{Al}_2(\text{SO}_4)_3\,(aq) + 3\text{H}_2\,(g)$$

Determine the volume of hydrogen gas produced at STP when a $2.00\ \text{g}$ piece of aluminum completely reacts.

Solution:

Step 1: List the known quantities and plan the problem.

Known

- Given: $2.00\ \text{g}\ \text{Al}$
- Molar mass $\text{Al} = 26.98\ \text{g/mol}$
- $2\ \text{mol}\ \text{Al} = 3\ \text{mol}\ \text{H}_2$

Unknown

- volume H_2 = ?

The grams of aluminum will first be converted to moles. Then the mole ratio will be applied to convert to moles of hydrogen gas. Finally, the molar volume of a gas will be used to convert to liters of hydrogen.

$$\text{g Al} \rightarrow \text{mol Al} \rightarrow \text{mol H}_2 \rightarrow \text{L H}_2$$

Step 2: Solve.

$$2.00 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \times \frac{3 \text{ mol H}_2}{2 \text{ mol Al}} \times \frac{22.4 \text{ L H}_2}{1 \text{ mol H}_2} = 2.49 \text{ L H}_2$$

Step 3: Think about your result.

The volume result is in liters. For much smaller amounts, it may be convenient to convert to milliliters. The answer here has three significant figures. Because the molar volume is a measured quantity of 22.4 L/mol, three is the maximum number of significant figures for this type of problem.

Example $13.19.2$**: Volume-Mass Stoichiometry**

Calcium oxide is used to remove sulfur dioxide generated in coal-burning power plants, according to the following reaction.

$$2\text{CaO}\,(s) + 2\text{SO}_2\,(g) + \text{O}_2\,(g) \rightarrow 2\text{CaSO}_4\,(s)$$

What mass of calcium oxide is required to react completely with $1.4 \times 10^3 \text{ L}$ of sulfur dioxide?

Solution:

Step 1: List the known quantities and plan the problem.

Known

- Given: $1.4 \times 10^3 \text{ L} = \text{SO}_2$
- $2 \text{ mol } \text{SO}_2 = 2 \text{ mol} \text{CaO}$
- Molar mass $\text{CaO} = 56.08 \text{ g/mol}$

Unknown

- mass CaO = ? g

The volume of SO_2 will be converted to moles, followed by the mole ratio, and finally a conversion of moles of CaO to grams.

$$\text{L SO}_2 \rightarrow \text{mol SO}_2 \rightarrow \text{mol CaO} \rightarrow \text{g CaO}$$

Step 2: Solve.

$$1.4 \times 10^3 \text{ L SO}_2 \times \frac{1 \text{ mol SO}_2}{22.4 \text{ L SO}_2} \times \frac{2 \text{ mol CaO}}{2 \text{ mol SO}_2} \times \frac{56.08 \text{ g CaO}}{1 \text{ mol CaO}} = 3.5 \times 10^3 \text{ g CaO}$$

Step 3: Think about your result.

The resultant mass could be reported as 3.5 kg, with two significant figures. Even though the 2:2 mole ratio does not mathematically affect the problem, it is still necessary for unit conversion.

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Summary

- Calculations are described for determining the amount of gas formed in a reaction.
- Calculations are described for determining amounts of a material needed to react with a gas.

Review

1. What are the conditions for all gases in these calculations?
2. How can you tell if all the ratios were set up correctly?
3. Why was 2 mol CaO/2mol SO_2 included in the second example if it did not affect the final number?

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13.20: Mole Road Map



Figure 13.20.1\ (Public Domain; Doug Smith (Wikimedia: Xsmith) via Commons Wikimedia, Currituck Sound Map [commons.wikimedia.org])

How do I get from here to there?

If I want to visit the town of Manteo, North Carolina, out on the coast, I will need a map of how to get there. I may have a printed map or I may download directions from the internet, but I need something to get me going in the right direction. Chemistry road maps serve the same purpose. How do I handle a certain type of calculation? There is a process and a set of directions to help.

Mole Road Map

Previously, we saw how the conversions between mass and number of particles required two steps, with moles as the intermediate. This concept can now be extended to also include gas volume at STP. The resulting diagram is referred to as a mole road map (see figure below).

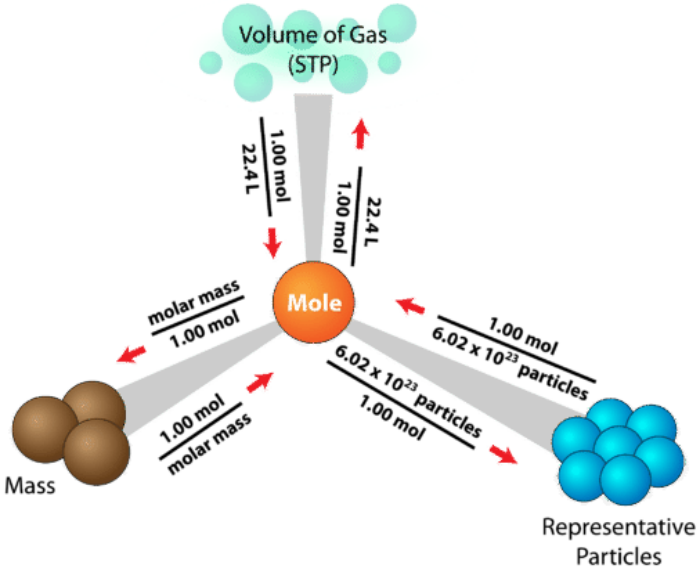


Figure 13.20.2: The mole road map shows the conversion factors needed to interconvert mass, number of particles, and volume of a gas. (CC BY-NC 3.0; Christopher Auyeung via CK-12 Foundation)

The mole is at the center of any calculation involving amount of a substance. The sample problem below is one of many different problems that can be solved using the mole road map.

Example 13.20.1: Mole Road Map

What is the volume of $79.3\ \text{g}$ of neon gas at STP?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\text{Ne} = 20.18\ \text{g/mol}$
- $1\ \text{mol} = 22.4\ \text{L}$

Unknown

- volume = ? L

The conversion factors will be grams → moles → gas volume.

Step 2: Calculate.

$$79.3\ \text{g Ne} \times \frac{1\ \text{mol Ne}}{20.18\ \text{g Ne}} \times \frac{22.4\ \text{L Ne}}{1\ \text{mol Ne}} = 88.0\ \text{L Ne}$$

Step 3: Think about your result.

The given mass of neon is equal to about 4 moles, resulting in a volume that is about 4 times larger than molar volume.

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Summary

- An overall process is given for calculations involving moles, grams, and gas volume.

Review

1. In the problem above, what is the formula weight of neon?
2. What value is at the center of all the calculations?
3. If we had 79.3 grams of Xe, would we expect a volume that is greater than or less than that obtained with neon?

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14: THERMOCHEMISTRY

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14.1: Chemical Potential Energy

Gunpowder was originally developed by the Chinese in the ninth century AD, primarily for rockets. This material is composed of charcoal, sulfur, and saltpeter (potassium nitrate). The explosive reaction that occurs involves the conversion of the charcoal to carbon dioxide, with the potassium nitrate providing the extra oxygen needed for a rapid reaction. Although sulfur was included to stabilize the product, gunpowder is still highly explosive.

Types of Energy

Two basic types of energy exist: potential energy and kinetic energy. **Potential energy** is stored energy. It has not yet been released, but is ready to go. **Kinetic energy** is the energy of motion. It causes work to be done through movement.

Chemical Potential Energy

Energy is the capacity for doing work or supplying heat. When you fill your car with gasoline, you are providing it with potential energy. **Chemical potential energy** is the energy stored in the chemical bonds of a substance. The various chemicals that make up gasoline contain a large amount of chemical potential energy that is released when the gasoline is burned in a controlled way in the engine of the car. The release of that energy does two things: some of the potential energy is transformed into work, which is used to move the car; at the same time, some of the potential energy is converted to heat and makes the car's engine very hot. The energy changes of a system occur as either heat or work, or some combination of both.



Figure 14.1.1: A dragster is able to accelerate because of the chemical potential energy of its fuel. The burning of the fuel also produces large amounts of heat. (CC BY-NC; CK-12)

Dynamite is another example of chemical potential energy. The major component of dynamite is nitroglycerin, a very unstable material. By mixing it with diatomaceous earth, the stability is increased and it is less likely to explode if it receives a physical shock. When ignited, the nitroglycerin explodes rapidly, releasing large amounts of nitrogen and other gases along with a massive amount of heat.



Figure 14.1.2: Dynamite explosion. (CC BY-NC; CK-12)

Summary

- Potential energy is stored energy.
- Kinetic energy is the energy of motion.
- Chemical potential energy is energy available in the chemical bonds of a compound.

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14.2: Heat

Blacksmiths heat solid iron in order to shape it into a variety of different objects. Iron is a rigid, solid metal. At room temperature, it is extremely difficult to bend iron. However, when heated to a high enough temperature, iron can be easily worked. The heat energy in the forge is transferred to the metal, making the iron atoms vibrate more and move around more readily.

Heat is energy that is transferred from one object or substance to another because of a difference in temperature between the two. Heat always flows from an object at a higher temperature to an object at a lower temperature (see figure below). The flow of heat will continue until the two objects are at the same temperature.

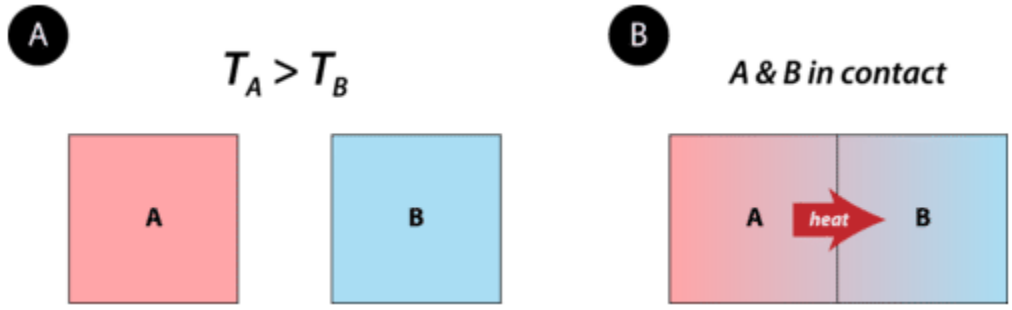


Figure 14.2.1***:*** *Object A starts with a higher temperature than object B. No heat flows when the objects are isolated from each other. When brought into contact, heat flows from A to B until the temperatures of the two objects are the same.*

Thermochemistry is the study of energy changes that occur during chemical reactions and during changes of state. When chemical reactions occur, some chemical bonds are broken, while new chemical bonds form. As a result of the rearrangement of atoms, the total chemical potential energy of the system either increases or decreases.

Summary

- Heat is transferred energy from a site of higher energy to a site of lower energy.
- Thermochemistry is the study of energy changes that occur during chemical reactions and during changes of state.

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14.3: Exothermic and Endothermic Processes

A campfire is an example of basic thermochemistry. 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, roast marshmallows, or simply to keep warm when it's cold outside.

Exothermic and Endothermic Processes

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: *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 are 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.

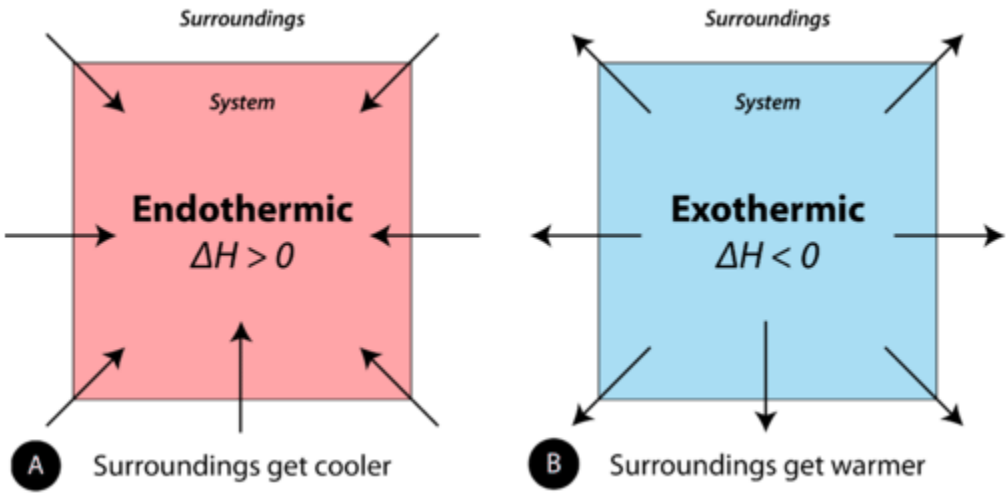


Figure 14.3.1: (A) Endothermic reaction. (B) Exothermic reaction. (CC BY-NC; CK-12)

Units of Heat

Heat flow is measured in one of two common units: the calorie and the joule. The joule (J) is the SI unit of energy. The calorie is familiar because it is commonly used when referring to the amount of energy contained within food. A **calorie** (cal) is the quantity of heat required to raise the temperature of 1 gram of water by $1^{\circ}\mathrm{C}$. For example, raising the temperature of $100\ \mathrm{g}$ of water from $20^{\circ}\mathrm{C}$ to $22^{\circ}\mathrm{C}$ would require $100 \times 2 = 200\ \mathrm{cal}$.

Calories contained within food are actually kilocalories (kcal). In other words, if a certain snack contains 85 food calories, it actually contains $85\ \mathrm{kcal}$ or $85,000\ \mathrm{cal}$. In order to make the distinction, the dietary calorie is written with a capital C.

$$1\ \text{kilocalorie} = 1\ \text{Calorie} = 1000\ \text{calories}$$

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To say that the snack "contains" 85 Calories means that $85\ \text{kcal}$ of energy are released when that snack is processed by the human body.

Heat changes in chemical reactions are typically measured in joules rather than calories. The conversion between a joule and a calorie is shown below.

$$1\ \text{J} = 0.2390\ \text{cal}\ \text{or}\ 1\ \text{cal} = 4.184\ \text{J}$$

We can calculate the amount of heat released in kilojoules when a 400 Calorie hamburger is digested:

$$400.\ \text{Cal} = 400.\ \text{kcal} \times \frac{4.184\ \text{kJ}}{1\ \text{kcal}} = 1.67 \times 10^3\ \text{kJ}$$

Summary

- The law of conservation of energy states that in any physical or chemical process, energy is neither created nor destroyed.
- A specific portion of matter in a given space that is being studied during an experiment or an observation is the *system*.
- The *surroundings* is everything in the universe that is not part of the system.
- A chemical reaction or physical change is endothermic if heat is absorbed by the system from the surroundings.
- A reaction or change is exothermic if heat is released by the system into the surroundings.

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14.4: Heat Capacity and Specific Heat

If a swimming pool and wading pool, both full of water at the same temperature, were subjected to the same input of heat energy, the wading pool would certainly rise in temperature more quickly than the swimming pool. The heat capacity of an object depends both on its mass and its chemical composition. Because of its much larger mass, the swimming pool of water has a larger heat capacity than the wading pool.

Heat Capacity and Specific Heat

Different substances respond to heat in different ways. If a metal chair sits in the bright sun on a hot day, it may become quite hot to the touch. An equal mass of water under the same sun exposure will not become nearly as hot. This means that water has a high **heat capacity** (the amount of heat required to raise the temperature of an object by $1^{\circ}\mathrm{C}$). Water is very resistant to changes in temperature, while metals generally are not. The **specific heat** of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by $1^{\circ}\mathrm{C}$. The table below lists the specific heats of some common substances. The symbol for specific heat is c_p, with the p subscript referring to the fact that specific heats are measured at constant pressure. The units for specific heat can either be joules per gram per degree $(\mathrm{J/g^{\circ}C})$ or calories per gram per degree $(\mathrm{cal/g^{\circ}C})$. This text will use $\mathrm{J/g^{\circ}C}$ for specific heat.

Table 14.4.1*: Specific Heats of Some Common Substances*

Substance	Specific Heat $(\mathrm{J/g^{\circ}C})$
Water (l)	4.18
Water (s)	2.06
Water (g)	1.87
Ammonia (g)	2.09
Ethanol (l)	2.44
Aluminum (s)	0.897
Carbon, graphite (s)	0.709
Copper (s)	0.385
Gold (s)	0.129
Iron (s)	0.449
Lead (s)	0.129
Mercury (l)	0.140
Silver (s)	0.233

Notice that water has a very high specific heat compared to most other substances. 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 14.4.1: 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. (CC BY-NC; CK-12)

Summary

- Heat capacity is the amount of heat required to raise the temperature of an object by $1^{\circ}\mathrm{C}$.
- The specific heat of a substance is the amount of energy required to raise the temperature of 1 gram of the substance by $1^{\circ}\mathrm{C}$.

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14.5: Specific Heat Calculations

Water has a high capacity for absorbing heat. In a car radiator, it serves to keep the engine cooler than it would otherwise run. As the water circulates through the engine, it absorbs heat from the engine block. When it passes through the radiator, the cooling fan and the exposure to the outside environment allow the water to cool somewhat before it makes another passage through the engine.

Specific Heat Calculations

The specific heat of a substance can be used to calculate the temperature change that a given substance will undergo when it is either heated or cooled. The equation that relates heat (q) to specific heat (c_p), mass (m), and temperature change (ΔT) is shown below.

$$q = c_p \times m \times \Delta T$$

The heat that is either absorbed or released is measured in joules. The mass is measured in grams. The change in temperature is given by $\Delta T = T_f - T_i$, where T_f is the final temperature and T_i is the initial temperature.

Example 14.5.1

A $15.0\ \text{g}$ piece of cadmium metal absorbs $134\ \text{J}$ of heat while rising from 24.0°C to 62.7°C. Calculate the specific heat of cadmium.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Heat $= q = 134\ \text{J}$
- Mass $= m = 15.0\ \text{g}$
- $\Delta T = 62.7^\circ\text{C} - 24.0^\circ\text{C} = 38.7^\circ\text{C}$

Unknown

The specific heat equation can be rearranged to solve for the specific heat.

Step 2: Solve.

$$c_p = \frac{q}{m \times \Delta T} = \frac{134\ \text{J}}{15.0\ \text{g} \times 38.7^\circ\text{C}} = 0.231\ \text{J/g}^\circ\text{C}$$

Step 3: Think about your result.

The specific heat of cadmium, a metal, is fairly close to the specific heats of other metals. The result has three significant figures.

Since most specific heats are known, they can be used to determine the final temperature attained by a substance when it is either heated or cooled. Suppose that $60.0\ \text{g}$ of water at 23.52°C was cooled by the removal of $813\ \text{J}$ of heat. The change in temperature can be calculated using the specific heat equation:

$$\Delta T = \frac{q}{c_p \times m} = \frac{813\ \text{J}}{4.18\ \text{J/g}^\circ\text{C} \times 60.0\ \text{g}} = 3.24^\circ\text{C}$$

Since the water was being cooled, the temperature decreases. The final temperature is:

$$T_f = 23.52^\circ\text{C} - 3.24^\circ\text{C} = 20.28^\circ\text{C}$$

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Summary

- The specific heat of a substance can be used to calculate the temperature change of the substance when it is heated or cooled.
- Specific heat calculations are illustrated.

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14.6: Enthalpy

The factors influencing a reaction are complicated and varied. Since a catalyst affects activation energy, we might assume it would have some sort of impact on the amount of heat that is absorbed or released by the reaction—but it does not. The change in heat content of a reaction depends solely on the chemical compositions of the reactants and products, not on the path taken to get from one to the other.

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 these cases, the system is at a constant pressure. **Enthalpy** (H) is the heat content of a system at constant pressure. Chemists routinely measure changes in enthalpy of chemical systems as reactants are converted into products. The heat that is absorbed or released by a reaction at constant pressure is the same as the enthalpy change, and is given the symbol ΔH. Unless otherwise specified, all reactions in this material are assumed to take place at constant pressure.

The change in enthalpy of a reaction is a measure of the differences in enthalpy of the reactants and products. The enthalpy of a system is determined by the energies needed to break chemical bonds and the energies needed to form chemical bonds. Energy needs to be put into the system in order to break chemical bonds—they do not come apart spontaneously in most cases. Bond formation to produce products will involve release of energy. The change in enthalpy shows the trade-offs made in these two processes. Does it take more energy to break bonds than that needed to form bonds? If so, the reaction is endothermic and the enthalpy change is positive. If more energy is produced in bond formation than that needed for bond breaking, the reaction is exothermic and the enthalpy is negative.

Several factors influence the enthalpy of a system. Enthalpy is an extensive property, determined in part by the amount of material being dealt with. 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.

Summary

- Enthalpy, the heat content of a system at constant pressure, is related to the heat of reaction.
- The enthalpy of a system is determined by the energies needed to break chemical bonds and the energies needed to form chemical bonds.
- Factors that influencing enthalpy include: amount of materials, the state of reactants and products, and direction of the reaction.

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14.7: Calorimetry

At one time, calories in foods were measured with a bomb calorimeter. A weighed amount of the food would be placed in the calorimeter and the system was then sealed and filled with oxygen. An electric spark ignited the food-oxygen mixture. The amount of heat released when the food burned gave an idea of the calories present within the food. Today, calories are calculated from the protein, carbohydrate, and fat content of food (all determined by chemical analysis).

Calorimetry

Calorimetry is the measurement of the transfer of heat into or out of a system during a chemical reaction or physical process. A **calorimeter** is an insulated container that is used to measure heat changes. The majority of reactions that can be analyzed in a calorimetry experiment are either liquids or aqueous solutions. A frequently used and inexpensive calorimeter is a set of nested foam cups fitted with a lid to limit the heat exchange between the liquid in the cup and the air in the surroundings (see figure below). In a typical calorimetry experiment, specific volumes of the reactants are dispensed into separate containers and the temperature of each is measured. They are then mixed into the calorimeter, which starts the reaction. The reactant mixture is stirred until the reaction is complete, while the temperature of the reaction is continuously monitored.

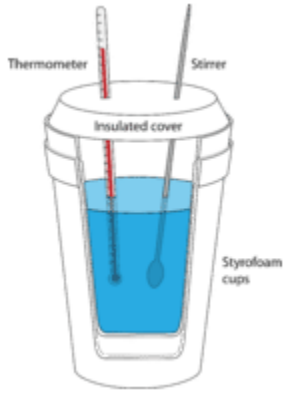


Figure 14.7.1: A simple constant-pressure calorimeter. (CC BY-NC; CK-12)

The key to all calorimetry experiments is the assumption that there is no heat exchange between the insulated calorimeter and the room. Consider the case of a reaction taking place between aqueous reactants: the water in which the solids have been dissolved is the surroundings, while the dissolved substances are the system. The temperature change that is measured is the temperature change that is occurring in the surroundings. If the temperature of the water increases as the reaction occurs, the reaction is exothermic. Heat was released by the system into the surrounding water. An endothermic reaction absorbs heat from the surroundings, so the temperature of the water decreases as heat leaves the surroundings to enter the system.

The temperature change of the water is measured in the experiment and the specific heat of water can be used to calculate the heat absorbed by the surroundings (q_{surr}).

$$q_{\text{surr}} = m \times c_p \times \Delta T$$

In the equation, m is the mass of the water, c_p is the specific heat of the water, and ΔT is $T_f - T_i$. The heat absorbed by the surroundings is equal, but opposite in sign, to the heat released by the system. Because the heat change is determined at constant pressure, the heat released by the system (q_{sys}) is equal to the enthalpy change (ΔH).

$$q_{\text{sys}} = \Delta H = -q_{\text{surr}} = -(m \times c_p \times \Delta T)$$

The sign of ΔH is positive for an endothermic reaction and negative for an exothermic reaction.

Example 14.7.1

In an experiment, $25.0\ \text{mL}$ of $1.00\ \text{M}\ \text{HCl}$ at 25.0°C is added to $25.0\ \text{mL}$ of $1.00\ \text{M}\ \text{NaOH}$ at 25.0°C in a foam cup calorimeter. A reaction occurs and the temperature rises to 32.0°C. Calculate the enthalpy change (ΔH) in kJ for this reaction. Assume the densities of the solutions are $1.00\ \text{g/mL}$ and that their specific heat is the same as that of water.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $c_p = 4.18\ \text{J/g}^\circ\text{C}$
- $V_{\text{final}} = 25.0\ \text{mL} + 25.0\ \text{mL} = 50.0\ \text{mL}$
- $\Delta T = 32.0^\circ\text{C} - 25.0^\circ\text{C} = 7.0^\circ\text{C}$
- Density $= 1.00\ \text{g/mL}$

Unknown

The volume and density can be used to find the mass of the solution after mixing. Then calculate the change in enthalpy by using $\Delta H = q_{\text{sys}} = -q_{\text{surr}} = -(m \times c_p \times \Delta T)$.

Step 2: Solve.

$$m = 50.0\ \text{mL} \times \frac{1.00\ \text{g}}{\text{mL}} = 50.0\ \text{g}$$

$$\Delta H = -(m \times c_p \times \Delta T) = -(50.0\ \text{g} \times 4.18\ \text{J/g}^\circ\text{C} \times 7.0^\circ\text{C}) = -1463\ \text{J} = -1.5\ \text{kJ}$$

Step 3: Think about the result.

The enthalpy change is negative because the reaction releases heat to the surroundings, resulting in an increase in the temperature of the water.

Summary

- Calorimetry is the measurement of the transfer of heat into or out of a system during a chemical reaction or physical process.
- A calorimeter is an insulated container that is used to measure heat changes.
- Calculations involving enthalpy changes are illustrated.

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14.8: Thermochemical Equations

Heating a home is becoming more and more expensive. The decision to use gas, oil, electricity, or wood can be multi-faceted. Part of the decision is based on which fuel will provide the highest amount of energy release when burned. Studies of thermochemistry can be very useful in getting reliable information for making these important choices.

Thermochemical Equation

When methane gas is combusted, heat is released, making the reaction exothermic. Specifically, the combustion of $1\ \text{mol}$ of methane releases 890.4 kilojoules of heat energy. This information can be shown as part of a balanced equation:

$$\text{CH}_4\,(g) + 2\text{O}_2\,(g) \rightarrow \text{CO}_2\,(g) + 2\text{H}_2\text{O}\,(l) + 890.4\ \text{kJ}$$

The equation tells us that $1\ \text{mol}$ of methane combines with $2\ \text{mol}$ of oxygen to produce $1\ \text{mol}$ of carbon dioxide and $2\ \text{mol}$ of water. In the process, $890.4\ \text{kJ}$ is released and so it is written as a product of the reaction. A **thermochemical equation** is a chemical equation that includes the enthalpy change of the reaction. The process in the above thermochemical equation can be shown visually in the figure below.

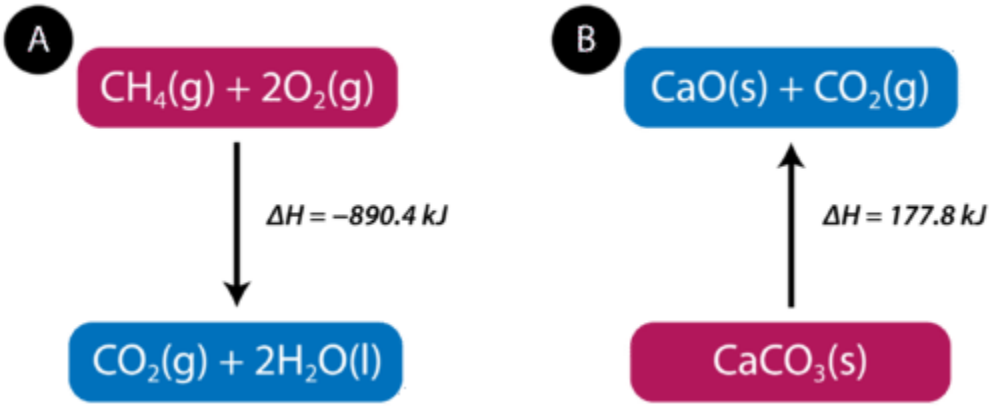


Figure $14.8.1$: *(A) As reactants are converted to products in an exothermic reaction, enthalpy is released into the surroundings. The enthalpy change of the reaction is negative. (B) As reactants are converted to products in an endothermic reaction, enthalpy is absorbed from the surroundings. The enthalpy change of the reaction is positive.* (CC BY-NC; CK-12)

In the combustion of methane example, the enthalpy change is negative because heat is being released by the system. Therefore, the overall enthalpy of the system decreases. The **heat of reaction** is the enthalpy change for a chemical reaction. In the case above, the heat of reaction is $-890.4\ \text{kJ}$. The thermochemical reaction can also be written in this way:

$$\text{CH}_4\,(g) + 2\text{O}_2\,(g) \rightarrow \text{CO}_2\,(g) + 2\text{H}_2\text{O}\,(l) \quad \Delta H = -890.4\ \text{kJ}$$

Heats of reaction are typically measured in kilojoules. It is important to include the physical states of the reactants and products in a thermochemical equation, as the value of the ΔH depends on those states.

Endothermic reactions absorb energy from the surroundings as the reaction occurs. When $1\ \text{mol}$ of calcium carbonate decomposes into $1\ \text{mol}$ of calcium oxide and $1\ \text{mol}$ of carbon dioxide, $177.8\ \text{kJ}$ of heat is absorbed. The process is shown visually in the figure above (B). The thermochemical reaction is shown below.

$$\text{CaCO}_3\,(s) + 177.8\ \text{kJ} \rightarrow \text{CaO}\,(s) + \text{CO}_2\,(g)$$

Because the heat is absorbed by the system, the $177.8\ \text{kJ}$ is written as a reactant. The heat of reaction is positive for an endothermic reaction.

$$\text{CaCO}_3\,(s) \rightarrow \text{CaO}\,(s) + \text{CO}_2\,(g) \quad \Delta H = 177.8\ \text{kJ}$$

The way in which a reaction is written influences the value of the enthalpy change for the reaction. Many reactions are reversible,

meaning that the product(s) of the reaction are capable of combining and reforming the reactant(s). If a reaction is written in the reverse direction, the sign of the ΔH changes. For example, we can write an equation for the reaction of calcium oxide with carbon dioxide to form calcium carbonate:

$$\mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g) \rightarrow \mathrm{CaCO_3}\,(s) + 177.8\,\mathrm{kJ}$$

The reaction is exothermic and thus the sign of the enthalpy change is negative.

$$\mathrm{CaO}\,(s) + \mathrm{CO_2}\,(g) \rightarrow \mathrm{CaCO_3}\,(s) \quad \Delta H = -177.8\,\mathrm{kJ}$$

Summary

- A thermochemical equation is a chemical equation that includes the enthalpy change of the reaction.
- The heat of reaction is the enthalpy change for a chemical reaction.

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14.9: Stoichiometric Calculations and Enthalpy Changes

There is a growing concern about damage to the environment done by emissions from manufacturing plants. Many companies are taking steps to reduce these harmful emissions by adding equipment that will trap the pollutants. In order to know what equipment (and the quantity) to order, studies are done to measure the amount of product currently produced. Since pollution is often both particulate and thermal, energy changes need to be determined in addition to the amounts of products released.

Stoichiometric Calculations and Enthalpy Changes

Chemistry problems that involve enthalpy changes can be solved by techniques similar to stoichiometry problems. Refer again to the combustion reaction of methane. Since the reaction of $1\ \text{mol}$ of methane released $890.4\ \text{kJ}$, the reaction of $2\ \text{mol}$ of methane would release $2 \times 890.4\ \text{kJ} = 1781\ \text{kJ}$. The reaction of $0.5\ \text{mol}$ of methane would release $\frac{890.4\ \text{kJ}}{2} = 445.2\ \text{kJ}$. As with other stoichiometry problems, the moles of a reactant or product can be linked to mass or volume.

Example 14.9.1

Sulfur dioxide gas reacts with oxygen to form sulfur trioxide in an exothermic reaction, according to the following thermochemical equation.

$$2\text{SO}_2\,(g) + \text{O}_2\,(g) \rightarrow 2\text{SO}_3\,(g) + 198\ \text{kJ}$$

Calculate the enthalpy change that occurs when $58.0\ \text{g}$ of sulfur dioxide is reacted with excess oxygen.

Solution:

Step 1: List the known quantities and plan the problem.

- Mass $\text{SO}_2 = 58.0\ \text{g}$
- Molar mass $\text{SO}_2 = 64.07\ \text{g/mol}$
- $\Delta H = -198\ \text{kJ}$ for the reaction of $2\ \text{mol}\ \text{SO}_2$

Unknown

The calculation requires two steps. The mass of SO_2 is converted to moles. Then the moles of SO_2 is multiplied by the conversion factor of $\left(\frac{-198\ \text{kJ}}{2\ \text{mol}\ \text{SO}_2}\right)$.

Step 2: Solve.

$$\Delta H = 58.0\ \text{g}\ \text{SO}_2 \times \frac{1\ \text{mol}\ \text{SO}_2}{64.07\ \text{g}\ \text{SO}_2} \times \frac{-198\ \text{kJ}}{2\ \text{mol}\ \text{SO}_2} = 89.6\ \text{kJ}$$

Step 3: Think about your result.

The mass of sulfur dioxide is slightly less than $1\ \text{mol}$. Since $198\ \text{kJ}$ is released for every $2\ \text{mol}$ of SO_2 that reacts, the heat released when about $1\ \text{mol}$ reacts is one half of 198. The $89.6\ \text{kJ}$ is slightly less than half of 198. The sign of ΔH is negative because the reaction is exothermic.

Summary

- Chemistry problems that involve enthalpy changes can be solved by techniques similar to stoichiometry problems.
- Calculations of energy changes in enthalpy equations are described.

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14.10: Heats of Fusion and Solidification

Suppose that you are holding an ice cube in your hand. It feels cold because heat energy leaves your hand and enters the ice cube. What happens to the ice cube? It melts. However, the temperature during a phase change remains constant. So the heat that is being lost by your hand does not raise the temperature of the ice above its melting temperature of $0^{\circ}\mathrm{C}$. Rather, all the heat goes into the change of state. Energy is absorbed during the process of changing ice into water. The water that is produced also remains at $0^{\circ}\mathrm{C}$ until all of the ice is melted.

Heats of Fusion and Solidification

All solids absorb heat as they melt to become liquids. The gain of heat in this endothermic process goes into changing the state, rather than changing the temperature. The **molar heat of fusion** (ΔH_{fus}) of a substance is the heat absorbed by one mole of that substance as it is converted from a solid to a liquid. Since the melting of any substance absorbs heat, it follows that the freezing of any substance releases heat. The **molar heat of solidification** $(\Delta H_{\text{solid}})$ of a substance is the heat released by one mole of that substance as it is converted from a liquid to a solid. Since fusion and solidification of a given substance are the exact opposite processes, the numerical value of the molar heat of fusion is the same as the numerical value of the molar heat of solidification, but opposite in sign. In other words, $\Delta H_{\text{fus}} = -\Delta H_{\text{solid}}$. The figure below shows all of the possible changes of state along with the direction of heat flow during each process.

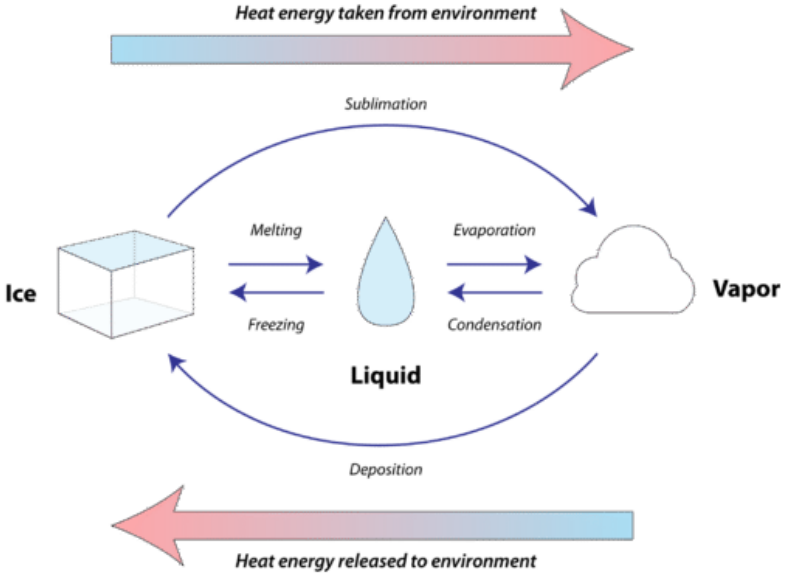


Figure 14.10.1: From left to right, heat is absorbed from the surroundings during melting, evaporation, and sublimation. From right to left, heat is released to the surroundings during freezing, condensation, and deposition. (CC BY-NC; CK-12)

Every substance has a unique value for its molar heat of fusion, depending on the amount of energy required to disrupt the intermolecular forces present in the solid. When $1\ \text{mol}$ of ice at $0^{\circ}\mathrm{C}$ is converted to $1\ \text{mol}$ of liquid water at $0^{\circ}\mathrm{C}$, $6.01\ \text{kJ}$ of heat are absorbed from the surroundings. When $1\ \text{mol}$ of water at $0^{\circ}\mathrm{C}$ freezes to ice at $0^{\circ}\mathrm{C}$, $6.01\ \text{kJ}$ of heat is released into the surroundings.

$$\begin{aligned} \mathrm{H_2O}\,(s) &\rightarrow \mathrm{H_2O}\,(l) \quad \Delta H_{\text{fus}} = 6.01\ \text{kJ/mol} \\ \mathrm{H_2O}\,(l) &\rightarrow \mathrm{H_2O}\,(s) \quad \Delta H_{\text{solid}} = -6.01\ \text{kJ/mol} \end{aligned}$$

The molar heats of fusion and solidification of a given substance can be used to calculate the heat absorbed or released when

various amounts are melted or frozen.

Example 14.10.1

Calculate the heat absorbed when $31.6\ \text{g}$ of ice at 0°C is completely melted.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass $= 31.6\ \text{g}$ ice
- Molar mass $\text{H}_2\text{O} = 18.02\ \text{g/mol}$
- Molar heat of fusion $= 6.01\ \text{kJ/mol}$

Unknown

The mass of ice is first converted to moles. This is then multiplied by the conversion factor of $\left(\frac{6.01\ \text{kJ}}{1\ \text{mol}}\right)$ in order to find the kJ of heat absorbed.

Step 2: Solve.

$$31.6\ \text{g ice} \times \frac{1\ \text{mol ice}}{18.02\ \text{g ice}} \times \frac{6.01\ \text{kJ}}{1\ \text{mol ice}} = 10.5\ \text{kJ}$$

Step 3: Think about your result.

The given quantity is a bit less than 2 moles of ice, and so just less than $12\ \text{kJ}$ of heat is absorbed by the melting process.

Summary

- The molar heat of fusion (ΔH_{fus}) of a substance is the heat absorbed by one mole of that substance as it is converted from a solid to a liquid.
- The molar heat of solidification $(\Delta H_{\text{solid}})$ of a substance is the heat released by one mole of that substance as it is converted from a liquid to a solid.
- Calculations of heat changes during fusion and solidification are described.

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14.11: Heats of Vaporization and Condensation

Natural resources for electric power generation have traditionally been waterfalls, oil, coal, or nuclear power. Research is being carried out to look for other renewable sources to run the generators. Geothermal sites (such as geysers) are being considered because of the steam they produce. Capabilities can be estimated by knowing how much steam is released in a given time at a particular site.

Heat of Vaporization and Condensation

Energy is absorbed in the process of converting a liquid at its boiling point into a gas. As with the melting point of a solid, the temperature of a boiling liquid remains constant and the input of energy goes into changing the state. The **molar heat of vaporization** (ΔH_{vap}) of a substance is the heat absorbed by one mole of that substance as it is converted from a liquid to a gas. As a gas condenses to a liquid, heat is released. The **molar heat of condensation** (ΔH_{cond}) of a substance is the heat released by one mole of that substance as it is converted from a gas to a liquid. Since vaporization and condensation of a given substance are the exact opposite processes, the numerical value of the molar heat of vaporization is the same as the numerical value of the molar heat of condensation, but opposite in sign. In other words, $\Delta H_{\text{vap}} = -\Delta H_{\text{cond}}$.

When $1\ \text{mol}$ of water at 100°C and $1\ \text{atm}$ pressure is converted to $1\ \text{mol}$ of water vapor at 100°C, $40.7\ \text{kJ}$ of heat is absorbed from the surroundings. When $1\ \text{mol}$ of water vapor at 100°C condenses to liquid water at 100°C, $40.7\ \text{kJ}$ of heat is released into the surroundings.

$$\begin{aligned} \text{H}_2\text{O}\,(l) &\rightarrow \text{H}_2\text{O}\,(g) \quad \Delta H_{\text{vap}} = 40.7\,\text{kJ/mol} \\ \text{H}_2\text{O}\,(g) &\rightarrow \text{H}_2\text{O}\,(l) \quad \Delta H_{\text{cond}} = -40.7\,\text{kJ/mol} \end{aligned}$$

Other substances have different values for their molar heats of fusion and vaporization; these substances are summarized in the table below.

Table 14.11.1*: Molar Heats of Fusion and Vaporization*

Substance	$\Delta H_{\text{fus}}\ (\text{kJ/mol})$	$\Delta H_{\text{vap}}\ (\text{kJ/mol})$
Ammonia (NH_3)	5.65	23.4
Ethanol $(\text{C}_2\text{H}_5\text{OH})$	4.60	43.5
Methanol (CH_3OH)	3.16	35.3
Oxygen (O_2)	0.44	6.82
Water (H_2O)	6.01	40.7

Notice that for all substances, the heat of vaporization is substantially higher than the heat of fusion. Much more energy is required to change the state from a liquid to a gas than from a solid to a liquid. This is because of the large separation of the particles in the gas state. The values of the heats of fusion and vaporization are related to the strength of the intermolecular forces. All of the substances in the table above, with the exception of oxygen, are capable of hydrogen bonding. Consequently, the heats of fusion and vaporization of oxygen are far lower than the others.

Example 14.11.1

What mass of methanol vapor condenses to a liquid as $20.0\ \text{kJ}$ of heat is released?

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\Delta H = 20.0\ \text{kJ}$

- $\Delta H_{\text{cond}} = -35.3\ \text{kJ/mol}$
- Molar mass $\text{CH}_3\text{OH} = 32.05\ \text{g/mol}$

Unknown

First the kJ of heat released in the condensation is multiplied by the conversion factor $\left(\frac{1\ \text{mol}}{-35.3\ \text{kJ}}\right)$ to find the moles of methanol that condensed. Then, moles are converted to grams.

Step 2: Solve.

$$-20.0\ \text{kJ} \times \frac{1\ \text{mol}\ \text{CH}_3\text{OH}}{-35.3\ \text{kJ}} \times \frac{32.05\ \text{g}\ \text{CH}_3\text{OH}}{1\ \text{mol}\ \text{CH}_3\text{OH}} = 18.2\ \text{g}\ \text{CH}_3\text{OH}$$

Step 3: Think about your result.

Condensation is an exothermic process, so the enthalpy change is negative. Slightly more than one-half mole of methanol is condensed.

Summary

- The molar heat of vaporization (ΔH_{vap}) is the heat absorbed by one mole of a substance as it is converted from a liquid to a gas.
- The molar heat of condensation (ΔH_{cond}) is the heat released by one mole of a substance as it is converted from a gas to a liquid.
- Examples of calculations involving the molar heat of vaporization and condensation are illustrated.

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14.12: Multi-Step Problems with Changes of State

If you have a cube of ice, which process will take more energy—the melting of that ice cube or the conversion of the water to steam? The short answer is that more energy is needed to convert the water to steam. The long answer is really a series of questions: How do you get from one point to the other? What is the temperature of the ice? What is the mass of that ice cube? A long process is involved to take the material from the starting point to the end point.

Multi-Step Problems with Changes of State

Heating curves show the phase changes that a substance undergoes as heat is continuously absorbed.

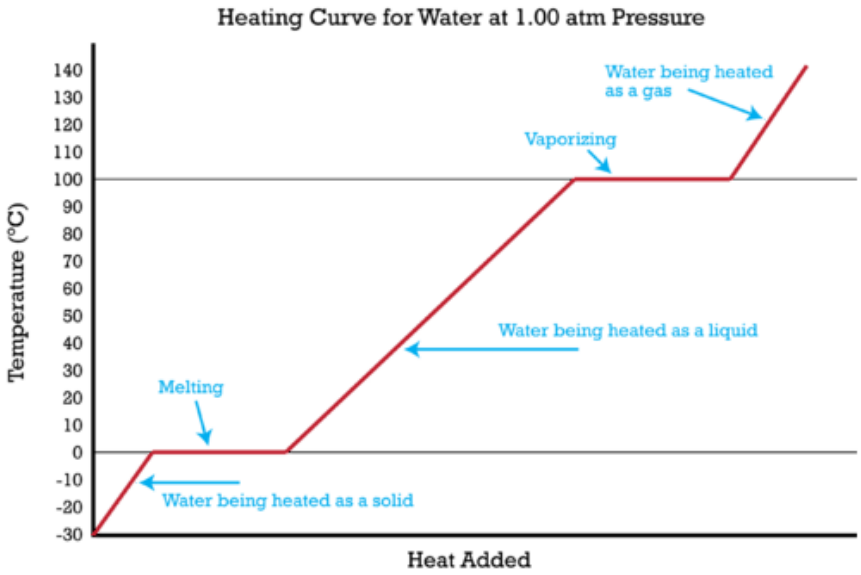


Figure 14.12.1: Heating curve of water. (CC BY-NC; CK-12)

The specific heat of a substance allows us to calculate the heat absorbed or released as the temperature of the substance changes. It is possible to combine that type of problem with a change of state to solve a problem involving multiple steps. The figure above shows ice at -30°C being converted in a five-step process to gaseous water (steam) at 140°C. It is now possible to calculate the heat absorbed during that entire process. The process and the required calculations are summarized below.

1. Ice is heated from -30°C to 0°C. The heat absorbed is calculated by using the specific heat of ice and the equation $\Delta H = c_p \times m \times \Delta T$.

2. Ice is melted at 0°C. The heat absorbed is calculated by multiplying the moles of ice by the molar heat of fusion.

3. Water at 0°C is heated to 100°C. The heat absorbed is calculated by using the specific heat of water and the equation $\Delta H = c_p \times m \times \Delta T$.

4. Water is vaporized to steam at 100°C. The heat absorbed is calculated by multiplying the moles of water by the molar heat of vaporization.

5. Steam is heated from 100°C to 140°C. The heat absorbed is calculated by using the specific heat of steam and the equation $\Delta H = c_p \times m \times \Delta T$.

Example 14.12.1

Calculate the total amount of heat absorbed (in kJ) when $2.00\ \text{mol}$ of ice at -30°C is converted to steam at 140.0°C. The required specific heats can be found in the table in "Heat Capacity and Specific Heat".

Solution

Step 1: List the known quantities and plan the problem.

Known

- $2.00\ \text{mol ice} = 36.04\ \text{g ice}$
- $c_p\ \text{(ice)} = 2.06\ \text{J/g}^\circ\text{C}$
- $c_p\ \text{(water)} = 4.18\ \text{J/g}^\circ\text{C}$
- $c_p\ \text{(steam)} = 1.87\ \text{J/g}^\circ\text{C}$
- $\Delta H_{\text{fus}} = 6.01\ \text{kJ/mol}$
- $\Delta H_{\text{vap}} = 40.7\ \text{kJ/mol}$

Unknown

Follow the steps previously described. Note that the mass of the water is needed for the calculations that involve the specific heat, while the moles of water is needed for the calculations that involve changes of state. All heat quantities must be in kilojoules so that they can be added together to get a total for the five-step process.

Step 2: Solve.

1. $$\Delta H_1 = 2.06\ \text{J/g}^\circ\text{C} \times 36.04\ \text{g} \times 30^\circ\text{C} \times \frac{1\ \text{kJ}}{1000\ \text{J}} = 2.23\ \text{kJ}$$
2. $$\Delta H_2 = 2.00\ \text{mol} \times \frac{6.01\ \text{kJ}}{1\ \text{mol}} = 12.0\ \text{kJ}$$
3. $$\Delta H_3 = 4.18\ \text{J/g}^\circ\text{C} \times 36.04\ \text{g} \times 100^\circ\text{C} \times \frac{1\ \text{kJ}}{1000\ \text{kJ}} = 15.1\ \text{kJ}$$
4. $$\Delta H_4 = 2.00\ \text{mol} \times \frac{40.7\ \text{kJ}}{1\ \text{mol}} = 81.4\ \text{kJ}$$
5. $$\Delta H_5 = 1.87\ \text{J/g}^\circ\text{C} \times 36.04\ \text{g} \times 40^\circ\text{C} \times \frac{1\ \text{kJ}}{1000\ \text{J}} = 2.70\ \text{kJ}$$

$$\Delta H_{\text{total}} = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4 + \Delta H_5 = 113.4\ \text{kJ}$$

Step 3: Think about your result.

The total heat absorbed as the ice at -30°C is heated to steam at 140°C is $133.4\ \text{kJ}$. The largest absorption of heat comes during the vaporization of the liquid water.

Summary

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15: WATER

15.1: Structure of Water

15.2: Structure of Ice

15.3: Physical Properties of Water

15.4: Solute and Solvent

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15.6: Liquid-Liquid Solutions

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15.1: Structure of Water

In his well-known poem "The Rime of the Ancient Mariner", Samuel Coleridge wrote "water, water everywhere, nor any drop to drink." Coleridge was talking about being out on the ocean, but not having any water because he had killed an albatross (apparently bringing bad luck to everyone on the ship). About 75% of the Earth's surface is water. The major constituent of the human body (over 60%) is water. This simple molecule plays important roles in all kinds of processes.

Structure of Water

Water is a simple molecule consisting of one oxygen atom bonded to two different hydrogen atoms. Because of the higher **electronegativity** of the oxygen atom, the bonds are polar covalent (**polar bonds**). The oxygen atom attracts the shared electrons of the covalent bonds to a significantly greater extent than the hydrogen atoms. As a result, the oxygen atom acquires a partial negative charge $(\delta-)$, while the hydrogen atoms each acquire a partial positive charge $(\delta+)$. The molecule adopts a bent structure because of the two lone pairs of electrons on the oxygen atom. The $\mathrm{H-O-H}$ bond angle is about 105°, slightly smaller than the ideal 109.5° of an sp^3 hybridized atomic orbital.

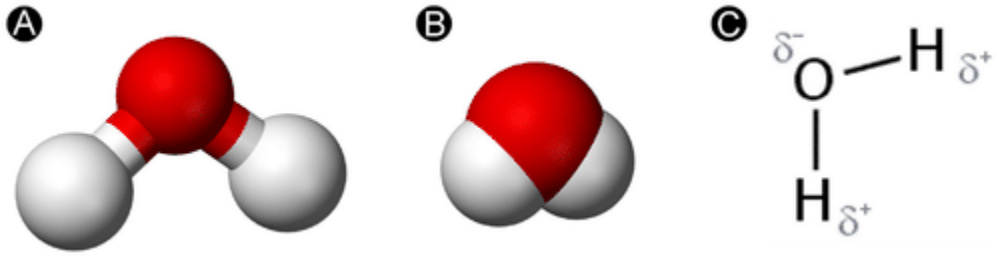


Figure $15.1.1$: The water molecule, visualized three different ways: ball-and-stick model, space-filling model, and structural formula with partial charges.

The bent shape of the water molecule is critical because the polar $\mathrm{O-H}$ bonds do not cancel one another and the molecule as a whole is polar. The figure below illustrates the net polarity of the water molecule. The oxygen is the negative end of the molecule, while the area between the hydrogen atoms is the positive end of the molecule.

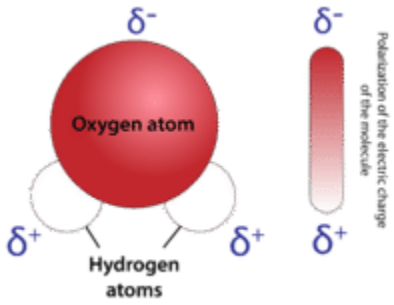


Figure $15.1.2$: Water is a polar molecule, as greater electron density is found around the more electronegative oxygen atom.

Polar molecules attract one another by dipole-dipole forces, as the positive end of one molecule is attracted to the negative end of the nearby molecule. In the case of water, the highly polar $\mathrm{O-H}$ bonds results in very little electron density around the hydrogen atoms. Each hydrogen atom is strongly attracted to the lone-pair electrons on an adjacent oxygen atom. These are called hydrogen bonds and are stronger than conventional dipole-dipole forces.

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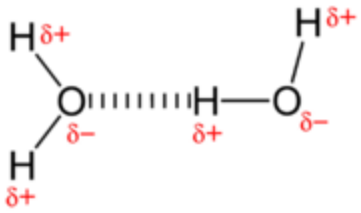


Figure 15.1.3: A hydrogen bond is the attraction between a lone pair of electrons on the oxygen atom of one molecule and the electron-deficient hydrogen atom of a nearby molecule.

Because each oxygen atom has two lone pairs, it can make hydrogen bonds to the hydrogen atoms of two separate other molecules. The figure below shows the result—an approximately tetrahedral geometry around each oxygen atom, consisting of two covalent bonds and two hydrogen bonds.

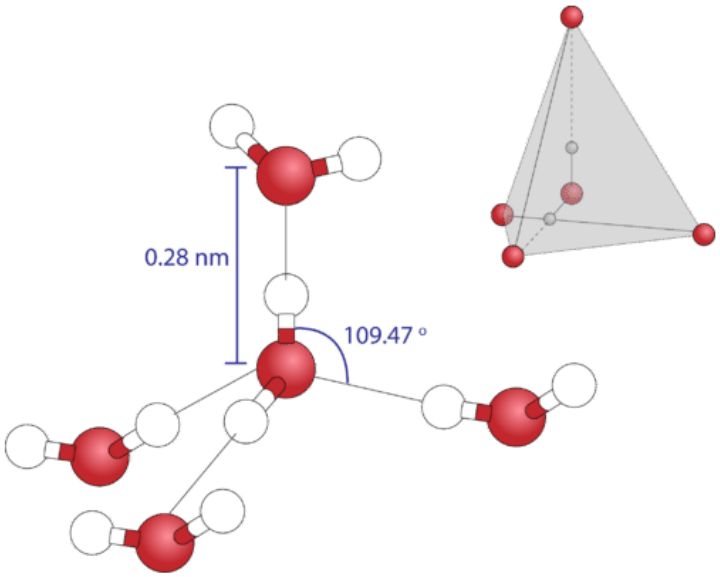


Figure 15.1.4: As a result of two covalent bonds and two hydrogen bonds, the geometry around each oxygen atom is approximately tetrahedral.

Summary

- Water is a molecular compound consisting of polar molecules that have a bent shape.
- The oxygen atom acquires a partial negative charge, while the hydrogen atom acquires a partial positive charge.

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15.2: Structure of Ice

Ice is an interesting and useful material. It can be used to cool food and keep it fresh. Ice can provide recreation, such as in the case of ice-skating. It can do great damage when it freezes—roads can buckle, houses can be damaged, water pipes can burst. All of this happens because of a unique property of water: when water freezes, it expands in volume as ice is formed.

Structure of Ice

Liquid water is a fluid. The hydrogen bonds in liquid water constantly break and reform as the water molecules tumble past one another. As water cools, its molecular motion slows and the molecules move gradually closer to one another. The **density** of any liquid increases as its temperature decreases. For most liquids, this continues as the liquid freezes; the solid state is denser than the liquid state. However, water behaves differently. It actually reaches its highest density at about $4^{\circ}\mathrm{C}$.

Table $15.2.1$*: Density of Water and Ice*

Temperature $(^{\circ}\mathrm{C})$	Density $\left(\mathrm{g/cm^3}\right)$
100 (liquid)	0.9584
50	0.9881
25	0.9971
10	0.9997
4	1.0000
0 (liquid)	0.9998
0 (solid)	0.9168

Between $4^{\circ}\mathrm{C}$ and $0^{\circ}\mathrm{C}$, the density gradually decreases as the hydrogen bonds begin to form a network characterized by a generally **hexagonal** structure with open spaces in the middle of the hexagons (see figure below).

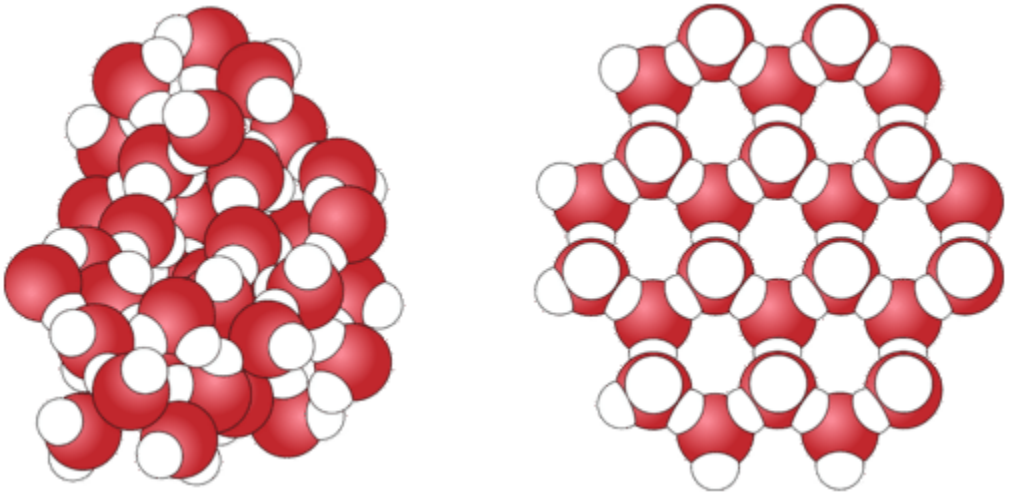


Figure $15.2.1$: The structure of liquid water (left) consists of molecules connected by short-lived hydrogen bonds because water is a fluid. In ice (right), the hydrogen bonds become permanent, resulting in an interconnected hexagonally-shaped framework of molecules.

Ice is less dense than liquid water and so it floats. Ponds or lakes begin to freeze at the surface, closer to the cold air. A layer of ice forms, but does not sink as it would if water did not have this unique structure dictated by its shape, polarity, and hydrogen bonding. If the ice were to sink as it froze, entire lakes would freeze solid. Since the ice does not sink, liquid water remains under the ice all winter long. This is important, as fish and other organisms are capable of surviving through winter. Ice is one of few solids that is less dense than its liquid form.

Summary

- The density of any liquid increases as its temperature decreases. As an exception, ice is less dense than liquid water.
- The intermolecular structure of ice has spaces that are not present in liquid water.

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15.3: Physical Properties of Water

Water loss to the atmosphere is a significant problem in many parts of the world. When water supplies are low, anything that can be done to decrease water loss is important for farmers. An evaporation pan can be used to measure how fast water evaporates in a given location. This information can be used as part of project development to cut down on evaporation, and increase the amount of usable water in a region.

Properties of Water

Compared to other molecular compounds of relatively low molar mass, ice melts at a very high temperature. A great deal of energy is required to break apart the hydrogen-bonded network of ice and return it to the liquid state. Likewise, the boiling point of water is very high. Most molecular compounds of similar molar mass are gases at room temperature.

Surface Tension

Water has a high **surface tension** (attraction between molecules at the surface of a liquid) because of its hydrogen bonding. Liquids that cannot hydrogen bond do not exhibit nearly as much surface tension. Surface tension can be seen by the curved meniscus that forms when water is in a thin column, such as in a graduated cylinder or a buret.



Figure 15.3.1: The meniscus of water in a graduated cylinder forms because of water's hydrogen bonding.

Vapor Pressure

The **vapor pressure** of a liquid is the pressure of the vapor produced by evaporation of a liquid or solid above the liquid or solid in a closed container. The hydrogen bonding between liquid water molecules explains why water has an unusually low vapor pressure. Relatively few molecules of water are capable of escaping the surface of the liquid and entering the vapor phase. Evaporation is slow, and thus the vapor exerts a low pressure in a closed container. Low vapor pressure is an important physical property of water, since lakes, oceans, and other large bodies of water would otherwise evaporate much more quickly.

Vapor pressure is influenced by temperature. As the temperature increases, more molecules are released from the surface of the liquid. This increases movement above the liquid surface, increasing the pressure in the vapor stage. The image below illustrates the effect of temperature on vapor pressure.

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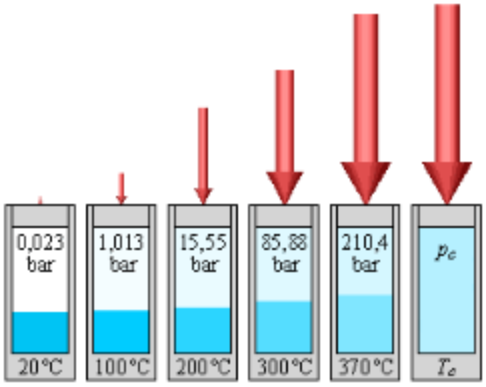


Figure 15.3.2: Vapor pressure of water at different temperatures.

Summary

- Water has high surface tension because of extensive hydrogen bonding.
- The vapor pressure of water is low due to hydrogen bonding.
- Vapor pressure increases as temperature increases.

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15.4: Solute and Solvent

In the winter, the temperature often gets well below the freezing point of water. This condition can create problems in car radiators. If the water freezes, water hoses will break, the engine block can crack, and significant damage can be done to the car.

Solute and Solvent

When one substance dissolves into another, a **solution** is formed. A solution is a homogeneous mixture consisting of a **solute** dissolved into a **solvent**. The solute is the substance that is being dissolved, while the solvent is the dissolving medium. Solutions can be formed with many different types and forms of solutes and solvents.

We know of many types of solutions. Check out a few examples in the table below.

Table 15.4.1*: Types of Solutions*

Type	Solvent	Solute	Example
gas/gas	nitrogen	oxygen	air
gas/liquid	water	carbon dioxide	soda pop
liquid/liquid	water	ethylene glycol	antifreeze
solid/liquid	water	salts	seawater

This section's focus is on solutions where the solvent is water. An aqueous solution is water that contains one or more dissolved substances. The dissolved substances in an aqueous solution many be solids, gases, or other liquids. Some examples are listed in the table above. Other examples include vinegar (acetic acid in water), alcoholic beverages (ethanol in water), and liquid cough medicines (various drugs in water).



Figure 15.4.1: Typical solution: clear and stable.

In order to be a true solution, a mixture must be stable. When sugar is fully dissolved into water, it can stand for an indefinite amount of time and the sugar will not settle out of the solution. Further, if the sugar-water solution is passed through a filter, it will be unchanged. The dissolved sugar particles will pass through the filter along with the water. This is because the dissolved particles in a solution are very small, usually less than 1 nm in diameter. Solute particles can be atoms, ions, or molecules, depending on the type of substance that has been dissolved.

Summary

- A solution is a homogeneous mixture of a solute in a solvent.
- A solute is the material present in the smaller amount in the solution.
- A solvent is the material present in the larger amount in the solution.

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15.5: Dissolving Process

Many people start their day with a cup of coffee (others drink coffee all day long). The coffee typically consumed is a type of solution—sometimes a very complex one. The coffee itself has been brewed so that material from the coffee bean will dissolve in hot water. Some people add sugar, while others add milk or cream. In some areas of the country, lattés are popular; one or more special flavors might be added along with the milk. Whatever the situation, the end result is an enjoyable solution to drink.

The Process of Dissolution

Water typically dissolves many ionic compounds and polar molecules. Nonpolar molecules such as those found in grease or oil do not dissolve in water. We will first examine the process that occurs when an ionic compound such as table salt (sodium chloride) dissolves in water.

Water molecules move about continuously due to their kinetic energy. When a crystal of sodium chloride is placed into water, the water's molecules collide with the crystal lattice. Recall that the crystal lattice is composed of alternating positive and negative ions. Water is attracted to the sodium chloride crystal because water is polar and has both a positive and a negative end. The positively charged sodium ions in the crystal attract the oxygen end of the water molecules because they are partially negative. The negatively charged chloride ions in the crystal attract the hydrogen end of the water molecules because they are partially positive. The action of the polar water molecules takes the crystal lattice apart (see image below).

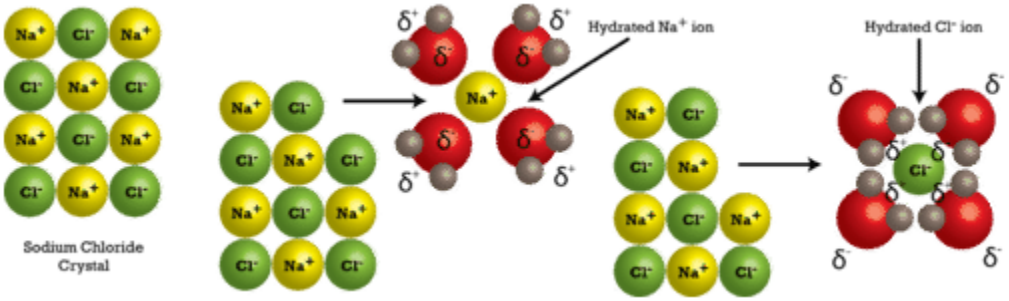


Figure 15.5.1: The dissolving of sodium chloride in water.

After coming apart from the crystal, the individual ions are then surrounded by solvent particles in a process called **solvation**. Note that the individual $\mathrm{Na^+}$ ions are surrounded by water molecules with the oxygen atom oriented near the positive ion. Likewise, the chloride ions are surrounded by water molecules with the opposite orientation. **Hydration** is the process of solute particles being surrounded by water molecules arranged in a specific manner. Hydration helps to stabilize aqueous solutions by preventing the positive and negative ions from coming back together and forming a precipitate.

Table sugar is sucrose $(\mathrm{C_{12}H_{22}O_{11}})$, and is an example of a molecular compound. Solid sugar consists of individual sugar molecules held together by intermolecular attractive forces. When water dissolves sugar, it separates the individual sugar molecules by disrupting the attractive forces, but does not break the covalent bonds between the carbon, hydrogen, and oxygen atoms. Dissolved sugar molecules are also hydrated, but without as distinct an orientation to the water molecules as in the case of the ions. The sugar molecules contain many $-\mathrm{OH}$ groups that can form hydrogen bonds with the water molecules, helping to form the sucrose solution.

Summary

- Motion of water molecules helps break up interactions between solid ions or molecules.
- Solvation involves surrounding ions with solvent particles.
- Ionic solute molecules are hydrated (surrounded by water molecules in a specific orientation).

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15.6: Liquid-Liquid Solutions

In 2010, a major oil spill occurred when an explosion on a drilling rig in the Gulf of Mexico released millions of gallons of crude oil into the Gulf. Oil is primarily a mixture of hydrocarbons (organic compounds composed of only carbon and hydrogen atoms). Because of its composition, oil does not dissolve in water. As a result, much of the Gulf of Mexico was contaminated, as was a great deal of shoreline in the affected area.

Liquid-Liquid Solutions

Nonpolar compounds do not dissolve in water. The attractive forces that operate between the particles in a nonpolar compound are weak dispersion forces. However, the nonpolar molecules are more attracted to themselves than they are to the **polar** water molecules. When a nonpolar liquid such as oil is mixed with water, two separate layers form because the liquids will not dissolve into each other (figure below). When another polar liquid such as ethanol is mixed with water, they completely blend and dissolve into one another. Liquids that dissolve in one another in all proportions are said to be **miscible**. Liquids that do not dissolve in one another are deemed **immiscible**. The general rule for deciding if one substance is capable of dissolving another is "like dissolves like". A nonpolar solid such as iodine will dissolve in nonpolar lighter fluid, but will not dissolve in polar water.



Figure 15.6.1: Water and oil form separate layers when they are mixed because the nonpolar oil will not dissolve into the polar water. The oil forms the top layer because it is less dense than the water.

For molecular compounds, the major factor that contributes to the material dissolving in water is the ability to form hydrogen bonds with the water solvent. Small compounds such as methanol, ethanol, acetic acid, and acetone have polar groups that can interact with the polar H of water. However, as the nonpolar portion of the molecule gets larger, solubility with water drops off. The nonpolar portion of the molecule increasingly repels the water and eventually overrides the interaction of the polar component with water.

Summary

- Liquids that dissolve in one another in all proportions are miscible; liquids that do not dissolve in one another are immiscible.
- The general rule for deciding if one substance is capable of dissolving another is "like dissolves like".
- Nonpolar molecules are usually insoluble in water.
- A non-ionized molecule must be relatively polar to interact with water molecules.

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15.7: Electrolytes and Nonelectrolytes

People around the world jog for exercise. For the most part, jogging can be a healthy way to stay fit. However, problems can develop for those who jog in the heat. Excessive sweating can lead to electrolyte loss, which can be life-threatening. Early symptoms of electrolyte deficiency can include nausea, fatigue, and dizziness. If not treated, individuals can experience muscle weakness and increased heart rate (which could lead to a heart attack). Sports drinks can be consumed to restore electrolytes quickly in the body.

Electrolytes and Nonelectrolytes

An **electrolyte** is a compound that conducts an electric current when it is in an aqueous solution or melted. In order to conduct a current, a substance must contain mobile ions that can move from one electrode to the other. All ionic compounds are electrolytes. When ionic compounds dissolve, they break apart into ions which are then able to conduct a current (**conductivity**). Even insoluble ionic compounds such as $\mathrm{CaCO_3}$ are electrolytes because they can conduct a current in the molten (melted) state.

A **nonelectrolyte** is a compound that does not conduct an electric current in either aqueous solution or in the molten state. Many molecular compounds, such as sugar or ethanol, are nonelectrolytes. When these compounds dissolve in water, they do not produce ions. The figure below illustrates the difference between an electrolyte and a nonelectrolyte.

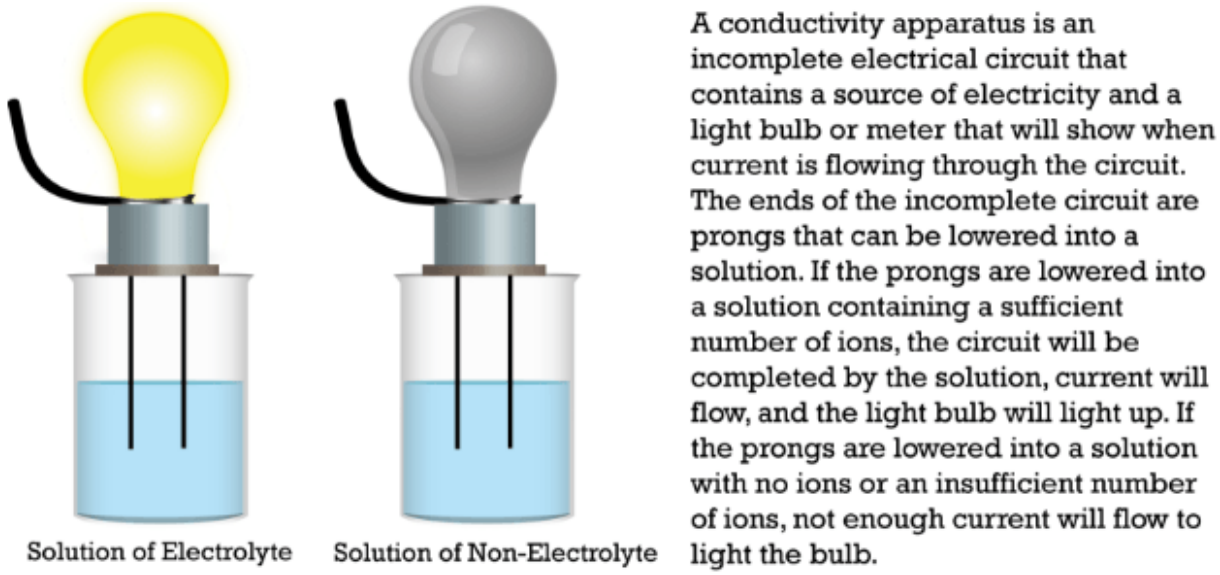


Figure 15.7.1: Conductivity apparatus.

Roles of Electrolytes in the Body

Several electrolytes play important roles in the body. Here are a few significant electrolytes:

1. Calcium - in bones and teeth. Also important for muscle contraction, blood clotting, and nerve function.

2. Sodium - found outside the cell. Mainly involved in water balance and nerve signaling.

3. Potassium - major cation inside the cell. Important for proper functioning of heart, muscles, kidneys, and nerves.

4. Magnesium - in bone and cells. Involved in muscle, bone, nervous system, and takes part in many biochemical reactions.

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Summary

- Electrolytes conduct electric current when in solution or melted.
- Nonelectrolytes do not conduct electric current when in solution or melted.
- Some electrolytes play important roles in the body.

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15.8: Dissociation

In many areas, winter ice on the streets and sidewalks represents a serious walking and driving hazard. One common approach to melting the ice is to put some form of deicing salt on the surface. Materials such as sodium chloride or calcium chloride are frequently employed for this purpose. In order to be effective, the solid material must first dissolve and break up into the ions that make up the compound.

Dissociation

An ionic crystal lattice breaks apart when it is dissolved in water. **Dissociation** is the separation of ions that occurs when a solid ionic compound dissolves. It is important to be able to write dissociation equations. Simply undo the crisscross method that you learned when writing chemical formulas of ionic compounds. The subscripts for the ions in the chemical formulas become the coefficients of the respective ions on the product side of the equation. Shown below are dissociation equations for NaCl, $\mathrm{Ca(NO_3)_2}$, and $\mathrm{(NH_4)_3PO_4}$.

$$\mathrm{NaCl}\,(s) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq)$$

$$\mathrm{Ca(NO_3)_2}\,(s) \rightarrow \mathrm{Ca^{2+}}\,(aq) + 2\mathrm{NO_3{}^-}\,(aq)$$

$$\mathrm{(NH_4)_3PO_4}\,(s) \rightarrow 3\mathrm{NH_4}\,(aq) + \mathrm{PO_4}\,(aq)$$

The formula unit of sodium chloride dissociates into one sodium ion and one chloride ion. The calcium nitrate formula unit dissociates into one calcium ion and two nitrate ions. This is because of the $2+$ charge of the calcium ion. Two nitrate ions, each with a $1-$ charge are required to make the equation balance electrically. The ammonium phosphate formula unit dissociates into three ammonium ions and one phosphate ion. Note that the polyatomic ions themselves do not dissociate further, but remain intact.

Do not confuse the subscripts of the atoms within the polyatomic ion for the subscripts that result from the crisscrossing of the charges that make up the original compound neutral. The 3 subscript of the nitrate ion and the 4 subscript of the ammonium ion are part of the polyatomic ion and simply remain as part of its formula after the compound dissociates. Notice that the compounds are solids (s) which then become ions in aqueous solution (aq).



Figure 15.8.1: Calcium nitrate is a typical ionic compound. In an aqueous solution, it dissociates into calcium ions and nitrate ions.

Nonelectrolytes do not dissociate when forming an aqueous solution. An equation can still be written that simply shows the solid going into solution. For the dissolution of sucrose:

$$\mathrm{C_{12}H_{22}O_{11}}\,(s) \rightarrow \mathrm{C_{12}H_{22}O_{11}}\,(aq)$$

Summary

- Dissociation is the separation of ions that occurs when a solid ionic compound dissolves.
- Nonionic compounds do not dissociate in water.

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15.9: Strong and Weak Electrolytes

Car batteries are used around the world to provide the power to start car engines. One essential component of car batteries is the strong electrolyte sulfuric acid. In the battery, this material ionizes into hydrogen ions and sulfate ions. As the battery is used, the concentrations of these ions decreases. Older batteries had openings in the top where new sulfuric acid could be added to replenish the supply. Today, batteries are sealed to prevent leakage of hazardous sulfuric acid.

Strong and Weak Electrolytes

Some polar molecular compounds are nonelectrolytes when they are in their pure state, but become electrolytes when they are dissolved in water. Hydrogen chloride (HCl) is a gas in its pure molecular state and is a nonelectrolyte. However, when HCl is dissolved in water, it conducts a current well because the HCl molecule ionizes into hydrogen and chloride ions.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H}^{+}\,(aq) + \mathrm{Cl}^{-}\,(aq)$$

HCl dissolved into water is called hydrochloric acid. A **strong electrolyte** is a solution in which a large fraction of the dissolved solute exists as ions. Ionic compounds, and some polar compounds, are completely broken apart into ions and thus conduct a current very well—this makes them strong electrolytes.

Some other polar molecular compounds become electrolytes upon being dissolved into water, but do not ionize to very great extent. Gaseous nitrous acid ionizes in solution to hydrogen ions and nitrite ions, but does so very weakly. Aqueous nitrous acid is composed of only about 5% ions and 95% intact nitrous acid molecules. A **weak electrolyte** is a solution in which only a small fraction of the dissolved solute exists as ions. The equation showing the ionization of a weak electrolyte utilizes a double arrow indicating an equilibrium between the reactants and products.

$$\mathrm{HNO_2}\,(g) \rightleftharpoons \mathrm{H}^{+}\,(aq) + \mathrm{NO_2}^{-}\,(aq)$$

Summary

- A strong electrolyte exists mainly as ions in solution.
- A weak electrolyte has only a small amount of ionization in solution.

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15.10: Suspensions

On a calm day, the Caribbean Sea is clear and smooth. You can see deep into the water—deep enough to see fish and underwater plant life without obstruction. The situation changes markedly when a storm blows up. The sand on the bottom of the sea is stirred up, the water becomes turbid, and you can't see anything below the surface. The sand forms a suspension in the water that slowly clears up again after the storm blows over.

Suspensions

Take a glass of water and throw in a handful of sand or dirt. Stir it and stir it and stir it. The water may become **turbid**, or unclear. Have you made a solution? Sand and dirt do not dissolve in water and though it may look homogenous for a few moments, the sand or dirt gradually sinks to the bottom of the glass.

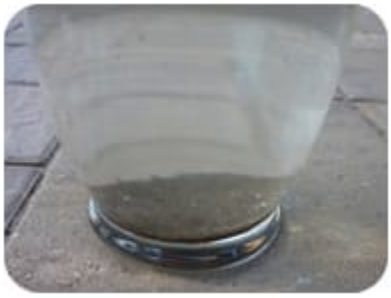


Figure 15.10.1: A suspension of dirt in water.

A **suspension** is a heterogeneous mixture in which some of the particles settle out of the mixture upon standing. The particles in a suspension are far larger than those of a solution and thus gravity is able to pull them down out of the **dispersion medium** (water). The typical diameter for the dispersed particles (the sand) of a suspension is about 1000 times greater than those of a solution (less than approximately two nanometers for particles in solution, compared to greater than 1000 nanometers for particles in suspension). Unlike in a solution, the dispersed particles can be separated from the dispersion medium by filtering. Suspensions are heterogeneous because at least two different substances in the mixture can be identified.

Some Typical Aqueous Suspensions

Milk

Milk is a complex mixture of water, proteins, fats, carbohydrates, and minerals. While the minerals and carbohydrates are water-soluble, the fats and some of the proteins do not dissolve, but are held in suspension.

Paint

Paint can either be water-based or use an organic solvent. We will limit this discussion to water-based paints. The materials will contain binders (organic polymers that help hold the paint to the surface of the material being painted). In addition, pigments are included to provide the desired color. Both the binders and the pigments are not water-soluble, so the paint must be stirred every time it is used. Paint that is allowed to sit for a long period of time will slowly begin to separate as the suspension of binders and pigments settles out.

Summary

- Suspensions are heterogeneous mixtures.
- Some of the material in a suspension will settle out on standing.
- Solid material in a suspension can be removed by filtration.

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15.11: Colloids

Imagine that you are sailing on a yacht. The engine suddenly breaks down and you are stranded in the middle of the ocean. You call the Coast Guard on your radio, but cannot give them an exact location, because your GPS isn't working. Fortunately, you have a smoke flare, which you fire off. The dense colored smoke shows the Coast Guard where you are so that they can rescue you. In using the flare, you are taking advantage of a mixture called a colloid.

Colloids

A **colloid** is a heterogeneous mixture whose particle size is intermediate between those of a solution and a suspension. The dispersed particles are spread evenly throughout the **dispersion medium**, which can be a solid, liquid, or gas. Because the dispersed particles of a colloid are not as large as those of a suspension, they do not settle out upon standing. The table below summarizes the properties and distinctions between solutions, colloids, and suspensions.

Properties of Solutions, Colloids, and Suspensions

Table 15.11.1: Properties of Solutions, Colloids, and Suspensions

Solutions	Colloids	Suspensions
Homogeneous.	Heterogeneous.	Heterogeneous.
Particle size: 0.01-$1\ \text{nm}$; atoms, ions, or molecules.	Particle size: 1-$1000\ \text{nm}$, dispersed; large molecules or aggregates.	Particle size: over $1000\ \text{nm}$, suspended; large particles or aggregates.
Do not separate on standing.	Do not separate on standing.	Particles settle out.
Cannot be separated by filtration.	Cannot be separated by filtration.	Can be separated by filtration.
Do not scatter light.	Scatter light (Tyndall effect).	May either scatter light or be opaque.

Colloids are unlike solutions because their dispersed particles are much larger than those of a solution. The dispersed particles of a colloid cannot be separated by filtration, but they scatter light—a phenomenon called the **Tyndall effect**.

The Tyndall Effect

When light is passed through a true solution, the dissolved particles are too small to deflect the light. However, the dispersed particles of a colloid, being larger, do deflect light. The Tyndall effect is the scattering of visible light by colloidal particles. You have undoubtedly "seen" a light beam as it passes through fog, smoke, or a scattering of dust particles suspended in air. All three are examples of colloids. Suspensions may scatter light, but if the number of suspended particles is sufficiently large, the suspension may simply be opaque, and the light scattering will not occur.



Figure 15.11.1: The Tyndall effect allows sunlight to be seen as it passes through a fine mist.

Examples of Colloids

The table below lists examples of colloidal systems, most of which are very familiar. The dispersed phase describes the particles, while the dispersion medium is the material in which the particles are distributed.

Classes of Colloids

Table 15.11.2: Classes of Colloids

Class of Colloid	Dispersed Phase	Dispersion Medium	Examples
Solid gel	solid	liquid	paint, jellies, blood, gelatin, mud
Solid aerosol	solid	gas	smoke, dust in air
Solid emulsion	liquid	solid	cheese, butter
Liquid emulsion	liquid	liquid	milk, mayonnaise
Liquid aerosol	liquid	gas	fog, mist, clouds, aerosol spray
Foam	gas	solid	marshmallow
Foam	gas	liquid	whipped cream, shaving cream

Another property of a colloidal system is observed when the colloids are studied under a light microscope. The colloids scintillate, reflecting brief flashes of light because the colloidal particles move in a rapid and random fashion. This phenomenon, called Brownian motion, is caused by collisions between the small colloidal particles and the molecules of the dispersion medium.

Emulsions

Butter and mayonnaise are examples of a class of colloids called **emulsions**. An emulsion is a colloidal dispersion of a liquid in either a liquid or a solid. A stable emulsion requires an emulsifying agent to be present. Mayonnaise is made in part of oil and vinegar. Since oil is nonpolar and vinegar is an aqueous solution and polar, the two do not mix, and quickly separate into layers. However, the addition of egg yolk causes the mixture to become stable and not separate. Egg yolk is capable of interacting with both the polar vinegar and the nonpolar oil. The egg yolk is called the emulsifying agent. Soap acts as an emulsifying agent between grease and water. Grease cannot be simply rinsed off your hands, or another surface, because it is insoluble. However, the soap stabilizes a grease-water mixture because one end of a soap molecule is polar, and the other end is nonpolar. This allows the grease to be removed from your hands or your clothing by washing with soapy water.

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Summary

- A colloid is a heterogeneous mixture whose particle size is intermediate between those of a solution and a suspension.
- The Tyndall effect is the scattering of visible light by colloidal particles.
- An emulsion is a colloidal dispersion of a liquid in either a liquid or a solid.

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16: SOLUTIONS

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16.1: Solute-Solvent Combinations

Solute-Solvent Combinations

The focus of "Chapter 15: Water" was water's role in the formation of aqueous solutions. We examined the primary characteristics of a solution and how water is able to dissolve solid solutes; we differentiated between a solution, a suspension, and a colloid. There are many examples of solutions that do not involve water at all, or that involve solutes that are not solids. The table below summarizes the possible combinations of solute-solvent states, along with examples of each.

Table 16.1.1*: Solute-Solvent Combinations*

Solute State	Solvent State	Example
liquid	gas	water in air
gas	gas	oxygen in nitrogen (gas mixture)
solid	liquid	salt in water
liquid	liquid	alcohol in water
gas	liquid	carbon dioxide in water
solid	solid	zinc in copper (brass alloy)
liquid	solid	mercury in silver and tin (dental amalgam)

Gas-Gas Solutions

Our air is a homogenous mixture of many different gases and therefore qualifies as a solution. Approximately 78% of the atmosphere is nitrogen, making it the solvent for this solution. The next major constituent is oxygen (about 21%), followed by the inert gas argon (0.9%), carbon dioxide (0.03%), and trace amounts of neon, methane, helium, and other gases.

Solid-Solid Solutions

Solid-solid solutions such as brass, bronze, and sterling silver are called alloys. Bronze (composed mainly of copper with added tin) was widely used in making weapons in times past, dating back to at least 2400 B.C. This metal alloy was hard and tough, but was eventually replaced by iron.

Liquid-Liquid Solutions

Perhaps the most familiar liquid-solid solution is dental amalgam, used to fill teeth when there is a cavity. Approximately 50% of the amalgam material is liquid mercury to which a powdered alloy of silver, tin, and copper is added. Mercury is used because it binds well with the solid metal alloy. However, the use of mercury-based dental amalgam has gone under question in recent years, because of concerns regarding the toxicity of mercury.

Summary

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16.2: Rate of Dissolution

Many people enjoy a cold glass of iced tea on a hot summer day. Some like it unsweetened, while others like to put sugar in it. How sugar dissolves in the tea depends on two factors: how much sugar was put into the tea, and how cold it is. Tea usually has to be stirred for a while to get all the sugar dissolved.

Rate of Dissolution

Dissolution is the process by which a solute dissolves into a solvent and forms a solution. We know that the dissolution of a solid by water depends upon the collisions that occur between the solvent molecules and the particles in the solid crystal. Anything that can be done to increase the frequency of those collisions and/or to give those collisions more energy will increase the rate of dissolution. Imagine that you were trying to dissolve some sugar in a glassful of tea. A packet of granulated sugar would dissolve faster than a cube of sugar. The rate of dissolution would be increased by stirring, or agitating the solution. Finally, the sugar would dissolve faster in hot tea than it would in cold tea.

Surface Area

The rate at which a solute dissolves depends upon the size of the solute particles. Dissolution is a surface phenomenon, since it depends on solvent molecules colliding with the outer surface of the solute. A given quantity of solute dissolves faster when it is ground into small particles, rather than in the form of large pieces, because more **surface area** is exposed. A packet of granulated sugar exposes far more surface area to the solvent and dissolves more quickly than a sugar cube.

Agitation of the Solution

Dissolving sugar in water will occur more quickly if the water is stirred. The stirring allows fresh solvent molecules to continually be in contact with the solute. If it is not stirred, then the water right at the surface of the solute becomes saturated with dissolved sugar molecules, meaning that it is more difficult for additional solute to dissolve. The sugar cube would eventually dissolve because random motions of the water molecules would bring enough fresh solvent into contact with the sugar, but the process would take much longer. It is important to realize that neither stirring nor breaking up a solute affect the overall amount of solute that dissolves—these actions only affect the rate of dissolution.

Temperature

Heating up a solvent gives the molecules more kinetic energy. The increased rapid motion means that the solvent molecules collide with the solute with greater frequency, and that the collisions occur with more force. Both factors increase the rate at which the solute dissolves. As we will see in the next section, a temperature change not only affects the rate of dissolution, but also affects the amount of solute that dissolves.

Summary

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16.3: Saturated and Unsaturated Solutions

When compounds are synthesized, contaminating materials are often mixed in with them. The process of recrystallization can be used to remove these impurities. The crystals are dissolved in a hot solvent, forming a solution. When the solvent is cooled, the compound is no longer as soluble and will precipitate out of solution, leaving contaminating materials still dissolved.

Saturated and Unsaturated Solutions

Table salt (NaCl) readily dissolves in water. Suppose that you have a beaker of water to which you add some salt, stirring until it dissolves. You add more and that dissolves. You keep adding more and more salt, eventually reaching a point at which no more of the salt will dissolve—no matter how long or how vigorously you stir it. Why? On the molecular level, we know that action of the water causes the individual ions to break apart from the salt crystal and enter the solution, where they remain hydrated by water molecules. What also happens is that some of the dissolved ions collide back again with the crystal and remain there. **Recrystallization** is the process of dissolved solute returning to the solid state. At some point, the rate at which the solid salt is dissolving becomes equal to the rate at which the dissolved solute is recrystallizing. When that point is reached, the total amount of dissolved salt remains unchanged. **Solution equilibrium** is the physical state described by the opposing processes of dissolution and recrystallization occurring at the same rate. The solution equilibrium for the dissolution of sodium chloride can be represented by one of two equations:

$$\mathrm{NaCl}\,(s) \rightleftharpoons \mathrm{NaCl}\,(aq)$$

While the above equation shows the change of state back and forth between solid and aqueous solution, the preferred equation also shows the dissociation that occurs as an ionic solid dissolves:

$$\mathrm{NaCl}\,(s) \rightleftharpoons \mathrm{Na}^{+}\,(aq) + \mathrm{Cl}^{-}\,(aq)$$

When the solution equilibrium point is reached and no more solute will dissolve, the solution is said to be saturated. A **saturated solution** is a solution that contains the maximum amount of solute that is capable of dissolving. At $20^{\circ}\mathrm{C}$, the maximum amount of NaCl that will dissolve in \100. \: \text{g}\) of water is $36.0\ \mathrm{g}$. If any more NaCl is added past this point, it will not dissolve, because the solution is saturated. What if more water is added to the solution instead? In that case, more NaCl would be capable of dissolving in the additional solvent. An **unsaturated solution** is a solution that contains less than the maximum amount of solute that is capable of being dissolved. The figure below illustrates the above process and shows the distinction between unsaturated and saturated.

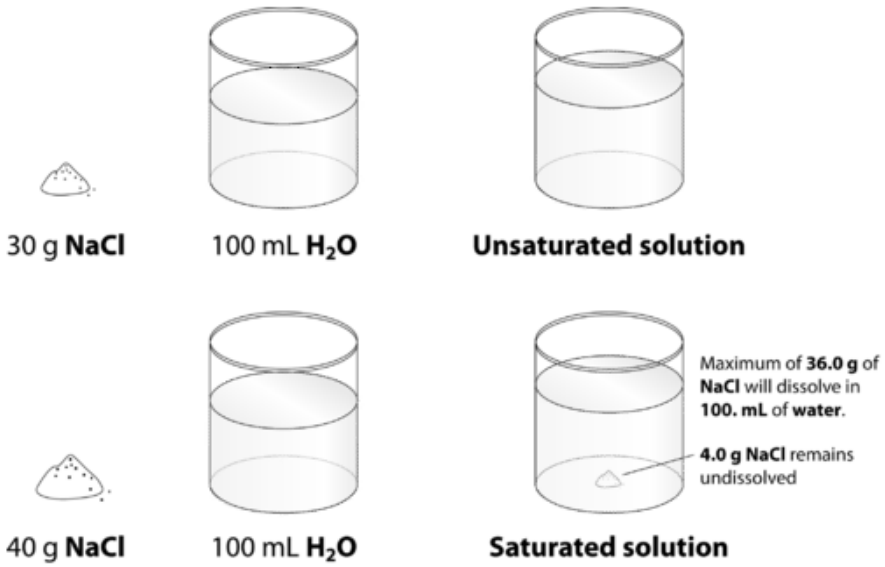


Figure 16.3.1: When $30.0\ \mathrm{g}$ of NaCl is added to $100\ \mathrm{mL}$, it all dissolves, forming an unsaturated solution. When $40.0\ \mathrm{g}$ is added, $36.0\ \mathrm{g}$ dissolves and $4.0\ \mathrm{g}$ remains undissolved, forming a saturated solution.

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How can you tell if a solution is saturated or unsaturated? If more solute is added and it does not dissolve, then the original solution was saturated. If the added solute dissolves, then the original solution was unsaturated. A solution that has been allowed to reach equilibrium, but which has extra undissolved solute at the bottom of the container, must be saturated.

Summary

- Recrystallization is the process of dissolved solute returning to the solid state.
- A saturated solution is a solution that contains the maximum amount of solute that is capable of dissolving.
- An unsaturated solution is a solution that contains less than the maximum amount of solute that is capable of being dissolved.
- Solution equilibrium exists when the rate of dissolution equals the rate of recrystallization.

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16.4: How Temperature Influences Solubility

Nuclear power plants require large amounts of water to generate steam for turbines and to cool equipment. They are usually situated near bodies of water to use that water as a coolant, returning the warmer water back to the lake or river. This increases the overall temperature of the water, which lowers the quantity of dissolved oxygen, affecting the survival of fish and other organisms.

How Temperature Influences Solubility

The **solubility** of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature. Solubility is often measured as the grams of solute per $100\ \text{g}$ of solvent. The solubility of sodium chloride in water is $36.0\ \text{g}$ per $100\ \text{g}$ water at 20°C. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubility is specific for a particular solvent. In this section, we will consider solubility of material in water as solvent.

The solubility of the majority of solid substances increases as the temperature increases. However, the effect is difficult to predict and varies widely from one solute to another. The temperature dependence of solubility can be visualized with the help of a **solubility curve**, a graph of the solubility vs. temperature (see figure below).

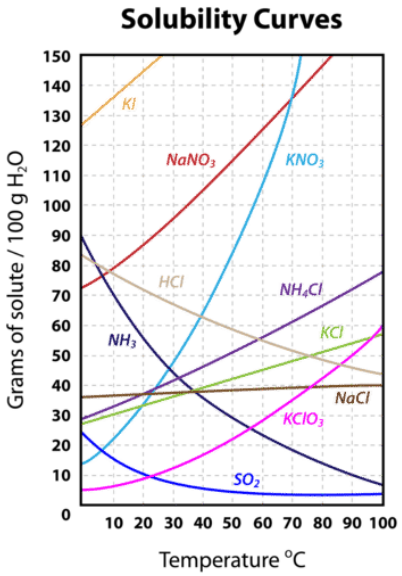


Figure $16.4.1$: Solubility curves for several compounds.

Notice how the temperature dependence of NaCl is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of NaCl. The curve for KNO_3, on the other hand, is very steep, and so an increase in temperature dramatically increases the solubility of KNO_3.

Several substances—HCl, NH_3, and SO_2—have solubility that decreases as temperature increases. They are all gases at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increase. As the kinetic energy of the gaseous solute increases, its molecules have a greater tendency to escape the attraction of the solvent molecules and return to the gas phase. Therefore, the solubility of a gas decreases as the temperature increases.

Solubility curves can be used to determine if a given solution is saturated or unsaturated. Suppose that $80\ \text{g}$ of KNO_3 is added to $100\ \text{g}$ of water at 30°C. According to the solubility curve, approximately $48\ \text{g}$ of KNO_3 will dissolve at 30°C. This means that the solution will be saturated since $48\ \text{g}$ is less than $80\ \text{g}$. We can also determine that there will be $80 - 48 = 32\ \text{g}$ of undissolved KNO_3 remaining at the bottom of the container. In a second scenario, suppose that this saturated solution is heated to 60°C. According to the curve, the solubility of KNO_3 at 60°C is about $107\ \text{g}$. The solution, in this case, is unsaturated since it contains only the original $80\ \text{g}$ of dissolved solute. Suppose in a third case, that the solution is cooled all the way down to 0°C. The solubility at 0°C is about $14\ \text{g}$, meaning that $80 - 14 = 66\ \text{g}$ of the KNO_3 will recrystallize.

Summary

- The solubility of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature.
- A solubility curve is a graph of the solubility vs. temperature
- The solubility of a solid in water increases with an increase in temperature.
- Gas solubility decreases as the temperature increases.

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16.5: Supersaturated Solutions

In the picture below, a thermal pack is applied to the back. Small packs can be used either for heating or cooling, depending on the material used. A heat pack contains a supersaturated solution of material such as sodium acetate. The solution is clear until a small metal trigger is activated. The sodium acetate then crystallizes out of solution and generates heat in the process.



Figure 16.5.1: A thermal pack is applied to heat or cool muscles.

Supersaturated Solutions

Some solutes, such as sodium acetate, do not recrystallize easily. Suppose an exactly saturated solution of sodium acetate is prepared at $50^{\circ}\mathrm{C}$. As it cools back to room temperature, no crystals appear in the solution, even though the solubility of sodium acetate is lower at room temperature. A **supersaturated solution** is a solution that contains more than the maximum amount of solute that is capable of being dissolved at a given temperature. The recrystallization of the excess dissolved solute in a supersaturated solution can be initiated by the addition of a tiny crystal of solute, called a seed crystal. The seed crystal provides a nucleation site on which the excess dissolved crystals can begin to grow. Recrystallization from a supersaturated solution is typically very fast.

Summary

- A supersaturated solution is a solution that contains more than the maximum amount of solute that is capable of being dissolved at a given temperature.
- When a seed crystal is added to the solution, a supersaturated solution can recrystallize.

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16.6: Henry's Law

Having a soft drink in outer space poses some special problems. Under microgravity, the carbonation can quickly dissipate if not kept under pressure. You can't open the can, or you will lose carbonation. So, a special pressurized container has been developed to get around the problem of gas loss at low gravity.

Henry's Law

Pressure has very little effect on the solubility of solids or liquids, but has a significant effect on the solubility of gases. Gas solubility increases as the partial pressure of a gas above the liquid increases. Suppose a certain volume of water is in a closed container with the space above it occupied by carbon dioxide gas at standard pressure. Some of the $\mathrm{CO_2}$ molecules come into contact with the surface of the water and dissolve into the liquid. Now suppose that more $\mathrm{CO_2}$ is added to the space above the container, causing a pressure increase. In this case, more $\mathrm{CO_2}$ molecules are in contact with the water and so more of them dissolve. Thus, the solubility increases as the pressure increases. As with a solid, the $\mathrm{CO_2}$ that is undissolved reaches an equilibrium with the dissolved $\mathrm{CO_2}$, represented by the equation:

$$\mathrm{CO_2}\,(g) \rightleftharpoons \mathrm{CO_2}\,(aq)$$

At equilibrium, the rate of gaseous $\mathrm{CO_2}$ dissolution is equal to the rate of dissolved $\mathrm{CO_2}$ coming out of the solution.

When carbonated beverages are packaged, they are done so under high $\mathrm{CO_2}$ pressure so that a large amount of carbon dioxide dissolves in the liquid. When the bottle is open, the equilibrium is disrupted because the $\mathrm{CO_2}$ pressure above the liquid decreases. Immediately, bubbles of $\mathrm{CO_2}$ rapidly exit the solution and escape out of the top of the open bottle. The amount of dissolved $\mathrm{CO_2}$ decreases. If the bottle is left open for an extended period of time, the beverage becomes "flat" as more and more $\mathrm{CO_2}$ comes out of the liquid.

The relationship of gas solubility to pressure is described by Henry's law, named after English chemist William Henry (1774-1836). **Henry's Law** states that the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid. Henry's law can be written as follows:

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

S_1 and P_1 are the solubility and the pressure at an initial set of conditions; S_2 and P_2 are the solubility and pressure at another changed set of conditions. The solubility of a gas is typically reported in $\mathrm{g/L}$.

Example 16.6.1

The solubility of a certain gas in water is $0.745\ \mathrm{g/L}$ at standard pressure. What is its solubility when the pressure above the solution is raised to $4.50\ \mathrm{atm}$? The temperature is constant at $20^{\circ}\mathrm{C}$.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $S_1 = 0.745\ \mathrm{g/L}$
- $P_1 = 1.00\ \mathrm{atm}$
- $P_2 = 4.50\ \mathrm{atm}$

Unknown

Substitute into Henry's law and solve for S_2.

Step 2: Solve.

$$S_2 = \frac{S_1 \times P_2}{P_1} = \frac{0.745\ \text{g/L} \times 4.50\ \text{atm}}{1.00\ \text{atm}} = 3.35\ \text{g/L}$$

Step 3: Think about your result.

The solubility is increased to 4.5 times its original value, according to the direct relationship.

Summary

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16.7: Percent Solutions

There are human cultures that do not recognize numbers above three. Anything greater than that is simply referred to as "much" or "many". Although this form of calculation may seem very limited, American culture does the same thing to a certain degree. For example, there are several ways to express the amount of solute in a solution in a quantitative manner. The **concentration** of a solution is a measure of the amount of solute that has been dissolved in a given amount of solvent or solution. A **concentrated** solution is one that has a relatively large amount of dissolved solute. A **dilute** solution is one that has a relatively small amount of dissolved solute. However, those terms are vague, and it is often necessary to express concentration with numbers.

Percent Solutions

One way to describe the concentration of a solution is by the percent of a solute in the solvent. The percent can further be determined in one of two ways: (1) the ratio of the mass of the solute divided by the mass of the solution or (2) the ratio of the volume of the solute divided by the volume of the solution.

Mass Percent

When the solute in a solution is a solid, a convenient way to express the concentration is by mass percent $\left(\frac{\text{mass}}{\text{mass}}\right)$, which is the grams of solute per $100\ \text{g}$ of solution.

$$\text{Percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

Suppose that a solution was prepared by dissolving $25.0\ \text{g}$ of sugar into $100\ \text{g}$ of water. The percent by mass would be calculated by:

$$\text{Percent by mass} = \frac{25\ \text{g sugar}}{125\ \text{g solution}} \times 100\% = 20\%\ \text{sugar}$$

Sometimes you may want to make up a particular mass of solution of a given percent by mass, and need to calculate what mass of the solvent to use. For example, you need to make $3000\ \text{g}$ of a 5% solution of sodium chloride. You can rearrange and solve for the mass of solute:

$$\text{mass of solute} = \frac{\text{percent by mass}}{100\%} \times \text{mass of solution} = \frac{5\%}{100\%} \times 3000\ \text{g} = 150\ \text{g NaCl}$$

You would need to weigh out $150\ \text{g}$ of NaCl and add it to $2850\ \text{g}$ of water. Notice that it was necessary to subtract the mass of the $\text{NaCl}\ (150\ \text{g})$ from the mass of solution $(3000\ \text{g})$ to calculate the mass of the water that would need to be added.

Volume Percent

The percentage of solute in a solution can more easily be determined by volume when the solute and solvent are both liquids. The volume of the solute divided by the volume of the solution, expressed as a percent, yields the percent by volume $\left(\frac{\text{volume}}{\text{volume}}\right)$ of the solution. If a solution is made by adding $40\ \text{mL}$ of ethanol to $200\ \text{mL}$ of water, the percent by volume is:

$$\begin{aligned}\text{Percent by volume} &= \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \\ &= \frac{40\ \text{mL ethanol}}{240\ \text{mL solution}} \times 100\% \\ &= 16.7\%\ \text{ethanol}\end{aligned}$$

Frequently, ingredient labels on food products and medicines have amounts listed as percentages (see figure below).

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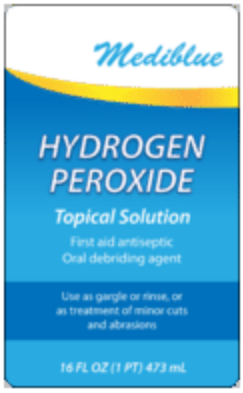


Figure 16.7.1: Hydrogen peroxide is commonly labeled as a 3% by volume solution for use as a disinfectant.

Summary

- The concentration of a solution is a measure of the amount of solute that has been dissolved in a given amount of solvent or solution.
- A concentrated solution has a relatively large amount of dissolved solute.
- A dilute solution has a relatively small amount of dissolved solute.
- Techniques for calculation of percent mass and percent volume solution concentrations are described.

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16.8: Molarity

Chemists deal with amounts of molecules every day. Chemical reactions are described as so many molecules of compound A reacting with so many molecules of compound B to form so many molecules of compound C. When we determine how much reagent to use, we need to know the number of molecules in a given volume of the reagent. Percent solutions only tell us the number of grams, not molecules. A $100\ \text{mL}$ solution of $2\%\ \text{NaCl}$ will have a very different number of molecules than a 2% solution of CsCl. So, we need another way to talk about numbers of molecules.

Molarity

Chemists primarily need the concentration of solutions to be expressed in a way that accounts for the number of particles that react according to a particular chemical equation. Since percentage measurements are based on either mass or volume, they are generally not useful for chemical reactions. A concentration unit based on moles is preferable. The **molarity** (M) of a solution is the number of moles of solute dissolved in one liter of solution. To calculate the molarity of a solution, divide the moles of solute by the volume of the solution expressed in liters:

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{\text{mol}}{\text{L}}$$

Note that the volume is in liters of solution and not liters of solvent. When a molarity is reported, the unit is the symbol M and is read as "molar". For example a solution labeled as $1.5\ \text{M}\ \text{NH}_3$ is read as "1.5 molar ammonia solution".

Example 16.8.1

A solution is prepared by dissolving $42.23\ \text{g}$ of NH_4Cl into enough water to make $500.0\ \text{mL}$ of solution. Calculate its molarity.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass $= 42.23\ \text{g}\ \text{NH}_4\text{Cl}$
- Molar mass $\text{NH}_4\text{Cl} = 53.50\ \text{g/mol}$
- Volume solution $= 500.0\ \text{mL} = 0.5000\ \text{L}$

Unknown

The mass of the ammonium chloride is first converted to moles. Then the molarity is calculated by dividing by liters. Note that the given volume has been converted to liters.

Step 2: Solve.

$$42.23\ \text{g}\ \text{NH}_4\text{Cl} \times \frac{1\ \text{mol}\ \text{NH}_4\text{Cl}}{53.50\ \text{g}\ \text{NH}_4\text{Cl}} = 0.7893\ \text{mol}\ \text{NH}_4\text{Cl}$$

$$\frac{0.7893\ \text{mol}\ \text{NH}_4\text{Cl}}{0.5000\ \text{L}} = 1.579\ \text{M}$$

Step 3: Think about your result.

The molarity is $1.579\ \text{M}$, meaning that a liter of the solution would contain $1.579\ \text{mol}\ \text{NH}_4\text{Cl}$. Four significant figures are appropriate.

In a laboratory situation, a chemist must frequently prepare a given volume of solutions of a known molarity. The task is to calculate the mass of the solute that is necessary. The molarity equation can be rearranged to solve for moles, which can then be converted to grams.

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Example 16.8.2

A chemist needs to prepare $3.00\ \text{L}$ of a $0.250\ \text{M}$ solution of potassium permanganate (KMnO_4). What mass of KMnO_4 does she need to make the solution?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Molarity $= 0.250\ \text{M}$
- Volume $= 3.00\ \text{L}$
- Molar mass $\text{KMnO}_4 = 158.04\ \text{g/mol}$

Unknown

Moles of solute is calculated by multiplying molarity by liters. Then, moles is converted to grams.

Step 2: Solve.

$$\text{mol KMnO}_4 = 0.250\ \text{M KMnO}_4 \times 3.00\ \text{L} = 0.750\ \text{mol KMnO}_4$$

$$0.750\ \text{mol KMnO}_4 \times \frac{158.04\ \text{g KMnO}_4}{1\ \text{mol KMnO}_4} = 119\ \text{g KMnO}_4$$

Step 3: Think about your result.

When $119\ \text{g}$ of potassium permanganate is dissolved into water to make $3.00\ \text{L}$ of solution, the molarity is $0.250\ \text{M}$.

Summary

- The molarity (M) of a solution is the number of moles of solute dissolved in one liter of solution.
- Calculations using the concept of molarity are described.

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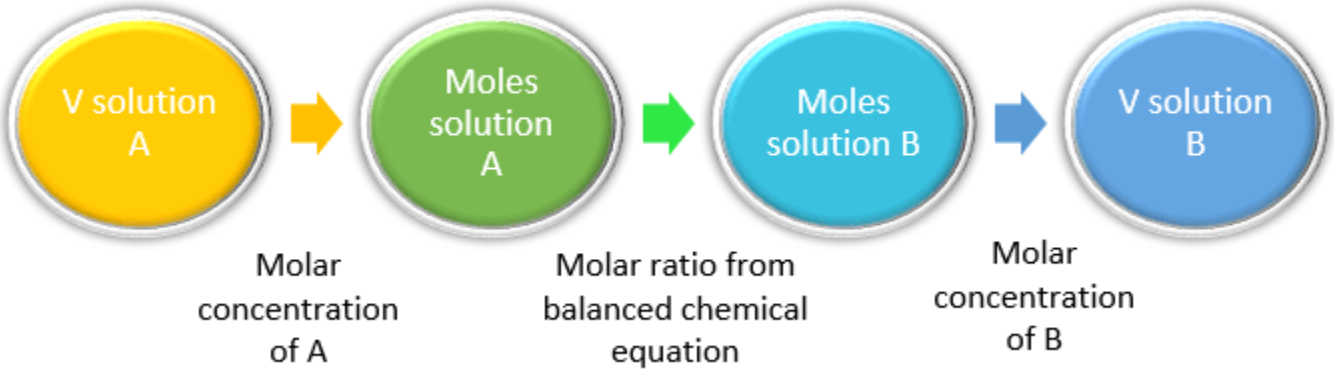
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16.9: Solution Stoichiometry

Learning Objectives

- Determine amounts of reactants or products in aqueous solutions.

As we learned previously, double replacement reactions involve the reaction between ionic compounds in solution and, in the course of the reaction, the ions in the two reacting compounds are "switched" (they *replace* each other). Because these reactions occur in aqueous solution, we can use the concept of molarity to directly calculate the number of moles of reactants or products that will be formed, and therefore their amounts (i.e. volume of solutions or mass of precipitates).



As an example, lead (II) nitrate and sodium chloride react to form sodium nitrate and the *insoluble* compound, lead (II) chloride.

$$\mathrm{Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow PbCl_2\,(s) + 2\,NaNO_3\,(aq)} \tag{16.9.1}$$

In the reaction shown above, if we mixed 0.123 L of a 1.00 M solution of NaCl with 1.50 M solution of $\mathrm{Pb(NO_3)_2}$, we could calculate the volume of $\mathrm{Pb(NO_3)_2}$ solution needed to completely precipitate the $\mathrm{Pb^{2+}}$ ions.

The molar concentration can also be expressed as the following:

$$1.00\,M\,\mathrm{NaCl} = \frac{1.00\ mol\ \mathrm{NaCl}}{1\ L\ \mathrm{NaCl\ solution}}$$

and

$$1.50\,M\,\mathrm{Pb(NO_3)_2} = \frac{1.50\ mol\ \mathrm{Pb(NO_3)_2}}{1\ L\ \mathrm{Pb(NO_3)_2 solution}}$$

First, we must examine the reaction stoichiometry in the balanced reaction (Equation 16.9.1). In this reaction, one mole of $\mathrm{Pb(NO_3)_2}$ reacts with two moles of NaCl to give one mole of $\mathrm{PbCl_2}$ precipitate. Thus, the concept map utilizing the stoichiometric ratios is:



so the volume of lead (II) nitrate that reacted is calculated as:

$$0.123\ L\ \text{NaCl solution} \times \frac{1.00\ mol\ \text{NaCl}}{1\ L\ \text{NaCl solution}} \times \frac{1\ mol\ \text{Pb(NO}_3)_2}{2\ mol\ \text{NaCl}} \times \frac{1\ L\ \text{Pb(NO}_3)_2\ \text{solution}}{1.5\ mol\ \text{Pb(NO}_3)_2} = 0.041\ \text{Pb(NO}_3)_2\ L\ \text{solution}$$

This volume makes intuitive sense for two reasons: (1) the number of moles of $\text{Pb(NO}_3)_2$ required is half of the number of moles of NaCl, based off of the stoichiometry in the balanced reaction (Equation 16.9.1); (2) the concentration of $\text{Pb(NO}_3)_2$ solution is 50% greater than the NaCl solution, so less volume is needed.

Example 16.9.1

What volume (in L) of 0.500 M sodium sulfate will react with 275 mL of 0.250 M barium chloride to completely precipitate all Ba^{2+} in the solution?

Solution

Solutions to Example 13.8.1

Steps for Problem Solving	Example 16.9.1
Identify the "given" information and what the problem is asking you to "find."	Given: 275 mL $BaCl_2$ 0.250 M BaCl_2 or $\frac{0.250\ mol BaCl_2}{1\ L\ BaCl_2\ solution}$ 0.500 M Na_2SO_4 or $\frac{0.500\ mol Na_2SO_4}{1\ L\ Na_2SO_4\ solution}$ Find: Volume Na_2SO_4 solution.
Set up and balance the chemical equation.	$\text{Na}_2\text{SO}_4\,(\text{aq}) + \text{BaCl}_2\,(\text{aq}) \longrightarrow \text{BaSO}_4\,(\text{s}) + \underline{2}\text{NaCl}\,(\text{aq})$ An insoluble product is formed after the reaction.
List other known quantities.	1 mol of Na_2SO_4 to 1 mol $BaCl_2$ 1000 mL = 1 L
Prepare a concept map and use the proper conversion factor.	mL $BaCl_2$ solution → L $BaCl_2$ solution → mol $BaCl_2$ solution → mol Na_2SO_4 solution $\frac{1\ L}{1000\ mL}$ $\frac{0.250\ mol\ BaCl_2}{1\ L\ BaCl_2 solution}$ $\frac{1\ mol\ Na_2SO_4}{1\ mol\ BaCl_2}$ $\frac{1\ L\ Na_2SO_4\ solution}{0.500\ mol}$

Steps for Problem Solving	Example 16.9.1
Cancel units and calculate.	$275\ \cancel{mL\ BaCl_2\ solution} \times \frac{1\ \cancel{L}}{1000\ \cancel{mL}} \times \frac{0.250\ \cancel{mol\ BaCl_2}}{1\ \cancel{L\ BaCl_2\ solution}} \times \frac{1\ \cancel{mol\ Na_2SO}}{1\ \cancel{mol\ BaCl}} \times \frac{1\ L\ Na_2SO_4\ solution}{0.500\ \cancel{molNa_2SO_4}}$ = 0.1375 L sodium sulfate
Think about your result.	The lesser amount (almost half) of sodium sulfate is to be expected as it is more concentrated than barium chloride. Also, the units are correct.

Exercise 16.9.1

What volume of 0.250 M lithium hydroxide will completely react with 0.500 L of 0.250 M of sulfuric acid solution?

Answer available at the end of this section.

Answers

Exercise 16.9.1

0.250 L $LiOH$ solution

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16.10: Preparing Solutions

Back in the "good ol' days" (a variable time frame, dependent on who you ask), many cooks did not bother with careful measurements. They simply "knew" how much flour to use, or how much water to add. Most of us need somewhat more precise ways to measure when we cook. Chemists are very careful when they prepare solutions, because the results of their experiments need to be quantitative. Simply "knowing" is not accurate enough for scientific purposes.

Preparing Solutions

If you are attempting to prepare $1.00\ \mathrm{L}$ of a $1.00\ \mathrm{M}$ solution of NaCl, you would obtain $58.44\ \mathrm{g}$ of sodium chloride. However, you cannot simply add the sodium chloride to $1.00\ \mathrm{L}$ of water. After the solute dissolves, the volume of the solution will be slightly greater than one liter, because the hydrated sodium and chloride ions take up space in the solution. Instead, a **volumetric flask** needs to be used. Volumetric flasks come in a variety of sizes (see image below) and are designed to allow a chemist to prepare a solution of only one specific volume.



Figure 16.10.1: Volumetric flasks.

In other words, you cannot use a 1-liter volumetric flask to make $500\ \mathrm{mL}$ of a solution. It can only be used to prepare 1 liter of a solution. The steps to follow when preparing a solution with a 1-liter volumetric flask are outlined below, and shown in the figure below.

1. The appropriate mass of solute is weighed out and added to a volumetric flask that has been about half-filled with distilled water.

2. The solution is swirled until all of the solute dissolves.

3. More distilled water is carefully added up to the line etched on the neck of the flask.

4. The flask is capped and inverted several times to completely mix.

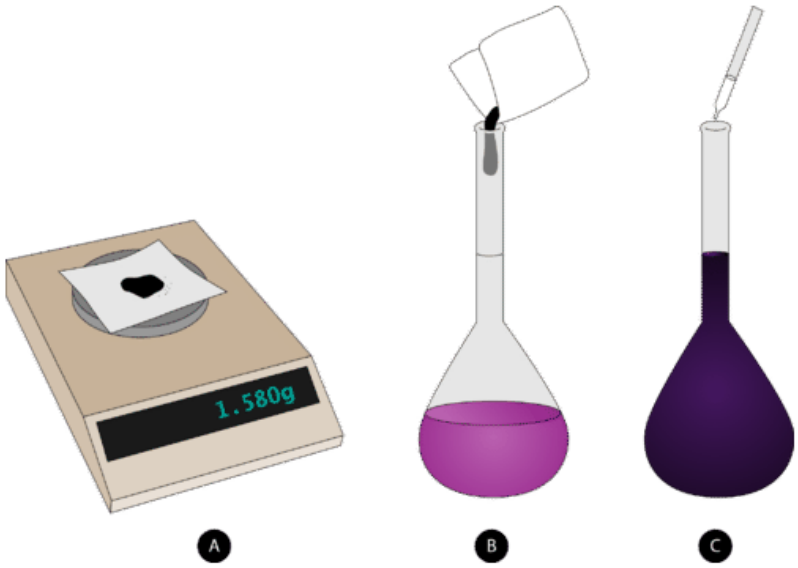


Figure 16.10.2: Steps to follow in preparing a solution of known molarity: (A) weigh out correct mass of solute, (B) dissolve into solvent in a volumetric flask, and (C) add solvent to the fill line on the flask and mix.

Summary

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16.11: Dilution

Muriatic acid (another name for HCl) is widely used for cleaning concrete and masonry surfaces. The acid must be diluted before use to get it down to a safer strength. Commercially available at concentrations of about 18%, muriatic acid can be used to remove scales and deposits (usually composed of basic materials).

Dilutions

When additional water is added to an aqueous solution, the concentration of that solution decreases. This is because the number of moles of the solute does not change, while the volume of the solution increases. We can set up an equality between the moles of the solute before the dilution (1) and the moles of the solute after the dilution (2).

$$\mathrm{mol}_1 = \mathrm{mol}_2$$

Since the moles of solute in a solution is equal to the molarity multiplied by the liters, we can set those equal.

$$M_1 \times L_1 = M_2 \times L_2$$

Finally, because the two sides of the equation are set equal to one another, the volume can be described in any unit that we choose, as long as that unit is the same on both sides. Our equation for calculating the molarity of a diluted solution becomes:

$$M_1 \times V_1 = M_2 \times V_2 \tag{16.11.1}$$

Suppose that you have $100.\ \mathrm{mL}$ of a $2.0\ \mathrm{M}$ solution of HCl. You dilute the solution by adding enough water to make the solution volume $500.\ \mathrm{mL}$. The new molarity can easily be calculated by using the above equation and solving for M_2:

$$M_2 = \frac{M_1 \times V_1}{V_2} = \frac{2.0\ \mathrm{M} \times 100.\ \mathrm{mL}}{500.\ \mathrm{mL}} = 0.40\ \mathrm{M\ HCl}$$

The solution has been diluted by one-fifth since the new volume is five times as great as the original volume. Consequently, the molarity is one-fifth of its original value.

Another common dilution problem involves deciding how much of a highly concentrated solution is required to make a desired quantity of solution of lesser concentration. The highly concentrated solution is typically referred to as the stock solution.

Example 16.11.1

Nitric acid $(\mathrm{HNO_3})$ is a powerful and corrosive acid. When ordered from a chemical supply company, its molarity is $16\ \mathrm{M}$. How much of the stock solution of nitric acid needs to be used to make $8.00\ \mathrm{L}$ of a $0.50\ \mathrm{M}$ solution?

Solution

Step 1: List the known quantities and plan the problem.

Known

- Stock $\mathrm{HNO_3} = 16\ \mathrm{M}$
- $V_2 = 8.00\ \mathrm{L}$
- $M_2 = 0.50\ \mathrm{M}$

Unknown

The unknown in the equation is V_1, the volume of the concentrated stock solution.

Step 2: Solve

Using Equation 16.11.1 we can solve for V_1:

$$V_1 = \frac{M_2 \times V_2}{M_1} = \frac{0.50\ \text{M} \times 8.00\ \text{L}}{16\ \text{M}} = 0.25\ \text{L} = 250\ \text{mL}$$

Step 3: Think about your result

$250\ \text{mL}$ of the stock HNO_3 needs to be diluted with water to a final volume of $8.00\ \text{L}$. The dilution is by a factor of 32 to go from $16\ \text{M}$ to $0.5\ \text{M}$.

Dilutions can be performed in the laboratory with various tools, depending on the volumes required and the desired accuracy. The images below illustrate the use of two different types of pipettes. In the first figure, a glass **pipette** is being used to transfer a portion of a solution to a graduated cylinder. Use of a pipette rather than a graduated cylinder for the transfer improves accuracy. The second figure shows a **micropipette**, which is designed to quickly and accurately dispense small volumes. Micropipettes are adjustable and come in a variety of sizes.



Figure 16.11.1: Volumetric pipette and micropipette.

Summary

- A process is described to calculate dilutions.
- Pipettes and micropipettes are used to transfer and dispense solution.

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16.12: Molality

For many purposes, the use of molarity is very convenient. However, when we want to know the concentration of solute present in situations where there are temperature changes, molarity will not work. The volume of the solution will change somewhat with temperature, enough to affect the accuracy of data observations and calculations. Another parameter is needed, one not affected by the temperature of the material we are studying.

Molality

A final way to express the concentration of a solution is by its molality. The **molality** (m) of a solution is the moles of solute divided by the kilograms of solvent. A solution that contains $1.0\ \text{mol}$ of NaCl dissolved into $1.0\ \text{kg}$ of water is a "one-molal" solution of sodium chloride. The symbol for molality is a lower-case m written in italics.

$$\text{Molality}\,(m) = \frac{\text{moles of solute}}{\text{kilograms of solvent}} = \frac{\text{mol}}{\text{kg}}$$

Molality differs from molarity only in the denominator. While molarity is based on the liters of solution, molality is based on the kilograms of solvent. Concentrations expressed in molality are used when studying properties of solutions related to vapor pressure and temperature changes. Molality is used because its value does not change with changes in temperature. The volume of a solution, on the other hand, is slightly dependent upon temperature.

Example 16.12.1

Determine the molality of a solution prepared by dissolving $28.60\ \text{g}$ of glucose $(\text{C}_6\text{H}_{12}\text{O}_6)$ into $250\ \text{g}$ of water.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass solute $= 28.60\ \text{g}\ \text{C}_6\text{H}_{12}\text{O}_6$
- Mass solvent $= 250\ \text{g} = 0.250\ \text{kg}$
- Molar mass $\text{C}_6\text{H}_{12}\text{O}_6 = 180.18\ \text{g/mol}$

Unknown

Convert grams of glucose to moles and divide by the mass of the water in kilograms.

Step 2: Solve.

$$28.60\ \text{g}\ \text{C}_6\text{H}_{12}\text{O}_6 \times \frac{1\ \text{mol}\ \text{C}_6\text{H}_{12}\text{O}_6}{180.18\ \text{g}\ \text{C}_6\text{H}_{12}\text{O}_6} = 0.1587\ \text{mol}\ \text{C}_6\text{H}_{12}\text{O}_6$$

$$\frac{0.1587\ \text{mol}\ \text{C}_6\text{H}_{12}\text{O}_6}{0.250\ \text{kg}\ \text{H}_2\text{O}} = 0.635\ m\ \text{C}_6\text{H}_{12}\text{O}_6$$

Step 3: Think about your result.

The answer represents the moles of glucose per kilogram of water and has three significant figures.

Molality and molarity are closely related in value for dilute aqueous solutions because the density of those solutions is relatively close to $1.0\ \text{g/mL}$. This means that $1.0\ \text{L}$ of solution has nearly a mass of $1.0\ \text{kg}$. As the solution becomes more concentrated, its density will not be as close to $1.0\ \text{g/mL}$ and the molality value will be different than the molarity. For solutions with solvents other than water, the molality will be very different than the molarity. Make sure that you are paying attention to which quantity is being used in a given problem.

Summary

- The molality (m) of a solution is the moles of solute divided by the kilograms of solvent.
- The value of molality does not change with temperature.

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16.13: The Lowering of Vapor Pressure

A knowledge of Latin can help us to better understand the English language—specifically, scientific terminology. Consider the word "colligative". Where did that come from? If you know a little Latin, you know that it comes from two Latin words meaning "to tie together".

Lowering Vapor Pressure

A **colligative property** is a property of a solution that depends only on the number of solute particles dissolved in the solution, and not on their identity. Recall that the vapor pressure of a liquid is determined by how easily its molecules are able to escape the surface of the liquid and enter the gaseous phase. When a liquid evaporates easily, it will have a relatively large number of its molecules in the gas phase and thus will have a high vapor pressure. Liquids that do not evaporate easily have a lower vapor pressure. The figure below shows the surface of a pure solvent compared to a solution. In the picture on the left, the surface is entirely occupied by liquid molecules, some of which will evaporate and form a vapor pressure. On the right, a nonvolatile solute has been dissolved into the solvent. **Nonvolatile** means that the solute itself has little tendency to evaporate. Because some of the surface is now occupied by solute particles, there is less room for solvent molecules. This results in less solvent being able to evaporate. The addition of a nonvolatile solute results in a lowering of the vapor pressure of the solvent.



Figure 16.13.1: The solution on the right has had some of its solvent particles replaced by solute particles. Since the solute particles do not evaporate, the vapor pressure of the solution is lower than that of the pure solvent.

The lowering of the vapor pressure depends on the number of solute particles that have been dissolved. The chemical nature of the solute is not important because the vapor pressure is merely a physical property of the solvent. The only requirement is that the solute was *only* dissolved and that it did not undergo a chemical reaction with the solvent.

While the chemical nature of the solute is not a factor, it is necessary to take into account whether the solute is an electrolyte or a nonelectrolyte. Recall that ionic compounds are strong electrolytes, and thus dissociate into ions when they dissolve. This results in a larger number of dissolved particles. For example, consider two different solutions of equal concentration: one is made from the ionic compound sodium chloride, while the other is made from the molecular compound glucose. The following equations show what happens when these solutions dissolve.

$$\begin{aligned} &\mathrm{NaCl}\,(s) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq) && 2 \text{ dissolved particles} \\ &\mathrm{C_6H_{12}O_6}\,(s) \rightarrow \mathrm{C_6H_{12}O_6}\,(aq) && 1 \text{ dissolved particle} \end{aligned}$$

The sodium chloride dissociates into two ions, while the glucose does not dissociate. Therefore, equal concentrations of each solution will result in twice as many dissolved particles in the case of the sodium chloride. The vapor pressure of the sodium chloride solution will be lowered twice the amount of the glucose solution.

Summary

- A colligative property is a property of a solution that depends only on the number of solute particles dissolved in the solution, and not on their identity.
- Nonvolatile means that the solute itself has little tendency to evaporate.

- Addition of a nonvolatile solute to a solution lowers the vapor pressure of the solution.

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16.14: Freezing Point Depression

Colligative properties have practical applications, such as the salting of roads in cold-weather climates. By applying salt to an icy road, the melting point of the ice is decreased, and the ice will melt more quickly, making driving safer. Sodium chloride (NaCl), and either calcium chloride $(\mathrm{CaCl_2})$ or magnesium chloride $(\mathrm{MgCl_2})$ are used most frequently, either alone or in a mixture. Sodium chloride is the least expensive option, but is less effective because it only dissociates into two ions, instead of three.

Freezing Point Depression

The figure below shows the phase diagram for a pure solvent and how it changes when a solute is added to it. The solute lowers the vapor pressure of the solvent, resulting in a lowering of the freezing point of the solution compared to the solvent. The **freezing point depression** is the difference in temperature between the freezing point of the pure solvent and that of the solution. On the graph, the freezing point depression is represented by ΔT_f.

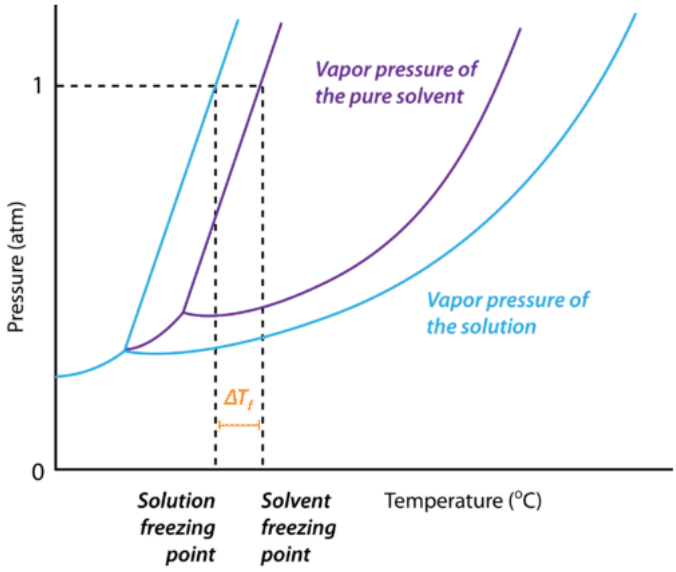


Figure $16.14.1$: The vapor pressure of a solution (blue) is lower than the vapor pressure of a pure solvent (pink). As a result, the freezing point of a solvent is lower when any solute is dissolved into it.

When a pure solvent freezes, its particles become more ordered as the intermolecular forces that operate between the molecules become permanent. In the case of water, the hydrogen bonds make the hexagonally-shaped network of molecules that characterizes the structure of ice. By dissolving a solute into the liquid solvent, this ordering process is disrupted. As a result, more energy must be removed from the solution in order to freeze it, and the freezing point of the solution is lower than that of the pure solvent.

The magnitude of the freezing point depression is directly proportional to the molality of the solution. The equation is:

$$\Delta T_f = K_f \times m$$

The proportionality constant, K_f, is called the **molal freezing-point depression constant**. It is a constant that is equal to the change in the freezing point for a 1-molal solution of a nonvolatile molecular solute. For water, the value of K_f is $-1.86^\circ\mathrm{C}/m$. So, the freezing temperature of a 1-molal aqueous solution of any nonvolatile molecular solute is $-1.86^\circ\mathrm{C}$. Every solvent has a unique molal freezing-point depression constant. These are shown in the table below, along with a related value for the boiling point called K_b.

Molal Freezing-Point and Boiling-Point Constants

Table 16.14.1: Molal Freezing-Point and Boiling-Point Constants

Solvent	Normal Freezing Point ($^\circ C$)	Molal Freezing-Point Depression Constant, K_f ($^\circ C/m$)	Normal Boiling Point ($^\circ C$)	Molal Boiling-Point Elevation Constant, K_b ($^\circ C/m$)
Acetic acid	16.6	-3.90	117.9	3.07
Camphor	178.8	-39.7	207.4	5.61
Naphthalene	80.2	-6.94	217.7	5.80
Phenol	40.9	-7.40	181.8	3.60
Water	0.00	-1.86	100.00	0.512

Example 16.14.1

Ethylene glycol ($C_2H_6O_2$) is a molecular compound that is used in many commercial antifreezes. A water solution of ethylene glycol is used in vehicle radiators to lower its freezing point, and thus prevent the water in the radiator from freezing. Calculate the freezing point of a solution of $400\ g$ of ethylene glycol in $500\ g$ of water.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass $C_2H_6O_2 = 400\ g$
- Molar mass $C_2H_6O_2 = 62.08\ g/mol$
- Mass $H_2O = 500\ g = 0.500\ kg$
- $K_f\,(H_2O) = -1.86^\circ C/m$

Unknown

This is a three-step problem. First, calculate the moles of ethylene glycol. Then, calculate the molality of the solution. Finally, calculate the freezing point depression.

Step 2: Solve.

$$400.\ g\ C_2H_6O_2 \times \frac{1\ mol\ C_2H_6O_2}{62.08\ g\ C_2H_6O_2} = 6.44\ mol\ C_2H_6O_2$$

$$\frac{6.44\ mol\ C_2H_6O_2}{0.500\ kg\ H_2O} = 12.9\ m\ C_2H_6O_2$$

$$\Delta T_f = K_f \times m = -1.86^\circ C/m = -24.0^\circ C$$

$$T_f = -24.0^\circ C$$

The normal freezing point of water is $0.0^\circ C$. Therefore, since the freezing point decreases by $24.0^\circ C$, the freezing point of the solution is $-24.0^\circ C$.

Step 3: Think about your result.

The freezing point of the water decreases by a large amount, protecting the radiator from damage due to the expansion of water when it freezes. There are three significant figures in the result.

Summary

- The freezing point depression is the difference in temperature between the freezing point of the pure solvent and that of the solution.
- The molal freezing-point depression constant is equal to the change in the freezing point for a 1-molal solution of a nonvolatile molecular solute.

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- Calculations involving freezing point depression are described.

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16.15: Boiling Point Elevation

Salt is often added to boiling water when preparing spaghetti or other pasta. One reason is to add flavor to the food. Some people believe that the addition of salt increases the boiling point of the water. Technically, they are correct, but the increase is rather small. You would need to add over 100 grams of NaCl to a liter of water to increase the boiling point by a couple of degrees, which is likely unhealthy.

Boiling Point Elevation

The figure below shows the phase diagram of a solution and the effect that the lowered vapor pressure has on the boiling point of the solution compared to the solvent. In this case, the solution has a higher boiling point than the pure solvent. Since the vapor pressure of the solution is lower, more heat must be supplied to the solution to bring its vapor pressure up to the pressure of the external atmosphere. The **boiling point elevation** is the difference in temperature between the boiling point of the pure solvent and that of the solution. On the graph, the boiling point elevation is represented by ΔT_b.

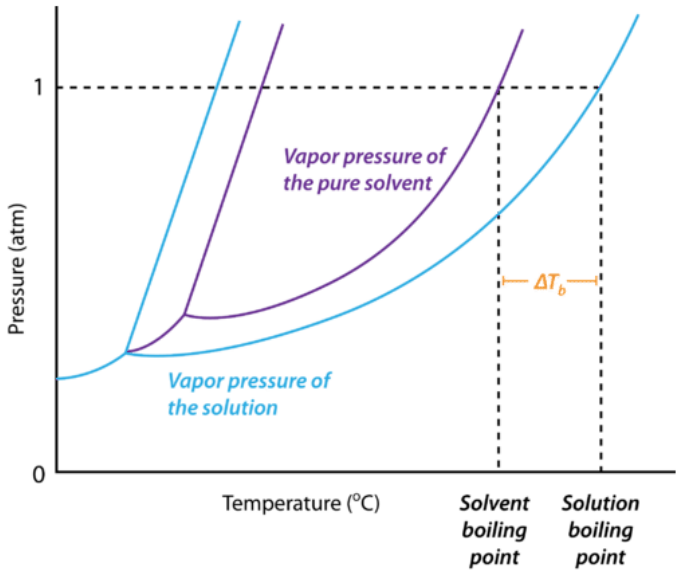


Figure $16.15.1$: The lowering of the vapor pressure in a solution causes the boiling point of the solution to be higher than that of the pure solvent.

The magnitude of the boiling point elevation is also directly proportional to the molality of the solution. The equation is:

$$\Delta T_b = K_b \times m$$

The proportionality constant, K_b, is called the **molal boiling-point elevation constant**. It is a constant that is equal to the change in the boiling point for a 1-molal solution of a nonvolatile molecular solute. For water, the value of K_b is $0.512^\circ\mathrm{C}/m$. So, the boiling temperature of a 1-molal aqueous solution of any nonvolatile molecular solute is $100.512^\circ\mathrm{C}$.

Summary

- Boiling point elevation is the difference in temperature between the boiling point of the pure solvent and that of the solution.
- The molal boiling-point elevation constant is equal to the change in the boiling point for a 1-molal solution of a nonvolatile molecular solute.
- Calculations involving the molal boiling point elevation constant are outlined.

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16.16: Electrolytes and Colligative Properties

The addition of ions creates significant changes in properties of solutions. Water molecules surround the ions and are somewhat tightly bound to them. Colligative properties are affected because the solvent properties are no longer the same as those in the pure solvent.

Electrolytes and Colligative Properties

Ionic compounds are electrolytes and dissociate into two or more ions as they dissolve. This must be taken into account when calculating the freezing and boiling points of electrolyte solutions. The sample problem below demonstrates how to calculate the freezing point and boiling point of a solution of calcium chloride. Calcium chloride dissociates into three ions according to the equation:

$$\mathrm{CaCl_2}\,(s) \rightarrow \blacksquare\,(aq) + 2\mathrm{Cl^-}\,(aq)$$

The values of the freezing point depression and the boiling point elevation for a solution of $\mathrm{CaCl_2}$ will be three times greater than they would be for an equal molality of a nonelectrolyte.

Example 16.16.1

Determine the freezing point and boiling point of a solution prepared by dissolving $82.20\ \mathrm{g}$ of calcium chloride into $400\ \mathrm{g}$ of water.

Solution

Step 1: List the known quantities and plan the problem.

Known

- Mass $\mathrm{CaCl_2} = 82.20\ \mathrm{g}$
- Molar mass $\mathrm{CaCl_2} = 110.98\ \mathrm{g/mol}$
- Mass $\mathrm{H_2O} = 400\ \mathrm{g} = 0.400\ \mathrm{kg}$
- $K_f\,(\mathrm{H_2O}) = -1.86^{\circ}\mathrm{C}/m$
- $K_b\,(\mathrm{H_2O}) = 0.512^{\circ}\mathrm{C}/m$
- $\mathrm{CaCl_2}$ dissociates into 3 ions

Unknown

- $T_f = ?\ ^{\circ}\mathrm{C}$
- $T_b = ?\ ^{\circ}\mathrm{C}$

The moles of $\mathrm{CaCl_2}$ is first calculated, followed by the molality of the solution. The freezing and boiling points are then determined, including multiplying by 3 for the three ions.

Step 2: Solve.

$$82.20\ \mathrm{g\ CaCl_2} \times \frac{1\ \mathrm{mol\ CaCl_2}}{110.98\ \mathrm{g\ CaCl_2}} = 0.7407\ \mathrm{mol\ CaCl_2}$$

$$\frac{0.7407\ \mathrm{mol\ CaCl_2}}{0.400\ \mathrm{kg\ H_2O}} = 1.85\ m\ \mathrm{CaCl_2}$$

$$\Delta T_f = K_f \times m \times 3 = -1.86^{\circ}\mathrm{C}/m \times 1.85\ m \times 3 = -10.3^{\circ}\mathrm{C} \quad T_f = -10.3^{\circ}\mathrm{C}$$

$$\Delta T_b = K_b \times m \times 3 = 0.512^{\circ}\mathrm{C}/m \times 1.85\ m \times 3 = 2.84^{\circ}\mathrm{C} \quad T_b = 102.84^{\circ}\mathrm{C}$$

Step 3: Think about your result.

Since the normal boiling point of water is $100.00^{\circ}\mathrm{C}$, the calculated result for ΔT_b must be added to 100.00 to find the new boiling point.

Summary

- Ionic compounds are electrolytes and dissociate into two or more ions as they dissolve.
- The effect of ionization on colligative properties is described.

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16.17: Osmosis and Osmotic Pressure

Learning Objectives

- Describe osmosis and how it relates to osmotic pressure.

Osmotic Pressure

The last colligative property of solutions we will consider is a very important one for biological systems. It involves **osmosis**, the process by which solvent molecules can pass through certain membranes but solute particles cannot. When two solutions of different concentration are present on either side of these membranes (called *semipermeable membranes*), there is a tendency for solvent molecules to move from the more dilute solution to the more concentrated solution until the concentrations of the two solutions are equal. This tendency is called **osmotic pressure**. External pressure can be exerted on a solution to counter the flow of solvent; the pressure required to halt the osmosis of a solvent is equal to the osmotic pressure of the solution.

Consider the apparatus illustrated in Figure 16.17.1, in which samples of pure solvent and a solution are separated by a membrane that only solvent molecules may permeate. Solvent molecules will diffuse across the membrane in both directions. Since the concentration of *solvent* is greater in the pure solvent than the solution, these molecules will diffuse from the solvent side of the membrane to the solution side at a faster rate than they will in the reverse direction. The result is a net transfer of solvent molecules from the pure solvent to the solution. Diffusion-driven transfer of solvent molecules through a semipermeable membrane is a process known as osmosis.

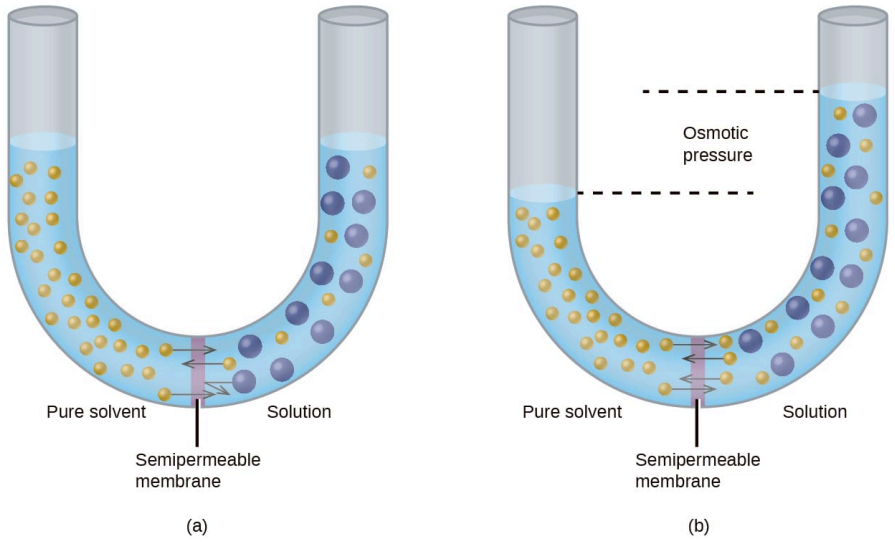


Figure 16.17.1: Osmosis results in the transfer of solvent molecules from a sample of low (or zero) solute concentration to a sample of higher solute concentration.

When osmosis is carried out in an apparatus like that shown in Figure 16.17.1, the volume of the solution increases as it becomes diluted by accumulation of solvent. This causes the level of the solution to rise, increasing its hydrostatic pressure (due to the weight of the column of solution in the tube) and resulting in a faster transfer of solvent molecules back to the pure solvent side. When the pressure reaches a value that yields a reverse solvent transfer rate equal to the osmosis rate, bulk transfer of solvent ceases. This pressure is called the osmotic pressure (π) of the solution. The osmotic pressure of a dilute solution can be determined in a similar way the pressure of an ideal gas is calculated using the ideal gas law:

$$\pi = \left(\frac{n}{V}\right) RT$$

where n is the number of moles of particles in solution, V is the volume, and R is the universal gas constant.

Osmolarity (osmol) is a way of reporting the total number of particles in a solution to determine osmotic pressure. It is defined as the molarity of a solute times the number of particles a formula unit of the solute makes when it dissolves (represented by i):

$$osmol = M \times i$$

If more than one solute is present in a solution, the individual osmolarities are additive to get the total osmolarity of the solution. Solutions that have the same osmolarity have the same osmotic pressure. If solutions of differing osmolarities are present on opposite sides of a semipermeable membrane, solvent will transfer from the lower-osmolarity solution to the higher-osmolarity solution. Counterpressure exerted on the high-osmolarity solution will reduce or halt the solvent transfer. An even higher pressure can be exerted to force solvent from the high-osmolarity solution to the low-osmolarity solution, a process called *reverse osmosis*. Reverse osmosis is used to make potable water from saltwater where sources of fresh water are scarce.

Example 16.17.1

A 0.50 M NaCl aqueous solution and a 0.30 M $Ca(NO_3)_2$ aqueous solution are placed on opposite sides of a semipermeable membrane. Determine the osmolarity of each solution and predict the direction of solvent flow.

Solution

The solvent will flow into the solution of higher osmolarity. The NaCl solute separates into two ions—Na^+ and Cl^-—when it dissolves, so its osmolarity is as follows:

$$\text{osmol (NaCl)} = 0.50\ \text{M} \times 2 = 1.0\ \text{osmol}$$

The $Ca(NO_3)_2$ solute separates into three ions—one Ca^{2+} and two NO_3^-—when it dissolves, so its osmolarity is as follows:

$$\text{osmol } [Ca(NO_3)_2] = 0.30\ \text{M} \times 3 = 0.90\ \text{osmol}$$

The osmolarity of the $Ca(NO_3)_2$ solution is lower than that of the NaCl solution, so water will transfer through the membrane from the $Ca(NO_3)_2$ solution to the NaCl solution.

Exercise 16.17.1

A 1.5 M $C_6H_{12}O_6$ aqueous solution and a 0.40 M $Al(NO_3)_3$ aqueous solution are placed on opposite sides of a semipermeable membrane. Determine the osmolarity of each solution and predict the direction of solvent flow.

Answer available at the end of this section.

Examples of osmosis are evident in many biological systems because cells are surrounded by semipermeable membranes. Carrots and celery that have become limp because they have lost water can be made crisp again by placing them in water. Water moves into the carrot or celery cells by osmosis. A cucumber placed in a concentrated salt solution loses water by osmosis and absorbs some salt to become a pickle. Osmosis can also affect animal cells. Solute concentrations are particularly important when solutions are injected into the body. Solutes in body cell fluids and blood serum give these solutions an osmotic pressure of approximately 7.7 atm. Solutions injected into the body must have the same osmotic pressure as blood serum; that is, they should be isotonic with blood serum. If a *less* concentrated solution, a **hypotonic** solution, is injected in sufficient quantity to dilute the blood serum, water from the diluted serum passes into the blood cells by osmosis, causing the cells to expand and rupture (Figure $16.17.2a$). This process is called hemolysis. When a *more* concentrated solution, a **hypertonic** solution, is injected, the cells lose water to the more concentrated solution, shrivel, and possibly die in a process called *crenation* (Figure $16.17.2c$). Only if red blood cells are placed in **isotonic** solutions that have the same osmolarity as exists inside the cells are they unaffected by negative effects of osmotic pressure (Figure $16.17.2b$). Glucose solutions of about 0.31 M, or sodium chloride (NaCl) solutions of about 0.16 M, are isotonic with blood plasma.

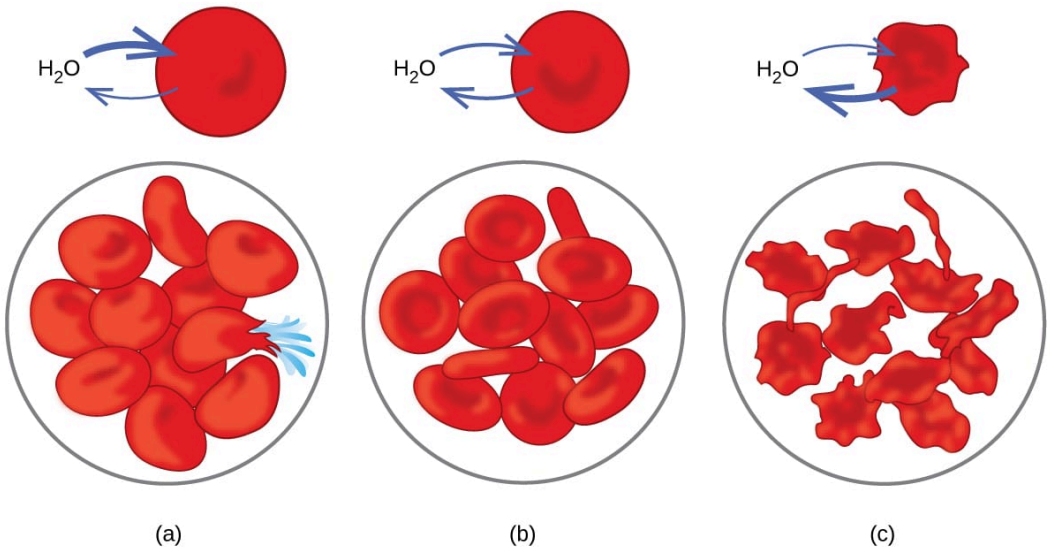


Figure 16.17.2: Red blood cell membranes are water permeable and will (a) swell and possibly rupture in a **hypotonic** solution; (b) maintain normal volume and shape in an **isotonic** solution; and (c) shrivel and possibly die in a **hypertonic** solution. (credit a/b/c: modifications of work by "LadyofHats"/Wikimedia commons)

Note: Isotonic Solutions for Red Blood Cells

The concentration of an red blood cell isotonic solution made with sodium chloride (NaCl) is half that of an isotonic solution made with glucose (0.16 M and 0.31 M respectively). This is because NaCl produces *two ions* when a formula unit dissolves, while molecular glucose produces only *one particle* when a formula unit dissolves. The osmolarities are therefore the same even though the concentrations of the two solutions are different.

isotonic NaCl solution: $osmol = 0.16\ M \times 2 = 0.32\ osmol/L$

isotonic glucose solution: $osmol = 0.31\ M \times 1 = 0.31\ osmol/L$

Osmotic pressure explains why you should not drink seawater if you are abandoned in a life raft in the middle of the ocean. Its osmolarity is about three times higher than most bodily fluids. You would actually become thirstier as water from your cells was drawn out to dilute the salty ocean water you ingested. Our bodies do a better job coping with hypotonic solutions than with hypertonic ones. The excess water is collected by our kidneys and excreted.

Osmotic pressure effects are used in the food industry to make pickles from cucumbers and other vegetables and in brining meat to make corned beef. It is also a factor in the mechanism of getting water from the roots to the tops of trees!

Answers

Exercise 16.17.1

osmol $C_6H_{12}O_6$ = 1.5; osmol $Al(NO_3)_3$ = 1.6

The solvent flows from $C_6H_{12}O_6$ solution (lower osmolarity) to $\mathrm{Al(NO_3)_3}$ solution (higher osmolarity).

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16.18: Calculating Molar Mass

Putting antifreeze into a radiator will keep an engine from freezing. By knowing how cold the weather will get and how much water is in the radiator, one can determine how much antifreeze to add to achieve a specific desired freezing point depression. This is made possible by knowing what antifreeze is. Can things be switched around, so as to get some information about the properties of the antifreeze (such as its molecular weight) from the freezing point decrease? As it turns out, this can be done fairly easily and accurately.

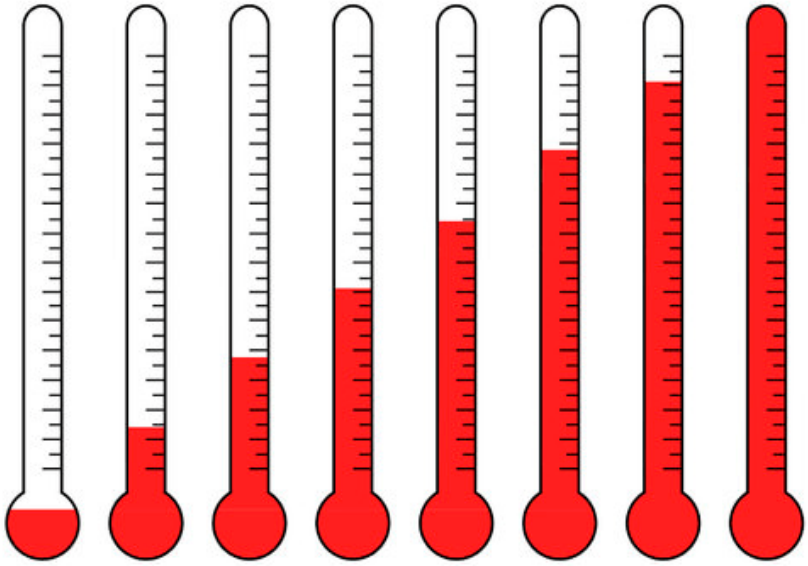


Figure 16.18.1: Changes in temperature.

Calculating Molar Mass

In the laboratory, freezing point or boiling point data can be used to determine the molar mass of an unknown solute. Since we know the relationship between a decrease in freezing point and the concentration of solute, if we dissolve a known mass of our unknown solute into a known amount of solvent, we can calculate the molar mass of the solute. The K_f or K_b of the solvent must be known. We also need to know if the solute is an electrolyte or a nonelectrolyte. If the solvent is an electrolyte, we would need to know the number of ions produced when it dissociates.

Example 16.18.1

$38.7\ \text{g}$ of a nonelectrolyte is dissolved into $218\ \text{g}$ of water. The freezing point of the solution is measured to be -5.53°C. Calculate the molar mass of the solute.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $\Delta T_f = -5.53^{\circ}\text{C}$
- Mass $\text{H}_2\text{O} = 218\ \text{g} = 0.218\ \text{kg}$
- Mass solute $= 38.7\ \text{g}$
- $K_f\,(\text{H}_2\text{O}) = -1.86^{\circ}\text{C}/m$

Unknown

Use the freezing point depression (ΔT_f) to calculate the molality of the solution. Then use the molality equation to calculate the moles of solute. Then divide the grams of solute by the moles to determine the molar mass.

Step 2: Solve.

$$m = \frac{\Delta T_f}{K_f} = \frac{-5.53^\circ\text{C}}{-1.86^\circ\text{C}/m} = 2.97\ m$$
$$\text{mol solute} = m \times \text{kg}\ \text{H}_2\text{O} = 2.97\ m \times 0.218\ \text{kg} = 0.648\ \text{mol}$$
$$\frac{38.7\ \text{g}}{0.648\ \text{mol}} = 59.7\ \text{g/mol}$$

Step 3: Think about your result.

The molar mass of the unknown solute is $59.7\ \text{g/mol}$. Knowing the molar mass is an important step in determining the identity of an unknown. A similar problem could be done with the change in boiling point.

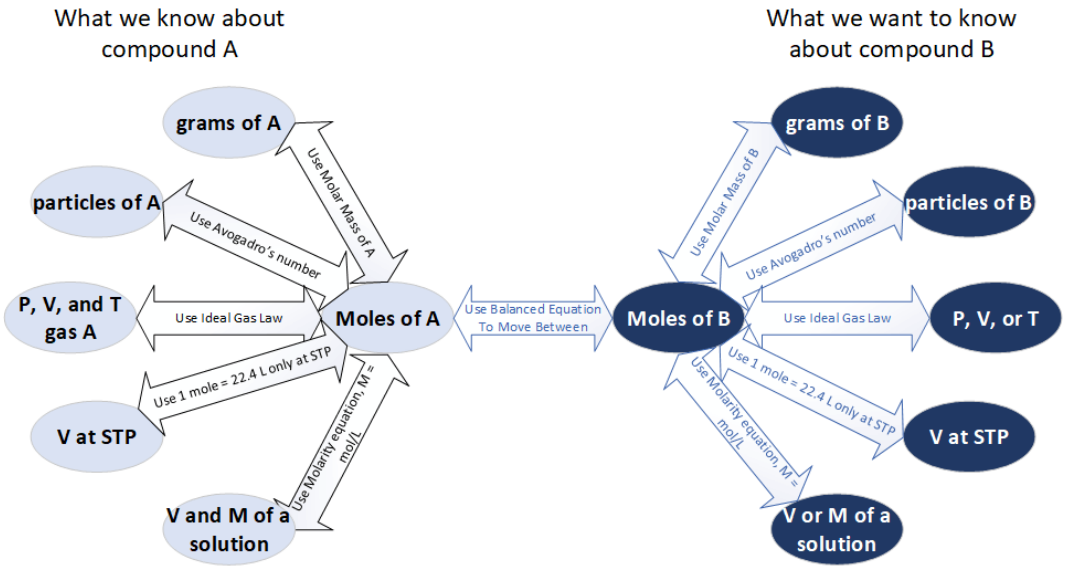
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16.19: Stoichiometry Map

Solving problems in chemistry requires choosing the right types of equations and equivalencies to get from what we know to what we want to know. In stoichiometry, we have often complex calculations, with many steps, so we need to organized to make sure we're choosing the right steps. Below, you can find a summarized map of steps for solving various types of stoichiometric problems. For all of these problems, we have some type of measurement or measurements of one substance, but we want to know something about a different substance. We also know or can find the balanced equation that relates these two substances to each other. Notice that no matter what information we start or end with, we need to find the **moles!** Each of the relationships in this map is explained further below.



Finding moles of A

We can find the moles of A from a wide range of information. You do not need the balanced equation to find the moles of A, but you would need to know the molar mass of the substance!

If we're given the grams of A:

This was covered in previous chapter, but it helps to have a reminder of the steps:

1. Find the chemical formula of A, if it isn't given to you
2. Use the chemical formula to find the molar mass (remember, molar mass is measured in g/mol)
3. Use the molar mass and dimensional analysis to convert the g A to mol A

Example 16.19.1

Convert 2.34 g of sodium chloride to moles.

Solution

1. the formula for sodium chloride is NaCl
2. Use the Periodic Table to find the molar mass of NaCl.
   - Mass of Na = 22.989769 g/mol
   - Mass of Cl = 35.453 g/mol

- Add the masses = 22.989769 g/mol + 35.453 g/mol = 58.442769 g/mol
- Round for proper significant figures = 58.443 g/mol (remember - we're using addition, so we round based on the least decimal places)

3. Use dimensional analysis to perform the conversion

$$2.34 \ \cancel{g \ NaCl}\left(\frac{1 \ mol \ NaCl}{58.443 \ \cancel{g \ NaCl}}\right) = 0.400 \ mol \ NaCl$$

If we're given the particles of A:

This is review, too! Here's a reminder of your steps:

1. Make sure the information given matches A. For example, if you are given that there are 1.2 * 10^{24} molecules of CO_2, and the balanced equation includes CO_2, then you'll be able to go to the next step directly. If, instead, you are given that there are 1.2 * 10^{24} ions from NaCl and the balanced equation is for NaCl, you'll first have to convert your number of ions to your formula units of NaCl.
2. Use Avogadro's number and dimensional analysis to convert from particles to moles, $6.022*10^{23}$ particles = 1 mole of particle.

Example 16.19.2

If there are $1.2*10^{24}$ ions from NaCl, how many moles of NaCl is this?

Solution

1. Since you are given ions, not the full NaCl, you will first need to convert the number of ions to the number of formula units of NaCl. If you recall, each NaCl will break up into one Na^+ ion and one Cl^- ion. This gives the following relationship:
   - 1 NaCl = 2 ions
   - Using this relationship, we can convert the ions to full NaCl:
   - $$1.2 * 10^{24} \ \cancel{ions}\left(\frac{1 \ NaCl}{2 \ \cancel{ions}}\right) = 6.0 * 10^{23} \ NaCl$$
2. Convert the formula unit of NaCl to moles
   - $$6.0 * 10^{23} \ \cancel{formula \ units \ NaCl}\left(\frac{1 \ mole \ NaCl}{6.022 * 10^{23} \ \cancel{formula \ units \ NaCl}}\right) = 1.0 \ mol \ NaCl$$

Converting A to moles, if it's a Gas

In the previous 2 examples, the substance we're converting could be a liquid, a gas, or a solid. There are specific, additional calculations we can do if A is a gas. **We can only use these if A is a gas!** Check the balanced equation - is there a (g) after the substance? This is also review, but there are two different conditions we look at for gases, so be careful to use the right one! Here's our steps:

1. Check if the gas is at STP (standard temperature and pressure). STP is only at a temperature of 273 K (or 0 degrees C) **and** a pressure of 1 atm. You will be told that the gas is at STP or given both the temperature and pressure. If the gas is at STP, then you can use the molar volume to find the moles of gas. This relationship is given below:
   - 1 mole of a gas = 22.4 L at STP
2. If the gas is not at STP, you'll typically need to use the ideal gas law and solve for the moles of gas (n from the equation). Ideal gas law:
   - PV = nRT

Example 16.19.3

Convert 85.3 L of a gas at STP to moles.

Solution

This gas is at STP, so we can use the molar volume to convert the volume of the gas to the moles:

$$85.3 \ \cancel{L\ gas}\left(\frac{1\ mol\ gas}{22.4\ \cancel{L\ gas}}\right) = 3.81\ mol\ gas$$

Example 16.19.4

How many moles of hydrogen gas is there if the sample of gas is at a pressure of 750 mm Hg, a volume of 1.35 L, and a temperature of 45 degrees C?

Solution

This gas is not at STP, so we'll need to use the ideal gas law. Remember, to use the ideal gas law, we need the units to match the units of R, the ideal gas constant. We'll use the following R value:

$$\frac{0.0821\ L * atm}{mol * K}$$

That means the units need to match this R value.

From our problem:

P = 750 mm Hg, which does **not** match the pressure units of atm from the R value, so we'll need to convert it before we can continue.

$$750\ \cancel{mm\ Hg}\left(\frac{1\ atm}{760\ \cancel{mm\ Hg}}\right) = 0.99\ atm$$

V = 1.35 L, which matches the units of R, so we don't need to convert it

T = 45 degrees C, which does **not match**, so we'll need to convert to K. Remember, we always need temperature in kelvin to be able to do gas calculations.

45 C + 273.15 = 318 K

We'll also need to use the ideal gas law, PV=nRT, and rearrange to solve for n.

$$n = \frac{PV}{RT}$$

Substitute for the values from the problem and the known R value:

$$n = \frac{0.99\ atm * 1.35\ L}{318K * \frac{0.0821\ L * atm}{mol * K}}$$

= 0.051 moles of hydrogen gas

Converting A to moles, if it's a Solution

The previous examples all assume you have a direct measurement of the pure substance, such as pure solid NaCl or pure

hydrogen gas, but many substances are easier to handle and more common as a mixture, particularly as a solution. With these types of situations, we would need to use the information we have about the solution to find the moles of the component we are interested in. Be careful about these - if your mixture has multiple solutes in it, then you'll need to make sure you focus on the right solute. Common solutions with many solutes in them include foods, beverages, air, and blood. Here's how we approach these problems:

1. Identify if you are working with a solution. There are a variety of hints you might get about this - you won't always be told directly. Some ways to tell:
   - In the balanced equation, solutes that are dissolved in water are marked with (aq) - if the equation is fully labeled, this is a quick way to tell
   - We know or can more easily measure the volume (L or mL) of the solution and its Molarity
   - Rarely - you may know other measurements of the solution, such as the concentration in percent by mass, percent by volume, or percent mass/volume and either the volume or mass of the solution. These require more steps to solve, but are still solvable. We'll focus on Molarity for this course, but do be aware you could encounter other types of information in lab or clinical settings.
2. Once we know the volume and Molarity of a solute in a solution, we can find the moles using the Molarity equation

Example 16.19.5

You measure out 12.32 mL of 0.1085 M NaOH in lab. How many moles of NaOH is this?

Solution

We are given a volume, 12.32 mL, and a concentration, 0.1085 M NaOH. The concentration is M, which is Molarity, so we can use the Molarity equation.

$$M = \frac{moles\ solute}{L\ solution}$$

We need to rearrange the equation to solve for moles solute:

M * L solution = moles solute

Notice that this problem requires L of solution, but we know the volume of the solution in mL. That means we'll need to convert the volume to L **before** we do the problem:

$$12.32\ \cancel{mL}\left(\frac{10^{-3}\ L}{1\ \cancel{mL}}\right) = 0.01232\ L$$

Plug in our M and the volume in L to solve for moles solute:

moles solute = 0.1085 M * 0.01232 L = 0.00133672 moles NaOH = 0.001337 moles NaOH (once we round for significant figures)

You can verify the units by remembering that the units of M can also be rewritten as mol/L.

Converting moles A to moles B

We have one, single piece of information that allows us to convert from moles of one compound to moles of another compound - that's our balanced equation! Here's how we do that:

1. Check that your equation is complete and balanced.
   - If you don't know what the formulas are for the products, you'll need to figure that out first.
   - Make sure the equation is balanced - remember, you can **only** change the coefficients to balance a chemical equation, you cannot change the formulas of the reactants or products.
2. Use the molar ratio from the coefficients in the balanced equation to convert from moles of A to moles of B.

Example 16.19.6

If 0.351 mol Zn fully reacts with excess aqueous hydrochloric acid, how many moles of $ZnCl_2$ would be made?

Solution

Since this is a mole to mole conversion (notice we are starting with a different substance than we end with), we'll need to use the balanced equation. That means writing it out. Based on the information given, we can determine a few things:

- Our first reactant is Zn, which is in its element form. Looking at the Periodic Table, or going from the experience of seeing Zinc in lab or lab videos, we can determine that Zn would be a solid.
- Our second reactant is aqueous hydrochloric acid. Based on the name, we know the formula is HCl and that it is (aq)
- One of our products has to be $ZnCl_2$ since we're told it is "made."
- We'll need to write out the equation to determine what, if anything, is missing.

$$\mathrm{Zn}\,(s) + \mathrm{HCl}\,(aq) \rightarrow \mathrm{ZnCl_2} + \mathrm{unknown}$$

When we see this equation, we need to know what is going on to predict the products. Notice that the "H" in HCl from the reactants is replaced by "Zn" in $ZnCl_2$ for the products. This tells us that we're probably working with a single replacement reaction. That tells us that the unknown missing compound would be hydrogen. We know from previous chapters that when hydrogen is by itself it forms H_2, and it is found as a gas. Our $ZnCl_2$ is also missing its physical state. We can use our solubility rules to find that it would be soluble in water, so it must be (aq).

This gives the equation:

$$\mathrm{Zn}\,(s) + \mathrm{HCl}\,(aq) \rightarrow \mathrm{ZnCl_2}\ \mathrm{(aq)} + \mathrm{H_2}\,(g)$$

This equation isn't balanced, so we need to balance it:

$$\mathrm{Zn}\,(s) + 2\ \mathrm{HCl}\,(aq) \rightarrow \mathrm{ZnCl_2}\ \mathrm{(aq)} + \mathrm{H_2}\,(g)$$

Now we can use the balanced equation to find our moles of $ZnCl_2$ formed.

$$0.351\ \cancel{mol\ Zn}\left(\frac{1\ mol\ ZnCl_2}{1\ \cancel{mol\ Zn}}\right) = 0.351\ mol\ ZnCl_2$$

Notice that the 1's in this conversion factor come from the coefficients in the balanced equation. These could be **any** value, based on what they are in the balanced equation. Notice also that we **only** needed the balanced equation to convert moles to moles.

Using moles B to find other values for B

Generally, once we have the moles of B, we can simply reverse the processes we did to find the moles of A to convert the moles of B back into grams, particles, pressure, volume, temperature, or Molarity. Be very careful to track what you know so you pick an appropriate process! Make a plan!

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17: EQUILIBRIUM

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17.1: Reversible Reaction

A solution of cobalt chloride in water is pink due to the presence of the solvated $\mathrm{Co^{2+}}$ ion. If sufficient HCl is added, the solution turns blue as the $\mathrm{CoCl_4^{2-}}$ ion forms. The reaction can be shifted back to the pink form if more water is added to the solution.



Figure 17.1.1: Cobalt chloride in various stages of equilibrium with hydrochloric acid. (Public Domain; Chemicalinterest via wikipedia)

Chemical Reversibility

Up until this point, we have written the equations for chemical reactions in a way that would seem to indicate that all reactions proceed completely until all the reactants have been converted into products. In reality, a great many chemical reactions do not proceed entirely to completion. A **reversible reaction** is a reaction in which the conversion of reactants to products and the conversion of products to reactants occur simultaneously. One example of a reversible reaction is the reaction of hydrogen gas and iodine vapor to form hydrogen iodide. The forward and reverse reactions can be written as follows.

$$\text{Forward reaction:}\quad \mathrm{H_2\,(g) + I_2\,(g) \rightarrow 2HI\,(g)}$$

$$\text{Reverse reaction:}\quad \mathrm{2HI\,(g) \rightarrow H_2\,(g) + I_2\,(g)}$$

In the forward reaction, hydrogen and iodine combine to form hydrogen iodide. In the reverse reaction, hydrogen iodide decomposes back into hydrogen and iodine. The two reactions can be combined into one equation by the use of a double arrow:

$$\mathrm{H_2\,(g) + I_2\,(g) \rightleftharpoons 2HI\,(g)}$$

The double arrow is the indication that the reaction is reversible.

When hydrogen and iodine gases are mixed in a sealed container, they begin to react and form hydrogen iodide. At first, only the forward reaction occurs because no HI is present. As the forward reaction proceeds, it begins to slow down as the concentrations of the $\mathrm{H_2}$ and the $\mathrm{I_2}$ decrease. As soon as some HI has formed, it begins to decompose back into $\mathrm{H_2}$ and $\mathrm{I_2}$. The rate of the reverse reaction starts out slow because the concentration of HI is low. Gradually, the rate of the forward reaction decreases while the rate of the reverse reaction increases. Eventually the rate of combination of $\mathrm{H_2}$ and $\mathrm{I_2}$ to produce HI becomes equal to the rate of decomposition of HI into $\mathrm{H_2}$ and $\mathrm{I_2}$. When the rates of the forward and reverse reactions have become equal to one another, the reaction has achieved a state of balance.

Summary

- A reversible reaction is defined.

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17.2: Chemical Equilibrium

A tug of war involves two teams at the ends of a rope. The goal is to pull the other team over a line in the middle. At first, there is a great deal of tension on the rope, but no apparent movement. A bystander might think that there is nothing happening. In reality, there is a great deal of tension on the rope as the two teams pull in opposite directions at the same time.

Chemical Equilibrium

Hydrogen and iodine gases react to form hydrogen iodide, according to the following reaction:

$$\mathrm{H_2}\,(g) + \mathrm{I_2}\,(g) \rightleftharpoons 2\mathrm{HI}\,(g)$$

$$\text{Forward reaction:}\quad \mathrm{H_2}\,(g) + \mathrm{I_2}\,(g) \rightarrow 2\mathrm{HI}\,(g)$$

$$\text{Reverse reaction:}\quad 2\mathrm{HI}\,(g) \rightarrow \mathrm{H_2}\,(g) + \mathrm{I_2}\,(g)$$

Initially, only the forward reaction occurs because no HI is present. As soon as some HI has formed, it begins to decompose back into $\mathrm{H_2}$ and $\mathrm{I_2}$. Gradually, the rate of the forward reaction decreases while the rate of the reverse reaction increases. Eventually, the rate of combination of $\mathrm{H_2}$ and $\mathrm{I_2}$ to produce HI becomes equal to the rate of decomposition of HI into $\mathrm{H_2}$ and $\mathrm{I_2}$. When the rates of the forward and reverse reactions have become equal to one another, the reaction has achieved a state of balance. **Chemical equilibrium** is the state of a system in which the rate of the forward reaction is equal to the rate of the reverse reaction.

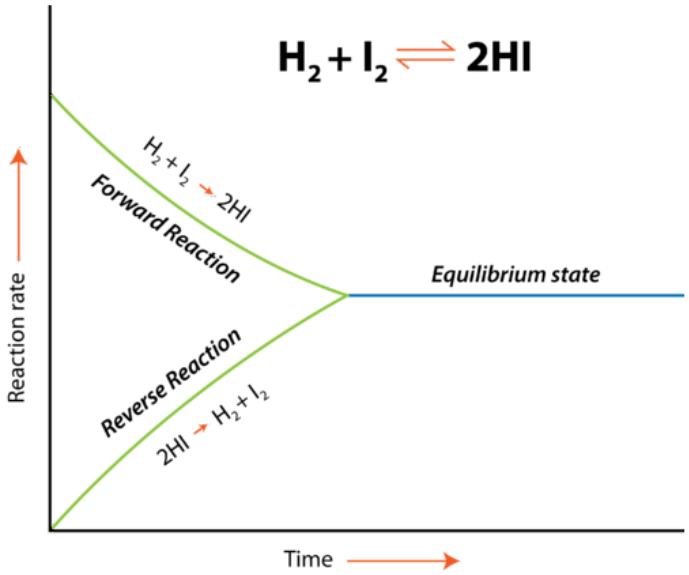


Figure 17.2.1: Equilibrium in reaction: $\mathrm{H_2}\,(g) + \mathrm{I_2}\,(g) \rightleftharpoons 2\mathrm{HI}\,(g)$.

Chemical equilibrium can be attained whether the reaction begins with all reactants and no products, all products and no reactants, or some of both. The figure below shows changes in concentration of $\mathrm{H_2}$, $\mathrm{I_2}$, and HI for two different reactions. In the reaction depicted by the graph on the left (A), the reaction begins with only $\mathrm{H_2}$ and $\mathrm{I_2}$ present. There is no HI initially. As the reaction proceeds toward equilibrium, the concentrations of the $\mathrm{H_2}$ and $\mathrm{I_2}$ gradually decrease, while the concentration of the HI gradually increases. When the curve levels out and the concentrations all become constant, equilibrium has been reached. At equilibrium, concentrations of all substances are constant. In reaction B, the process begins with only HI and no $\mathrm{H_2}$ or $\mathrm{I_2}$. In this case, the concentration of HI gradually decreases while the concentrations of $\mathrm{H_2}$ and $\mathrm{I_2}$ gradually increase until equilibrium is again reached. Notice that in both cases, the relative position of equilibrium is the same, as shown by the relative concentrations of reactants and products. The concentration of HI at equilibrium is significantly higher than the concentrations of $\mathrm{H_2}$ and $\mathrm{I_2}$. This is true whether the reaction began with all reactants or all products. The position of equilibrium is a property of the particular reversible reaction and does not depend upon how equilibrium was achieved.

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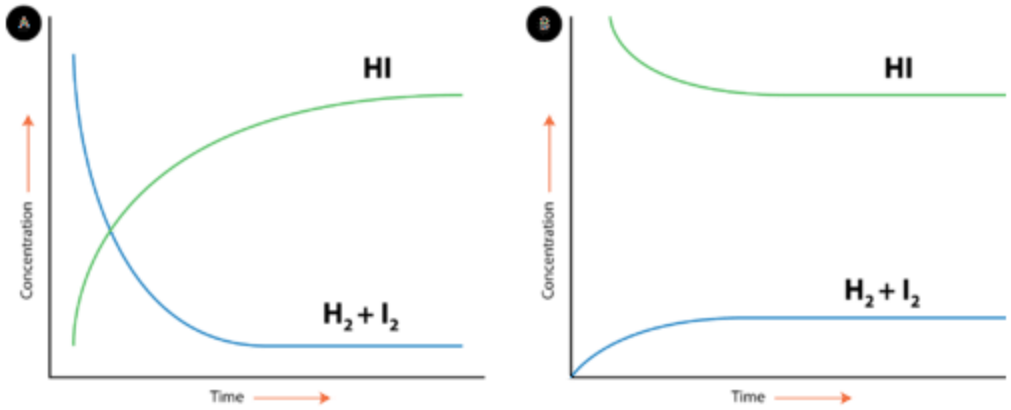


Figure 17.2.2: Equilibrium between reactants and products.

Conditions for Equilibrium

It may be tempting to think that once equilibrium has been reached, the reaction stops. Chemical equilibrium is a dynamic process. The forward and reverse reactions continue to occur even after equilibrium has been reached. However, because the rates of the reactions are the same, there is no change in the relative concentrations of reactants and products for a reaction that is at equilibrium. The conditions and properties of a system at equilibrium are summarized below.

1. The system must be closed, meaning no substances can enter or leave the system.
2. Equilibrium is a dynamic process. Even though we don't necessarily see the reactions, both the forward and reverse reactions are taking place.
3. The rates of the forward and reverse reactions must be equal.
4. The amount of reactants and products do not have to be equal. However, after equilibrium is attained, the amounts of reactants and products will be constant.

The description of equilibrium in this concept refers primarily to equilibrium between reactants and products in a chemical reaction. Other types of equilibrium include phase equilibrium and solution equilibrium. A **phase equilibrium** occurs when a substance is in equilibrium between two states. For example, a stoppered flask of water attains equilibrium when the rate of evaporation is equal to the rate of condensation. A **solution equilibrium** occurs when a solid substance is in a saturated solution. At this point, the rate of dissolution is equal to the rate of recrystallization. Although these are all different types of transformations, most of the rules regarding equilibrium apply to any situation in which a process occurs reversibly.

Summary

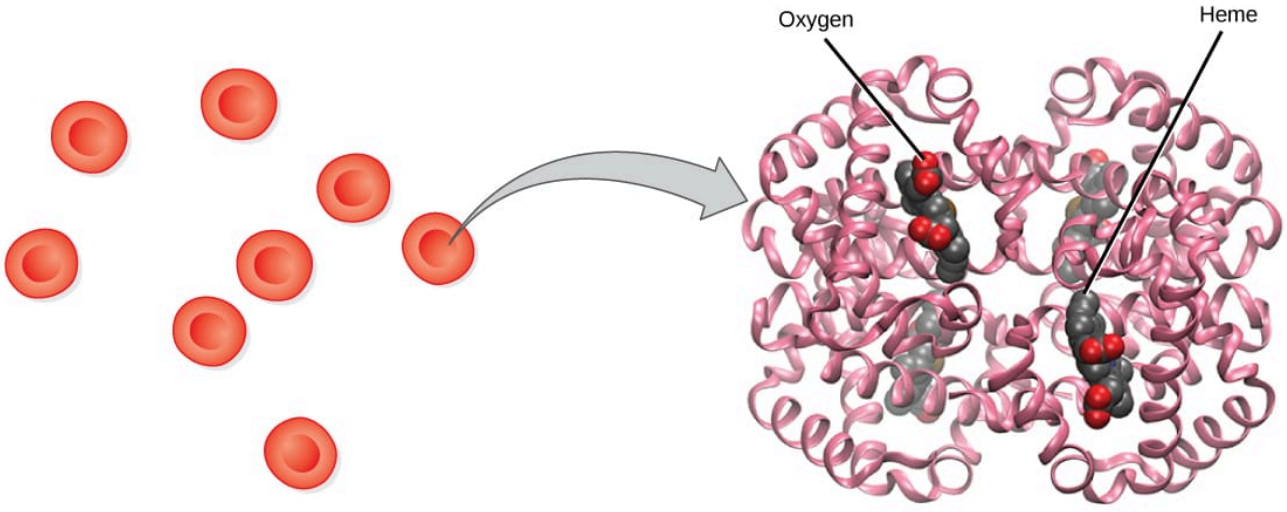
- Chemical equilibrium is the state of a system in which the rate of the forward reaction is equal to the rate of the reverse reaction.
- Conditions for chemical equilibrium are listed.
- A phase equilibrium occurs when a substance is in equilibrium between two states.
- A solution equilibrium occurs when a solid substance is in a saturated solution.

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17.3: Equilibrium Constant

Red blood cells transport oxygen to tissues so that they can function. In the absence of oxygen, cells cannot carry out their biochemical responsibilities. Oxygen moves to the cells attached to hemoglobin, a protein found in red blood cells.



The protein inside (a) red blood cells that carries oxygen to cells and carbon dioxide to the lungs is (b) hemoglobin. Hemoglobin is made up of four symmetrical subunits and four heme groups. Iron associated with the heme binds oxygen. It is the iron in hemoglobin that gives blood its red color. (CC_BY Openstax)

In cases of carbon monoxide poisoning, CO binds much more strongly to the hemoglobin, blocking oxygen attachment and lowering the amount of oxygen reaching the cells. Treatment involves the patient breathing pure oxygen to displace the carbon monoxide. The equilibrium reaction shown below illustrates the shift toward the right when excess oxygen is added to the system:

$$\mathrm{Hb(CO)_4}\,(aq) + 4\mathrm{O_2}\,(g) \rightleftharpoons \mathrm{Hb(O_2)_4}\,(aq) + 4\mathrm{CO}\,(g)$$

Equilibrium Constant

Consider the hypothetical reversible reaction in which reactants A and B react to form products C and D. This equilibrium can be shown as demonstrated below, where the lower case letters represent the coefficients of each substance.

$$a\mathrm{A} + b\mathrm{B} \rightleftharpoons c\mathrm{C} + d\mathrm{D}$$

As we have established, the rates of the forward and reverse reactions are the same at equilibrium, and so the concentrations of all of the substances are constant. Since that is the case, it stands to reason that a ratio of the concentrations for any given reaction at equilibrium maintains a constant value. The **equilibrium constant** (K_{eq}) is the ratio of the mathematical product of the concentrations of the products of a reaction to the mathematical product of the concentrations of the reactants of the reaction. Each concentration is raised to the power of its coefficient in the balanced chemical equation. For the general reaction above, the equilibrium constant expression is written as follows:

$$K_{\mathrm{eq}} = \frac{[\mathrm{C}]^c\,[\mathrm{D}]^d}{[\mathrm{A}]^a\,[\mathrm{B}]^b}$$

The concentrations of each substance, indicated by the square brackets around the formula, are measured in molarity units $(\mathrm{mol/L})$.

The value of the equilibrium constant for any reaction is only determined by experiment. As detailed in the previous section, the

position of equilibrium for a given reaction does not depend on the starting concentrations, and so the value of the equilibrium constant is truly constant. It does, however, depend on the temperature of the reaction. This is because equilibrium is defined as a condition resulting from the rates of forward and reverse reactions being equal. If the temperature changes, the corresponding change in those reaction rates will alter the equilibrium constant. For any reaction in which a K_{eq} is given, the temperature should be specified.

Summary

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17.4: Calculations with Equilibrium Constants

Iron is an important component of red blood cells. Patients who have low iron will usually be anemic and have a lower than normal number of red blood cells. One way to assess serum iron concentration is with the use of Ferrozine, a complex organic molecule. Ferrozine forms a product with Fe^{3+}, producing a pink color. In order to determine factors affecting the reaction, we need to measure the equilibrium constant. If the equilibrium does not lie far in the direction of products, precautions need to be taken when using this material to measure iron in serum.

Calculations with Equilibrium Constants

The general value of the equilibrium constant gives us information about whether the reactants or the products are favored at equilibrium. Since the product concentrations are in the numerator of the equilibrium expression, a $K_{\mathrm{eq}} > 1$ means that the products are favored over the reactants. A $K_{\mathrm{eq}} < 1$ means that the reactants are favored over the products.

Though it would often seem that the K_{eq} value would have various units depending on the values of the exponents in the expression, the general rule is that any units are dropped. All K_{eq} values will be reported as having no units.

Example 17.4.1

Equilibrium occurs when nitrogen monoxide gas reacts with oxygen gas to form nitrogen dioxide gas:

$$2\mathrm{NO}\,(g) + \mathrm{O_2}\,(g) \rightleftharpoons 2\mathrm{NO_2}\,(g)$$

At equilibrium at $230^{\circ}\mathrm{C}$, the concentrations are measured to be $[\mathrm{NO}] = 0.0542\ \mathrm{M}$, $[\mathrm{O_2}] = 0.127\ \mathrm{M}$, and $[\mathrm{NO_2}] = 15.5\ \mathrm{M}$. Calculate the equilibrium constant at this temperature.

Solution:

Step 1: List the known quantities and plan the problem.

Known

- $[\mathrm{NO}] = 0.0542\ \mathrm{M}$
- $[\mathrm{O_2}] = 0.127\ \mathrm{M}$
- $[\mathrm{NO_2}] = 15.5\ \mathrm{M}$

Unknown

The equilibrium expression is first written according to the general form in the text. The equilibrium values are substituted into the expression and the value calculated.

Step 2: Solve.

$$K_{\mathrm{eq}} = \frac{[\mathrm{NO_2}]^2}{[\mathrm{NO}]^2\,[\mathrm{O_2}]}$$

Substituting in the concentrations at equilibrium:

$$K_{\mathrm{eq}} = \frac{(15.5)^2}{(0.0542)^2\,(0.127)} = 6.44 \times 10^5$$

Step 3: Think about your result.

The equilibrium concentration of the product $\mathrm{NO_2}$ is significantly higher than the concentrations of the reactants NO and $\mathrm{O_2}$. As a result, the K_{eq} value is much larger than 1, an indication that the product is favored at equilibrium.

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The equilibrium expression only shows those substances whose concentrations are variable during the reaction. A pure solid or a pure liquid does not have a concentration that will vary during a reaction. Therefore, an equilibrium expression omits pure solids and liquids, and only shows the concentrations of gases and aqueous solutions. The decomposition of mercury (II) oxide can be shown by the following equation, followed by its equilibrium expression.

$$2\text{HgO}\,(s) \rightleftharpoons 2\text{Hg}\,(l) + \text{O}_2\,(g) \qquad K_{\text{eq}} = [\text{O}_2]$$

The stoichiometry of an equation can also be used in a calculation of an equilibrium constant. At 40°C, solid ammonium carbamate decomposes to ammonia and carbon dioxide gases.

$$\text{NH}_4\text{CO}_2\text{NH}_2\,(s) \rightleftharpoons 2\text{NH}_3\,(g) + \text{CO}_2\,(g)$$

At equilibrium, the $[\text{CO}_2]$ is found to be $4.71 \times 10^{-3}\ \text{M}$. Can the K_{eq} value be calculated from that information alone? Because the ammonium carbamate is a solid, it is not present in the equilibrium expression.

$$K_{\text{eq}} = [\text{NH}_3]^2\,[\text{CO}_2]$$

The stoichiometry of the chemical equation indicates that as the ammonium carbamate decomposes, $2\ \text{mol}$ of ammonia gas is produced for every $1\ \text{mol}$ of carbon dioxide. Therefore, at equilibrium, the concentration of the ammonia will be twice the concentration of carbon dioxide. So,

$$[\text{NH}_3] = 2 \times \left(4.71 \times 10^{-3}\right) = 9.42 \times 10^{-3}\ \text{M}$$

Substituting these values into the K_{eq} expression:

$$K_{\text{eq}} = \left(9.42 \times 10^{-3}\right)^2 \left(4.71 \times 10^{-3}\right) = 4.18 \times 10^{-7}$$

Using Equilibrium Constants

The equilibrium constants are known for a great many reactions. Hydrogen and bromine gases combine to form hydrogen bromide gas. At 730°C, the equation and K_{eq} are given below.

$$\text{H}_2\,(g) + \text{Br}_2\,(g) \rightleftharpoons 2\text{HBr}\,(g) \qquad K_{\text{eq}} = 2.18 \times 10^6$$

A certain reaction is begun with only HBr. When the reaction mixture reaches equilibrium at 730°C, the concentration of bromine gas is measured to be $0.00243\ \text{M}$. What is the concentration of the H_2 and the HBr at equilibrium?

Since the reaction begins with only HBr and the mole ratio of H_2 to Br_2 is 1:1, the concentration of H_2 at equilibrium is also $0.00243\ \text{M}$. The equilibrium expression can be rearranged to solve for the concentration of HBr at equilibrium:

$$\begin{aligned} K_{\text{eq}} &= \frac{[\text{HBr}]^2}{[\text{H}_2]\,[\text{Br}_2]} \\ [\text{HBr}] &= \sqrt{K_{\text{eq}}\,[\text{H}_2]\,[\text{Br}_2]} \\ &= \sqrt{2.18 \times 10^6\,(0.00243)\,(0.00243)} = 3.59\ \text{M} \end{aligned}$$

Since the value of the equilibrium constant is very high, the concentration of HBr is much greater than that of H_2 and Br_2 at equilibrium.

Summary

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17.5: Le Châtelier's Principle

There are some who enjoy going up in an airplane, strapping on a parachute, and diving out the door to free-fall before opening the chute, and dropping to the ground. This stressful activity (so they say) relieves the stress of everyday life. The release of adrenaline caused by this stressful behavior is said to promote a mood enhancement that helps you deal better with other stresses in your daily life.

Le Chatelier's Principle

Chemical equilibrium was studied by French chemist Henri Le Chatelier (1850-1936) and his description of how a system responds to a stress on equilibrium has become known as **Le Chatelier's principle:** When a chemical system that is at equilibrium is disturbed by a stress, the system will respond in order to relieve the stress. Stresses to a chemical system involve changes in the concentration of reactants or products, changes in the temperature of the system, or changes in the pressure of the system. We will discuss each of these stresses separately. The change to the equilibrium position in every case is either a favoring of the forward reaction or a favoring of the reverse reaction. When the forward reaction is favored, the concentrations of products increase, while the concentrations of reactants decrease. When the reverse reaction is favored, the concentrations of products decrease, while the concentrations of reactants increase.

Le Chatelier's Principle

Table 17.5.1

Original Equilibrium	Favored Reaction	Result
$\mathrm{A} \rightleftharpoons \mathrm{B}$	Forward: $\mathrm{A} \rightarrow \mathrm{B}$	$[\mathrm{A}]$ decreases; $[\mathrm{B}]$ increases
$\mathrm{A} \rightleftharpoons \mathrm{B}$	Reverse: $\mathrm{A} \leftarrow \mathrm{B}$	$[\mathrm{A}]$ increases; $[\mathrm{B}]$ decreases



Figure 17.5.1: Henri Le Chatelier.

Summary

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17.6: Effect of Concentration

Phenolphthalein is a chemical that has one structure in a high acid environment, and another structure in a low acid environment. If the hydrogen ion concentration is high, the compound is colorless, but turns red if the hydrogen ion concentration is low. By adding hydrogen ions to the solution, or removing them through a chemical reaction, we can vary the color of the dye.

Effect of Concentration

A change in concentration of one of the substances in an equilibrium system typically involves either the addition or the removal of one of the reactants or products. Consider the Haber-Bosch process for the industrial production of ammonia from nitrogen and hydrogen gases:

$$\mathrm{N_2}\,(g) + 3\mathrm{H_2}\,(g) \rightleftharpoons 2\mathrm{NH_3}\,(g)$$

If the concentration of one substance in a system is increased, the system will respond by favoring the reaction that removes that substance. When more $\mathrm{N_2}$ is added, the forward reaction will be favored because the forward reaction uses up $\mathrm{N_2}$ and converts it to $\mathrm{NH_3}$. The forward reaction speeds up temporarily as a result of the addition of a reactant. The position of equilibrium shifts as more $\mathrm{NH_3}$ is produced. The concentration of $\mathrm{NH_3}$ increases, while the concentrations of $\mathrm{N_2}$ and $\mathrm{H_2}$ decrease. After some time passes, equilibrium is reestablished with new concentrations of all three substances. As can be seen in the figure below, if more $\mathrm{N_2}$ is added, a new equilibrium is achieved by the system. The new concentration of $\mathrm{NH_3}$ is higher because of the favoring of the forward reaction. The new concentration of the $\mathrm{H_2}$ is lower. The concentration of $\mathrm{N_2}$ is higher than in the original equilibrium, but went down slightly following the addition of the $\mathrm{N_2}$ that disturbed the original equilibrium. By responding in this way, the value of the equilibrium constant for the reaction, K_{eq}, does not change as a result of the stress to the system.

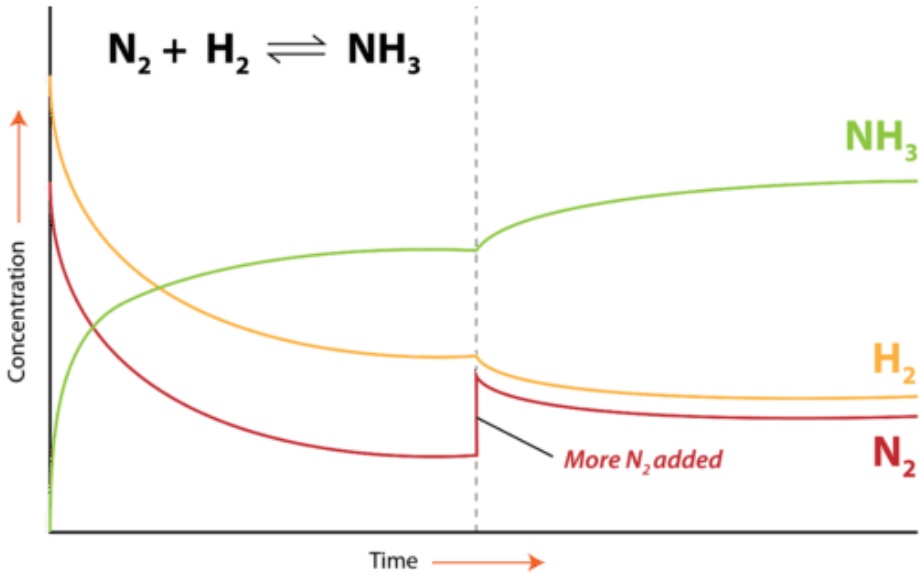


Figure 17.6.1: The Haber-Bosch process is an equilibrium between reactant $\mathrm{N_2}$ and $\mathrm{H_2}$ and product $\mathrm{NH_3}$.

If more $\mathrm{NH_3}$ were added, the reverse reaction would be favored. This "favoring" of a reaction means temporarily speeding up the reaction in that direction until equilibrium is reestablished. Recall that once equilibrium is reestablished, the rates of the forward and reverse reactions are again equal. The addition of $\mathrm{NH_3}$ would result in increased formation of the reactants, $\mathrm{N_2}$ and $\mathrm{H_2}$.

An equilibrium can also be disrupted by the removal of one of the substances. If the concentration of a substance is decreased, the system will respond by favoring the reaction that replaces that substance. In the industrial Haber-Bosch process, $\mathrm{NH_3}$ is removed from the equilibrium system as the reaction proceeds. As a result, the forward reaction is favored so that more $\mathrm{NH_3}$ will be produced. The concentrations of $\mathrm{N_2}$ and $\mathrm{H_2}$ decrease. Continued removal of $\mathrm{NH_3}$ will eventually force the reaction to go to completion until all of the reactants are used up. If either $\mathrm{N_2}$ or $\mathrm{H_2}$ were removed from the equilibrium system, the reverse reaction would be favored, and the concentration of $\mathrm{NH_3}$ would decrease.

The effects of changes in concentration on an equilibrium system, according to Le Chatelier's principle, are summarized in the table below.

Effects of changes in concentration on an equilibrium system, according to Le Chatelier's principle

Table 17.6.1	
Stress	**Response**
Addition of reactant	Forward reaction favored.
Addition of product	Reverse reaction favored.
Removal of reactant	Reverse reaction favored.
Removal of product	Forward reaction favored.

Summary

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17.7: Effect of Temperature

Carbon monoxide is often thought of as nothing more than a hazardous gas produced from incomplete combustion of carbon products. However, there is a large market for industrially-manufactured carbon monoxide that is used to synthesize most of the acetic acid produced in the world. One reaction that leads to CO formation involves its formation by passing air over excess carbon at high temperatures. The initial product (carbon dioxide) equilibrates with the remaining hot carbon, forming carbon monoxide. At lower temperatures, $\mathrm{CO_2}$ formation is favored, while CO is the predominant product above $800^{\circ}\mathrm{C}$.

Effect of Temperature

Increasing or decreasing the temperature of a system at equilibrium is also a stress to the system. The equation for the Haber-Bosch process is written again below, as a thermochemical equation.

$$\mathrm{N_2}\,(g) + 3\mathrm{H_2}\,(g) \rightleftharpoons 2\mathrm{NH_3}\,(g) + 91\ \mathrm{kJ}$$

The forward reaction is the exothermic direction: the formation of $\mathrm{NH_3}$ releases heat. The reverse reaction is the endothermic direction: as $\mathrm{NH_3}$ decomposes to $\mathrm{N_2}$ and $\mathrm{H_2}$, heat is absorbed. An increase in the temperature of a system favors the direction of the reaction that absorbs heat, the endothermic direction. Absorption of heat in this case is a relief of the stress provided by the temperature increase. For the Haber-Bosch process, an increase in temperature favors the reverse reaction. The concentration of $\mathrm{NH_3}$ in the system decreases, while the concentrations of $\mathrm{N_2}$ and $\mathrm{H_2}$ increase.

A decrease in the temperature of a system favors the direction of the reaction that releases heat, the exothermic direction. For the Haber-Bosch process, a decrease in temperature favors the forward reaction. The concentration of $\mathrm{NH_3}$ in the system increases, while the concentrations of $\mathrm{N_2}$ and $\mathrm{H_2}$ decrease.

For changes in concentration, the system responds in such a way that the value of the equilibrium constant, K_{eq}, is unchanged. However, a change in temperature shifts the equilibrium, and the K_{eq} value either increases or decreases. As discussed in the previous section, values of K_{eq} are dependent on the temperature. When the temperature of the system for the Haber-Bosch process is increased, the resultant shift in equilibrium toward the reactants means that the K_{eq} value decreases. When the temperature is decreased, the shift in equilibrium toward the products means that the K_{eq} value increases.

Le Chatelier's principle, as related to temperature changes, can be illustrated by the reaction in which dinitrogen tetroxide is in equilibrium with nitrogen dioxide:

$$\mathrm{N_2O_4}\,(g) + \mathrm{heat} \rightleftharpoons 2\mathrm{NO_2}\,(g)$$

Dinitrogen tetroxide $(\mathrm{N_2O_4})$ is colorless, while nitrogen dioxide $(\mathrm{NO_2})$ is dark brown in color. When $\mathrm{N_2O_4}$ breaks down into $\mathrm{NO_2}$, heat is absorbed according to the forward reaction above. Therefore, an increase in temperature of the system will favor the forward reaction. Conversely, a decrease in temperature will favor the reverse reaction.

Summary

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17.8: Effect of Pressure

The ammonia storage tank in the picture below does two things. First, it stores ammonia at high pressure to minimize the reverse reaction that would lead to less ammonia, and more nitrogen and hydrogen. Secondly, it sends an important message. Ammonia is used to make methamphetamine, a dangerous drug of abuse. Locks and other safety mechanisms built into the tanks help stop the theft of ammonia to be used for this illicit activity.



Figure 17.8.1: An ammonia storage tank.

Effect of Pressure

Changing the pressure of an equilibrium system in which gases are involved is also a stress to the system. A change in the pressure on a liquid or a solid has a negligible effect. We will return again to the equilibrium for the Haber-Bosch process. Imagine the gases are contained in a closed system in which the volume of the system is controlled by an adjustable piston, as shown in the figure below.

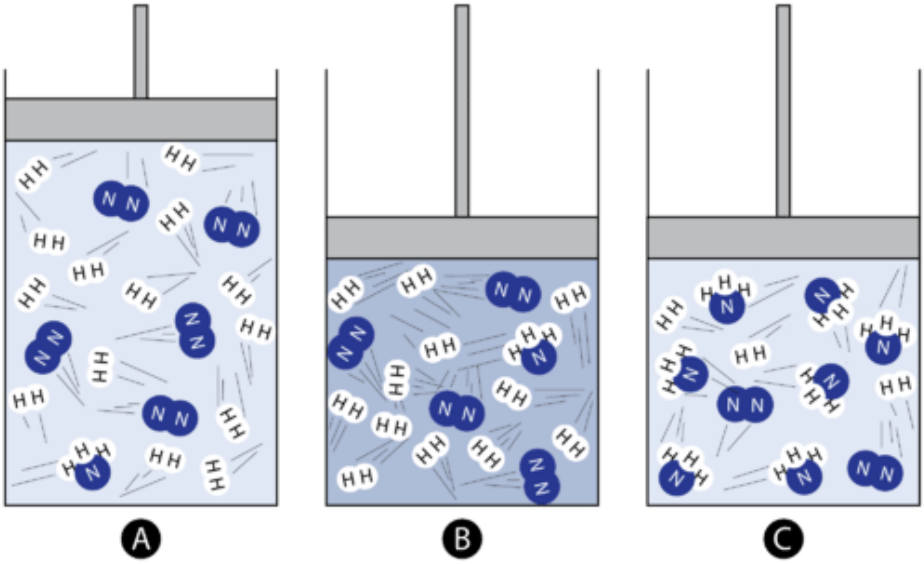


Figure 17.8.2: Effect of pressure on ammonia formation.

On the far left, the reaction system contains primarily $\mathrm{N_2}$ and $\mathrm{H_2}$, with only one molecule of $\mathrm{NH_3}$ present. As the piston is pushed inwards, the pressure of the system increases according to Boyle's Law. This is a stress to the equilibrium. In the middle image, the same number of molecules is now confined in a smaller space and so the pressure has increased. According to Le Chatelier's principle, the system responds in order to relieve the stress. In the image on the right, the forward reaction has been favored, in which one molecule of $\mathrm{N_2}$ combines with three molecules of $\mathrm{H_2}$ to form two molecules of $\mathrm{NH_3}$. The overall result is a decrease in the number of gas molecules in the entire system. This in turn decreases the pressure and provides a relief to the original stress of a pressure increase. An increase in pressure on an equilibrium system favors the reaction which produces fewer total moles of gas. In this case, it is the forward reaction that is favored.

A decrease in pressure on the above system could be achieved by pulling the piston outward, increasing the container volume. The equilibrium would respond by favoring the reverse reaction in which $\mathrm{NH_3}$ decomposes to $\mathrm{N_2}$ and $\mathrm{H_2}$. This is because the overall number of gas molecules would increase, and so would the pressure. A decrease in pressure on an equilibrium system favors the reaction that produces more total moles of gas. This is summarized in the table below.

Table 17.8.1

Stress	Response
Pressure increase	Reaction produces fewer gas molecules.
Pressure decrease	Reaction produces more gas molecules.

Like changes in concentration, the K_{eq} value for a given reaction is unchanged by a change in pressure.

It is important to remember when analyzing the effect of a pressure change on equilibrium that only gases are affected. If a certain reaction involves liquids or solids, they should be ignored. Calcium carbonate decomposes according to the equilibrium reaction:

$$\mathrm{CaCO_3}\,(s) \rightleftharpoons \mathrm{CaO}\,(s) + \mathrm{O_2}\,(g)$$

Oxygen is the only gas in the system. An increase in the pressure of the system slows the rate of decomposition of $\mathrm{CaCO_3}$ because the reverse reaction is favored. When a system contains equal moles of gas on both sides of the equation, pressure has no effect on the equilibrium position, as in the formation of HCl from $\mathrm{H_2}$ and $\mathrm{Cl_2}$.

$$\mathrm{H_2}\,(g) + \mathrm{Cl_2}\,(g) \rightleftharpoons 2\mathrm{HCl}\,(g)$$

Summary

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17.9: Nonreversible Reactions

Fires are a part of life. Some fires clear the land and allow new growth. Other fires provide warmth on a cold night. Unfortunately, many fires are destructive, leaving damage in their wake. All fires leave the environment changed, never to revert back to its original state. The carbon dioxide and water generated by a fire go off into the atmosphere and do not return. The change is permanent and irreversible.

Going to Completion

When one of the products of a reaction is removed from the chemical equilibrium system as soon as it is produced, the reverse reaction cannot establish itself and equilibrium is never reached. Reactions such as these are said to go to completion. These processes are often referred to as **non-reversible reactions**. Reactions which go to completion tend to produce one of three types of products: (1) an insoluble precipitate, (2) a gas, (3) a molecular compound such as water. Examples of these reactions are shown below.

1. Formation of a precipitate:

$$\mathrm{AgNO_3}\,(aq) + \mathrm{NaCl}\,(aq) \rightarrow \mathrm{NaNO_3}\,(aq) + \mathrm{AgCl}\,(s)$$

2. Formation of a gas:

$$\mathrm{Mg}\,(s) + 2\mathrm{HCl}\,(aq) \rightarrow \mathrm{MgCl_2}\,(aq) + \mathrm{H_2}\,(g)$$

3. Formation of water:

$$\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$$

If we look at these reactions in more detail, we can see some things that are not apparent the way the equations are written. Looking at the first equation, we do not see a double arrow between reactants and products, because the reaction is considered essentially irreversible. However, if we consider the net ionic equation

$$\mathrm{Ag^+} + \mathrm{Cl^-} \rightarrow \mathrm{AgCl}$$

then the reverse reaction would be:

$$\mathrm{AgCl} \rightarrow \mathrm{Ag^+} + \mathrm{Cl^-}$$

The K_{eq} for the reverse reaction is 1.8×10^{-10}. For all practical purposes, the reaction goes to completion.

Formation of a gas in an open system is essentially irreversible since the gas escapes into the atmosphere. Looking at the activity series, we see that Mg is much higher in the series than hydrogen. So the reaction would be expected to go strongly in the indicated direction.

The third reaction gets a little more complicated. In solution, the reactants HCl and NaOH will be ionized completely, as does the NaCl product. Water exists in an equilibrium with $\mathrm{H^+}$ and $\mathrm{OH^-}$, with the dissociation constant for water being 1×10^{-14}. So, in the solution resulting from the reaction given here, the $[\mathrm{H^+}]$ is $1 \times 10^{-7}\ \mathrm{M}$, a very insignificant amount. For all practical purposes, this reaction can be said to go to completion.

Summary

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17.10: Le Châtelier's Principle and the Equilibrium Constant

With online banking, management of personal finances can become less complicated in some ways. You can automatically deposit paychecks, pay bills, and designate how much goes into savings or other special accounts each month. If you want to maintain 10% of your bank account in savings, you can set up a program that moves money in and out of the account when you get a paycheck or pay bills. The amount of money in savings will change as the money comes in and out of the bank, but the ratio of savings to checking will always be constant.

Le Chatelier's Principle and the Equilibrium Constant

Occasionally, when students apply Le Chatelier's principle to an equilibrium problem involving a change in concentration, they assume that K_{eq} must change. This seems logical since we talk about "shifting" the equilibrium in one direction or the other. However, K_{eq} is a constant, for a given equilibrium at a given temperature, so it must not change. Here is an example of how this works. Consider the simplified equilibrium below:

$$\text{A} \rightleftharpoons \text{B}$$

Let's say we have a 1.0 liter container. At equilibrium, the following amounts are measured:

$$\begin{aligned} \text{A} &= 0.50\ \text{mol} \\ \text{B} &= 1.0\ \text{mol} \end{aligned}$$

The value of K_{eq} is given by:

$$K_{\text{eq}} = \frac{[\text{B}]}{[\text{A}]} = \frac{1.0\ \text{M}}{0.50\ \text{M}} = 2.0$$

Now we will disturb the equilibrium by adding 0.50 moles of A to the mixture. The equilibrium will shift towards the right, forming more B. Immediately after the addition of A and before any response, we how have $1.0\ \text{mol}$ of A and $1.0\ \text{mol}$ of B. The equilibrium then shifts in the forward direction. We will introduce a variable (x), which will represent the change in concentrations as the reaction proceeds. Since the mole ratio of A:B is 1:1, as $[\text{A}]$ decreases by the amount x, the $[\text{B}]$ increases by the amount x. We set up an analysis called *ICE*, which stands for Initial, Change, and Equilibrium. The values in the table below represent molar concentrations.

	A	B
Initial	1.0	1.0
Change	$-x$	$+x$
Equilibrium	$1.0 - x$	$1.0 + x$

At the new equilibrium position, the values for A and B as a function of x can be set equal to the value of the K_{eq}. Then, one can solve for x.

$$K_{\text{eq}} = 2.0 = \frac{[\text{B}]}{[\text{A}]} = \frac{1.0 + x}{1.0 - x}$$

Solving for x:

$$\begin{aligned} 2.0\,(1.0 - x) &= 1.0 + x \\ 2.0 - 2.0x &= 1.0 + x \\ 3.0x &= 1.0 \\ x &= 0.33 \end{aligned}$$

This value for x is now plugged back in to the Equilibrium line of the table, and used to find the final concentrations of A and B after the reaction.

$$[\text{A}] = 1.0 - x = 0.67\ \text{M}$$

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$$[\mathrm{B}] = 1.0 + x = 1.33\ \mathrm{M}$$

The value of K_{eq} has been maintained since $\frac{1.33}{0.67} = 2.0$. This shows that even though a change in concentration of one of the substances in equilibrium causes a shift in the equilibrium position, the value of the equilibrium constant does not change.

Summary

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18: ACIDS AND BASES

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18.1: Properties of Acids

A cup of coffee first thing in the morning can help start the day. However, keeping the coffee maker clean can be a problem. Lime deposits build up after a while and slow down the brewing process. The best cure for this is to put vinegar (dilute acetic acid) in the pot and run it through the brewing cycle. The vinegar dissolves the deposits and cleans the maker, which will speed up the brewing process back to its original rate. Just be sure to run water through the brewing process after the vinegar, or you will get some really horrible coffee!

Acids

Acids are very common in some of the foods that we eat. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid, which is better known as vitamin C. Carbonated sodas contain phosphoric acid. Vinegar contains acetic acid. Your own stomach utilizes hydrochloric acid to digest food.

Acids are a distinct class of compounds because of the properties of their aqueous solutions. These properties are:

1. Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current. Some acids are strong electrolytes because they ionize completely in water, yielding a great many ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.
2. Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.
3. Acids change the color of certain acid-base indicates. Two common indicators are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein turns colorless.
4. Acids react with active metals to yield hydrogen gas. Recall that an activity series is a list of metals in descending order of reactivity. Metals that are above hydrogen in the activity series will replace the hydrogen from an acid in a single-replacement reaction, as shown below:

$$\mathrm{Zn}\,(s) + \mathrm{H_2SO_4}\,(aq) \rightarrow \mathrm{ZnSO_4}\,(aq) + \mathrm{H_2}\,(g)$$

5. Acids react with bases to produce a salt compound and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. The products of this reaction are an ionic compound, which is labeled as a salt, and water.

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18.2: Properties of Bases

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The base in the antacid will react with the HCl in the stomach and neutralize it, taking care of that unpleasant feeling.

Bases

Bases have properties that mostly contrast with those of acids.

1. Aqueous solutions of bases are also electrolytes. Bases can be either strong or weak, just as acids can.
2. Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
3. Bases also change the color of indicators. Litmus turns blue in the presence of a base while phenolphthalein turns pink.
4. Bases do not react with metals in the way that acids do.
5. Bases react with acids to produce a salt and water.



Figure $18.2.1$***:*** *Phenolphthalein indicator in presence of base.*

Please note that tasting chemicals and touching them are NOT good lab practices and should be avoided—in other words, don't do this at home.

Bases are less common as foods, but they are nonetheless present in many household products. Many cleaners contain ammonia, a base. Sodium hydroxide is found in drain cleaner. **Antacids**, which combat excess stomach acid, are comprised of bases such as magnesium hydroxide or sodium hydrogen carbonate.

Summary

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18.3: Arrhenius Acids

Venus is the planet nearest to Earth, but has a very different and hostile environment. It has a surface temperature that averages around $450^{\circ}\mathrm{C}$. The atmosphere is composed of carbon dioxide, but clouds of sulfuric acid move through the upper atmosphere, helping to create the extremely unfriendly conditions. Because of these conditions, Venus is not a place you want to visit on vacation.

Arrhenius Acids

Swedish chemist Svante Arrhenius (1859-1927) was the first to propose a theory to explain the observed behavior of acids and bases. Because of their ability to conduct a current, he knew that both acids and bases contained ions in solution. An **Arrhenius acid** is a compound that ionizes to yield hydrogen ions (H^{+}) in aqueous solution.

Acids are molecular compounds with ionizable hydrogen atoms. Only hydrogen atoms that are part of a highly polar covalent bond are ionizable. Hydrogen chloride (HCl) is a gas at room temperature and under normal pressure. The $\mathrm{H}-\mathrm{Cl}$ bond in hydrogen chloride is a polar bond. The hydrogen atom is electron deficient because of the higher electronegativity of the chlorine atom. Consequently, the hydrogen atom is attracted to the lone pair of electrons in a water molecule when HCl is dissolved in water. The result is that the $\mathrm{H}-\mathrm{Cl}$ bond breaks, with both bonding electrons remaining with the Cl, forming a chloride ion. The H^{+} ion attaches to the water molecule, forming a polyatomic ion called the hydronium ion. The **hydronium ion** $(\mathrm{H_3O^{+}})$ can be thought of as a water molecule with an extra attached hydrogen ion.

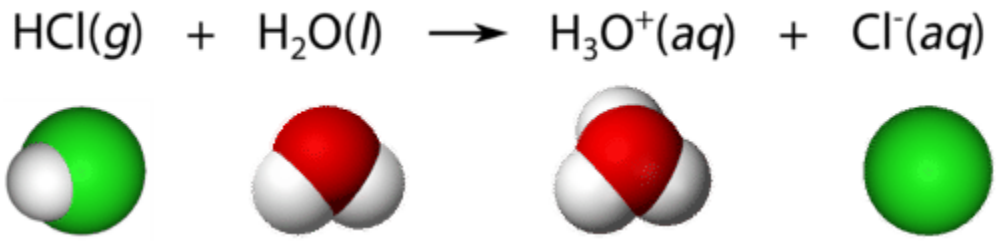


Figure $18.3.1$: Formation of a hydronium ion. (CC BY-NC; CK-12)

Equations showing the ionization of an acid in water are frequently simplified by omitting the water molecule:

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H}^{+}\,(aq) + \mathrm{Cl}^{-}\,(aq)$$

This is merely a simplification of the previous equation, but it is commonly used. Any hydrogen ions in an aqueous solution will be attached to water molecules as hydronium ions.

Not all hydrogen atoms in molecular compounds are ionizable. In methane $(\mathrm{CH_4})$, the hydrogen atoms are covalently bonded to carbon in bonds that are only slightly polar. The hydrogen atoms are not capable of ionizing, and methane has no acidic properties. Acetic acid $(\mathrm{CH_3COOH})$ belongs to a class of acids called organic acids. There are four hydrogen atoms in the molecule, but only the one hydrogen that is bonded to an oxygen atom is ionizable.

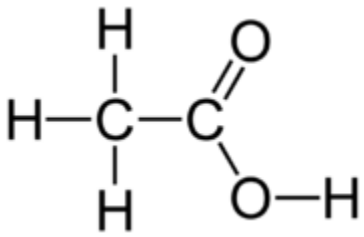


Figure $18.3.2$: The $\mathrm{O}-\mathrm{H}$ bond can be ionized to yield the H^{+} ion and the acetate ion. The other hydrogen atoms in this molecule are not acidic. (CC BY-NC; CK-12)

The table below lists some of the more common acids:

Common Acids

Table 18.3.1: Common Acids

Acid Name	Formula
Hydrochloric acid	HCl
Nitric acid	HNO_3
Sulfuric acid	H_2SO_4
Phosphoric acid	H_3PO_4
Acetic acid	CH_3COOH
Hypochlorous acid	$HClO$

A **monoprotic acid** is an acid that contains only one ionizable hydrogen. Hydrochloric acid and acetic acid are monoprotic acids. A **polyprotic acid** is an acid that contains multiple ionizable hydrogens. Most common polyprotic acids are either diprotic (such as H_2SO_4) or triprotic (such as H_3PO_4).

Summary

- An Arrhenius acid is a compound that ionizes to yield hydrogen ions (H^+) in aqueous solution.
- Examples of Arrhenius acids are given.

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18.4: Arrhenius Bases

Sodium hydroxide is a versatile chemical. It can be used for such mundane purposes as cleaning clogged drains. Several commercial preparations contain sodium hydroxide for this purpose. It also has a number of applications in the food processing field. Ice cream is thickened using NaOH. If olives are soaked in a solution containing sodium hydroxide and other chemicals, the olives will turn black. Soft pretzels are made chewy by the application of sodium hydroxide to the food. This compound has been widely used in the synthesis of plastics, for etching aluminum, for paint removal, and is employed in the dehorning of cattle.

Arrhenius Bases

An **Arrhenius base** is a compound that ionizes to yield hydroxide ions (OH^-) in aqueous solution. The table below lists several of the more common bases.

Table $18.4.1$*: Common Bases*

Base Name	Formula
Sodium hydroxide	NaOH
Potassium hydroxide	KOH
Magnesium hydroxide	$\mathrm{Mg(OH)_2}$
Calcium hydroxide	$\mathrm{Ca(OH)_2}$

All of the bases listed in the table are solids at room temperature. Upon dissolving in water, each dissociates into a metal cation and the hydroxide ion.

$$\mathrm{NaOH}\,(s) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

Sodium hydroxide is a very caustic substance also known as lye. Lye is used as a rigorous cleaner and is an ingredient in the manufacture of soaps. Care must be taken with strong bases like sodium hydroxide, as exposure can lead to severe burns (see figure below).



Figure $18.4.1$: This foot has severe burns due to prolonged contact with a solution of sodium hydroxide, also known as lye. (CC BY-NC; CK-12)

Sodium belongs to the group of elements called the alkali metals. An **alkaline solution** is another name for a solution that is basic. All alkali metals react readily with water to produce the metal hydroxide and hydrogen gas. The resulting solutions are basic.

$$2\mathrm{K}\,(s) + 2\mathrm{H_2O}\,(l) \rightarrow 2\mathrm{KOH}\,(aq) + \mathrm{H_2}\,(g)$$

Bases that consist of an alkali metal cation and the hydroxide anion are all very soluble in water. Compounds of the Group 2 metals (the alkaline earth metals) are also basic. However, these compounds are generally not as soluble in water. Therefore, the dissociation reactions for these compounds are shown as equilibrium reactions:

$$\mathrm{Mg(OH)_2}\,(s) \overset{\mathrm{H_2O}}{\rightleftharpoons} \mathrm{Mg^{2+}}\,(aq) + 2\mathrm{OH^-}\,(aq)$$

These relatively insoluble hydroxides were some of the compounds discussed in the context of the solubility product constant (K_{sp}). The solubility of magnesium hydroxide is $0.0084\ \mathrm{g}$ per liter of water at $25^{\circ}\mathrm{C}$. Because of its low solubility, magnesium hydroxide is not as dangerous as sodium hydroxide. In fact, magnesium hydroxide is the active ingredient in a product called milk of magnesia, which is used as an antacid or a mild laxative.

Summary

- Arrhenius base is a compound that ionizes to yield hydroxide ions $(\mathrm{OH^-})$ in aqueous solution.
- Examples of Arrhenius bases are given.

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18.5: Brønsted-Lowry Acids and Bases

The Arrhenius concept of acids and bases was a significant contribution to the scientific understanding of acids and bases. It replaced and expanded Lavoisier's original idea that all acids contained oxygen. However, the Arrhenius theory also had its shortcomings. It did not take into account the role of the solvent. In addition, this concept did not deal with acid-base behavior in solvents such as benzene, where there could be no ionization. So, formation of a new theory was imperative, which built on the findings of Arrhenius, but also went beyond them.

Brønsted-Lowry Acids and Bases

The Arrhenius definition of acids and bases is somewhat limited. There are some compounds whose properties suggest that they are either acidic or basic, but which do not qualify according to the Arrhenius definition. An example is ammonia $(\mathrm{NH_3})$. Its aqueous solution turns litmus blue, it reacts with acids, and displays all the other properties of a base. However, it does not contain the hydroxide ion. In 1923, a broader definition of acids and bases was independently proposed by Danish chemist Johannes Brønsted (1879-1947) and English chemist Thomas Lowry (1874-1936). A **Brønsted-Lowry acid** is a molecule or ion that donates a hydrogen ion in a reaction. A **Brønsted-Lowry base** is a molecule or ion that accepts a hydrogen ion in a reaction. A hydrogen ion is commonly referred to as a proton, and so acids and bases are proton donors and proton acceptors, respectively, according to the Brønsted-Lowry definition. All substances that are categorized as acids and bases under the Arrhenius definition are also defined as such under the Brønsted-Lowry definition. The new definition, however, includes some substances that are left out according to the Arrhenius definition.

What kind of molecule would qualify as a Brønsted-Lowry base? These molecules need to be able to accept a hydrogen ion (or proton). Two possibilities come to mind: an anion that can form a neutral compound with a proton, or a molecule in which one or more atoms has lone-pair electrons. The most obvious anion is the Arrhenius base $\mathrm{OH^-}$. This ion can form a water molecule with a proton by accepting the proton. The acetate anion $\mathrm{CH_3COO^-}$ is another anion that can combine with a proton to form acetic acid. Lone-pair electron groups would include the nitrogen atom (see figure below). The two electrons at the top of the nitrogen atom are not connected in any type of bond, but they do interact readily with a bare proton.

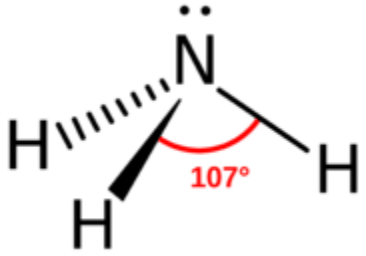


Figure $18.5.1$: *Lone pair electrons on nitrogen.* (CC BY-NC; CK-12)

Oxygen is another atom with lone pair electrons that can function as a Brønsted-Lowry base.



Figure $18.5.2$: Electrons in an oxygen atom. (CC BY-NC; CK-12)

The two single electrons (to the left and bottom of the atom) can form single covalent bonds with other atoms, while the two pairs of double electrons (top and right) are available to interact with a hydrogen ion.

Summary

- A Brønsted-Lowry acid is a molecule or ion that donates a hydrogen ion in a reaction.
- A Brønsted-Lowry base is a molecule or ion that accepts a hydrogen ion in a reaction.

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18.6: Brønsted-Lowry Acid-Base Reactions

The Roman god Janus was considered the god of gates, doors, beginnings, and endings. He is portrayed with two faces, looking in two directions at once. Janus would have been in a good position to look at the acid-base reactions we see in this concept, since they are equilibrium reactions involving two different forms of both acids and bases.

Brønsted-Lowry Acid-Base Reactions

An acid-base reaction, according to the Brønsted-Lowry definition, is a transfer of a proton from one molecule or ion to another. When ammonia is dissolved in water, it undergoes the following reversible reaction.

$$\underset{\text{base}}{\mathrm{NH_3}\,(aq)} + \underset{\text{acid}}{\mathrm{H_2O}\,(l)} \rightleftharpoons \underset{\text{acid}}{\mathrm{NH_4^+}\,(aq)} + \underset{\text{base}}{\mathrm{OH^-}\,(aq)}$$

In this reaction, the water molecule is donating a proton to the ammonia molecule. The resulting products are the ammonium ion and the hydroxide ion. The water is acting as a Brønsted-Lowry acid, while the ammonia is acting as a Brønsted-Lowry base. The hydroxide ion that is produced causes the solution to be basic.

We can also consider the reverse reaction in the above equation. In that reaction, the ammonia ion donates a proton to the hydroxide ion. The ammonium ion is a Brønsted-Lowry acid, while the hydroxide ion is a Brønsted-Lowry base. Most Brønsted-Lowry acid-base reactions can be analyzed in this way. One acid and one base are reactants, and one acid and one base are products.

In the above reaction, water acted as an acid, which may seem a bit unexpected. Water can also act as a base in a Brønsted-Lowry acid-base reaction, as long as it reacts with a substance that is a better proton donor. Shown below is the reaction of water with the hydrogen sulfate ion.

$$\underset{\text{acid}}{\mathrm{HSO_4^-}\,(aq)} + \underset{\text{base}}{\mathrm{H_2O}\,(l)} \rightleftharpoons \underset{\text{acid}}{\mathrm{H_3O^+}\,(aq)} + \underset{\text{base}}{\mathrm{SO_4^{2-}}\,(aq)}$$

So, water is capable of being either an acid or a base, a characteristic called amphoterism. An **amphoteric substance** is one that is capable of acting as either an acid or a base by donating or accepting hydrogen ions.

Conjugate Acids and Bases

When a substance that is acting as a Brønsted-Lowry acid donates its proton, it becomes a base in the reverse reaction. In the reaction above, the hydrogen sulfate ion $(\mathrm{HSO_4^-})$ donates a proton to water and becomes a sulfate ion $(\mathrm{SO_4^{2-}})$. The $\mathrm{HSO_4^-}$ and the $\mathrm{SO_4^{2-}}$ are linked to one another by the presence or absence of the $\mathrm{H^+}$ ion. A **conjugate acid-base pair** is a pair of substances related by the loss or gain of a single hydrogen ion. A **conjugate acid** is the particle produced when a base accepts a proton. The hydrogen sulfate ion is the conjugate acid of the sulfate ion. A **conjugate base** is the particle produced when an acid donates a proton. The sulfate ion is the conjugate base of the hydrogen sulfate ion.

In the reaction illustrated below, water serves both as acid and base simultaneously. One water molecule serves as an acid and donates a proton. The other water molecule functions as a base by accepting the proton.

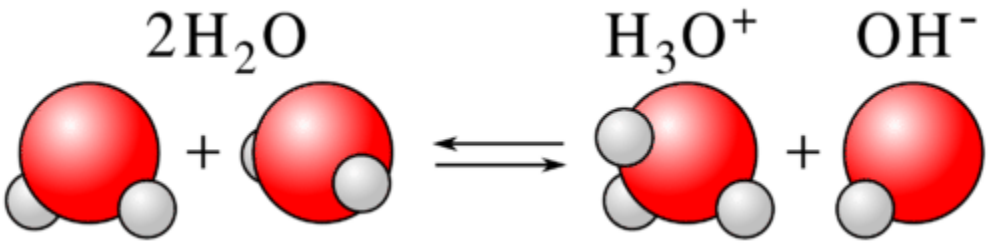


Figure 18.6.1: Water molecules as a conjugate acid-base pair. (CC BY-NC; CK-12)

A typical Brønsted-Lowry acid-base reaction contains two conjugate acid-base pairs, as shown below.

$$\mathrm{HNO_2}\,(aq) + \mathrm{PO_4}^{3-}\,(aq) \rightleftharpoons \mathrm{NO_2}^{-}\,(aq) + \mathrm{HPO_4}^{2-}\,(aq)$$

One conjugate acid-base pair is $\mathrm{NHO_2/NO_2}^{-}$, while the other pair is $\mathrm{HPO_4}^{2-}/\mathrm{PO_4}^{3-}$.

Summary

- An acid-base reaction, according to the Brønsted-Lowry definition, is a transfer of a proton from one molecule or ion to another.
- A conjugate acid-base pair is a pair of substances related by the loss or gain of a single hydrogen ion.
- A conjugate acid is the particle produced when a base accepts a proton.
- A conjugate base is the particle produced when an acid donates a proton.
- Examples of conjugate acid-base pairs are given.

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18.7: Lewis Acids and Bases

Ideas in science do not stay static. One discovery builds upon another. The concept of acids and bases has grown from the fundamental ideas of Arrhenius to Brønsted-Lowry to Lewis. Each step adds to our understanding of the surrounding world, and makes the "big picture" even bigger.

Lewis Acids and Bases

Gilbert Lewis (1875-1946) proposed a third theory of acids and bases that is even more general than either the Arrhenius or Brønsted-Lowry theories. A **Lewis acid** is a substance that accepts a pair of electrons to form a covalent bond. A **Lewis base** is a substance that donates a pair of electrons to form a covalent bond. So, a Lewis acid-base reaction is represented by the transfer of a pair of electrons from a base to an acid. A hydrogen ion, which lacks any electrons, accepts a pair of electrons. It is an acid under both the Brønsted-Lowry and Lewis definitions. Ammonia consists of a nitrogen atom as the central atom with a lone pair of electrons. The reaction between ammonia and the hydrogen ion can be depicted as shown in the figure below.

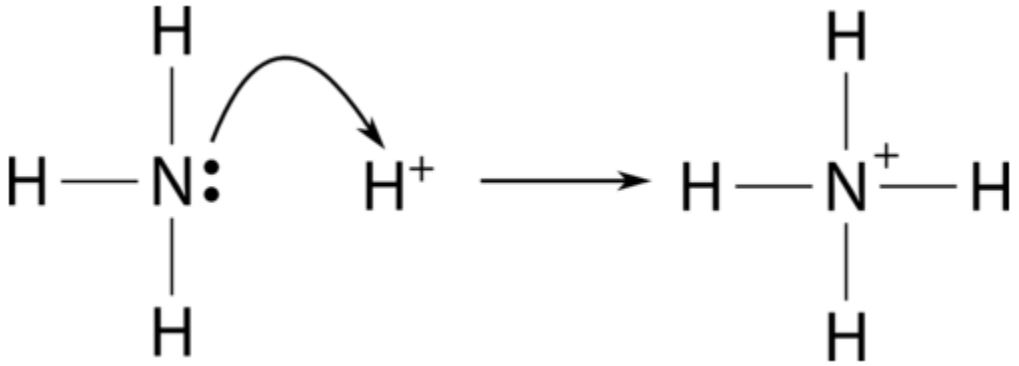


Figure 18.7.1: Reaction between ammonia and a proton. (CC BY-NC; CK-12)

The lone pair on the nitrogen atom is transferred to the hydrogen ion, making the $\mathrm{NH_3}$ a Lewis base while the $\mathrm{H^+}$ is a Lewis acid.

Some reactions that do not qualify as acid-base reactions under the other definitions do so under only the Lewis definition. An example is the reaction of ammonia with boron trifluoride.

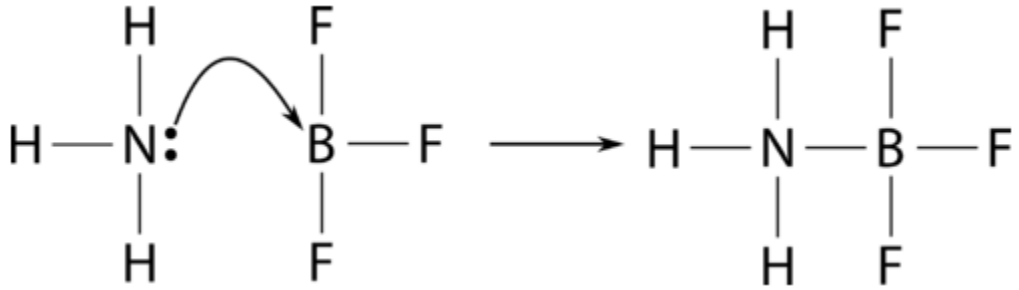


Figure 18.7.2: Ammonia and boron trifluoride. (CC BY-NC; CK-12)

Boron trifluoride is the Lewis acid, while ammonia is again the Lewis base. As there is no hydrogen ion involved in this reaction, it qualifies as an acid-base reaction only under the Lewis definition. The table below summarizes the three acid-base theories.

Acid-Base Definitions

Table 18.7.1: Acid-Base Definitions

Type	Acid	Base
Arrhenius	$\mathrm{H^+}$ ions in solution	$\mathrm{OH^-}$ ions in solution

Table 18.7.1: Acid-Base Definitions		
Brønsted-Lowry	H^+ donor	H^+ acceptor
Lewis	electron-pair acceptor	electron-pair donor

Summary

- A Lewis acid is a substance that accepts a pair of electrons to form a covalent bond.
- A Lewis base is a substance that donates a pair of electrons to form a covalent bond.
- Examples of Lewis acids and bases are given.

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18.8: Ion-Product of Water

At one time, you could take the little caps off the top of a car battery and check the condition of the sulfuric acid inside. If it got low, you could add more acid. But, sulfuric acid is hazardous stuff, so the batteries are now sealed to protect people. Because of the acid's dangerous nature, it is not a good idea to cut into a battery to see what it looks like—you could get acid burns.

The Ion-Product of Water

The **self-ionization of water** (the process in which water ionizes to hydronium ions and hydroxide ions) occurs to a very limited extent. When two molecules of water collide, there can be a transfer of a hydrogen ion from one molecule to the other. The products are a positively charged hydronium ion and a negatively charged hydroxide ion.

$$\mathrm{H_2O}\,(l) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H_3O^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

We often use the simplified form of the reaction:

$$\mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

The equilibrium constant for the self-ionization of water is referred to as the ion-product for water and is given the symbol K_w.

$$K_\mathrm{w} = [\mathrm{H^+}]\,[\mathrm{OH^-}]$$

The **ion-product of water** (K_w) is the mathematical product of the concentration of hydrogen ions and hydroxide ions. Note that $\mathrm{H_2O}$ is not included in the ion-product expression because it is a pure liquid. The value of K_w is very small, in accordance with a reaction that favors the reactants. At $25^\circ\mathrm{C}$, the experimentally determined value of K_w in pure water is 1.0×10^{-14}.

$$K_\mathrm{w} = [\mathrm{H^+}]\,[\mathrm{OH^-}] = 1.0 \times 10^{-14}$$

In pure water, the concentrations of hydrogen and hydroxide ions are equal to one another. Pure water or any other aqueous solution in which this ratio holds is said to be neutral. To find the molarity of each ion, the square root of K_w is taken.

$$[\mathrm{H_+}] = [\mathrm{OH^-}] = 1.0 \times 10^{-7}$$

An **acidic solution** is a solution in which the concentration of hydrogen ions is greater than the concentration of hydroxide ions. For example, hydrogen chloride ionizes to produce $\mathrm{H^+}$ and $\mathrm{Cl^-}$ ions upon dissolving in water.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H^+}\,(aq) + \mathrm{Cl^-}\,(aq)$$

This increases the concentration of $\mathrm{H^+}$ ions in the solution. According to Le Chatelier's principle, the equilibrium represented by $\mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq)$ is forced to the left, towards the reactant. As a result, the concentration of the hydroxide ion decreases.

A **basic solution** is a solution in which the concentration of hydroxide ions is greater than the concentration of hydrogen ions. Solid potassium hydroxide dissociates in water to yield potassium ions and hydroxide ions.

$$\mathrm{KOH}\,(s) \rightarrow \mathrm{K^+}\,(aq) + \mathrm{OH^-}\,(aq)$$

The increase in concentration of the $\mathrm{OH^-}$ ions causes a decrease in the concentration of the $\mathrm{H^+}$ ions and the ion-product of $[\mathrm{H^+}]\,[\mathrm{OH^-}]$ remains constant.

Example 18.8.1

Hydrochloric acid (HCl) is a strong acid, meaning it is 100% ionized in solution. What is the $[\mathrm{H^+}]$ and the $[\mathrm{OH^-}]$ in a solution of $2.0 \times 10^{-3}\ \mathrm{M}\ \mathrm{HCl}$?

Solution

Step 1: List the known values and plan the problem.

Known

- $[\mathrm{HCl}] = 2.0 \times 10^{-3}\ \mathrm{M}$
- $K_\mathrm{w} = 1.0 \times 10^{-14}$

Unknown

- $[\mathrm{H^+}] = ?\ \mathrm{M}$
- $[\mathrm{OH^-}] = ?\ \mathrm{M}$

Because HCl is 100% ionized, the concentration of $\mathrm{H^+}$ ions in solution will be equal to the original concentration of HCl. Each HCl molecule that was originally present ionizes into one $\mathrm{H^+}$ ion and one $\mathrm{Cl^-}$ ion. The concentration of $\mathrm{OH^-}$ can then be determined from the $[\mathrm{H^+}]$ and K_w.

Step 2: Solve.

$$[\mathrm{H^+}] = 2.0 \times 10^{-3}\,\mathrm{M}$$

$$K_\mathrm{w} = [\mathrm{H^+}]\,[\mathrm{OH^-}] = 1.0 \times 10^{-14}$$

$$[\mathrm{OH^-}] = \frac{K_\mathrm{w}}{[\mathrm{H^+}]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-3}} = 5.0 \times 10^{-12}\,\mathrm{M}$$

Step 3: Think about your result.

The $[\mathrm{H^+}]$ is much higher than the $[\mathrm{OH^-}]$ because the solution is acidic. As with other equilibrium constants, the unit for K_w is customarily omitted.

Summary

- The self-ionization of water is described and an ionization constant for the process is stated.
- Acidic and basic solutions are defined.
- Calculations using K_w are illustrated.

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18.9: The pH Scale

Grapefruit juice has a pH of somewhere between 2.9 and 3.3, depending on the specific product. Excessive exposure to this juice can cause erosion of tooth enamel and can lead to tooth damage. The acids in grapefruit juice are carbon-based, with citric acid being one of the major constituents. This compound has three ionizable hydrogens on each molecule which contribute to the relatively low pH of the juice. Another component of grapefruit juice is malic acid, containing two ionizable hydrogens per molecule.

The pH Scale

Expressing the acidity of a solution by using the molarity of the hydrogen ion is cumbersome because the quantities are generally very small. Danish scientist Søren Sørensen (1868-1939) proposed an easier system for indicating the concentration of H^+ called the pH scale. The letters pH stand for the power of the hydrogen ion. The **pH** of a solution is the negative logarithm of the hydrogen-ion concentration:

$$\mathrm{pH} = -\log\ [\mathrm{H}^+]$$

In pure water or a neutral solution, the $[\mathrm{H}^+] = 1.0 \times 10^{-7}\ \mathrm{M}$. Substituting into the pH expression:

$$\mathrm{pH} = -\log\left[1.0 \times 10^{-7}\right] = -\left(-7.00\right) = 7.00$$

The pH of pure water or any neutral solution is thus 7.00. For recording purposes, the numbers to the right of the decimal point in the pH value are the significant figures. Since 1.0×10^{-7} has two significant figures, the pH can be reported as 7.00.

A logarithmic scale condenses the range of acidity to numbers that are easy to use. Consider a solution with $[\mathrm{H}^+] = 1.0 \times 10^{-4}\ \mathrm{M}$. That is a hydrogen-ion concentration that is 1000 times higher than the concentration in pure water. The pH of such a solution is 4.00, a difference of just 3 pH units. Notice that when the $[\mathrm{H}^+]$ is written in scientific notation and the coefficient is 1, the pH is simply the exponent with the sign changed. The pH of a solution with $[\mathrm{H}^+] = 1 \times 10^{-2}\ \mathrm{M}$ is 2 and the pH of a solution with $[\mathrm{H}^+] = 1 \times 10^{-10}\ \mathrm{M}$ is 10.

As we saw earlier, a solution with $[\mathrm{H}^+]$ higher than 1.0×10^{-7} is acidic, while a solution with $[\mathrm{H}^+]$ lower than 1.0×10^{-7} is basic. Consequently, solutions with a pH of less than 7 are acidic, while those with a pH higher than 7 are basic. Figure 18.9.1 illustrates this relationship, along with some examples of various solutions.

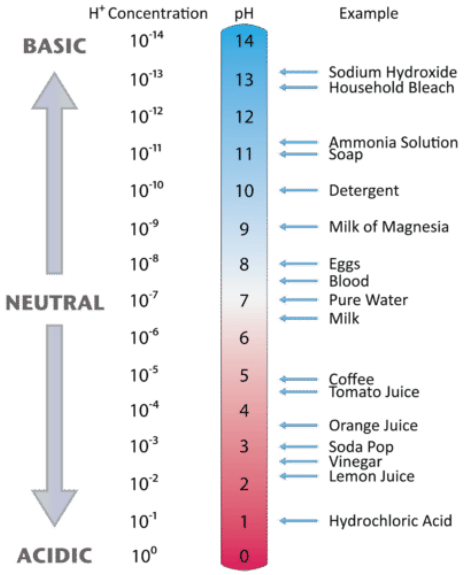


Figure 18.9.1: The pH values for several common materials. (CC BY-NC; CK-12)

Summary

- The pH of a solution is the negative logarithm of the hydrogen-ion concentration.
- Solutions with a pH of less than 7 are acidic, while those with a pH higher than 7 are basic.
- pH values for several common materials are listed.

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18.10: Calculating pH of Acids and Bases

It's not uncommon to see a tropical fish tank in homes or businesses. These brightly-colored creatures are relaxing to watch, but do require a certain amount of maintenance in order for them to survive. Tap water is usually too alkaline when it comes out of the faucet, so some adjustments need to be made. The pH of the water will change over time while it is in the tank, which means it needs to be tested every so often. Any fish tank caretaker has a chance to be a chemist for their fish!

Calculating pH of Acids and Bases

Calculation of pH is simple when there is a $1 \times 10^{\text{power}}$ problem. However, in real life, this is rarely the situation. If the coefficient is not equal to 1, a calculator must be used to find the pH. For example, the pH of a solution with $[\text{H}^+] = 2.3 \times 10^{-5}\ \text{M}$ can be found as shown below.

$$\text{pH} = -\log\left[2.3 \times 10^{-5}\right] = 4.64$$

When the pH of a solution is known, the concentration of the hydrogen ion can be calculated. The inverse of the logarithm (or antilog) is the 10^x key on a calculator.

$$[\text{H}^+] = 10^{-\text{pH}}$$

For example, suppose that you have a solution with a pH of 9.14. To find the $[\text{H}^+]$ use the 10^x key.

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-9.14} = 7.24 \times 10^{-10}\ \text{M}$$

Hydroxide Ion Concentration and pH

As we saw earlier, the hydroxide ion concentration of any aqueous solution is related to the hydrogen ion concentration through the value of K_w. We can use that relationship to calculate the pH of a solution of a base.

Example 18.10.1

Sodium hydroxide is a strong base. Find the pH of a solution prepared by dissolving $1.0\ \text{g}$ of NaOH into enough water to make $1.0\ \text{L}$ of solution.

Solution

Step 1: List the known values and plan the problem.

Known

- Mass $\text{NaOH} = 1.0\ \text{g}$
- Molar mass $\text{NaOH} = 40.00\ \text{g/mol}$
- Volume solution $= 1.0\ \text{L}$
- $K_\text{w} = 1.0 \times 10^{-14}$

Unknown

First, convert the mass of NaOH to moles. Second, calculate the molarity of the NaOH solution. Because NaOH is a strong base and is soluble, the $[\text{OH}^-]$ will be equal to the concentration of the NaOH. Third, use K_w to calculate the $[\text{H}^+]$ in the solution. Lastly, calculate the pH.

Step 2: Solve.

$$1.00 \ \cancel{\text{g NaOH}} \times \frac{1 \text{ mol NaOH}}{40.00 \ \cancel{\text{g NaOH}}} = 0.025 \text{ mol NaOH}$$

$$\text{Molarity} = \frac{0.025 \text{ mol NaOH}}{1.00 \text{ L}} = 0.025 \text{ M NaOH} = 0.025 \text{ M OH}^-$$

$$[\text{H}^+] = \frac{K_\text{w}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{0.025 \text{ M}} = 4.0 \times 10^{-13} \text{ M}$$

$$\text{pH} = -\log [\text{H}^+] = -\log \left(4.0 \times 10^{-13}\right) = 12.40$$

Step 3: Think about your result.

The solution is basic and so its pH is greater than 7. The reported pH is rounded to two decimal places because the original mass and volume has two significant figures.

Summary

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18.11: The pOH Concept

Soap Lake has a long history as a healing place. Indian tribes would put aside their rivalries when they came to the lake to enjoy the high mineral content of the water. In the days before good antibiotics, thousands of visitors would come and enjoy the soothing waters of the lake. Soap Lake is alkaline, with water quality though to be similar to that of the moons of the planet Jupiter.

The pOH Concept

As with the hydrogen-ion concentration, the concentration of the hydroxide ion can be expressed logarithmically by the pOH. The **pOH** of a solution is the negative logarithm of the hydroxide-ion concentration:

$$\text{pOH} = -\log\left[\text{OH}^-\right]$$

The pH of a solution can be related to the pOH. Consider a solution with a pH $= 4.0$. The $[\text{H}^+]$ of the solution would be 1.0×10^{-4} M. Dividing K_w by this yields a $[\text{OH}^-]$ of 1.0×10^{-10} M. Finally the pOH of the solution equals $-\log\left(1.0 \times 10^{-10}\right) = 10$. This example illustrates the following relationship.

$$\text{pH} + \text{pOH} = 14$$

The pOH scale is similar to the pH scale in that a pOH of 7 is indicative of a neutral solution. A basic solution has a pOH less than 7, while an acidic solution has a pOH of greater than 7. The pOH is convenient to use when finding the hydroxide ion concentration from a solution with a known pH.

Example 18.11.1

Find the hydroxide concentration of a solution with a pH of 4.42.

Solution

Step 1: List the known values and plan the problem.

Known

- pH $= 4.42$
- pH + pOH $= 14$

Unknown

First, the pOH is calculated, followed by the $[\text{OH}^-]$.

Step 2: Solve.

$$\text{pOH} = 14 - \text{pH} = 14 - 4.42 = 9.58$$
$$\left[\text{OH}^-\right] = 10^{-\text{pOH}} = 10^{-9.58} = 2.6 \times 10^{-10}\ \text{M}$$

Step 3: Think about your result.

The pH is that of an acidic solution, and the resulting hydroxide-ion concentration is less than 1×10^{-7} M. The answer has two significant figures because the given pH has two decimal places.

The diagram below shows all of the interrelationships between $[\text{H}^+]$, $[\text{OH}^-]$, pH, and pOH.

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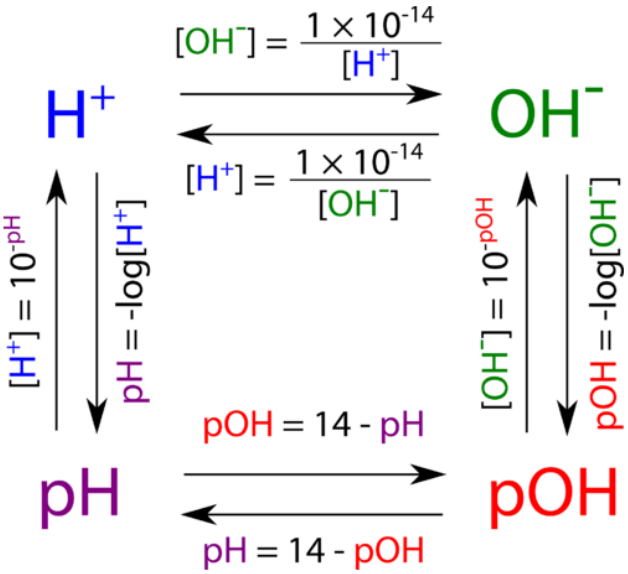


Figure 18.11.1: Relationships between hydrogen ion concentration, hydroxide ion concentration, pH and pOH. (CC BY-NC; CK-12)

Summary

- The pOH of a solution is the negative logarithm of the hydroxide-ion concentration.
- Calculations involving pOH are described.

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18.12: Strong and Weak Acids and Acid Ionization Constant (K_{a})

The etching of glass is a slow process that can produce beautiful artwork. Traditionally, the glass has been treated with dilute hydrofluoric acid which gradually dissolves the glass under it. Parts of the piece that should not be etched are covered with wax or some other non-reactive material. In more recent times, compounds such as ammonium bifluoride have been used. Whichever chemical is employed, the artist must be very careful not to get any on their skin.

Strong and Weak Acids and Acid Ionization Constant

Acids are classified as either strong or weak, based on their ionization in water. A **strong acid** is an acid which is completely ionized in an aqueous solution. Hydrogen chloride (HCl) ionizes completely into hydrogen ions and chloride ions in water.

$$\mathrm{HCl}\,(g) \rightarrow \mathrm{H}^{+}\,(aq) + \mathrm{Cl}^{-}\,(aq)$$

A **weak acid** is an acid that ionizes only slightly in an aqueous solution. Acetic acid (found in vinegar) is a very common weak acid. Its ionization is shown below.

$$\mathrm{CH_3COOH}\,(aq) \rightleftharpoons \mathrm{H}^{+}\,(aq) + \mathrm{CH_3COO}^{-}\,(aq)$$

The ionization of acetic acid is incomplete, and so the equation is shown with a double arrow. The extent of ionization of weak acids varies, but is generally less than 10%. A $0.10\ \mathrm{M}$ solution of acetic acid is only about 1.3% ionized, meaning that the equilibrium strongly favors the reactants.

Weak acids, like strong acids, ionize to yield the H^{+} ion and a conjugate base. Because HCl is a strong acid, its conjugate base (Cl^{-}) is extremely weak. The chloride ion is incapable of accepting the H^{+} ion and becoming HCl again. In general, the stronger the acid, the weaker its conjugate base. Likewise, the weaker the acid, the stronger its conjugate base.

Table $18.12.1$*: Relative Strengths of Acids and their Conjugate Bases*

Acid	Conjugate Base
Strong Acids	
HCl (hydrochloric acid) (strongest)	Cl^{-} (chloride ion) (weakest)
$\mathrm{H_2SO_4}$ (sulfuric acid)	$\mathrm{HSO_4}^{-}$ (hydrogen sulfate ion)
$\mathrm{HNO_3}$ (nitric acid)	$\mathrm{NO_3}^{-}$ (nitrate ion)
Weak Acids	
$\mathrm{H_3PO_4}$ (phosphoric acid)	$\mathrm{H_2PO_4}^{-}$ (dihydrogen phosphate ion)
$\mathrm{CH_3COOH}$ (acetic acid)	$\mathrm{CH_3COO}^{-}$ (acetate ion)
$\mathrm{H_2CO_3}$ (carbonic acid)	$\mathrm{HCO_3}^{-}$ (hydrogen carbonate ion)
HCN (hydrocyanic acid) (weakest)	CN^{-} (cyanide ion) (strongest)

Strong acids are 100% ionized in solution. Weak acids are only slightly ionized. Phosphoric acid is stronger than acetic acid, and so is ionized to a greater extent. Acetic acid is stronger than carbonic acid, and so on.

The Acid Ionization Constant, K_{a}

The ionization for a general weak acid, HA, can be written as follows:

$$\mathrm{HA}\,(aq) \rightarrow \mathrm{H}^{+}\,(aq) + \mathrm{A}^{-}\,(aq)$$

Because the acid is weak, an equilibrium expression can be written. An **acid ionization constant** (K_a) is the equilibrium constant for the ionization of an acid.

$$K_a = \frac{[H^+]\,[A^-]}{[HA]}$$

The acid ionization represents the fraction of the original acid that has been ionized in solution. Therefore, the numerical value of K_a is a reflection of the strength of the acid. Weak acids with relatively higher K_a values are stronger than acids with relatively lower K_a values. Because strong acids are essentially 100% ionized, the concentration of the acid in the denominator is nearly zero and the K_a value approaches infinity. For this reason, K_a values are generally reported for weak acids only.

The table below is a listing of acid ionization constants for several acids. Note that polyprotic acids have a distinct ionization constant for each ionization step, with each successive ionization constant being smaller than the previous one.

Table $18.12.2$*: Acid Ionization Constants at* $25°C$

Name of Acid	Ionization Equation	K_a
Sulfuric acid	$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$ $HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$	very large 1.3×10^{-2}
Oxalic acid	$H_2C_2O_4 \rightleftharpoons H^+ + HC_2O_4^-$ $HC_2O_4^- \rightleftharpoons H^+ + C_2O_4^{2-}$	6.5×10^{-2} 6.1×10^{-5}
Phosphoric acid	$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$ $H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2--}$ $HPO_4^{2-} \rightleftharpoons H^+ + PO_4^{3-}$	7.5×10^{-3} 6.2×10^{-8} 4.8×10^{-13}
Hydrofluoric acid	$HF \rightleftharpoons H^+ + F^-$	7.1×10^{-4}
Nitrous acid	$HNO_2 \rightleftharpoons H^+ + NO_2^-$	4.5×10^{-4}
Benzoic acid	$C_6H_5COOH \rightleftharpoons H^+ + C_6H_5COO^-$	6.5×10^{-5}
Acetic acid	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$	1.8×10^{-5}
Carbonic acid	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$ $HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	4.2×10^{-7} 4.8×10^{-11}
Hydrocyanic acid	$HCN \rightleftharpoons H^+ + CN^-$	4.9×10^{-10}

Summary

- A strong acid is an acid which is completely ionized in an aqueous solution.
- A weak acid is an acid that ionizes only slightly in an aqueous solution.
- The acid ionization constant (K_a) is defined.

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18.13: Strong and Weak Bases and Base Ionization Constant

All the complex electronics and apparatuses in a space shuttle generate heat, as do the astronauts. The shuttles have a complex arrangement of systems to dissipate that heat into outer space. One of the components of this system is a series of coils filled with ammonia that are located on the outside of the shuttle. Ammonia absorbs the heat and then releases it into space as the gas circulates through the coils. This approach is both inexpensive and effective.

Strong and Weak Bases and Base Ionization Constant, K_{b}

As with acids, bases can either be strong or weak, depending on the extent of their ionization. A **strong base** is a base that ionizes completely in an aqueous solution. The most common strong bases are soluble metal hydroxide compounds such as potassium hydroxide. Some metal hydroxides are not as strong, simply because they are not as soluble. Calcium hydroxide is only slightly soluble in water, but the portion that does dissolve also dissociates into ions.

A **weak base** is a base that ionizes only slightly in an aqueous solution. Recall that a base can be defined as a substance that accepts a hydrogen ion from another substance. When a weak base such as ammonia is dissolved in water, it accepts an H^{+} ion from water, forming the hydroxide ion and the conjugate acid of the base, the ammonium ion.

$$\mathrm{NH_3}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{NH_4}^{+}\,(aq) + \mathrm{OH}^{-}\,(aq) \tag{18.13.1}$$

The equilibrium greatly favors the reactants and the extent of ionization of the ammonia molecule is very small.

An equilibrium expression can be written for the reactions of weak bases with water. Because the concentration of water is extremely large and virtually constant, the water is not included in the expression. A **base ionization constant** (K_{b}) is the equilibrium constant for the ionization of a base. For ammonia, the expression is:

$$K_{\mathrm{b}} = \frac{[\mathrm{NH_4}^{+}]\,[\mathrm{OH}^{-}]}{[\mathrm{NH_3}]} \tag{18.13.2}$$

The numerical value of K_{b} is a reflection of the strength of the base. Weak bases with relatively high K_{b} values are stronger than bases with relatively low K_{b} values. The table below is a listing of base ionization constants for several weak bases.

Table 18.13.1: Base Ionization Constants at $25^{\circ}\mathrm{C}$

Name of Base	Ionization Equation	K_{b}
Methylamine	$\mathrm{CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3}^{+} + \mathrm{OH}^{-}$	5.6×10^{-4}
Ammonia	$\mathrm{NH_3 + H_2O \rightleftharpoons NH_4}^{+} + \mathrm{OH}^{-}$	1.8×10^{-5}
Pyridine	$\mathrm{C_5H_5N + H_2O \rightleftharpoons C_5H_5NH}^{+} + \mathrm{OH}^{-}$	1.7×10^{-9}
Acetate ion	$\mathrm{CH_3COO}^{-} + \mathrm{H_2O} \rightleftharpoons \mathrm{CH_3COOH} + \mathrm{OH}^{-}$	5.6×10^{-10}
Fluoride ion	$\mathrm{F}^{-} + \mathrm{H_2O} \rightleftharpoons \mathrm{HF} + \mathrm{OH}^{-}$	1.4×10^{-11}
Urea	$\mathrm{H_2NCONH_2 + H_2O \rightleftharpoons H_2NCONH_3}^{+} + \mathrm{OH}^{-}$	1.5×10^{-14}

Notice that the conjugate base of a weak acid is also a weak base. For example, the acetate ion has a small tendency to accept a hydrogen ion from water to form acetic acid and the hydroxide ion.

Summary

- A strong base is a base that ionizes completely in an aqueous solution.
- A weak base is a base that ionizes only slightly in an aqueous solution.
- A base ionization constant (K_{b}) is the equilibrium constant for the ionization of a base.
- The conjugate base of a weak acid is also a strong base.

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18.14: Relative Strengths of Acids and Bases

Learning Objectives

- Assess the relative strengths of acids and bases according to their ionization constants

We can rank the strengths of acids by the extent to which they ionize in aqueous solution. The reaction of an acid with water is given by the general expression:

$$\mathrm{HA}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{A^-}(aq) \tag{18.14.1}$$

Water is the base that reacts with the acid HA, $\mathrm{A^-}$ is the conjugate base of the acid HA, and the hydronium ion is the conjugate acid of water. A strong acid yields 100% (or very nearly so) of $\mathrm{H_3O^+}$ and $\mathrm{A^-}$ when the acid ionizes in water; Figure $18.14.1$ lists several strong acids. A weak acid gives small amounts of $\mathrm{H_3O^+}$ and $\mathrm{A^-}$.

Figure $18.14.1$*: Some of the common strong acids and bases are listed here.*

Six Strong Acids		Six Strong Bases	
$\mathrm{HClO_4}$	perchloric acid	LiOH	lithium hydroxide
HCl	hydrochloric acid	NaOH	sodium hydroxide
HBr	hydrobromic acid	KOH	potassium hydroxide
HI	hydroiodic acid	$\mathrm{Ca(OH)_2}$	calcium hydroxide
$\mathrm{HNO_3}$	nitric acid	$\mathrm{Sr(OH)_2}$	strontium hydroxide
$\mathrm{H_2SO_4}$	sulfuric acid	$\mathrm{Ba(OH)_2}$	barium hydroxide

The relative strengths of acids may be determined by measuring their equilibrium constants in aqueous solutions. In solutions of the same concentration, stronger acids ionize to a greater extent, and so yield higher concentrations of hydronium ions than do weaker acids.

We can rank the strengths of bases by their tendency to form hydroxide ions in aqueous solution. The reaction of a base with water is given by:

$$\mathrm{B}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{HB^+}(aq) + \mathrm{OH^-}(aq) \tag{18.14.2}$$

Water is the acid that reacts with the base, $\mathrm{HB^+}$ is the conjugate acid of the base B, and the hydroxide ion is the conjugate base of water. A strong base yields 100% (or very nearly so) of $\mathrm{OH^-}$ and $\mathrm{HB^+}$ when it reacts with water; Figure $18.14.1$ lists several strong bases. A weak base yields a small proportion of hydroxide ions. Soluble ionic hydroxides such as NaOH are considered strong bases because they dissociate completely when dissolved in water.

The Ionization of Weak Acids and Weak Bases

Many acids and bases are weak; that is, they do not ionize fully in aqueous solution. A solution of a weak acid in water is a mixture of the nonionized acid, hydronium ion, and the conjugate base of the acid, with the nonionized acid present in the greatest concentration. Thus, a weak acid increases the hydronium ion concentration in an aqueous solution (but not as much as the same amount of a strong acid).

Acetic acid $(\mathrm{CH_3CO_2H})$ is a weak acid. When we add acetic acid to water, it ionizes to a small extent according to the equation:

$$\mathrm{CH_3CO_2H}(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{H_3O^+}(aq) + \mathrm{CH_3CO_2{}^-}(aq) \tag{18.14.3}$$

giving an equilibrium mixture with most of the acid present in the nonionized (molecular) form. This equilibrium, like other equilibria, is dynamic; acetic acid molecules donate hydrogen ions to water molecules and form hydronium ions and acetate ions at the same rate that hydronium ions donate hydrogen ions to acetate ions to reform acetic acid molecules and water molecules.

We can tell by measuring the pH of an aqueous solution of known concentration that only a fraction of the weak acid is ionized at any moment (Figure $18.14.4$). The remaining weak acid is present in the nonionized form.



Figure $18.14.4$: pH paper indicates that a 0.1-M solution of HCl (beaker on left) has a pH of 1. The acid is fully ionized and $[\mathrm{H_3O^+}]$ = *0.1 M. A 0.1-M solution of CH_3CO_2H (beaker on right) has a pH of 3 (* $[\mathrm{H_3O^+}]$ = *0.001 M) because the weak acid CH_3CO_2H is only partially ionized. In this solution,* $[\mathrm{H_3O^+}] < [\mathrm{CH_3CO_2H}]$. *(credit: modification of work by Sahar Atwa)*

At equilibrium, a solution of a weak base in water is a mixture of the nonionized base, the conjugate acid of the weak base, and hydroxide ion with the nonionized base present in the greatest concentration. Thus, a weak base increases the hydroxide ion concentration in an aqueous solution (but not as much as the same amount of a strong base).

For example, a solution of the weak base trimethylamine, $(CH_3)_3N$, in water reacts according to the equation:

$$(\mathrm{CH_3})_3\mathrm{N}(aq) + \mathrm{H_2O}(l) \rightleftharpoons (\mathrm{CH_3})_3\mathrm{NH^+}(aq) + \mathrm{OH^-}(aq) \tag{18.14.4}$$

This gives an equilibrium mixture with most of the base present as the nonionized amine. This equilibrium is analogous to that described for weak acids.

We can confirm by measuring the pH of an aqueous solution of a weak base of known concentration that only a fraction of the base reacts with water (Figure 14.4.5). The remaining weak base is present as the unreacted form.



Figure 18.14.5: pH paper indicates that a 0.1-M solution of NH_3 (left) is weakly basic. The solution has a pOH of 3 ($[OH^-]$ = 0.001 M) because the weak base NH_3 only partially reacts with water. A 0.1-M solution of NaOH (right) has a pOH of 1 because NaOH is a strong base (credit: modification of work by Sahar Atwa).

Example 18.14.6: Formic Acid in Ant stings

Formic acid, HCO_2H, is the irritant that causes the body's reaction to ant stings.



The pain of an ant's sting is caused by formic acid. (credit: John Tann)

Summary

The strengths of acids and bases in aqueous solutions can be determined by their acid or base ionization constants. Strong acids are completely ionized in aqueous solution because their conjugate bases are weaker bases than water. Weak acids are only partially ionized because their conjugate bases are strong enough to compete successfully with water for possession of protons. Strong bases react with water to quantitatively form hydroxide ions. Weak bases give only small amounts of hydroxide ion.

Glossary

percent ionization
: ratio of the concentration of the ionized acid to the initial acid concentration, times 100

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18.15: Calculating K_{a} and K_{b}

The pH meter was invented because Florida orange growers needed a way to test the acidity of their fruit. The first meter was invented by Arnold Beckman, who went on to form Beckman Instruments. Beckman's business was very successful, and he used much of his fortune to fund science education and research. The Beckman family donated $40 million to build the Beckman Institute at the University of Illinois.

Calculating K_{a} and K_{b}

The numerical value of K_{a} and K_{b} can be determined from an experiment. A solution of known concentration is prepared and its pH is measured with an instrument called a **pH meter**.



Figure 18.15.1: A pH meter is a laboratory device that provides quick, accurate measurements of the pH of solutions. (CC BY-NC; CK-12)

Example 18.15.1

A $0.500\ \mathrm{M}$ solution of formic acid is prepared and its pH is measured to be 2.04. Determine the K_{a} for formic acid.

Solution

Step 1: List the known values and plan the problem.

Known

- Initial $[\mathrm{HCOOH}] = 0.500\ \mathrm{M}$
- pH $= 2.04$

Unknown

First, the pH is used to calculate the $[\mathrm{H}^+]$ at equilibrium. An ICE table is set up in order to determine the concentrations of HCOOH and HCOO^- at equilibrium. All concentrations are then substituted into the K_{a} expression and the K_{a} value is calculated.

Step 2: Solve.

$$[\mathrm{H}^+] = 10^{-\mathrm{pH}} = 10^{-2.04} = 9.12 \times 10^{-3}\ \mathrm{M}$$

Since each formic acid molecule that ionizes yields one H^+ ion and one formate ion (HCOO^-), the concentrations of H^+ and HCOO^- are equal at equilibrium. We assume that the initial concentrations of each ion are zero, resulting in the following ICE table.

	HCOOH	$\mathrm{H^+}$	$\mathrm{HCOO^-}$
Initial	0.500	0	0
Change	-9.12×10^{-3}	$+9.12 \times 10^{-3}$	$+9.12 \times 10^{-3}$
Equilibrium	0.491	9.12×10^{-3}	9.12×10^{-3}

Now, substituting into the K_a expression gives:

$$K_\mathrm{a} = \frac{[\mathrm{H^+}]\,[\mathrm{HCOO^-}]}{[\mathrm{HCOOH}]} = \frac{\left(9.12 \times 10^{-3}\right)\left(9.12 \times 10^{-3}\right)}{0.491} = 1.7 \times 10^{-4}$$

Step 3: Think about your result.

The value of K_a is consistent with that of a weak acid. Two significant figures are appropriate for the answer, since there are two digits after the decimal point in the reported pH.

Similar steps can be taken to determine the K_b of a base. For example, a $0.750\ \mathrm{M}$ solution of the weak base ethylamine $(\mathrm{C_2H_5NH_2})$ has a pH of 12.31.

$$\mathrm{C_2H_5NH_2 + H_2O \rightleftharpoons C_2H_5NH_3{}^+ + OH^-}$$

Since one of the products of the ionization reaction is the hydroxide ion, we need to first find the $[\mathrm{OH^-}]$ at equilibrium. The pOH is $14 - 12.31 = 1.69$. The $[\mathrm{OH^-}]$ is then found from $10^{-1.69} = 2.04 \times 10^{-2}\ \mathrm{M}$. The ICE table is then set up as shown below.

	$\mathrm{C_2H_5NH_2}$	$\mathrm{C_2H_5NH_3{}^+}$	$\mathrm{OH^-}$
Initial	0.750	0	0
Change	-2.04×10^{-2}	$+2.04 \times 10^{-2}$	$+2.04 \times 10^{-2}$
Equilibrium	0.730	2.04×10^{-2}	2.04×10^{-2}

Substituting into the K_b expression yields the K_b for ethylamine.

$$K_\mathrm{b} = \frac{[\mathrm{C_2H_5NH_3{}^+}]\,[\mathrm{OH^-}]}{[\mathrm{C_2H_5NH_2}]} = \frac{\left(2.04 \times 10^{-2}\right)\left(2.04 \times 10^{-2}\right)}{0.730} = 5.7 \times 10^{-4}$$

Summary

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18.16: Calculating pH of Weak Acid and Base Solutions

Bees are beautiful creatures that help plants flourish. They carry pollen from one plant to another to facilitate plant growth and development. However, they can be troublesome when they sting! For those who are allergic to bee venom, this can be a serious, life-threatening problem. For all other humans, it can be a painful experience. When stung by a bee, one first-aid treatment is to apply a paste of baking soda (sodium bicarbonate) to the stung area. This weak base helps with the itching and swelling that accompanies the bee sting.

Calculating pH of Weak Acid and Base Solutions

The K_a and K_b values have been determined for a great many acids and bases, as shown in Tables 21.12.2 and 21.13.1. These can be used to calculate the pH of any solution of a weak acid or base whose ionization constant is known.

Example 18.16.1

Calculate the pH of a $2.00\ \text{M}$ solution of nitrous acid (HNO_2). The K_a for nitrous acid is 4.5×10^{-4}.

Solution

Step 1: List the known values and plan the problem.

Known

- Initial $[\text{HNO}_2] = 2.00\ \text{M}$
- $K_a = 4.5 \times 10^{-4}$

Unknown

First, an ICE table is set up with the variable x used to signify the change in concentration of the substance due to ionization of the acid. Then the K_a expression is used to solve for x and calculate the pH.

Step 2: Solve.

	HNO_2	H^+	NO_2^-
Initial	2.00	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$2.00 - x$	x	x

The K_a expression and value are used to set up an equation to solve for x.

$$K_a = 4.5 \times 10^{-4} = \frac{(x)(x)}{2.00 - x} = \frac{x^2}{2.00 - x}$$

The quadratic equation is required to solve this equation for x. However, a simplification can be made of the fact that the extent of ionization of weak acids is small. The value of x will be significantly less than 2.00, so the "$-x$" in the denominator can be dropped.

$$4.5 \times 10^{-4} = \frac{x^2}{2.00 - x} \approx \frac{x^2}{2.00}$$

$$x = \sqrt{4.5 \times 10^{-4}\,(2.00)} = 2.9 \times 10^{-2}\ \text{M} = [\text{H}^+]$$

Since the variable x represents the hydrogen-ion concentration, the pH of the solution can now be calculated.

$$\text{pH} = -\log[\text{H}^+] = -\log\left[2.9 \times 10^{-2}\right] = 1.54$$

Step 3: Think about your result.

> The pH of a $2.00\ \mathrm{M}$ solution of a strong acid would be equal to $-\log(2.00) = -0.30$. The higher pH of the $2.00\ \mathrm{M}$ nitrous acid is consistent with it being a weak acid and therefore not as acidic as a strong acid would be.

The procedure for calculating the pH of a solution of a weak base is similar to that of the weak acid in the example. However, the variable x will represent the concentration of the hydroxide ion. The pH is found by taking the negative logarithm to get the pOH, followed by subtracting from 14 to get the pH.

Summary

- The procedure for calculating the pH of a weak acid or base is illustrated.

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18.17: Neutralization Reaction and Net Ionic Equations for Neutralization Reactions

Pouring concrete and working it are messy jobs. In the process, a lot of wastewater with an alkaline pH is generated. Often, regulations require that this wastewater be cleaned up at the site. One practical way to neutralize the basic pH is to bubble $\mathrm{CO_2}$ into the water. The carbon dioxide forms a weak acid (carbonic acid, $\mathrm{H_2CO_3}$) in solution which serves to bring the alkaline pH down to something closer to neutral.

Neutralization Reactions and Net Ionic Equations for Neutralization Reactions

A **neutralization reaction** is a reaction in which an acid and a base react in an aqueous solution to produce a salt and water. The aqueous sodium chloride that is produced in the reaction is called a salt. A **salt** is an ionic compound composed of a cation from a base and an anion from an acid. A salt is essentially any ionic compound that is neither an acid nor a base.

Strong Acid-Strong Base Reactions

When equal amounts of a strong acid such as hydrochloric acid are mixed with a strong base such as sodium hydroxide, the result is a neutral solution. The products of the reaction do not have the characteristics of either an acid or a base. The balanced molecular equation is:

$$\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$$

Chemical reactions occurring in aqueous solution are more accurately represented with a net ionic equation. The full ionic equation for the neutralization of hydrochloric acid by sodium hydroxide is written as follows:

$$\mathrm{H^+}\,(aq) + \mathrm{Cl^-}\,(aq) + \mathrm{Na^+}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{Na^+}\,(aq) + \mathrm{Cl^-}\,(aq) + \mathrm{H_2O}\,(l)$$

Since the acid and base are both strong, they are fully ionized and so are written as ions, as is the NaCl formed as a product. The sodium and chloride ions are spectator ions in the reaction, leaving the following as the net ionic reaction.

$$\mathrm{H^+}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{H_2O}\,(l)$$

All neutralization reactions of a strong acid with a strong base simplify to the net ionic reaction of hydrogen ion combining with hydroxide ion to produce water.

What if the acid is a diprotic acid such as sulfuric acid? The balanced molecular equation now involves a 1:2 ratio between acid and base.

$$\mathrm{H_2SO_4}\,(aq) + 2\mathrm{NaOH}\,(aq) \rightarrow \mathrm{Na_2SO_4}\,(aq) + \mathrm{H_2O}\,(l)$$

In order for the reaction to be a full neutralization, twice as many moles of NaOH must react with the $\mathrm{H_2SO_4}$. The sodium sulfate salt is soluble, and so the net ionic reaction is again the same. Different mole ratios occur for other polyprotic acids or bases with multiple hydroxides such as $\mathrm{Ca(OH)_2}$.

Reactions Involving a Weak Acid or Weak Base

Reactions where at least one of the components is weak do not generally result in a neutral solution. The reaction between weak nitrous acid and strong potassium hydroxide is shown below.

$$\mathrm{HNO_2}\,(aq) + \mathrm{KOH}\,(aq) \rightarrow \mathrm{KNO_2}\,(aq) + \mathrm{H_2O}\,(l)$$

In order to write the net ionic equation, the weak acid must be written as a molecule since it does not ionize to a great extent in water. The base and the salt are fully dissociated.

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$$\mathrm{HNO_2}\,(aq) + \mathrm{K^+}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{K^+}\,(aq) + \mathrm{NO_2{}^-}\,(aq) + \mathrm{H_2O}\,(l)$$

The only spectator ion is the potassium ion, resulting in the net ionic equation:

$$\mathrm{HNO_2}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{NO_2{}^-}\,(aq) + \mathrm{H_2O}\,(l)$$

The strong hydroxide ion essentially "forces" the weak nitrous acid to become ionized. The hydrogen ion from the acid combines with the hydroxide ion to form water, leaving the nitrite ion as the other product. The resulting solution is not neutral (pH $= 7$), but instead is slightly basic.

Reactions can also involve a weak base and strong acid, resulting in a solution that is slightly acidic. The molecular and net ionic equations for the reaction of hydrochloric acid and ammonia are shown below.

$$\begin{aligned}&\mathrm{HCl}\,(aq) + \mathrm{NH_3}\,(aq) \rightarrow \mathrm{NH_4Cl}\,(aq)\\&\mathrm{H^+}\,(aq) + \mathrm{NH_3}\,(aq) \rightarrow \mathrm{NH_4{}^+}\,(aq) \qquad (\mathrm{Cl^-\ is\ a\ spectator\ ion})\end{aligned}$$

Reactions between acids and bases that are both weak may result in solutions that are neutral, acidic, or basic.

Summary

- A neutralization reaction is a reaction in which an acid and a base react in an aqueous solution to produce a salt and water.
- Equations for acid-base neutralizations are given.
- Net ionic equations for neutralization reactions are given.

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18.18: Titration Experiment

A large amount of current research involves the development of biodiesel fuels. Often this material can be made from used vegetable oils. The vegetable oil is treated with lye to create the biofuel. In the oils is a variable amount of acid that needs to be determined, so that the workers will know how much lye to add to make the final fuel. Before the lye is added, the native vegetable oil is titrated to find out how much free acid is present. Then, the amount of lye added can be adjusted to take into account the amount needed to neutralize these free acids.

Titration Experiment

In the neutralization of hydrochloric acid by sodium hydroxide, the mole ratio of acid to base is 1:1.

$$\mathrm{HCl}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$$

One mole of HCl would be fully neutralized by one mole of NaOH. If instead the hydrochloric acid was reacted with barium hydroxide, the mole ratio would be 2:1.

$$2\mathrm{HCl}\,(aq) + \mathrm{Ba(OH)_2}\,(aq) \rightarrow \mathrm{BaCl_2}\,(aq) + 2\mathrm{H_2O}\,(l)$$

Now two moles of HCl would be required to neutralize one mole of $\mathrm{Ba(OH)_2}$. The mole ratio ensures that the number of moles of $\mathrm{H^+}$ ions supplied by the acid is equal to the number of $\mathrm{OH^-}$ ions supplied by the base. This must be the case for neutralization to occur. The **equivalence point** is the point in a neutralization reaction where the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions.

In the laboratory, it is useful to have an experiment where the unknown concentration of an acid or a base can be determined. This can be accomplished by performing a controlled neutralization reaction. A **titration** is an experiment where a volume of a solution of known concentration is added to a volume of another solution in order to determine its concentration. Many titrations are acid-base neutralization reactions, though other types of titrations can also be performed.

In order to perform an acid-base titration, the chemist must have a way to visually detect that the neutralization reaction has occurred. An **indicator** is a substance that has a distinctly different color when in an acidic or basic solution. A commonly used indicator for strong acid-strong base titrations is phenolphthalein. Solutions in which a few drops of phenolphthalein have been added turn from colorless to brilliant pink as the solution turns from acidic to basic. The steps in a titration reaction are outlined below.

1. A measured volume of an acid of unknown concentration is added to an Erlenmeyer flask.
2. Several drops of an indicator are added to the acid and mixed by swirling the flask.
3. A buret is filled with a base solution of known molarity.
4. The stopcock of the buret is opened and base is slowly added to the acid, while the flask is constantly swirled to ensure mixing. The stopcock is closed at the exact point at which the indicator just changes color.

The **standard solution** is the solution in a titration whose concentration is known. In the titration described above, the base solution is the standard solution. It is very important in a titration to add the solution from the buret slowly, so that the point at which the indicator changes color can be found accurately. The **end point** of a titration is the point at which the indicator changes color. When phenolphthalein is the indicator, the end point will be signified by a faint pink color.

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Figure 18.18.1: Phenolphthalein in basic solution. (CC BY-NC; CK-12)

Summary

- The equivalence point is the point in a neutralization reaction where the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions.
- A titration is an experiment where a volume of a solution of known concentration is added to a volume of another solution in order to determine its concentration. (Many titrations are acid-base neutralization reactions.)
- An indicator is a substance that has a distinctly different color when in an acidic or basic solution.
- The process for carrying out a titration is described.

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18.19: Titration Calculations

The manufacture of soap requires a number of chemistry techniques. One necessary piece of information is the saponification number. This is the amount of base needed to hydrolyze a certain amount of fat to produce the free fatty acids that are an essential part of the final product. The fat is heated with a known amount of base (usually NaOH or KOH). After hydrolysis is complete, the leftover base is titrated to determine how much was needed to hydrolyze the fat sample.

Titration Calculations

At the equivalence point in a neutralization, the moles of acid are equal to the moles of base.

$$\text{moles acid} = \text{moles base}$$

Recall that the molarity (M) of a solution is defined as the moles of the solute divided by the liters of solution (L). So the moles of solute are therefore equal to the molarity of a solution multiplied by the volume in liters.

$$\text{moles solute} = \mathrm{M} \times \mathrm{L}$$

We can then set the moles of acid equal to the moles of base.

$$\mathrm{M}_A \times \mathrm{V}_A = \mathrm{M}_B \times \mathrm{V}_B$$

M_A is the molarity of the acid, while M_B is the molarity of the base. V_A and V_B are the volumes of the acid and base, respectively.

Suppose that a titration is performed and $20.70\ \mathrm{mL}$ of $0.500\ \mathrm{M}\ \mathrm{NaOH}$ is required to reach the end point when titrated against $15.00\ \mathrm{mL}$ of HCl of unknown concentration. The above equation can be used to solve for the molarity of the acid.

$$\mathrm{M}_A = \frac{\mathrm{M}_B \times \mathrm{V}_B}{\mathrm{V}_A} = \frac{0.500\ \mathrm{M} \times 20.70\ \mathrm{mL}}{15.00\ \mathrm{mL}} = 0.690\ \mathrm{M}$$

The higher molarity of the acid compared to the base in this case means that a smaller volume of the acid is required to reach the equivalence point.

The above equation works only for neutralizations in which there is a 1:1 ratio between the acid and the base. The example below demonstrates the technique to solve a titration problem for a titration of sulfuric acid with sodium hydroxide.

Example 18.19.1

In a titration of sulfuric acid against sodium hydroxide, $32.20\ \mathrm{mL}$ of $0.250\ \mathrm{M}\ \mathrm{NaOH}$ is required to neutralize $26.60\ \mathrm{mL}$ of $\mathrm{H_2SO_4}$. Calculate the molarity of the sulfuric acid.

$$\mathrm{H_2SO_4}\,(aq) + 2\mathrm{NaOH}\,(aq) \rightarrow \mathrm{Na_2SO_4}\,(aq) + 2\mathrm{H_2O}\,(l)$$

Solution

Step 1: List the known values and plan the problem.

- Molarity $\mathrm{NaOH} = 0.250\ \mathrm{M}$
- Volume $\mathrm{NaOH} = 32.20\ \mathrm{mL}$
- Volume $\mathrm{H_2SO_4} = 26.60\ \mathrm{mL}$

Unknown

First determine the moles of NaOH in the reaction. From the mole ratio, calculate the moles of $\mathrm{H_2SO_4}$ that reacted. Finally, divide the moles of $\mathrm{H_2SO_4}$ by its volume to get the molarity.

Step 2: Solve.

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$$\text{mol NaOH} = \text{M} \times \text{L} = 0.250\ \text{M} \times 0.03220\ \text{L} = 8.05 \times 10^{-3}\ \text{mol NaOH}$$

$$8.05 \times 10^{-3}\ \text{mol NaOH} \times \frac{1\ \text{mol } H_2SO_4}{2\ \text{mol NaOH}} = 4.03 \times 10^{-3}\ \text{mol } H_2SO_4$$

$$\frac{4.03 \times 10^{-3}\ \text{mol } H_2SO_4}{0.02660\ \text{L}} = 0.151\ \text{M } H_2SO_4$$

Step 3: Think about your result.

The volume of H_2SO_4 required is smaller than the volume of $NaOH$ because of the two hydrogen ions contributed by each molecule.

Summary

- The process of calculating concentration from titration data is described and illustrated.

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18.20: Titration Curves

The x-y plot that we know of as a graph was the brainchild of the French mathematician-philosopher Rene Descartes (1596-1650). His studies in mathematics led him to develop what was known as "Cartesian geometry", including the concept of today's graphs. The coordinates are often referred to as Cartesian coordinates.

Titration Curves

As base is added to acid at the beginning of a titration, the pH rises very slowly. Nearer to the equivalence point, the pH begins to rapidly increase. If the titration is a strong acid with a strong base, the pH at the equivalence point is equal to 7. A bit past the equivalence point, the rate of change of the pH again slows down. A **titration curve** is a graphical representation of the pH of a solution during a titration. The figure below shows two different examples of a strong acid-strong base titration curve. On the left is a titration in which the base is added to the acid, and so the pH progresses from low to high. On the right is a titration in which the acid is added to the base. In this case, the pH starts out high and decreases during the titration. In both cases, the equivalence point is reached when the moles of acid and base are equal and the pH is 7. This also corresponds to the color change of the indicator.

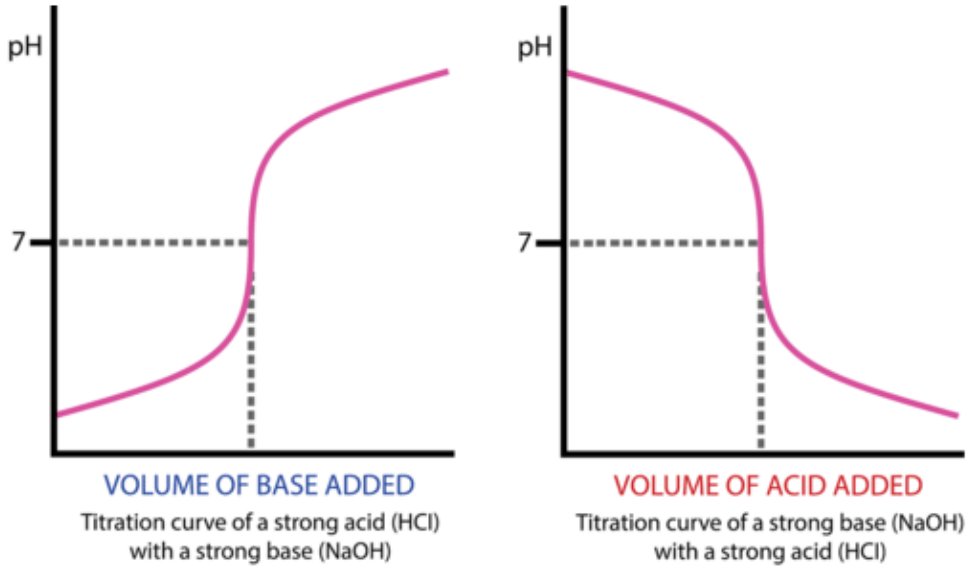


Figure 18.20.1: A titration curve shows the pH changes that occur during the titration of an acid with a base. On the left, base is being added to acid. On the right, acid is being added to base. In both cases, the equivalence point is at pH 7. (CC BY-NC; CK-12)

Titration curves can also be generated in the case of a weak acid-strong base titration or a strong acid-weak base titration. The general shape of the titration curve is the same, but the pH at the equivalence point is different. In a weak acid-strong base titration, the pH is greater than 7 at the equivalence point. In a strong acid-weak base titration, the pH is less than 7 at the equivalence point.

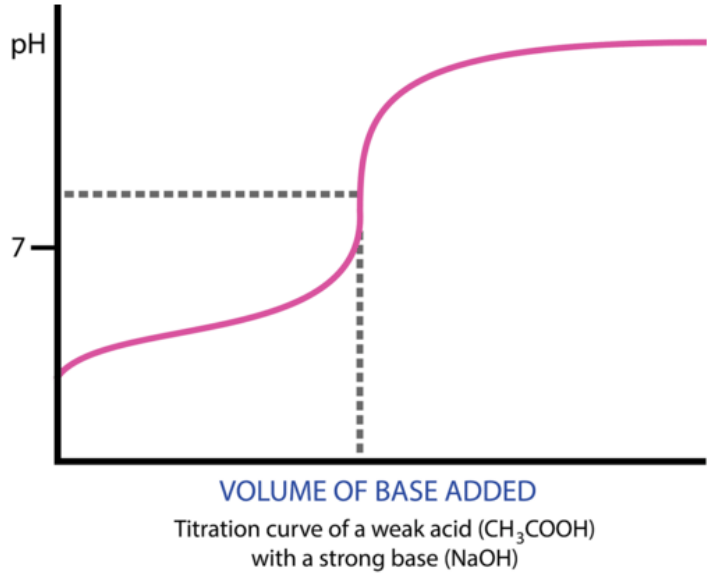


Figure 18.20.2: Titration curve of weak acid and strong base. (CC BY-NC; CK-12)

Summary

- A titration curve is a graphical representation of the pH of a solution during a titration.
- In a strong acid-strong base titration, the equivalence point is reached when the moles of acid and base are equal and the pH is 7.
- In a weak acid-strong base titration, the pH is greater than 7 at the equivalence point.
- In a strong acid-weak base titration, the pH is less than 7 at the equivalence point.

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18.21: Indicators

"Boil Them Cabbage Down" is an old bluegrass song. Many people enjoy the music, but chemistry students also enjoy the product of boiled cabbage. Extracting the anthocyanin dye from cabbage leaves with boiling water gives a solution that is red when acidic, purple when neutral, and green to yellow when basic.

Indicators

An **acid-base indicator** is a substance that displays different colors when in the presence of an acid or a base. How does that work? An indicator is a weak acid that ionizes within a known pH range, usually about 2 pH units. We can represent the protonated form of the indicator molecule as HIn and the deprotonated form as $\mathrm{In^-}$. The following equilibrium exists for the indicator:

$$\mathrm{HIn}\,(aq) \rightleftharpoons \mathrm{H^+}\,(aq) + \mathrm{In^-}\,(aq)$$

According to Le Chatelier's principle, the addition of $\mathrm{H^+}$ ions (as in a low pH solution) drives the equilibrium to the left and the protonated HIn predominates. The addition of $\mathrm{OH^-}$ (as in a high pH solution) decreases the $\mathrm{H^+}$ concentration and drives the equilibrium to the right and the deprotonated $\mathrm{In^-}$ predominates. To be useful as an indicator, the two forms must be different colors. In the case of phenolphthalein, the protonated form is colorless, while the deprotonated form is pink. The figure below shows a variety of acid-base indicators that can be used in titration experiments.

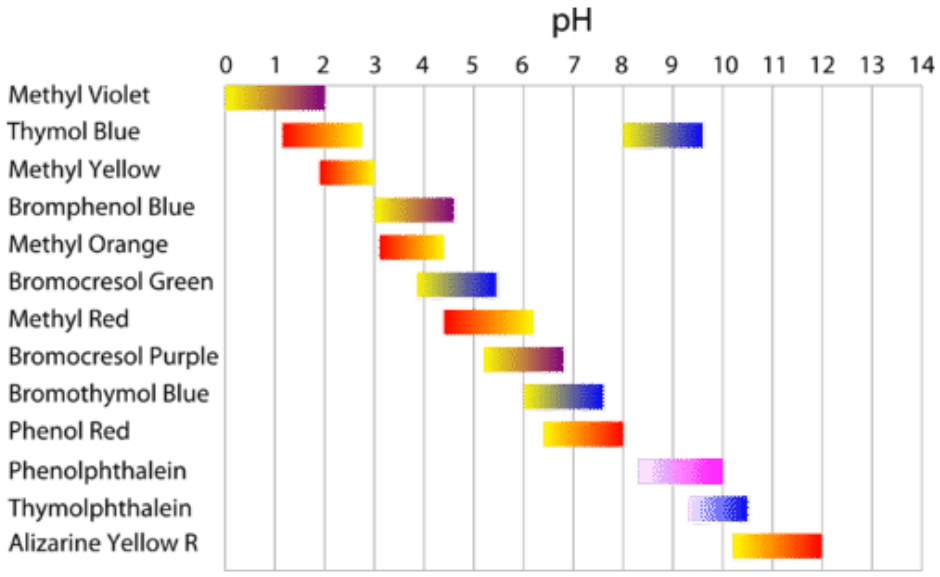


Figure $18.21.1$: pH indicators. (CC BY-NC; CK-12)

Depending on the pH at the equivalence point, the appropriate indicator must be chosen. For example, bromphenol blue has a yellow color below a pH of about 3 and a blue-violet color above a pH of about 4. Bromphenol blue would not be a good choice as the indicator for a strong acid-strong base titration, because the pH is 7 at the equivalence point. Instead, it could be used for a strong acid-weak base titration, where the pH at the equivalence point is lower.

Most indicators have two colored forms. Universal indicator displays the entire rainbow of colors from low pH to high pH (see figure below). Universal indicator is used to make pH paper, which can be used to quickly test solutions for their approximate pH.



Figure 18.21.2: Universal indicator tape. (CC BY-NC; CK-12)

Summary

- An acid-base indicator is a substance that displays different colors when in the presence of an acid or a base.
- Examples of acid-base indicators are given.

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18.22: Hydrolysis of Salts- Equations

Baking seems easy with all the pre-mixed items available. ("just add water and stir"). However, there is a good amount of chemistry involved in baking with ingredients that are measured out. One important ingredient is baking powder. The fluffiness in the final product of a non-yeast recipe is usually due to the carbon dioxide formed by baking powder. One popular brand uses a mix of sodium bicarbonate and sodium aluminum sulfate to produce the $\mathrm{CO_2}$. The reaction is:

$$3\mathrm{NaHCO_3} + \mathrm{NaAl(SO_4)_2} \rightarrow \mathrm{Al(OH)_3} + 2\mathrm{Na_2SO_4} + 3\mathrm{CO_2}$$

If all goes well, the biscuits rise, the pancakes are fluffy, and everybody is happy.

Hydrolysis of Salts: Equations

A salt is an ionic compound that is formed when an acid and a base neutralize each other. While it may seem that salt solutions are always neutral, they can frequently be either acidic or basic.

Consider the salt formed when the weak acid hydrofluoric acid is neutralized by the strong base sodium hydroxide. The molecular and net ionic equations are shown below.

$$\mathrm{HF}\,(aq) + \mathrm{NaOH}\,(aq) \rightarrow \mathrm{NaF}\,(aq) + \mathrm{H_2O}\,(l)$$
$$\mathrm{HF}\,(aq) + \mathrm{OH^-}\,(aq) \rightarrow \mathrm{F^-}\,(aq) + \mathrm{H_2O}\,(l)$$

Since sodium fluoride is soluble, the sodium ion is a spectator ion in the neutralization reaction. The fluoride ion is capable of reacting, to a small extent, with water, accepting a proton.

$$\mathrm{F^-}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{HF}\,(aq) + \mathrm{OH^-}\,(aq)$$

The fluoride ion is acting as a weak Brønsted-Lowry base. The hydroxide ion that is produced as a result of the above reaction makes the solution slightly basic. **Salt hydrolysis** is a reaction in which one of the ions from a salt reacts with water, forming either an acidic or basic solution.

Salts That Form Basic Solutions

When solid sodium fluoride is dissolved into water, it completely dissociates into sodium ions and fluoride ions. The sodium ions do not have any capability of hydrolyzing, but the fluoride ions hydrolyze to produce a small amount of hydrofluoric acid and hydroxide ion.

$$\mathrm{F^-}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{HF}\,(aq) + \mathrm{OH^-}\,(aq)$$

Salts that are derived from the neutralization of a weak acid (HF) by a strong base (NaOH) will always produce salt solutions that are basic.

Salts That Form Acidic Solutions

Ammonium chloride $(\mathrm{NH_4Cl})$ is a salt that is formed when the strong acid HCl is neutralized by the weak base $\mathrm{NH_3}$. Ammonium chloride is soluble in water. The chloride ion produced is incapable of hydrolyzing because it is the conjugate base of the strong acid HCl. In other words, the $\mathrm{Cl^-}$ ion cannot accept a proton from water to form HCl and $\mathrm{OH^-}$, as the fluoride ion did in the previous section. However, the ammonium ion is capable of reacting slightly with water, donating a proton and so acting as an acid.

$$\mathrm{NH_4{}^+}\,(aq) + \mathrm{H_2O}\,(l) \rightleftharpoons \mathrm{H_3O^+}\,(aq) + \mathrm{NH_3}\,(aq)$$

Salts That Form Neutral Solutions

A salt that is derived from the reaction of a strong acid with a strong base forms a solution that has a pH of 7. An example is

sodium chloride, formed from the neutralization of HCl by NaOH. A solution of NaCl in water has no acidic or basic properties, since neither ion is capable of hydrolyzing. Other salts that form neutral solutions include potassium nitrate $(\mathrm{KNO_3})$ and lithium bromide (LiBr). The table below summarizes how to determine the acidity or basicity of a salt solution.

Table 18.22.1

Salt formed from:	Salt Solution
Strong acid + Strong base	Neutral
Strong acid + Weak base	Acidic
Weak acid + Strong base	Basic

Salts formed from the reaction of a weak acid and a weak base are more difficult to analyze due to competing hydrolysis reactions between the cation and the anion. These salts are not considered in this chapter's concept.

Summary

- Salt hydrolysis is a reaction in which one of the ions from a salt reacts with water, forming either an acidic or basic solution.
- Salts that are derived from the neutralization of a weak acid by a strong base will always produce salt solutions that are basic.
- Salts that are derived from the neutralization of a strong acid by a weak base will always produce salt solutions that are acidic.
- A salt that is derived from the reaction of a strong acid with a strong base forms a solution that has a pH of 7.

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18.23: Buffers

Diabetes mellitus is a disorder of glucose metabolism in which insulin production by the pancreas is impaired. Since insulin helps glucose enter the cells, a decrease of this hormone means that glucose cannot be used in its normal fashion. When this happens, the body begins to break down fats, producing a decrease in blood pH. Chemical systems in the body can balance this pH shift for a while, but excessive acid production can create serious problems if not corrected by administering insulin to restore normal glucose use.

Buffers

If only $1.0\ \text{mL}$ of $0.10\ \text{M}$ hydrochloric acid is added to $1.0\ \text{L}$ of pure water, the pH drops drastically from 7.0 to 4.0. This is a 1000-fold increase in the acidity of the solution. For many purposes, it is desirable to have a solution which is capable of resisting such large changes in pH when relatively small amounts of acid or base are added to it. Such a solution is called a buffer. A **buffer** is a solution of a weak acid or a base and its salt. Both components must be present for the system to act as a buffer to resist changes in pH. Commercial buffer solutions, which have a wide variety of pH values, can be obtained.

Some common buffer systems are listed in the table below.

Table $18.23.1$*: Some Common Buffers*

Buffer System	Buffer Components	pH of buffer (equal molarities of both components)
Acetic acid/acetate ion	CH_3COOH/ CH_3COO^-	4.74
Carbonic acid/hydrogen carbonate ion	$\text{H}_2\text{CO}_3/\text{HCO}_3^-$	6.38
Dihydrogen phosphate ion/hydrogen phosphate ion	$\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$	7.21
Ammonia/ammonium ion	$\text{NH}_3/\text{NH}_4^+$	9.25

One example of a buffer is a solution made of acetic acid (the weak acid) and sodium acetate (the salt). The pH of a buffer consisting of $0.50\ \text{M}\ \text{CH}_3\text{COOH}$ and $0.50\ \text{M}\ \text{CH}_3\text{COONa}$ is 4.74. If $10.0\ \text{mL}$ of $1.0\ \text{M}\ \text{HCl}$ is added to $1.0\ \text{L}$ of the buffer, the pH only decreases to 4.73. This ability to "soak up" the additional hydrogen ions from the HCl that was added is due to the reaction below.

$$\text{CH}_3\text{COO}^-\,(aq) + \text{H}^+\,(aq) \rightarrow \text{CH}_3\text{COOH}\,(aq)$$

Since both the acetate ion and the acetic acid were already present in the buffer, the only thing that changes is the ratio of one to the other. Small changes in that ratio have only very minor effects on the pH.

If $10.0\ \text{mL}$ of $1.0\ \text{M}\ \text{NaOH}$ were added to another $1.0\ \text{L}$ of the same buffer, the pH would only increase to 4.76. In this case the buffer takes up the additional hydroxide ions.

$$\text{CH}_3\text{COOH}\,(aq) + \text{OH}^-\,(aq) \rightarrow \text{CH}_3\text{COO}^-\,(aq) + \text{H}_2\text{O}\,(l)$$

Again the ratio of acetate ion to acetic acid changes slightly, this time causing a very small increase in the pH.

It is possible to add so much acid or base to a buffer that its ability to resist a significant change in pH is overwhelmed. The **buffer capacity** is the amount of acid or base that can be added to a buffer solution before a large change in pH occurs. The buffer capacity is exceeded when the number of moles of H^+ or OH^- that are added to the buffer exceeds the number of moles of the buffer components.

Summary

- A buffer is a solution of a weak acid or a base and its salt.
- The buffer capacity is the amount of acid or base that can be added to a buffer solution before a large change in pH occurs.
- Reactions showing how buffers regulate pH are described.

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19: NUCLEAR CHEMISTRY

19.1: Discovery of Radioactivity

19.2: Nuclear Decay Processes

19.3: Detection of Radioactivity

19.4: Half-Life

19.5: Background Radiation

19.6: Nuclear Fission Processes

19.7: Nuclear Power Generation

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19.1: Discovery of Radioactivity



Figure 19.1.3: The radioactive danger symbol. (By Clkr-Free-Vector-Images; under Pixabay license)

If you visit the nuclear medicine department of a large hospital, you are very likely to see the symbol shown above. This sign means that radioactive materials are present, and special safety precautions need to be taken. These materials are used for the diagnosis and treatment of many diseases. The people using these materials are specially trained to handle them safely. Radioactive materials can be dangerous and should be respected, but need not be feared.

Discovery of Radioactivity

John Dalton first proposed his atomic theory in an 1804 lecture to the Royal Institution, a prestigious British scientific society. In this talk, he put forth the idea that all atoms of an element were identical, and that atoms were indestructible. In a little over 100 years, both of these ideas were shown to be incorrect.

In 1919, studies on atomic weights led Francis Aston (1877-1925) to the conclusion that some elements with different atomic weights were actually the same element in different isotope forms. Aston used a mass spectrograph to separate isotopes of different elements. He won the Nobel Prize in Chemistry for this work in 1922.

Natural Radioactivity

In 1895, Wilhelm Röntgen produced great interest around what caused the phenomenon of x-rays. One researcher of this phenomenon was Henri Becquerel. Becquerel studied the fluorescent properties of uranium salts, believing they had something to do with x-rays. He soon learned that uranium could expose a photographic plate without an external input of energy (thought to be needed to produce uranium fluorescence). Becquerel also showed that the uranium salt emissions could be affected by a magnetic field, which was not true for x-rays.

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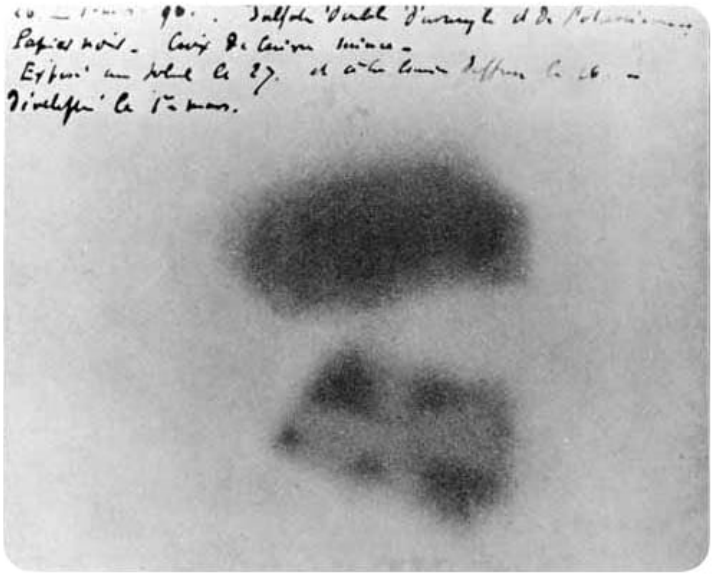


Figure 19.1.1: Becquerel's experiment: uranium exposure to photographic plate.

Pierre and Marie Curie studied the properties of uranium salts with the express purpose of identifying the details of these emissions. They were the first to coin the term "**radioactivity**", meaning the spontaneous emission of radiation in the form of particles or high energy photons, resulting from a nuclear reaction. The major contributions to the work came from Marie, who showed that the amount of radioactivity present was due to the amount of a specific element, and not due to some chemical reaction. She discovered the element polonium and named it after her native Poland. Madame Curie shared the 1903 Nobel Prize in Physics with her husband Pierre and Henri Becquerel. She won the Nobel Prize in Chemistry in 1911.

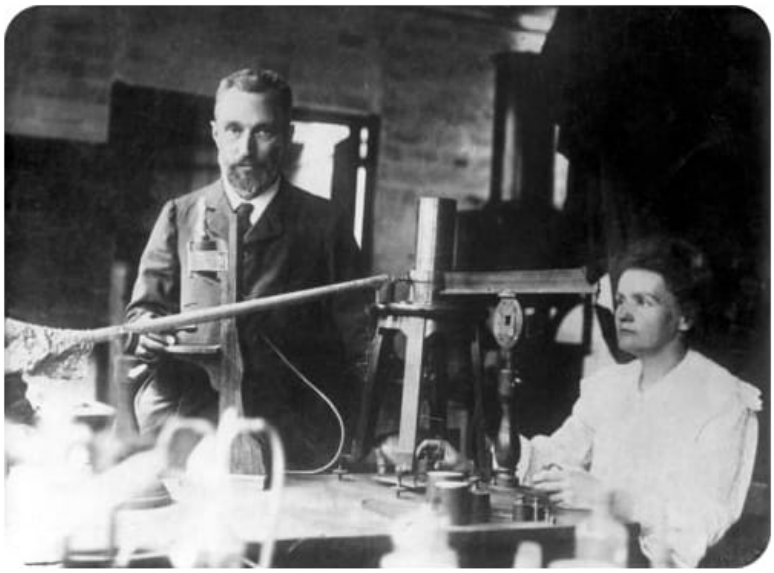


Figure 19.1.2: Pierre and Marie Curie in their lab.

Ernest Rutherford, a later researcher, was able to demonstrate three different types of radioactive emissions. These emission types differed in terms of mass, charge, and their ability to penetrate materials. He designated them simply as alpha (α) emissions, beta (β) emissions, and gamma (γ) emissions.

Radioactivity involves the spontaneous emission of material and/or energy from the nucleus of an atom. The most common radioactive atoms have high atomic numbers and contain a large excess of neutrons. Some typical radioactive elements are technetium (atomic number 43), promethium (atomic number 61), and all elements atomic number 84 (polonium) and higher. There are four primary types of emission, either involving the release of a particle from the nucleus or the release of energy. In many instances, both energy and a particle are produced by the radioactive event.

It should be noted that some elements considered to be stable do have radioactive isotopes. Carbon-14 is radioactive, but this isotope is only a small fraction of the total amount of carbon in existence (about one part per trillion carbon atoms). Hydrogen also has a radioactive isotope known as hydrogen-3 or tritium, again a very small fraction of the total elemental hydrogen present.

Summary

- Francis Aston discovered that some elements with different atomic weights were actually the same element in different isotope forms.
- Henri Becquerel showed that uranium salt emissions could be affected by a magnetic field, which was not true for x-rays.
- Marie and Pierre Curie were the first to coin the term *radioactivity*, meaning the spontaneous emission of radiation in the form of particles or high energy photons, resulting from a nuclear reaction.
- Ernest Rutherford, a later researcher, was able to demonstrate three different types of radioactive emissions: alpha (α) emissions, beta (β) emissions, and gamma (γ) emissions.

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19.2: Nuclear Decay Processes

Food irradiation is a sensitive subject for many people. The practice involves exposing food to ionizing radiation in order to kill harmful bacteria (such as salmonella) that cause sickness. The food is essentially unchanged and does not lose any nutritive value. Parasites and insect pests are easily destroyed by this process, while bacteria take longer to kill. Viruses are not affected by the radiation treatment. But don't worry—the food is not radioactive, and you will not glow in the dark if you eat it.

Nuclear Decay Processes

Radioactive decay involves the emission of a particle and/or energy as one atom changes into another. In most instances, the atom changes its identity to become a new element. There are four different types of emissions that occur.

Alpha Emission

Alpha (α) **decay** involves the release of helium ions from the nucleus of an atom. This ion consists of two protons and two neutrons and has a $2+$ charge. Release of an α-particle produces a new atom that has an atomic number two less than the original atom and an atomic weight that is four less. A typical alpha decay reaction is the conversion of uranium-238 to thorium:

$$^{238}_{92}\mathrm{U} \rightarrow {}^{234}_{90}\mathrm{Th} + {}^{4}_{2}\alpha^{+}$$

We see a decrease of two in the atomic number (uranium to thorium) and a decrease of four in the atomic weight (238 to 234). Usually the emission is not written with atomic number and weight indicated since it is a common particle whose properties should be memorized. Quite often the alpha emission is accompanied by gamma (γ) radiation, a form of energy release. Many of the largest elements in the periodic table are alpha-emitters.

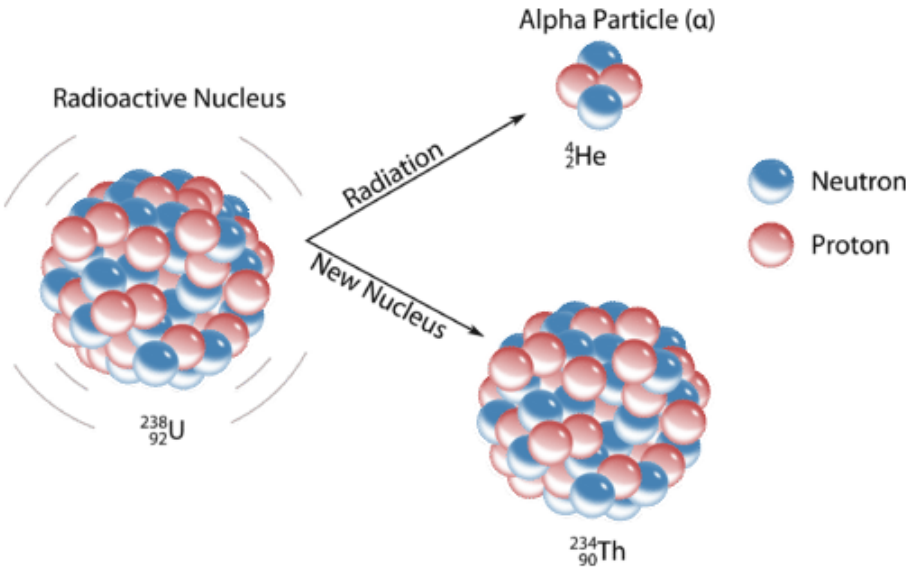


Figure 19.2.1: Emission of an alpha particle from the nucleus.

Beta Emission

Beta (β) **decay** is a more complicated process. Unlike the α-emission, which simply expels a particle, the β-emission involves the transformation of a neutron in the nucleus to a proton and an electron. The electron is then ejected from the nucleus. In the process, the atomic number increases by one while the atomic weight stays the same. As is the case with α-emissions, β -emissions are often accompanied by γ-radiation.

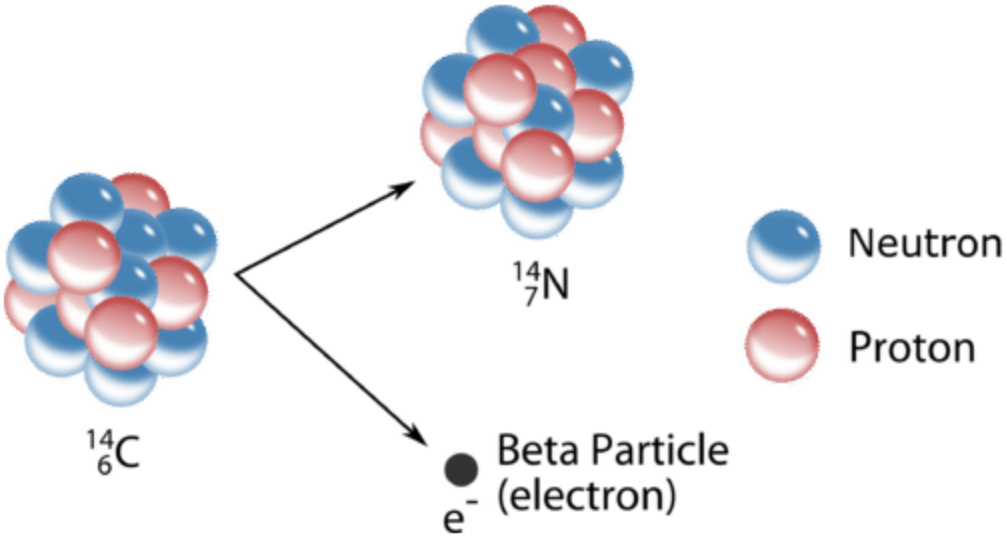


Figure 19.2.2: Beta emission.

A typical beta decay process involves carbon-14, often used in radioactive dating techniques. The reaction forms nitrogen-14 and an electron:

$$^{14}_{6}\mathrm{C} \rightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}\mathrm{e}$$

Again, the beta emission is usually simply indicated by the Greek letter β; memorization of the process is necessary in order to follow nuclear calculations in which the Greek letter β appears without further notation.

Positron Emission

A **positron** is a positive electron (a form of antimatter). This rare type of emission occurs when a proton is converted to a neutron and a positron in the nucleus, with ejection of the positron. The atomic number will decrease by one, while the atomic weight does not change. A positron is often designated by β^+.

Carbon-11 emits a positron to become boron-11:

$$^{11}_{6}\mathrm{C} \rightarrow {}^{11}_{5}\mathrm{B} + {}^{0}_{+1}\beta$$

Electron Capture

An alternate way for a nuclide to increase its neutron to proton ratio is by a phenomenon called electron capture. In electron capture, an electron from an inner orbital is captured by the nucleus of the atom, and combined with a proton to form a neutron. For example, silver-106 undergoes electron capture to become palladium-106.

$$^{106}_{47}\mathrm{Ag} + {}^{0}_{-1}\mathrm{e} \rightarrow {}^{106}_{46}\mathrm{Pd}$$

Note that the overall result of electron capture is identical to positron emission. The atomic number decreases by one while the mass number remains the same.

Gamma Emission

Gamma (γ) radiation is simply energy. It may be released by itself or, more commonly, in association with other radiation events. There is no change of atomic number or atomic weight in a simple γ-emission. Often, an isotope may produce γ-radiation as a result of a transition in a metastable isotope. This type of isotope may just "settle", with a shifting of particles in the nucleus. The composition of the atom is not altered, but the nucleus could be considered more "comfortable" after the shift. This shift increases the stability of the isotope from the energetically unstable (or "metastable") isotope to a more stable form of the nucleus.

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Summary

- Radioactive decay involves the emission of a particle and/or energy as one atom changes into another.
- Alpha (α) decay involves the release of helium ions from the nucleus of an atom.
- Beta (β) decay involves the transformation of a neutron in the nucleus to a proton and an electron.
- A positron is a positive electron (a form of antimatter). This rare type of emission occurs when a proton is converted to a neutron and a positron in the nucleus, with ejection of the positron.
- Gamma (γ) radiation is simply energy. It may be released by itself or, more commonly, in association with other radiation events.

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19.3: Detection of Radioactivity

Uranium is worth more on the open market than it was several years ago. So, people are out there looking for more uranium. However, uranium-hunters are not waving Geiger counters around to find this material—all the surface uranium has been found. Instead, they must study geology. Deep holes are drilled to explore the underground geology of a site, and the rocks obtained are chemically analyzed. No more clicking—just modern-day geology and chemistry.

Detection of Radioactivity

Radioactivity is determined by measuring the number of decay processes per unit time. Perhaps the easiest way is simply to determine the number of counts/minute, with each count measuring a single decay process, such as the emission of an α-particle. A particular isotope may have an activity of 5,000 counts/minute (cpm) while another isotope might only have $250\ \mathrm{cpm}$. The amount of activity gives a rough indication of the amount of the radioisotope present—the higher the activity, the more radioactive isotope in the sample.

Units of Measurement

The **curie** (Ci) is one measure of the rate of decay (named after Pierre and Marie Curie). One curie is equivalent to 3.7×10^{10} disintegrations per second. Since this is obviously a large and unwieldy number, radiation is often expressed in millicuries or microcuries (still very large numbers). Another measure is the **becquerel** (Bq), named after Henri Becquerel. The becquerel is defined as an activity of one disintegration/second. Both of these units are concerned with the disintegration rate of the radioactive isotope, and give no indication of dosage to the target material.

Personal Dosimeters

Measurement of exposure to radioactivity is important for anyone who deals with radioactive materials on a regular basis. Perhaps the simplest device is a personal **dosimeter**—a film badge that will fog up when exposed to radiation. The amount of fogging is proportional to the amount of radiation present. These devices are not very sensitive to low levels of radiation. More sensitive systems use crystals that respond in some way to radioactivity by registering the number of emissions in a given time. These systems tend to be more sensitive and more reliable than film badges.



Figure 19.3.1: Dosimeter.

Geiger Counters

A **Geiger counter** provides a sensitive means of detecting radioactivity. A tube is filled with an inert gas, which will conduct electricity when radiation enters it. When a charged particle comes into the tube, it changes the electrical potential between the anode and the cathode. This change in potential in the tube produces a change in voltage in the electrical circuit and registers as a count. Geiger counters are fairly inexpensive and reliable, so they are useful in a wide range of applications. More complicated types of counters are also available, but are generally used in sophisticated experiments.

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Figure $19.3.2$: Geiger counter.

Summary

- The curie (Ci) is one measure of the rate of radioactive decay; one curie is equivalent to 3.7×10^{10} disintegrations per second.
- Another measure of radioactive decay is the becquerel (Bq), defined as an activity of one disintegration/second.
- The dosimeter is a device to measure radioactivity—a film badge that will fog up when exposed to radiation.
- The Geiger counter provides a sensitive means of detecting radioactivity.

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19.4: Half-Life

Uranium isotopes produce plutonium-239 as a decay product. The plutonium can be used in nuclear weapons and is a power source for nuclear reactors, which generate electricity. This isotope has a half-life of 24,100 years, causing concern in regions where radioactive plutonium has accumulated and is stored. At some storage sites, the waste is slowly leaking into the groundwater and contaminating nearby rivers. The 24,100 year half-life means that it will be present in the earth for a very long time.

Half-Life

Radioactive materials lose some of their activity each time a decay event occurs. This loss of activity can be estimated by determining the **half-life** of an isotope. The half-life is defined as the period of time needed for one half of a given quantity of a substance to undergo a change. For a radioisotope, every time a decay event occurs, a count is detected on the Geiger counter, or other measuring device. A specific isotope might have a total count of $30,000\ \text{cpm}$. In one hour, the count could be $15,000\ \text{cpm}$ (half of the original count). So, the half-life of that isotope is one hour. Some isotopes have long half-lives—the half-life of U-234 is 245,000 years. Other isotopes have shorter half-lives. I-131, used in thyroid scans, has a half-life of 8.02 days.

Half-Life Calculations

Information on the half-life of an isotope can be used to calculate how much radioactivity of that isotope will be present after a certain period of time. There is a formula that allows calculation at any time after the initial count, but we are just going to look at loss of activity after different half-lives. The isotope I-125 is used in certain laboratory procedures and has a half-life of 59.4 days. If the initial activity of a sample of I-125 is $32,000\ \text{cpm}$, how much activity will be present in 178.2 days? Begin by determining how many half-lives are represented by 178.2 days:

$$\frac{178.2\ \text{days}}{59.4\ \text{days/half-life}} = 3\ \text{half-lives}$$

Then, simply count activity:

$$\begin{aligned}\text{initial activity } (t_0) &= 32,000\ \text{cpm}\\ \text{after one half-life} &= 16,000\ \text{cpm}\\ \text{after two half-lives} &= 8,000\ \text{cpm}\\ \text{after three half-lives} &= 4,000\ \text{cpm}\end{aligned}$$

Be sure to keep in mind that the initial count is at time zero (t_0), and to subtract from that count at the first half-life. The second half-life has an activity of half the previous count (not the initial count).

For the more mathematically inclined, the following formula can be used to calculate the amount of radioactivity remaining after a given time:

$$N_t = N_0 \times (0.5)^{\text{number of half-lives}}$$

where $N_t =$ activity at time t and $N_0 =$ initial activity at time $t = 0$.

If we have an initial activity of $42,000\ \text{cpm}$, what will the activity be after four half-lives?

$$\begin{aligned}N_t &= N_0\,(0.5)^4\\ &= (42,000)\,(0.5)\,(0.5)\,(0.5)\,(0.5)\\ &= 2625\ \text{cpm}\end{aligned}$$

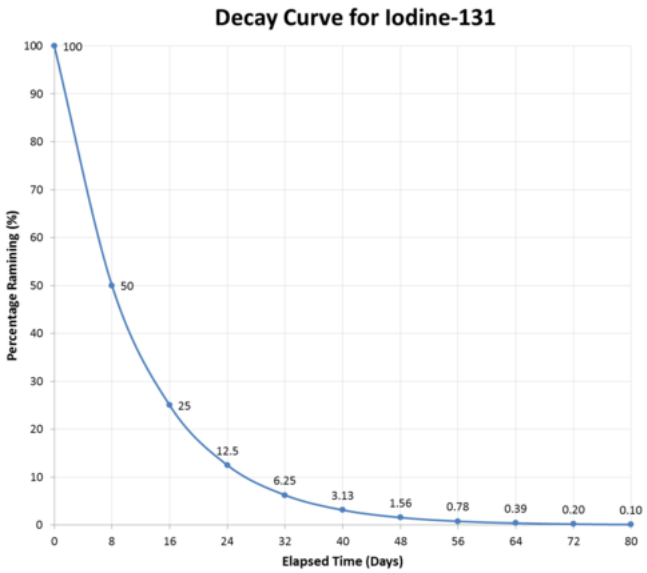


Figure 19.4.1: Typical radioactive decay curve.

The graph above illustrates a typical decay curve for a radioactive material. The activity decreases by one-half during each succeeding half-life. Half-lives of different elements vary considerably, as shown in the table below:

Table 19.4.1: Isotope Half-Lives

Isotope	Decay Mode	Half-Life
Cobalt-60	beta	5.3 years
Neptunium-237	alpha	2.1 million years
Polonium-214	alpha	0.00016 seconds
Radium-224	alpha	3.7 days
Tritium (H-3)	beta	12 years

We have talked about the activity and decay of individual isotopes. In the real world, there is a decay chain that takes place until a stable end-product is produced. For U-238, the chain is a long one, with a mix of isotopes having very different half-lives. The end of the chain resides in lead, a stable element that does not decay further.

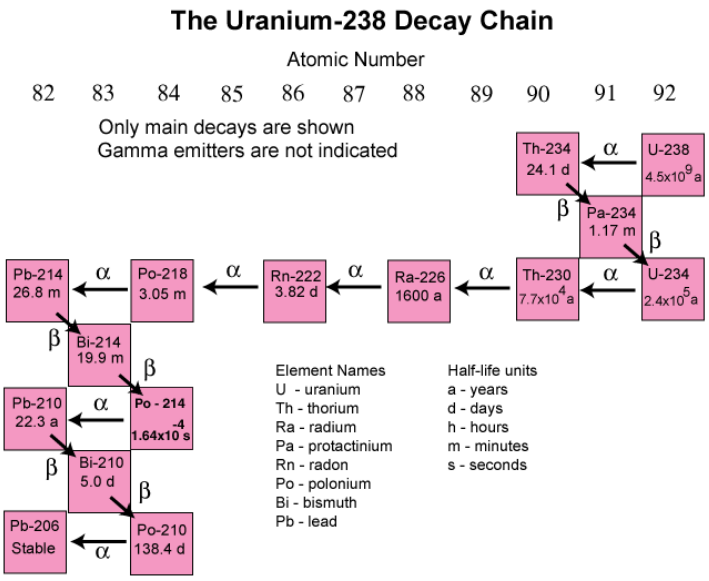


Figure 19.4.2: Decay of uranium to stable end-product.

Summary

- Radioactive materials lose some of their activity each time a decay event occurs.
- Radioactive loss of activity can be estimated by determining the half-life of an isotope.
- Half-life is defined as the period of time needed for one half of a given quantity of a substance to undergo a change.
- Half-lives of different elements vary considerably.

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19.5: Background Radiation

Sitting in a hot bath or spa has always been a great prescription for dealing with sore muscles. People used to believe that it was even more beneficial to immerse themselves in radioactive hot springs, and drink water containing radioactive materials (some still do). In the early 1900s, people spent millions of dollars on treatments and "radioactive water" with the belief that all of their ailments would cease. Radioactivity in the water was usually due to radon gas that leaked up from deep underground, formed by decay of other radioisotopes.

Background Radiation

We are all exposed to a small amount of radiation in our daily lives. This **background radiation** comes from naturally occurring sources and from human-produced radiation. Exposure to x-rays and nuclear medicine isotopes, ground sources, and cosmic radiation account for almost half of the background exposure of the average American. Radon gas, formed from the decay of uranium and thorium isotopes, is responsible for a little over half the total amount of background radiation. See the table below for background sources.

Sources of Background Radiation

Table 19.5.1: Sources of Background Radiation

radon	54%
consumer products	3%
nuclear medicine	4%
cosmic radiation	8%
terrestrial	8%
internal	11%
x-rays	11%
other	1%

The Problem of Radon

Small amounts of uranium and thorium are found in the soil of a large number of areas in the U.S. When radioactive isotopes of these elements decay, radon is one of the products formed. Radon is a colorless, odorless gas and is chemically inert, as it is one of the noble gases.

Radon is also radioactive and can easily be inhaled into the lungs. Over time, this internal radon exposure can lead to the development of lung cancer. The incidence of lung cancer in smokers exposed to radon is much higher than in non-smokers exposed to radon, since smoking has already produced some lung damage, and the radon simply makes the damage worse. Radon exposure is highest in homes lacking good air circulation to move the gas out of the residence. There are a number of inexpensive approaches to decreasing your exposure to radon. A good start is to test your living area for radon with a radon test kit.



Figure 19.5.1: Radon test kit.

Summary

- Background radiation is defined.
- Sources of background radiation are listed.

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19.6: Nuclear Fission Processes

Nuclear fission was first discovered by two German scientists, Fritz Strassman and Otto Hahn, in the 1930s. They began their work by bombarding uranium with neutrons, hoping to create larger elements. Instead, they were very surprised to find Ba-141, a much smaller element. They immediately contacted a fellow scientist in the field, Lise Meitner, who carried out calculations to demonstrate that fission had taken place.

Nuclear Fission

Radioactive decay by the release of alpha or beta particles is not the only way new isotopes are formed. When a neutron collides with a nucleus, the nucleus splits into two isotopes, each of which is roughly half the mass of the original atom. A small amount of mass is "left over" and released as energy, as predicted by Einstein's famous equation $E = mc^2$, that relates mass and energy. This process is known as **nuclear fission**. The neutron must be a "slow" neutron, traveling at a speed that is approximately that of the molecules of a gas at the same temperature in the system producing the neutrons. High-speed ("fast") neutrons will not result in nuclear fission.

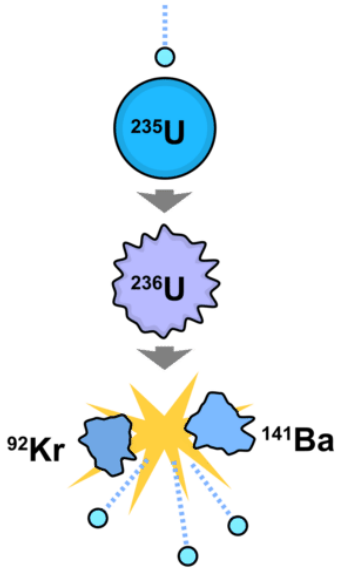


Figure 19.6.1: Fission of a uranium nucleus produced by collision with a neutron.

The example above illustrates the basic nuclear fission process. A neutron (generally produced by some controlled process, not usually a natural event) collides with an atom of U-235. Momentarily, a U-236 atom forms, which then splits into two smaller atoms (Kr-93 and Ba-141) in the diagram. This process results in the release of three new neutrons, which can then initiate fission reactions with more atoms. We will see later how this propagation of neutrons can be employed in a reactor for the generation of electricity.

An extended version of this process can be seen in the figure below. Not every collision of a neutron with U-235 results in a fission reaction. A neutron from the initial fission process may strike an atom of U-238, which does not continue the process. Another neutron may not collide with a nucleus, and be lost in the environment. However, a third neutron produced from the initial collision can collide with more U-235, and continue the chain reaction to produce more neutrons.

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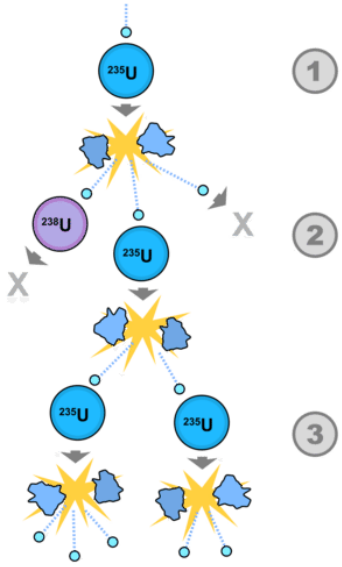


Figure 19.6.2: Fission reaction with U-235.

Typical nuclear fission reactions balance in terms of mass. The total mass of the reactants is equal to the total mass of the products:

$$^{235}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{92}_{36}\mathrm{Kr} + {}^{142}_{56}\mathrm{Ba} + 2{}^{1}_{0}\mathrm{n} + \text{energy}$$

There are a total of 236 mass units on the left of the equation and 236 mass units on the right. In the same manner, we see 92 protons on the left and 92 on the right. The energy that is released is the binding energy that holds the nucleus together.

Another set of fission products from U-235 can be seen in the following reaction:

$$^{235}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{95}_{42}\mathrm{Mo} + {}^{139}_{57}\mathrm{La} + 2{}^{1}_{0}\mathrm{n} + \text{energy}$$

Again, we see that the total number of mass units and of protons is equal on both sides of the equation.

Summary

- Nuclear fission is the process of a neutron colliding with a nucleus. The nucleus splits into two isotopes, each of which is roughly half the mass of the original atom. A small amount of mass is "left over" and released as energy.
- Examples of nuclear fission processes are illustrated.

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19.7: Nuclear Power Generation

On Wednesday, March 28, 1979, the residents of Middleton, PA woke up to a very scary situation. A nuclear power plant near the town had experienced a series of malfunctions that led to the release of some radioactive gases into the atmosphere, along with a partial meltdown of the reactor core. Fortunately, follow-up studies showed that there were no health effects on workers or the general public. A thorough investigation was conducted that led to significant improvements in safety and operation of these nuclear power plants. One of the two reactors was shut down completely, while the other remained in operation until its permanent deactivation in 2014.

Nuclear Power Generation

The generation of electricity is critical for operation of businesses, health care delivery, schools, homes, and other areas requiring the use of electrical power. According to 2011 statistics, coal is used for 42% of the total power generated, with natural gas being employed for another 25%. Nuclear power plants are employed in about 19% of the cases, with renewable energy sources supplying the last 13%. All of these fuels are used to heat water to generate steam. The steam then turns a turbine to generate electricity.

The diagram below shows the layout of a typical nuclear power plant. The radioactive rods are in the red container along with water, which is heated to steam. The energy for this heat comes from fission reactions of uranium. The steam passes through the turbine and causes the turbine to spin, generating electricity. As the steam condenses, it is run through a cooling tower to lower its temperature. The water then recirculates through the reactor core to be used again.

The control rods play an important role in the modulation of the nuclear chain reaction (usually a collision of a neutron with uranium). Each collision produces more neutrons than were present initially. If left unsupervised, the reaction would soon get out of control. Rods are commonly made of boron or a number of metals and metal alloys. The purpose of the control rods is to absorb neutrons to regulate the rate of the chain reaction, so that the water does not overheat and destroy the reactor.

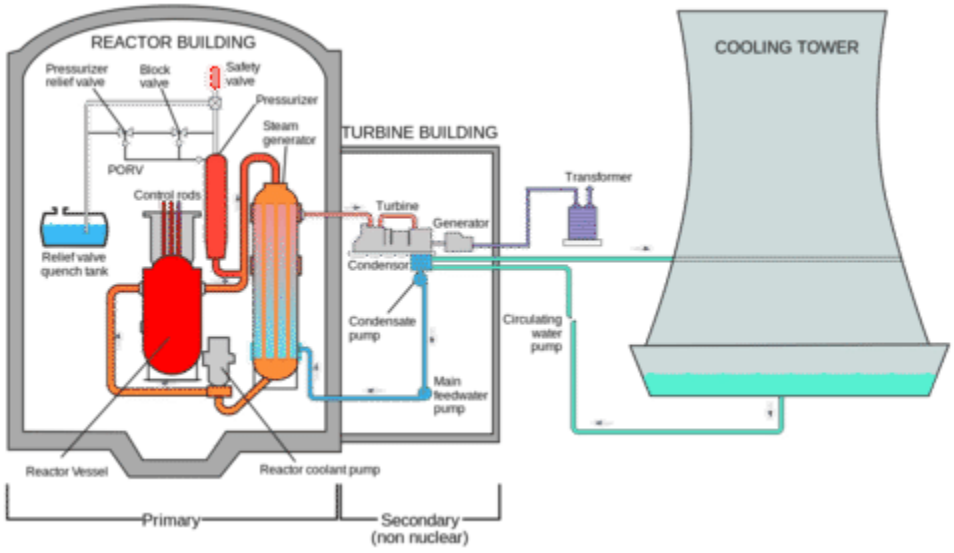


Figure 19.7.1: Schematic for a nuclear power plant.

Nuclear power is also used to propel ships. The turbine can be connected to a propeller system. The rotating turbine shaft will turn the propeller to move the ship.



Figure 19.7.2: Nuclear submarine.

Summary

- The importance of nuclear power in generating electricity is described.
- The operation of a nuclear power plant is described.

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19.8: Nuclear Fusion

A number of reactions take place in the sun that cannot be duplicated on Earth. Some of these reactions involve the formation of large elements from smaller ones. So far, we have only been able to observe formation of very small elements here on Earth. The reaction sequence observed appears to be the following: hydrogen-1 atoms collide to form the larger hydrogen isotopes, hydrogen-2 (deuterium) and hydrogen-3 (tritium). In the process, positrons and gamma rays are formed. The positrons will collide with any available electrons and annihilate, producing more gamma rays. In the process, tremendous amounts of energy are produced to keep us warm and continue supplying reactions.

Nuclear Fusion

In contrast to nuclear fission, which results in smaller isotopes being formed from larger ones, the goal of **nuclear fusion** is to produce larger materials from the collision of smaller atoms. The forcing of the smaller atoms together results in tighter packing, and the release of energy. As seen in the figure below, energy is released in the formation of the larger atom, helium (He) from the fusion of hydrogen-2 and hydrogen-3, as well as from the expulsion of a neutron.

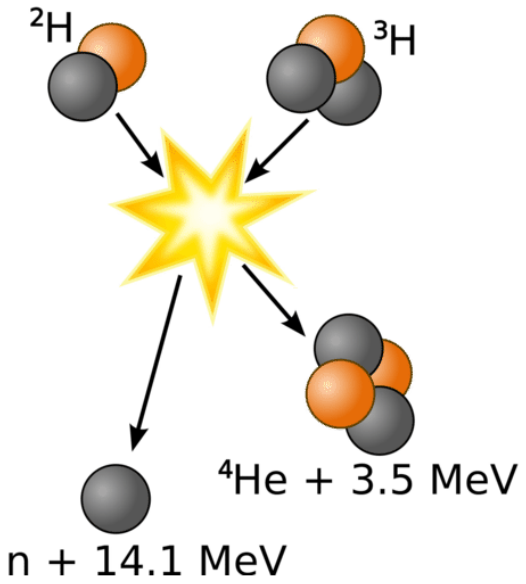


Figure $19.8.1$: Nuclear fusion reaction between deuterium and tritium.

This release of energy is what drives research on fusion reactors today. If such a reaction could be accomplished efficiently on Earth, it could provide a clean source of nuclear energy. Unlike fission reactions, nuclear fusion does not produce radioactive products that represent hazards to living systems.

Nuclear fusion reactions in the laboratory have been extraordinarily difficult to achieve. Extremely high temperatures (i.e., millions of degrees) are required. Methods must be developed to force the atoms together and hold them together long enough to react. The neutrons released during the fusion reactions can interact with atoms in the reactor, and convert them to radioactive materials. There has been some success in the field of nuclear fusion reactions, but the journey to feasible fusion power is still a long and uncertain one.

Summary

- Nuclear fusion produces larger materials from the collision of smaller atoms. The forcing of the smaller atoms together results in tighter packing, and the release of energy.
- Examples of nuclear fusion reactions are given.

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19.9: Penetrating Ability of Emissions

Designated containers can store radioisotopes that will be administered to patients as part of a medical treatment. Isotopes are stored in lead containers that block the radiation from escaping and exposing people. These lead containers are called a "pig", a term long used for referring to castings of lead or other metals. The origin of the term is not clear, but was probably first applied to these containers by the people who made them.

Penetrating Ability of Emissions

The various emissions differ considerably in their ability to go through matter, known as their **penetrating ability**. The α-particle has the least penetrating power since it is the largest and slowest emission. It can be blocked by a sheet of paper or a human hand. Beta particles are more penetrating than alpha particles, but can be stopped by a thin sheet of aluminum. Of the three basic types of emissions, gamma particles are the most penetrating. A thick lead shield is required to stop gamma emissions. Positrons represent a special case in that they annihilate when they come in contact with electrons. The collision of a positron and an electron results in the formation of two gamma emissions that go 180 degrees away from each other.

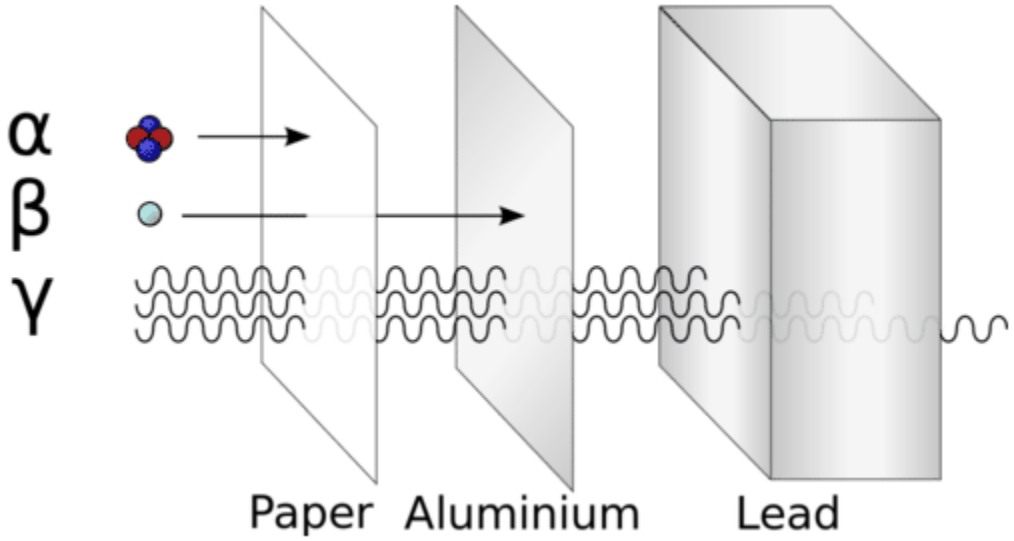


Figure 19.9.1: Penetrating ability of radioactive emissions.

Blocking of alpha particles can easily be accomplished by as little as $10\ \text{mm}$ plastic or paper. Beta emissions represent a somewhat different situation. The negative charge on a beta particle has the potential for activating the element being used to block the radiation. Lead and tungsten are large atoms with many protons and neutrons in their nuclei. While the beta electron may be blocked, the target material could become irradiated in the process.

High-density materials are much more effective protection against gamma emissions than low-density ones. Gamma rays are usually blocked effectively by lead shielding. The thickness of the shielding will determine the effectiveness of the protection offered by the lead.

Summary

- The penetrating ability of a radioactive emission is its ability to go through matter.
- The relative penetrating abilities of radioactive emissions are described.

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19.10: Effects of Radiation

Bacterial contamination in our food often makes the news. There are many bacteria present on raw food, especially raw meat. Campylobacter, salmonella, and other microorganisms can be present, even after cooking, if the meat has not been sufficiently exposed to the heat. Ionizing radiation can be used to disrupt the DNA-RNA-protein synthesis cycle that allows the bacteria to reproduce. Cobalt-60 is a common radiation source, as is cesium-137. But...just to be safe...order that burger well-done!

Effects of Radiation

In order to better understand how cellular radiation damage occurs, we need to take a quick review of how the cell functions. **DNA** in the nucleus is responsible for protein synthesis and for regulation of many cellular functions. In the process of protein synthesis, DNA partially unfolds to produce **messenger RNA** (mRNA). The mRNA leaves the nucleus and interacts with ribosomes, transfer RNA, amino acids, and other cellular constituents in the cytoplasm. Through a complex series of reactions, proteins are produced to carry out a number of specialized processes within the organism. Anything that disturbs this flow of reactions can produce damage to the cell.

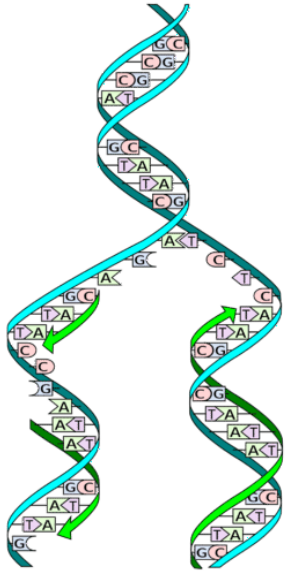


Figure 19.10.1: DNA replication.

The major effect of ionizing radiation on the cell is the disruption of the DNA strand. With the DNA structure damaged, the cell cannot reproduce in its normal fashion. Protein synthesis is affected, as are a number of processes necessary for proper cell function. One common effect is the generation of cancer cells. These cells have an abnormal structure due to the damaged DNA. In addition, they usually grow rapidly since the normal control processes regulating cell growth have been changed by the altered composition of the DNA. Tissue damage is also common in people with severe exposure to radiation.

Effects of Radiation on Humans

We can see two general types of effects when humans are exposed to radiation. Low-level exposure can lead to development of cancer. The regulatory processes regulating cell growth are disrupted, leading to uncontrolled growth of abnormal cells. Acute exposure can produce nausea, weakness, skin burns, and internal tissue damage. Cancer patients receiving radiation therapy experience these symptoms; the radiation is targeted to a specific site in the body so that the damage is primarily to the cancer cells, and the patient is able to recover from the exposure.

Summary

- Anything that affects DNA replication and protein synthesis can damage a cell.
- Effects of ionizing radiation on protein synthesis are listed.
- The impact of ionizing radiation on human health is discussed.

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19.11: Radioisotopes in Medical Diagnosis and Treatment

The molecule picture below is thyroxine, a compound produced by the thyroid gland. This molecule regulates how the body uses energy. In a condition known as hypothyroidism, the thyroid makes less thyroxine than normal. A person with this disease feels tired all the time and often puts on weight. Hypothyroidism is treated with thyroid hormone supplements.

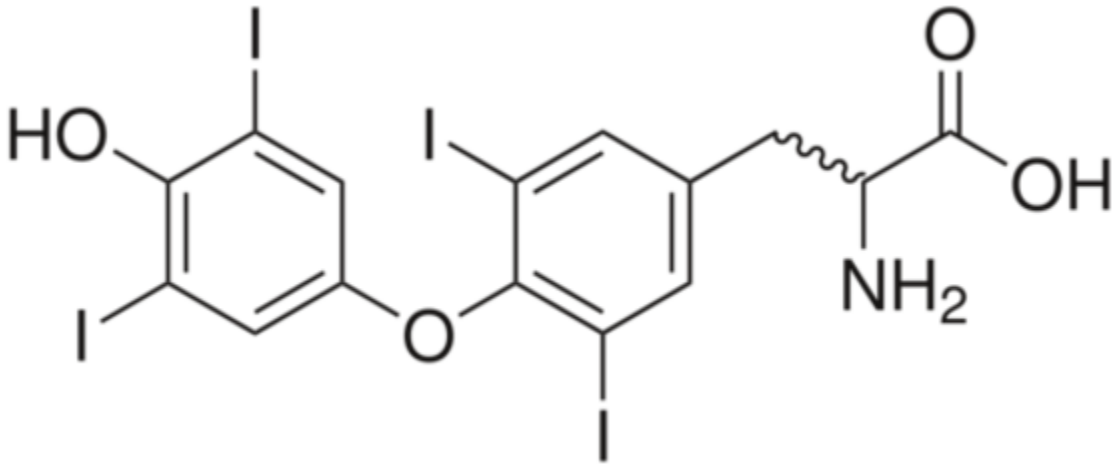


Figure 19.11.1: Thyroxine molecule.

Radioisotopes in Medical Diagnosis and Treatment

Radioisotopes are widely used to diagnose disease and as effective treatment tools. For diagnosis, the isotope is administered, and then located in the body using a scanner of some sort. The decay product (often gamma emission) can be located, and the intensity measured. The amount of isotope taken up by the body can then give information as to the extent of the medical problem.

An isotope of iodine (I-131) is used in both the diagnosis and treatment of thyroid cancer. The thyroid will normally absorb iodine to produce the iodine-containing thyroid hormones. An overactive thyroid gland will absorb the radioactive material, which can then destroy excess thyroid tissue or any cancer of the thyroid. The material is sometimes used to image cancers in other parts of the body.

Technetium-99m is perhaps the most widely used radioisotope in diagnosis and treatment (the "m" stands for *metastable*). This isotope decays to Tc-99 and a gamma emission of low intensity, making the radiation damage fairly negligible. The half-life is about six hours, so it will remain in the body for some time. Tc-99m can be used to look at cardiac damage. The isotope flows in the bloodstream; if there is less blood flow in the heart, there will be less isotope concentrated in the heart muscle. Similar information can be obtained for blood flow in the brain.

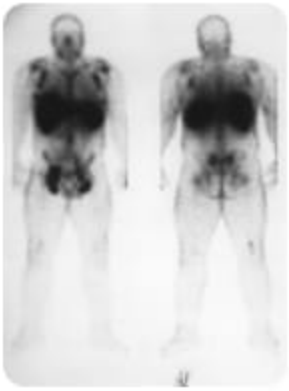


Figure 19.11.2: Radioisotope scan to detect tumors.

Isotopes can be very useful in scans to locate cancer cells. The patient in the above image has multiple tumors that have spread (metastasized) from the main tumor. A radioisotope has been attached to **antibodies** that bind to specific cancer cells. The very dark spots in the armpits, neck, and groin represent areas where tumor cells exist.

Many other examples could be presented. There are presently over 25 different isotopes in use for diagnosis and treatment. A very partial list can be seen in the table below.

Radioisotopes Employed in Diagnosis and/or Treatment

Table 19.11.1: Radioisotopes Employed in Diagnosis and/or Treatment

Isotope	Half-Life	Application
Cr-51	28 days	Label red blood cells.
Fe-59	446 days	Study iron metabolism in spleen.
Xe-133	5 days	Study lung function.
Ho-166	26 hours	Cancer treatment.

Summary

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19.12: PET Scans

Does your heart beat faster when you are scared? Do you have a tender feeling in your heart for that "special person"? Sorry to disappoint you, but that is actually simply a response generated by the brain. Science tells us that the seat of emotion is in the brain. Using PET scans and other techniques, scientists look at specific areas of the brain that process and store information dealing with strong emotions. They haven't localized the love site, but they're working on it.

PET Scans

One of the more interesting and useful approaches to radioisotope use in medicine is **positron emission topography (PET)**, often referred to as a PET scan. This technique is especially useful in studying the processes in the brain. Many compounds do not enter the brain because of what is called the "**blood-brain barrier**", a filtering system to block material from being transported into brain tissue. This mechanism serves to protect the brain from a wide variety of harmful substances.

In order to get a good picture of what is happening in the brain, radiolabels are attached to different compounds that will enter the brain. Since the brain uses about 25% of the glucose found in the body, this molecule is often labeled with a positron emitter such as F-18 (half-life of 109.8 minutes) to study brain function in general. Other labels are attached to specific compounds that will localize in certain areas of the brain to look at specific structures.

The PET scanner detects gamma emissions from the collision of a positron with an electron (see figure below). As the positron is released from the nucleus of the atom, it will collide with an electron. This meeting of matter (electron) with antimatter (positron) results in annihilation of both particles, and the release of two gamma emissions that are 180° apart from one another. The apparatus detects these gamma rays and stores the data in a computer. From this information, a detailed picture of the brain can be developed.



Figure $19.12.1$: Scanner used to detect positron emissions.

One useful application of PET scanning is in the diagnosis of Alzheimer's disease. This debilitating memory loss condition primarily occurs in elderly individuals. A protein known as beta-amyloid gradually forms deposits in the brain called plaque. Memory loss and impaired movement are the result of the plaque growth.

The compound known as Pittsburgh compound B is often used to identify areas of plaque in the brain. The radiolabel is C-11 (half-life of 20.38 minutes).

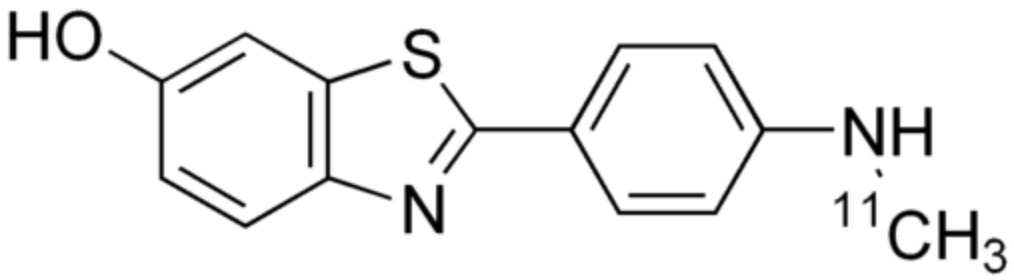


Figure 19.12.2: Pittsburgh compound B used in diagnosis of Alzheimer's disease.

The label attaches to plaque and can be observed using PET scans.

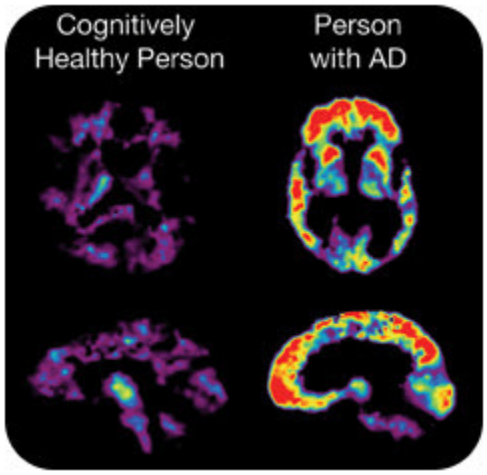


Figure 19.12.3: Brain scans using Pittsburgh compound B to locate plaque.

The computer translates the amount of isotope into a color scale, with red indicating a high level of radioactivity and yellow indicating somewhat less activity. We can see from the scans that the cognitively healthy individual shows the presence of very little plaque in the brain (see figure above). The individual with Alzheimer's demonstrates high concentrations of beta-amyloid in numerous areas of the brain.

Other studies have been done looking at brain function in drug addicts. One of the theories about drug addiction involves the amount of dopamine action in the brain (a chemical that is a part of the system to transport nerve impulses). Studies of dopamine action have been helpful in understanding addictive processes.

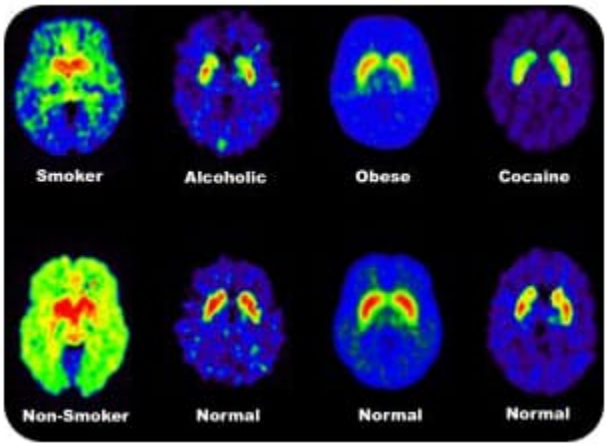


Figure 19.12.4: PET scan of dopamine binding in brains of normal and addicted individuals.

The figure above shows the binding of chemicals that attach to dopamine receptors. The non-addicted individuals have large numbers of receptors for dopamine. The addicted persons show less binding to these receptors, indicating that fewer receptors

are present. Since dopamine is somehow linked with the sense of pleasure, these data may help to bring a better understanding to the biochemical processes in drug addiction.

Summary

- Positron Emission Topography (PET) scans make use of radioisotopes in medicine.
- As an example, PET scans can be used to detect Alzheimer's Disease and to help scientists learn about addiction.

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