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LBS Science

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Contents

CHAPTER 1

Introduction to Chemistry

Chapter Outline

1.1 What Is Chemistry?

Lesson Objectives

The student will:

- give a brief history of how chemistry began.
- list some new materials produced by chemists.

Vocabulary

- chemistry

Introduction

During medieval times, a group of people known as alchemists began looking for ways to transform common metals, such as lead, copper, and iron, into gold. Can you imagine how much money you would make if you could go to the store, buy some iron nails, and turn them into gold? You'd be rich in no time!

The Origin and Evolution of Chemistry

Alchemists experimented with many different kinds of chemicals, searching for what they termed the "philosopher's stone" – a legendary substance necessary for transforming common metals into gold (see **Figure** 1.1). We now know that there is no such thing as a philosopher's stone, nor is there any chemical reaction that creates gold from another metal. We know this because we now have a much better understanding of how the matter in our universe behaves. Nevertheless, those early alchemists kindled interest in chemical transformations and inspired the development of modern chemistry.

Chemistry is the scientific study of matter and the changes that it undergoes. It's no coincidence that the word "chemistry" looks a lot like the word "alchemy." Early alchemists were commonly known as "chemists," and over time, people started referring to their work, particularly the more legitimate forms of it, as chemistry. While it may seem strange, in many ways it is appropriate that our word for the present-day study of matter comes from the early practice of alchemy. A lot of the techniques and equipment fundamental to modern chemistry were actually developed by early alchemists.

Alchemy was practiced in China and India and later in the Arab world and Europe. The origin of the word "alchemy," however, is something of a mystery. Certainly, early Europeans derived the term "alchemy" from the Arabic word "al-kimia," meaning "the art of transformation" (of course, the transformation that alchemists were primarily concerned with involved the creation of gold). Most of what we know today about early alchemy is based

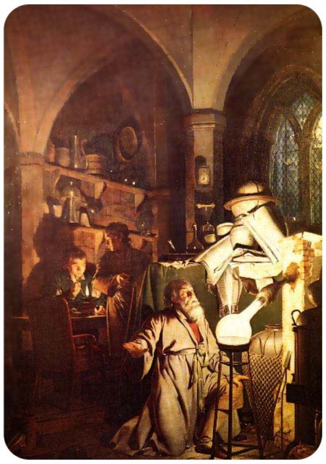


FIGURE 1.1

This painting shows an alchemist in search of the philosopher's stone.

on translations of Arabic documents, as the Muslim alchemists were among the first to keep careful notes about their experiments.

Even though our earliest records of alchemy come from the Arab Empire, some scholars believe that Arabs adopted alchemy and the word "al-kimia" from the Greeks around 650 AD. The Greeks, in turn, may have learned of alchemy from the Egyptians. Khem was an ancient name for Egypt, and the Egyptians of early history were known as masters at the art of working with gold. It's very likely that al-kimia is actually a distorted version of the word "al-kimiya," meaning "the art of the land of Khem," or the art of Egypt.

Medieval Europeans were similarly fascinated by alchemy. Unfortunately, many alchemists in Europe borrowed ideas from the more mystical of the Arabian alchemists, and, as a result, European alchemy quickly became associated with wizardry, magic, and the search for the philosopher's stone. It wasn't until the late 17th century that European chemists began applying the scientific method (see the lesson "The Process of Science" for a more detailed discussion about the scientific method). Robert Boyle (1627 – 1691) was the first European chemist to do so, using quantitative experiments to measure the relationship between the pressure and the volume of a gas. His use of scientific methods paved the way for other European chemists and helped to establish the modern science of chemistry.

The man who would greatly advance the development of modern chemistry was Antoine Lavoisier (1743 - 1794). Considered the father of modern chemistry, Lavoisier (seen in **Figure** 1.2) discovered that although matter may change its shape or form, its mass always remains the same. As a result, he would state the first version of the law of conservation of mass. Lavoisier also wrote the first extensive list of elements, including oxygen and hydrogen, and helped to reform chemical nomenclature.

What Chemists Do

You might wonder why the study of chemistry is so important if you can't use it to turn iron into gold or to develop a potion that will make you immortal. Why didn't chemistry die when scientists like Boyle and Lavoisier proved alchemy was nothing but a hoax? Although we can't use chemistry to make gold or to live forever, modern chemistry is still very powerful. There may be no such thing as a potion that cures all diseases, but many chemists today are working to develop cures for specific diseases, including HIV/AIDS and various forms of cancer.

Chemists apply information about matter and the changes it undergoes to improve our lives in many different ways. Modern chemists study not only chemicals that can help us, but also chemicals that can hurt us. For example,

FIGURE 1.2

Antoine Lavoisier is considered to be the father of modern chemistry due to his many contributions to chemistry.

environmental chemists test the air, soil, and water in our neighborhoods to make sure that we aren't exposed to heavy metals (such as mercury or lead) or chemical pesticides. Moreover, when environmental chemists do find dangerous substances, they use their knowledge of chemistry to clean up the contamination. Similarly, every time you buy packaged food from the grocery store, you can be sure that many tests have been done by chemists to ensure those foods don't contain any toxins or carcinogens (cancer-causing chemicals).

Chemists are also responsible for creating many important materials that we use today. In addition, many technologies rely on chemistry as well. In fact, flat-screen LCD televisions, cubic zirconium rings, and energy-efficient LED lights are all thanks to our improved understanding of chemistry.

One of the huge breakthroughs in recent history has been the discovery of plastics (see **Figure** 1.3). Initially, plastic was made by chemically modifying cellulose, a naturally occurring chemical found in plants. As chemical knowledge developed, however, scientists began to realize that plastics had special properties. On a microscopic scale, plastics are composed of thousands of tiny chains of molecules all tangled up together. Scientists reasoned that if they altered the chemicals in these chains but still managed to keep the chains intact, they could make new plastics with new properties. Thus began the plastic revolution!

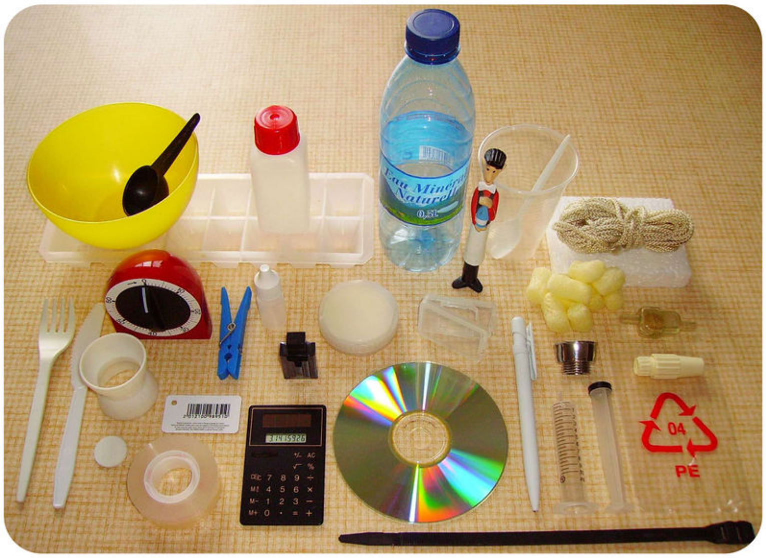


FIGURE 1.3

Some common objects made of plastic.

Semiconductors are another class of "new" materials whose development is largely based on our improved understanding of chemistry. Because scientists know how matter is put together, they can predict how to fine-tune the chemical composition of a semiconductor in order to make it absorb light and act as a solar cell, or to emit light and act as a light source.

We've come a long way from our early days of producing bronze and steel. Nevertheless, as our understanding of chemistry improves, we will be able to create even more useful materials than the ones we have today.

Lesson Summary

- Chemistry began as the study of alchemy. Most alchemists were searching for the philosopher's stone, a fabled substance that could turn common metals into gold.
- Chemistry is the scientific study of matter and the changes that it undergoes.
- The word "chemistry" comes from the Arabic word "al-kimia," meaning "the art of transformation."
- Chemists apply information about matter and the changes it undergoes in many different ways to improve our lives.

Further Reading / Supplemental Links

The learner.org

- http://learner.org/resources/series61.html

Review Questions

1. What was the origin of the word "chemistry"?
2. Name at least two new materials created by chemists in the last 40 years.

1.2 The Process of Science

Lesson Objectives

The student will:

- describe the scientific method of problem solving.
- list some values of the scientific method of problem solving.
- describe the difference between hypothesis, theory, and scientific law as scientific terms.
- explain the necessity for experimental controls.
- identify components in an experiment that represent experimental controls.
- explain the concept of a model and why scientists use models.
- explain the limitations of models as scientific representations of reality.

Vocabulary

- controlled experiment
- experiment
- hypothesis
- model
- problem
- scientific law
- scientific method
- theory

Introduction

Socrates (469 BC - 399 BC), Plato (427 BC - 347 BC), and Aristotle (384 BC - 322 BC) are among the most famous of the Greek philosophers. Plato was a student of Socrates, and Aristotle was a student of Plato. These three were probably the greatest thinkers of their time. Aristotle's views on physical science were, in particular, highly influential and widely accepted until well into the 1300s.

Unfortunately, many of Aristotle's opinions were wrong. Aristotle was a brilliant man without doubt, but he was using a method unsuitable for determining the nature of the physical world. The philosopher's method depended on logical thinking and not on observing the natural world. This led to many errors in Aristotle's views on nature. Let's consider two of Aristotle's ideas as examples.

In Aristotle's opinion, men were bigger and stronger than women, so it was logical to him that men would have more teeth than women do. Thus, Aristotle concluded this as a fact without actually counting the number of teeth in any mouths. Had he done so, he would have found that men and women have exactly the same number of teeth. As another example, Aristotle considered what would happen if he were to drop two balls identical in all ways but mass. In his mind, it was clear that the heavier ball would fall faster than a lighter one would, and he concluded that this must be a law of nature. Once again, he did not consider doing an experiment to see which ball would fall faster. This conclusion, however, was also incorrect. Eighteen centuries later, Galileo tried this experiment by dropping two balls of different masses off a building (the Leaning Tower of Pisa, according to legend). Galileo discovered, by experimental observation, that the two balls hit the ground at exactly the same time. Aristotle's logical conclusion was again wrong.

The Scientific Method of Problem Solving

In the 16th and 17th centuries, innovative thinkers were developing a new way to understand the nature of the world around them. They were developing a method that relied upon making observations of phenomena and drawing conclusions that corresponded to their observations.

The **scientific method** is a method of investigation involving experimentation and observation to acquire new knowledge, solve problems, and answer questions. Scientists frequently list the scientific method as a series of steps. Some scientists oppose this listing of steps because not all steps occur in every case, nor do they always occur in the same order. The scientific method is listed here as a series of steps, but you should remember that you are not required to rigidly follow this list. Instead, the scientific method is a valuable tool that provides a basic and adaptable strategy for tackling scientific questions.

The Steps in the Scientific Method

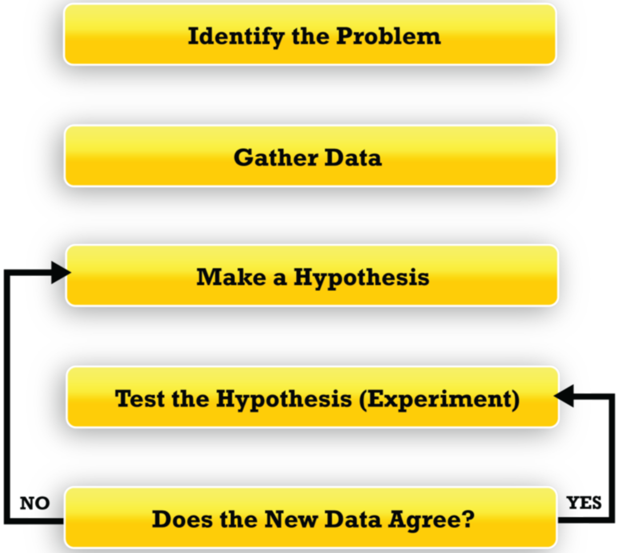
Step 1: Identify the **problem** or phenomenon that needs to be investigated. This is sometimes referred to as "defining the problem."

Step 2: Gather and organize data on the problem. This step is also known as "making observations."

Step 3: Suggest a possible solution or explanation. A suggested solution is called a **hypothesis**.

Step 4: Test the hypothesis by making new observations.

Step 5: If the new observations support the hypothesis, you accept the hypothesis for further testing. If the new observations do not agree with your hypothesis, add the new observations to your observation list and return to Step 3.



Suppose you are required to maintain a large campfire, but you are completely unfamiliar with the property that makes objects combustible (able to burn). The first step in the scientific method is to define the problem. The problem statement for this investigation is: What property makes objects combustible?

The next step is to gather data on the problem. At the beginning, you may be collecting objects at random to put into the fire. It is important to keep good records of what objects were tried and whether or not they burned. A list of organized data (observations) is shown in **Table** 1.1.

TABLE 1.1: **List of Items Tried in the Fire**

Will Burn	**Won't Burn**
tree limbs	rocks
chair legs	bricks
pencils	marbles
baseball bat	hubcaps

The list of organized observations helps because now you can focus on only collecting items on the "will burn" list and not waste any effort by dragging items that won't burn to the fire. However, you would soon use up all the items on the "will burn" list, making it necessary to guess what property the "will burn" objects have that allow them to burn. If you had that answer, you could keep the fire going by bringing objects that may not be on the "will burn" list but have the "will burn" property.

The third step in the scientific method is to suggest a hypothesis. A hypothesis is a tentative explanation that can be tested by further investigation. Your guess about what property makes the "will burn" objects combustible is a hypothesis. Suppose you notice that all the items on the "will burn" list are cylindrical in shape, so you hypothesize that "cylindrical objects burn."

The fourth step in the scientific method is to test your hypothesis. To test the hypothesis that cylindrical objects burn, you go out and collect a group of objects that are cylindrical, including iron pipes, soda bottles, broom handles, and tin cans. When these cylindrical objects are placed in the fire and most of them do not burn, you realize your hypothesis is not supported by these new observations. The new observations are the test, but your hypothesis has failed the test.

When the new observations fail to support your hypothesis, you reject your original hypothesis, add your new data to the table, and make a new hypothesis based on the updated observations list. A new updated table is seen in **Table** 1.2.

TABLE 1.2: **List of Items Tried in the Fire**

Will Burn	**Won't Burn**
tree limbs	rocks
chair legs	bricks

TABLE 1.2: (continued)

Will Burn	Won't Burn
pencils	marbles
baseball bat	hubcaps
broom handle	iron pipes
	soda bottles
	tin cans

According to the schematic diagram of the scientific method, if the new data does not support the hypothesis, the scientist returns to Step 3 and makes a new hypothesis. When the hypothesis is supported by the the results of several experiments, you might think that the work is finished. For a hypothesis to be useful, however, it must withstand repeated testing. Other scientists must be able to repeat the experiments using the same materials and conditions and get the same results. Scientists submit reports of research to other scientists, usually by publishing an article in a scientific journal, so that the work can be verified.

Scientific Hypotheses, Theories, and Laws

Hypotheses that have passed many supportive tests are often called theories. A **theory** is an explanation that summarizes a hypothesis or a set of hypotheses and has been supported with repeated testing. Theories have a great deal more supportive testing behind them than do hypotheses. Suppose your new hypothesis is "wooden objects burn." You will find this hypothesis more satisfactory since all of the wooden objects you try will burn. You can see from this example that the hypothesis does not become what we think of as a "fact" but rather a tentatively accepted theory, which must undergo continuous testing and possible adjustments. Even if your theory seems successful, you might be ignoring other types of combustible materials, such as a large stack of old car tires, objects made of fabric or paper, or perhaps containers of petroleum. You can see that even though you are quite satisfied with your theory because it does the job you want it to do, you actually do not have a complete statement on all the properties that make objects burn.

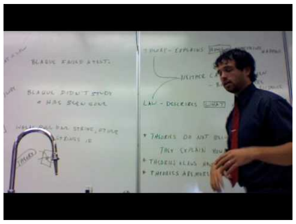
In science, theories can either be descriptive (qualitative) or mathematical (quantitative). However, a scientific theory must be falsifiable, or capable of being proved false, in order to be accepted as a theory. A theory is never proven true and is never a "fact." As long as a theory is consistent with all observations, scientists will continue to use it. When a theory is contradicted by observations, it is discarded and replaced.

A theory is also a possible explanation for a law. A **scientific law** is a statement that summarizes the results of many observations and experiments. A law describes an observed pattern in data that occurs without any known exception. A law that has withstood the test of time is incorporated into the field of knowledge. Because they explain the patterns described in laws, theories can be used to predict future events.

One example of a scientific law was discovered around the 1800s by a group of scientists who were working with gases to, among other reasons, improve the design of the hot air balloon. After many tests, Jacques Charles and other scientists discovered that patterns and regularities existed in their observations of gas behavior. They found that if the temperature of the gas increased, the volume of the gas also increased. This relationship has held true over time and is now a scientific law. Any scientific theory that describes gas behavior would need to reflect this law and predict that the volume of a gas increases whenever the temperature increases.

Around the same time, another scientist named J. W. Henry was trying to find a pattern between the pressure of a gas and the amount of the gas dissolved in water. Henry found that when one of these variables increased, the other variable increased in the same proportion. If you have ever gone scuba diving, you may already be familiar with this observation. During training, scuba divers learn about a problem known as "the bends." As scuba divers dive deeper, the increased pressure of the air they breathe causes more nitrogen to be dissolved in the diver's blood. Coming up too quickly from a dive causes the pressure to decrease rapidly and the nitrogen to leave the blood quickly, which could lead to joint pains known as "the bends." Henry's Law is a scientific law because it indicates a repeatedly observed relationship (regularity) between gas pressure and the amount of dissolved gas.

In this video a teacher discusses the difference between a theory and a law **(1f - IE Stand.)**: http://www.youtube.com/watch?v=eDED5fCY86s (4:18).



MEDIA

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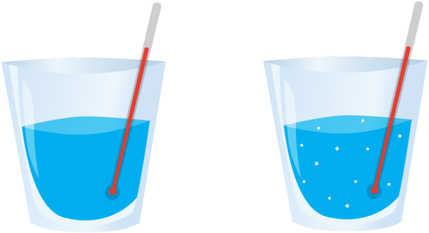
Experimentation

The scientific method requires that observations be made. Sometimes the phenomenon we wish to observe does not occur in nature or is inconvenient for us to observe. If we can arrange for the phenomenon to occur at our convenience, we can control other variables and have all of our measuring instruments on hand to help us make observations. An **experiment** is a controlled method of testing a hypothesis under the conditions we want at a time and place of our choosing. When scientists conduct experiments, they are usually seeking new information or trying to verify someone else's data. In comparison, classroom experiments often demonstrate and verify information that is already known to scientists but may be new to students.

Suppose a scientist observed two pools of water in bowl-shaped rocks that are located near each other while walking along the beach on a very cold day following a rainstorm. One of the pools was partially covered with ice, while the other pool had no ice on it. The unfrozen pool seemed to contain seawater that splashed up on the rock from the surf, but the other pool was too high up for seawater to splash in and was most likely filled with only rainwater.

Since both pools were at the same temperature, the scientist wondered why one pool was partially frozen and the other was not. By tasting the water (not a good idea), the scientist determined that the unfrozen pool tasted saltier than the partially frozen one. The scientist thought perhaps salt water had a lower freezing point than fresh water, so she decided to go home to test her hypothesis. In order to test this hypothesis, the scientist will conduct an experiment during which she can make accurate observations. So far, the scientist has identified a question, gathered a small amount of data, and suggested a hypothesis.

For the experiment, the scientist prepared two identical containers of fresh water and added some salt to one of them. A thermometer was placed in each container, and both containers were placed in a freezer. The scientist then observed the conditions and temperatures of the two liquids at regular intervals (see the tables below).



Temperature and Conditions of Fresh Water

Temperature and Conditions of Salt Water

time (**min**)	**temperature** (°**C**)	**condition**	**time** (**min**)	**temperature** (°**C**)	**condition**
0	25	liquid	0	25	liquid
5	20	liquid	5	20	liquid
10	15	liquid	10	15	liquid
15	10	liquid	15	10	liquid
20	5	liquid	20	5	liquid
25	0	frozen	25	0	liquid
30	-5	frozen	30	-5	frozen

The scientist found support for her hypothesis from this experiment: fresh water freezes at a higher temperature than salt water. Much more support would be needed before the scientist is confident in this hypothesis. She would perhaps ask other scientists to verify the work.

In the scientist's experiment, it was necessary that she freeze the salt water and fresh water under exactly the same conditions. Why? The scientist was testing whether or not the presence of salt in water would alter its freezing point. It is known that even changing the air pressure will alter the freezing point of water. In order to conclude that the presence of the salt was what caused the change in freezing point, all other conditions had to be identical. When doing an experiment, it is important to set up the experiment so that relationships can be seen clearly. A **controlled experiment** is one that compares the results of an experimental sample to a control sample. The control sample is identical to the experimental sample in all ways except for the one variable being tested. The fresh water sample is the control sample, while the sample containing salt is the experimental sample. The presence of salt is the only thing allowed to change in the two samples and is the effect being tested. In an experiment, there may be only one variable, and the purpose of the control is to guarantee that there is only one variable. Unless experiments are controlled, the results are not valid.

Suppose you wish to determine which brand of microwave popcorn leaves the fewest unpopped kernels. You will need a supply of various brands of microwave popcorn to test, and you will need a microwave oven. If you used different brands of microwave ovens with different brands of popcorn, the percentage of unpopped kernels could be caused by the different brands of popcorn or by the different brands of ovens. Under such circumstances, the experimenter would be unable to conclude confidently whether the popcorn or the oven caused the difference. To eliminate this problem, you must use the same microwave oven for every test. In order to reasonably conclude that the change in one variable was caused by the change in another specific variable, there must be no other variables in the experiment. By using the same microwave oven, you control the number of variables in the experiment.

This video presents both how errors can occur and how to avoid them when conducting experiments in the laboratory **(1b, 1c, 1j IE Stand.)**: http://www.youtube.com/watch?v=RU1DZeqY4To (3:57).

MEDIA

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Scientific Models

Chemists rely on both careful observation and well-known physical laws. By putting observations and laws together, chemists develop models. A **model** is a descriptive, graphic, or three-dimensional representation of a hypothesis or theory used to help enhance understanding. Scientists often use models when they need a way to communicate their understanding of what might be very small (such as an atom or molecule) or very large (such as the universe).

A model is any simulation, substitute, or stand-in for what you are actually studying and provide a way of predicting what will happen given a certain set of circumstances. A good model contains the essential variables that you are concerned with in the real system, explains all the observations on the real system, and is as simple as possible. A model may be as uncomplicated as a sphere representing the earth or billiard balls representing gaseous molecules, but it may also be as complex as mathematical equations representing light.

If you were asked to determine the contents of a box that cannot be opened, you could do a variety of experiments in order to develop an idea (or a model) of what the box contains. You would probably shake the box, perhaps put magnets near it, and possibly determine its mass. When you completed your experiments, you would develop an idea of what is inside; that is, you would propose a model of what is inside the box that cannot be opened. With your model, you could predict how the unopened box would behave under a different set of conditions.

However, even though your model may be capable of accurately predicting some behavior of the unopened box, you would find that the model does not always agree with new experimental results and observations. The model is only be as good as the data you have collected. Because you would never be able to open the box to see what is inside, you also would never be able to create a perfectly accurate model of the box. The model can only be modified and refined with further experimentation.

Chemists have created models about what happens when different chemicals are mixed together, heated up, cooled down, or compressed by using many observations from past experiments. They use these models to predict what might happen during future experiments. Once chemists have models that predict the outcome of experiments reasonably well, those working models can be applied for practical purposes, such as producing an especially strong plastic or detecting potential toxins in your food.

A good example of how a model is useful to scientists is to examine how models were used to develop the atomic theory. As you will learn in the chapter "The Atomic Theory," the concept of an atom has changed over many years. In order to understand the different theories of atomic structure proposed by various scientists, models were drawn to make the concepts easier to understand.

Lesson Summary

- The scientific method is a method of investigation involving experimentation and observation to acquire new knowledge, solve problems, and answer questions.
- The steps in the scientific method are:
    1. Identify the problem.
    2. Gather data (make observations).
    3. Suggest a hypothesis.
    4. Test the hypothesis (experiment).
    5. Accept the hypothesis for further testing, or reject the hypothesis and make a new one.
- A hypothesis is a tentative explanation that can be tested by further investigation.
- A theory is an explanation that summarizes a hypothesis or a set of hypotheses and has been supported with repeated testing.
- A scientific law is a statement that summarizes the results of many observations and experiments.
- An experiment is a controlled method of testing a hypothesis.
- A controlled experiment is one that compares the results of an experimental sample to a control sample.
- The control sample is identical to the experimental sample in all ways except for the one variable whose effect is being tested.
- A model is a descriptive, graphic, or three-dimensional representation of a hypothesis or theory used to help enhance understanding.
- Scientists often use models when they need a way to communicate their understanding of what might be very small (such as an atom or molecule) or very large (such as the universe).

Further Reading / Supplemental Links

The learner.org

- http://learner.org/resources/series61.html

This website has a video that explores the history of the scientific method.

- http://videos.howstuffworks.com/hsw/5881-scientific-method-history-video.htm

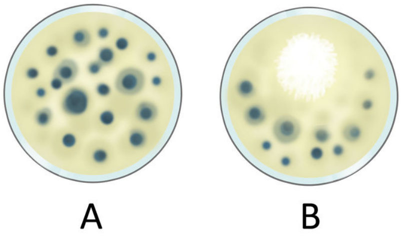
This video is a ChemStudy film called "High Temperature Research." The film is somewhat dated but the information is accurate.

- http://www.youtube.com/watch?v=Tt2JEwbOtq8

Review Questions

Use the following paragraph to answer questions 1 and 2.

In 1928, Sir Alexander Fleming was studying Staphylococcus*Penicillium* was also growing in some of the dishes. As shown in the illustration below, Petri dish A represents a dish containing only *Staphylococcus* bacteria. In dishes containing the *Penicillium* mold, represented by Petri dish B below, Fleming noticed that a clear area existed around the mold because all the bacteria in this area had died. In the culture dishes without the mold, no clear areas were present. Fleming suggested that the mold was producing a chemical that killed the bacteria. He decided to isolate this substance and test it to see if it would kill bacteria. Fleming grew some *Penicillium* mold in a nutrient broth. After the mold grew in the broth, he removed all the mold from the broth and added the broth to a culture of bacteria. All the bacteria died.



1. Which of the following statements is a reasonable expression of Fleming's hypothesis?
    a. Nutrient broth kills bacteria.
    b. There are clear areas around the Penicillium*Staphylococcus* doesn't grow.
    c. Mold kills bacteria.
    d. Penicillium*Staphylococcus*.
    e. Without mold in the culture dish, there were no clear areas in the bacteria.
2. Fleming grew Penicillium*Penicillium*, and poured the broth into culture dishes containing bacteria to see if the broth would kill the bacteria. What step in the scientific method does this represent?
    a. Collecting and organizing data

   b. Making a hypothesis
   c. Testing a hypothesis by experiment
   d. Rejecting the old hypothesis and making a new one
   e. None of these

3. A scientific investigation is not
   a. True
   b. False
4. Which of the following is closest in meaning to the word "hypothesis"?
   a. Fact
   b. Law
   c. Formula
   d. Suggested explanation
   e. Conclusion
5. Why do scientists sometimes discard theories?
   a. The steps in the scientific method were not followed in order.
   b. Public opinion disagrees with the theory.
   c. The theory is opposed by the church.
   d. Contradictory observations are found.
   e. Congress voted against it.

Use the following paragraph to answer questions 6 through 10.

Gary noticed that two plants of the same type were different in size after three weeks, even though they were initially the same size when his mother planted them on the same day. Since the larger plant was in the full sun all day and the smaller plant was in the shade of a tree for most of the day, Gary believed that the sunshine was responsible for the difference in plant size. In order to test this, Gary bought ten small plants of the same size and type. He also made sure they have the same amount and type of soil. Gary then built a frame to hold a canvas roof over five of the plants, while the other five were nearby but out in the sun. Gary was careful to make sure that each plant received exactly the same amount of water and plant food every day.

6. Which of the following is a reasonable statement of Gary's hypothesis?
   a. Different plants have different characteristics.
   b. Plants that get more sunshine grow larger than plants that get less sunshine.
   c. Plants that grow in the shade grow larger.
   d. Plants that don't receive water will die.
   e. Plants that receive the same amount of water and plant food will grow the same amount.
7. What scientific reason might Gary have for insisting that the container size for the all plants be the same?
   a. Gary wanted to determine if the size of the container would affect the plant growth.
   b. Gary wanted to make sure the size of the container did not affect differential plant growth in his experiment.
   c. Gary wanted to control how much plant food his plants received.
   d. Gary wanted his garden to look organized.
   e. There is no possible scientific reason for having the same-sized containers.
8. What scientific reason might Gary have for insisting that all plants receive the same amount of water every day?
   a. Gary wanted to test the effect of shade on plant growth, and therefore he wanted to have no variables other than the amount of sunshine on the plants.
   b. Gary wanted to test the effect of the amount of water on plant growth.

c. Gary's hypothesis was that water quality was affecting plant growth.
d. Gary was conserving water.
e. There is no possible scientific reason for having the same amount of water for each plant every day.

9. What was the variable being tested in Gary's experiment?

    a. The amount of water
    b. The amount of plant food
    c. The amount of soil
    d. The amount of sunshine
    e. The type of soil

10. Which of the following factors did Gary not control in his experimental setup that may be varying?

    a. Individual plant variation
    b. Soil temperature due to the different colors of the containers
    c. Water loss due to evaporation from the soil
    d. The effect of insects, which may attack one set of plants but not the other
    e. All of the above are possible factors that Gary did not control

11. When a mosquito sucks blood from its host, it penetrates the skin with its sharp beak and injects an anti-coagulant so that the blood will not clot. It then sucks some blood and removes its beak. If the mosquito carries disease-causing microorganisms, it injects these into its host along with the anti-coagulant. It was assumed for a long time that the typhus virus was injected by the louse (singular for lice) when sucking blood in a manner similar to the mosquito. This turned out not to be true. The infection is not in the saliva of the louse but in the feces. The disease is thought to be spread when louse feces come in contact with scratches or bite wounds on the host's skin. A test of this was carried out in 1922 when two workers fed infected lice on a monkey, taking great care that no louse feces came into contact with the monkey. After two weeks, the monkey had not

    a. To prove that the lice carried the typhus virus
    b. To prove the monkey was similar to man
    c. To prove that the monkey was not immune to typhus
    d. To prove that mosquitoes were not carriers of typhus
    e. To demonstrate that the workers were mean

12. When a theory has been known for a long time, it becomes a law.

    a. True
    b. False

13. During Pasteur's time, anthrax was a widespread and disastrous disease for livestock. Many people whose livelihood was raising livestock lost large portions of their herds to this disease. Around 1876, a horse doctor in eastern France named Louvrier claimed to have invented a cure for anthrax. The influential men of the community supported Louvrier's claim of having cured hundreds of cows of anthrax. Pasteur went to Louvrier's hometown to evaluate the cure. The cure was explained to Pasteur as a multi-step process during which: 1) the cow was rubbed vigorously to make her as hot as possible; 2) long gashes were cut into the cows skin; 3) turpentine was poured into the cuts; 4) an inch-thick coating of cow manure mixed with hot vinegar was plastered onto the cow; and 5) the cow was completely wrapped in a cloth. Since some cows recover from anthrax with no treatment, performing the cure on a single cow would not be conclusive, so Pasteur proposed an experiment to test Louvrier's cure. Four healthy cows were to be injected with anthrax microbes. After the cows became ill, Louvrier would pick two of the cows (A and B) and perform his cure on them, while the other two cows (C and D) would be left untreated. The experiment was performed, and after a few days, one of the untreated cows died and the other got better. Of the cows treated by Louvrier's cure, one cow died and the other got better. In this experiment, what was the purpose of infecting cows C and D?

    a. To give Louvrier more than two cows to choose from
    b. To make sure the injection actually contained anthrax

  c. To serve as experimental controls (a comparison of treated to untreated cows)
  d. To kill as many cows as possible

14. A hypothesis is:
    a. a description of a consistent pattern in observations.
    b. an observation that remains constant.
    c. a theory that has been proven.
    d. a tentative explanation for a phenomenon.
15. A scientific law is:
    a. a description of a consistent pattern in observations.
    b. an observation that remains constant.
    c. a theory that has been proven.
    d. a tentative explanation for a phenomenon.
16. A number of people became ill after eating oysters in a restaurant. Which of the following statements is a hypothesis about this occurrence?
    a. Everyone who ate oysters got sick.
    b. People got sick whether the oysters they ate were raw or cooked.
    c. Symptoms included nausea and dizziness.
    d. The cook felt really bad about it.
    e. Bacteria in the oysters may have caused the illness.
17. Which statement best describes the reason for using experimental controls?
    a. Experimental controls eliminate the need for large sample sizes.
    b. Experimental controls eliminate the need for statistical tests.
    c. Experimental controls reduce the number of measurements needed.
    d. Experimental controls allow comparison between groups that are different in only one variable.
18. A student decides to set up an experiment to determine the relationship between the growth rate of plants and the presence of detergent in the soil. He sets up ten seed pots. In five of the seed pots, he mixes a precise amount of detergent with the soil. The other five seed pots have no detergent in the soil. The five seed pots with detergent are placed in the sun, and the five seed pots with no detergent are placed in the shade. All ten seed pots receive the same amount of water as well as the same number and type of seeds. He grows the plants for two months and charts the growth every two days. What is wrong with his experiment?
    a. The student has too few pots.
    b. The student has two variables different between the groups.
    c. The student did not add detergent to all ten pots.
    d. The student has no experimental control on the soil.
19. A scientist plants two rows of corn for experimentation. She puts fertilizer on row 1 but does not put fertilizer on row 2. Both rows receive the same amount of sun and water. She checks the growth of the corn over the course of five months. What is acting as the control in this experiment?
    a. The corn without fertilizer
    b. The corn with fertilizer
    c. The amount of water
    d. The height of the corn plants
20. If you have a control group for your experiment, which of the following is true?
    a. There can be more than one difference between the control group and the test group but no more than three differences, or else the experiment is invalid.
    b. The control group and the test group may have many differences between them.
    c. The control group must be identical to the test group except for one variable.
    d. None of these are true.

21. If the hypothesis is rejected by the experiment, then:
    a. the experiment may have been a success.
    b. the experiment was a failure.
    c. the experiment must be poorly designed.
    d. the experiment didn't follow the scientific method.
22. A well-substantiated explanation of an aspect of the natural world is a:
    a. theory.
    b. law.
    c. hypothesis.
    d. none of the above.

1.3 References

1. Joseph Wright of Derby. The Alchymist, In Searchof the Philosopher's Stone. Public domain
2. Line engraving by Louis Delaistre, after a design by Juien Boilly. Antoine Lavoisier, The Father ofModern Chemistry. Public domain
3. Cjp24. Somecommon objects madeof plastic. Public domain

CHAPTER 2

Matter and Energy

Chapter Outline

2.1 What is Matter?

Lesson Objectives

The student will:

- define matter and explain how it is composed of building blocks known as atoms.
- explain the differences between substances and mixtures.
- classify mixtures as homogeneous or heterogeneous.
- identify the chemical symbols of common elements.
- explain the difference between an element and a compound by their symbols or formulas.
- demonstrate the proper use of parentheses and subscripts in writing chemical formulas.
- determine the number of atoms and name of each element in a compound.

Vocabulary

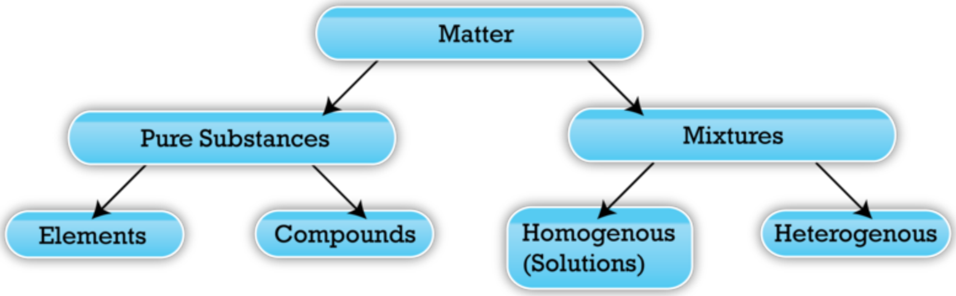
- atom
- compound
- element
- heterogeneous mixture
- homogeneous mixture
- law of constant composition
- matter
- molecule

Introduction

Matter is anything that has mass and volume. The entire universe is composed of matter, which is in turn composed of atoms. An **atom** is the basic building block of all matter. All matter in the universe, from a teaspoon of salt to the Pacific Ocean, has mass and occupies space. The salt and ocean, however, have very different properties and behaviors. Since everything in the universe is composed of matter, there are clearly many types of matter. In this lesson, you will learn about how scientists classify the different types of matter.

Categories of Matter

Matter can be classified into two broad categories: mixtures and pure substances, as illustrated below.



Mixtures are physical combinations of two or more substances. The term "physical combination" refers to mixing together different substances that do not chemically react with each other. The physical appearance of the substances may change, but the atoms in the substances do not.

In comparison, a pure substance is a form of matter that has a constant composition and constant properties throughout the sample. Elements and compounds are both example of pure substances.

Mixtures: Homogeneous and Heterogeneous

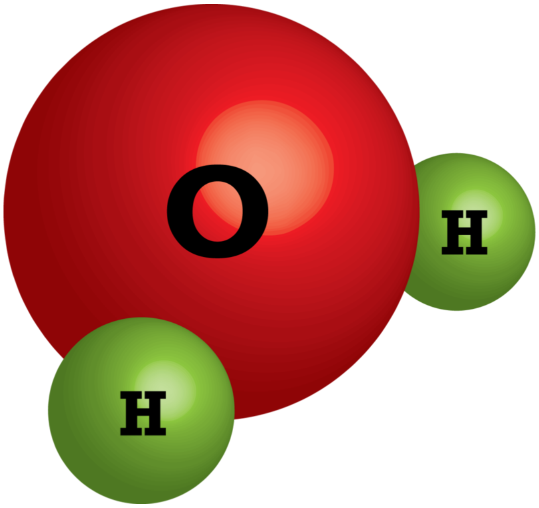
One example of a mixture is sand and gravel stirred together. In this case, you can see that there are two different substances present, each with the same properties that it had before it was mixed. When substances do not mix thoroughly and evenly (like sand and gravel), the mixture is said to be heterogeneous. A **heterogeneous mixture** consists of visibly different substances.

Another example of a mixture is salt dissolved in water. In this case, you cannot see the different substances, but you can test the solution to show that each substance (salt and water) has the same chemical properties it had before being mixed. When substances mix thoroughly and evenly (like salt in water), the mixture is said to be homogeneous. **Homogeneous mixtures** are often referred to as solutions. Solutions often may appear to be one pure substance, but some simple tests can show that the solutions are indeed mixtures.

Pure Substances: Elements and Compounds

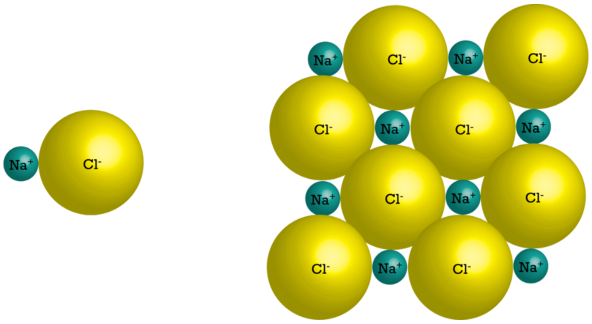
Elements are the simplest substances. An **element** is a substance that is made up of only one type of atom. It doesn't matter if the atoms are in groups, as in P_4 or S_8, or isolated, as in Na. As long as there is only one kind of atom, the substance is an element. Elements cannot be chemically broken down into anything smaller and still retain the properties of the element. For example, an atom of iron can be smashed into electrons, protons, and neutrons, but those pieces would not have the properties of iron.

Atoms from two or more elements can chemically combine to form a new substance. **Compounds** are substances that are made up of more than one type of atom. In other words, compounds are chemical combinations of elements. These combinations form new substances with completely different properties than the atoms from which they were formed.



The image above is a model of water. Water is a compound consisting of one atom of oxygen and two atoms of hydrogen. Hydrogen is an explosive gas, and oxygen is a gaseous substance that supports combustion. Yet, when these two elements are chemically combined to form water, the product neither burns nor supports combustion. In fact, water is used to put out fires.

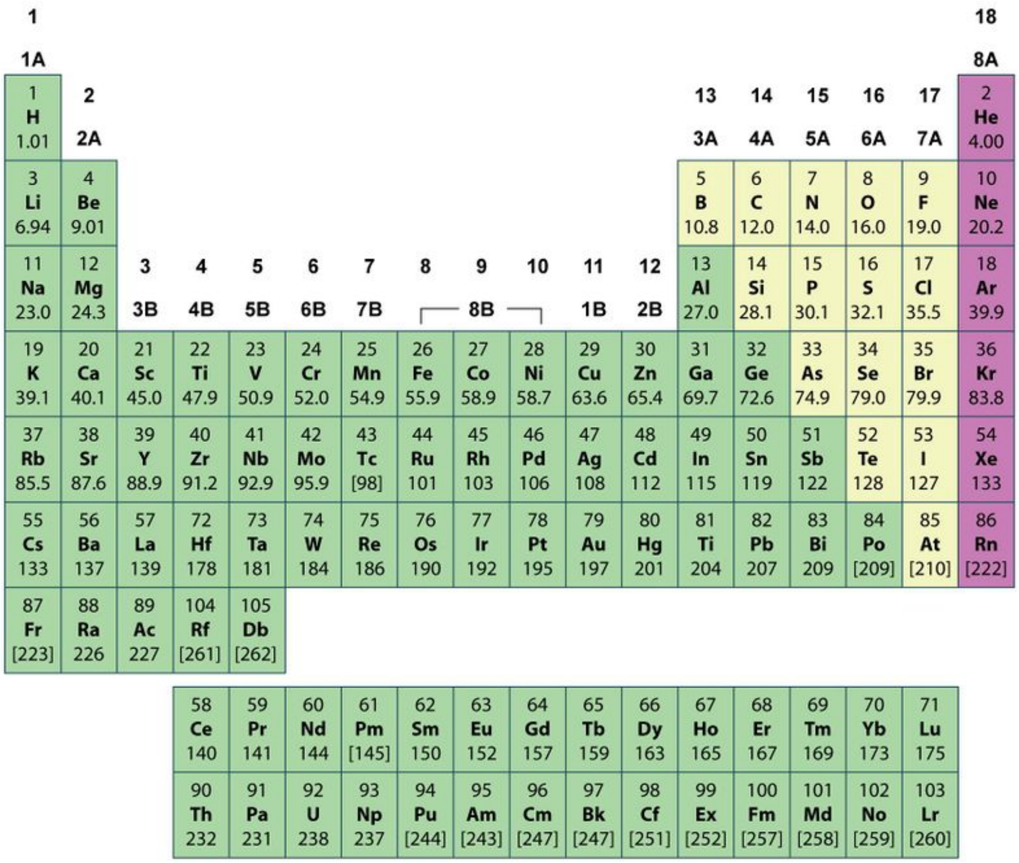
A **molecule** is the smallest particle of a compound. If you break up the molecule, you no longer have the properties of the compound. Molecules, like atoms, are too small to be seen. Even with the most powerful microscopes, we have only seen the very largest of molecules.



The illustration above shows a single unit of the compound called sodium chloride on the left. This single unit is made up of one sodium ion and one chloride ion. Sodium is a very reactive metal that explodes in water and burns in air, while chlorine is a very deadly, poisonous gas. When these two are combined, we get table salt (sodium chloride). When sodium chloride is in solid form, many units join together, as illustrated above on the right.

Elements: Names and Symbols

Everything, from ants to galaxies, is composed of atoms. So far, scientists have discovered or created 118 different types of atoms. Scientists have given a name to each different type of element and organized them into a chart called the periodic table. As you can see in the table below, each square contains one of the elements.



1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	8A
1 **H** 1.01																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.8	6 **C** 12.0	7 **N** 14.0	8 **O** 16.0	9 **F** 19.0	10 **Ne** 20.2
11 **Na** 23.0	12 **Mg** 24.3											13 **Al** 27.0	14 **Si** 28.1	15 **P** 30.1	16 **S** 32.1	17 **Cl** 35.5	18 **Ar** 39.9
19 **K** 39.1	20 **Ca** 40.1	21 **Sc** 45.0	22 **Ti** 47.9	23 **V** 50.9	24 **Cr** 52.0	25 **Mn** 54.9	26 **Fe** 55.9	27 **Co** 58.9	28 **Ni** 58.7	29 **Cu** 63.6	30 **Zn** 65.4	31 **Ga** 69.7	32 **Ge** 72.6	33 **As** 74.9	34 **Se** 79.0	35 **Br** 79.9	36 **Kr** 83.8
37 **Rb** 85.5	38 **Sr** 87.6	39 **Y** 88.9	40 **Zr** 91.2	41 **Nb** 92.9	42 **Mo** 95.9	43 **Tc** [98]	44 **Ru** 101	45 **Rh** 103	46 **Pd** 106	47 **Ag** 108	48 **Cd** 112	49 **In** 115	50 **Sn** 119	51 **Sb** 122	52 **Te** 128	53 **I** 127	54 **Xe** 133
55 **Cs** 133	56 **Ba** 137	57 **La** 139	72 **Hf** 178	73 **Ta** 181	74 **W** 184	75 **Re** 186	76 **Os** 190	77 **Ir** 192	78 **Pt** 195	79 **Au** 197	80 **Hg** 201	81 **Ti** 204	82 **Pb** 207	83 **Bi** 209	84 **Po** [209]	85 **At** [210]	86 **Rn** [222]
87 **Fr** [223]	88 **Ra** 226	89 **Ac** 227	104 **Rf** [261]	105 **Db** [262]													
			58 **Ce** 140	59 **Pr** 141	60 **Nd** 144	61 **Pm** [145]	62 **Sm** 150	63 **Eu** 152	64 **Gd** 157	65 **Tb** 159	66 **Dy** 163	67 **Ho** 165	68 **Er** 167	69 **Tm** 169	70 **Yb** 173	71 **Lu** 175	
			90 **Th** 232	91 **Pa** 231	92 **U** 238	93 **Np** 237	94 **Pu** [244]	95 **Am** [243]	96 **Cm** [247]	97 **Bk** [247]	98 **Cf** [251]	99 **Ex** [252]	100 **Fm** [257]	101 **Md** [258]	102 **No** [259]	103 **Lr** [260]	

Each element not only has its own name, it also has its own symbol. Scientists use abbreviations called chemical symbols to represent the elements. Many of these symbols are the first one or two letters of the modern name of the element. The first letter of a chemical symbol must always be a capital letter, and the second letter, when there is a second letter, must always be a lowercase letter. **Table** 2.1 shows some examples of elements and their symbols.

TABLE 2.1: **Examples of Elements**

Element	**Symbol**
Hydrogen	H
Oxygen	O
Carbon	C
Calcium	Ca
Aluminum	Al

As seen in **Table** 2.2, the symbols for some of the elements consist of the first letter of the name and another letter (not the second letter) that comes later in the name.

TABLE 2.2: **More Examples of Elements**

Element	**Symbol**
Zinc	Zn
Magnesium	Mg
Chlorine	Cl
Arsenic	As

TABLE 2.2: (continued)

Element	Symbol
Zirconium	Zr

For other elements, the symbols were already used for other elements. When trying to decide on a symbol for silver, for example, the symbol S was already used for sulfur, and the symbol Si was already used for silicon. Since silver has been known to man for over a thousand years, it had a Latin name from ancient times. The old Latin name for silver was argentum**Table** 2.3.

TABLE 2.3: **Examples of Elements Whose Symbol Comes from Latin**

Element	Ancient Name	Symbol
Silver	Argentum	Ag
Potassium	Kalium	K
Sodium	Natrium	Na
Gold	Aurum	Au
Lead	Plumbum	Pb
Copper	Cuprum	Cu
Iron	Ferrum	Fe

Compounds: Chemical Formulas

The chemical symbols are not only used to represent the elements, they are also used to write chemical formulas for the millions of different compounds. For a given chemical compound, the **law of constant composition** states that the ratio by mass of the elements in the compound is always the same, regardless of the source of the compound. The law of constant composition can be used to distinguish between compounds and mixtures. Compounds have a constant composition, and mixtures do not. Pure water is always 88.8% oxygen and 11.2% hydrogen by weight, regardless of the source of the water. Brass is an example of a mixture. Brass consists of two elements, copper and zinc, but it can contain as little as 10% or as much as 45% zinc.

The formula for a compound uses the symbols to indicate the type of atoms involved and uses subscripts to indicate the number of each atom in the formula. For example, aluminum combines with oxygen to form the compound aluminum oxide. Forming aluminum oxide requires two atoms of aluminum and three atoms of oxygen. Therefore, we write the formula for aluminum oxide as Al_2O_3. The symbol Al tells us that the compound contains aluminum, and the subscript 2 tells us that there are two atoms of aluminum in each molecule. The O tells us that the compound contains oxygen, and the subscript 3 tells us that there are three atoms of oxygen in each molecule. It was decided by chemists that when the subscript for an element is 1, no subscript needs to be used. Thus the chemical formula $MgCl_2$ tells us that one molecule of this substance contains one atom of magnesium and two atoms of chlorine. The formula for sodium chloride is NaCl, which indicates that the compound contains one atom each of sodium and chlorine. The formula for sodium carbonate, Na_2CO_3, indicates that there are two atoms of sodium, one atom of carbon, and three atoms of oxygen. In formulas that contain parentheses, the subscript outside of the parentheses applies to everything inside. For example, the subscript 2 in $Ca(OH)_2$, the subscript 2 applies to the (OH). Therefore, this molecule of calcium hydroxide contains one atom of calcium, two atoms of oxygen, and two atoms of hydrogen.

Lesson Summary

- All matter has mass and occupies space.
- Matter can be classified into two broad categories: pure substances and mixtures.

- A pure substance is a form of matter that has constant composition and constant properties throughout the sample.
- Mixtures are physical combinations of two or more substances.
- Elements and compounds are both example of pure substances.
- Compounds are substances that are made up of more than one type of atom.
- Elements are the simplest substances made up of only one type of atom.
- The elements are organized into a chart called the periodic table.
- Scientists use abbreviations called chemical symbols to represent the elements.
- The first letter of a chemical symbol is capitalized, and the second letter is not.

Further Reading / Supplemental Links

You may listen to Tom Lehrer's humorous song "The Elements" with animation at this website.

- http://www.privatehand.com/flash/elements.html

This website provides a review about matter and the categories of matter.

- http://www.thetech.org/exhibits/online/topics/50a.html

Review Questions

1. Pure substances contain only one type of
    a. atoms only.
    b. molecules only.
    c. atoms or molecules.
    d. mixture.
2. What type of mixture produces the same properties for every sample of the mixture?
    a. heterogeneous
    b. homogeneous
    c. mechanical
    d. environmental
3. Which of the following is a heterogeneous mixture?
    a. pure gold
    b. distilled water
    c. helium
    d. milk
4. Which of the following is not
    a. concrete
    b. pizza
    c. sugar water
    d. soup
5. If you can easily see the different parts that make up a mixture, you know that it is a ______ mixture.
    a. homogeneous

   b. heterogeneous
   c. biodegradable
   d. plasma

6. What do we call a material that is composed of two or more pure substances?
   a. a compound
   b. an element
   c. a mixture
   d. a heterogeneous mixture
7. Identify the following mixtures as homogeneous or heterogeneous.
   a. brass
   b. sugar dissolved in water
   c. vegetable soup
8. Identify which of the following pure substances are elements and which are compounds.
   a. table salt
   b. oxygen
   c. water
9. A pure substance composed of two or more elements chemically combined is a
   a. homogeneous mixture.
   b. compound.
   c. element.
   d. heterogeneous mixture.
10. The smallest piece of a compound that still has all the properties of the compound is a(n)
    a. atom.
    b. formula.
    c. mixture.
    d. molecule.
11. Identify the elements involved in the compound H_2SO_4.
12. How many phosphorus atoms are present in one molecule of H_3PO_4?

2.2 Properties and Changes of Matter

Lesson Objectives

The student will:

- explain the difference between physical and chemical properties of matter.
- list examples of physical properties.
- list examples of chemical properties.
- classify properties as chemical properties or physical properties.
- explain the difference between physical and chemical changes in matter.
- list examples of physical changes.
- list examples of chemical changes.
- classify changes as physical changes or chemical changes.

Vocabulary

- chemical change
- chemical property
- physical change
- physical property

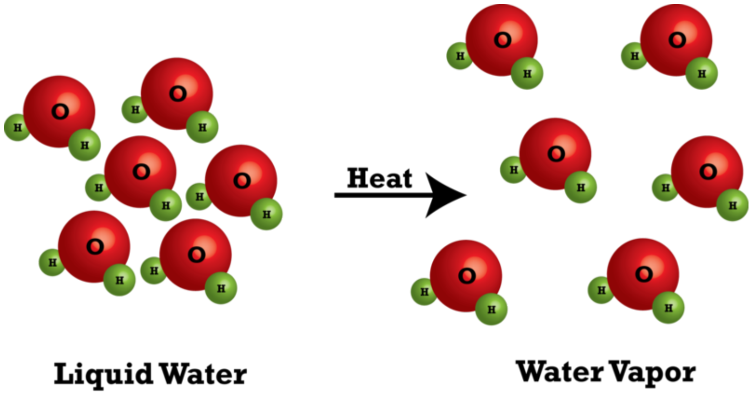
Introduction

What kinds of properties do chemists actually measure in the laboratory? Well, you can probably guess a few. Imagine that you are having dinner at a friend's house and are served something that you don't recognize. What types of observations might you make to determine what you've been given? You might note the smell or color of the food. You might observe whether the food is a liquid or a solid. You could also pick up a small amount of food with your fork and try to figure out how much it weighs. A light dessert might be something like an angel cake, while a heavy dessert is probably a pound cake. You might also want to know something about the food's texture. Is it hard and granular like sugar cubes, or soft and easy to spread like butter?

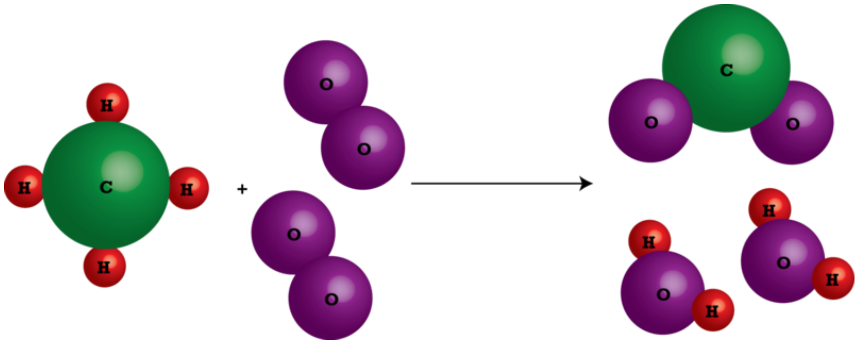
Believe it or not, the observations you are likely to make when trying to identify an unknown food are very similar to the observations that a chemists makes when trying to learn about a new material. In general, chemists are interested in characteristics that you can test and observe, such as a chemical's smell or color, and characteristics that are far too small to see, such as what the oxygen you breathe in or the carbon dioxide you breathe out looks like. Chemists rely on color, state (solid, liquid, or gas), temperature, volume, mass, and texture. There is, however, one property you might use to learn about a food but that you should definitely not use to learn about a chemical – taste!

Physical and Chemical Properties

There are two basic types of properties that are used to identify or describe matter: physical properties and chemical properties. **Physical properties** are properties that can be observed without changing the identity of the substance. In the image below, we have water molecules that are held in liquid form on the left. Each molecule contains two atoms of hydrogen chemically bounded with one atom of oxygen. When we heat the liquid water, it changes to water vapor. The physical properties change - we can see the liquid water, but the water vapor cannot be seen. Liquid water has a higher density than water vapor, and so on. But even though the physical properties have changed, the molecules are exactly the same as before. Each water molecule still contains two hydrogen atoms and one oxygen atom chemically bounded together.



On the other hand, **chemical properties** can only be observed when a substance is changed into a new substance. In the image below, on the left we have a molecule of methane (CH_4) and two molecules of oxygen (O_2). On the right, we have two molecules of water (H_2O) and one molecule of carbon dioxide (CO_2). In this case, not only has the appearance changed, but the structures of the molecules have also changed. The new substances do not have the same chemical properties as the original ones. Therefore, this is a chemical change. The chemical properties, such as how they react and what they react with, however, will still be the same as before.



Physical and Chemical Changes

Chemists make a distinction between two different types of changes that they study: physical changes and chemical changes. **Physical changes** are changes that do not alter the identity of a substance. Some types of physical changes include:

- changes of state (changes from a solid to a liquid or a gas, and vice versa)
- separation of a mixture
- physical deformation (cutting, denting, stretching)
- making solutions (special kinds of mixtures)

If you have a jar containing a mixture of pennies and nickels and you sort the mixture so that you have one pile of pennies and another pile of nickels, you have not altered the identity of either the pennies or the nickels. You've merely separated them into two groups. Similarly, if you have a piece of paper and you rip it up, you don't change the paper into something other than a piece of paper. These are examples of a physical change. For the most part, physical changes tend to be reversible, or capable of occurring in both directions. You can turn liquid water into solid water (ice) through cooling, and you can also turn solid water into liquid water through heating (**Figure** 2.1).

FIGURE 2.1

Melting snow is an example of a physical change.

Chemical changes are changes that occur when one substance is turned into another substance. Chemical changes are frequently harder to reverse than physical changes. One good example of a chemical change is burning paper. In contrast to the act of ripping paper, the act of burning paper actually results in the formation of new chemicals (carbon dioxide and water, to be exact). Notice that whereas ripped paper can be at least partially reassembled, burned paper cannot be "unburned." In other words, burning only goes in one direction. The fact that burning is not reversible is another good indication that it involves a chemical change. Another example of a chemical change, illustrated in **Figure** 2.2, is the explosion of fireworks.

FIGURE 2.2

Fireworks are an example of a chemical change.

Lesson Summary

- There are two basic types of properties that are used to identify or describe matter: physical properties and chemical properties.
- Physical properties are those that can be observed without changing the identity of the substance.
- Chemical properties are those that can be observed only when a substance is changed into a new substance.
- Chemists make a distinction between two different types of changes that they study: physical changes and chemical changes.
- Physical changes are changes that do not alter the identity of a substance
- Chemical changes are changes that occur when one substance is turned into another substance.
- Chemical changes are frequently harder to reverse than physical changes.

Further Reading / Supplemental Links

This website provides some free PowerPoint presentations. The presentation on “Matter and Energy" provides a review of some properties of matter, as well as provide examples of the topics covered in this lesson.

- http://science.pppst.com/energy.html

Review Questions

For questions 1-2, determine whether the description is of a physical property or a chemical property.

1. Water boils at 100°C.
   a. This is a physical property.
   b. This is a chemical property.
2. Diamonds will cut glass.
   a. This is a physical property.
   b. This is a chemical property.

For questions 3-7, determine whether the description is of a physical change or a chemical change.

3. Water can be separated by electrolysis into hydrogen gas and oxygen gas.
   a. This is a physical change.
   b. This is a chemical change.
4. Sugar dissolves in water.
   a. This is a physical change.
   b. This is a chemical change.
5. Vinegar and baking soda react to produce a gas.
   a. This is a physical change.
   b. This is a chemical change.
6. Yeast acts on sugar to form carbon dioxide and ethanol.

   a. This is a physical change.
   b. This is a chemical change.

7. Wood burns, producing several new substances.
   a. This is a physical change.
   b. This is a chemical change.

2.3 Energy

Lesson Objectives

The student will:

- explain the difference between kinetic and potential energy.
- state the law of conservation of matter and energy.
- define heat.
- define work.

Vocabulary

- chemical potential energy
- energy
- kinetic energy
- law of conservation of energy
- law of conservation of matter and energy
- potential energy
- work

Introduction

Just like matter, energy is a term that we are all familiar with and use on a daily basis. Before you go on a long hike, you eat an energy*energy* bill is paid; on TV, politicians argue about the *energy* crisis. But have you ever wondered what energy really is? If you stop to think about it, energy is very complicated. When you plug a lamp into an electric socket, you see energy in the form of light, but when you plug a heating pad into that same socket, you only feel warmth. Without energy, we couldn't turn on lights, we couldn't brush our teeth, we couldn't make our lunch, and we couldn't travel to school. In fact, without energy, we couldn't even wake up because our bodies require energy to function. We use energy for every single thing that we do, whether we're awake or asleep. Although we all use energy, very few of us understand what it is.

Types of Energy: Kinetic and Potential

Energy is the ability to do work or cause change. Machines use energy, our bodies use energy, energy comes from the sun, energy causes forest fires, and energy helps us to grow food. With all these seemingly different types of energy, it's hard to believe that there are really only two different forms of energy: kinetic energy and potential energy.

Kinetic energy is energy associated with motion. When an object is moving, it has kinetic energy, and when the object stops moving, it has no kinetic energy. Although all moving objects have kinetic energy, not all moving objects have the same amount of kinetic energy. The amount of kinetic energy possessed by an object is determined by its mass and its speed. The heavier an object is and the faster it is moving, the more kinetic energy it has.

Kinetic energy is very common and is easy to spot in the world around you. Sometimes we even capture kinetic energy and use it to power things like our home appliances. Have you ever seen windmills lining the slopes of a hill like the ones in **Figure** 2.3? These windmills capture the kinetic energy of the wind to provide power that people can use in their homes and offices. As wind rushes along the hills, the kinetic energy of the blowing air particles turns the windmills, which convert the wind's kinetic energy into electricity.

FIGURE 2.3

This is a photograph of a wind farm in Southern California. Kinetic energy from the rushing air particles turns the windmills, allowing us to capture the wind's kinetic energy and use it.

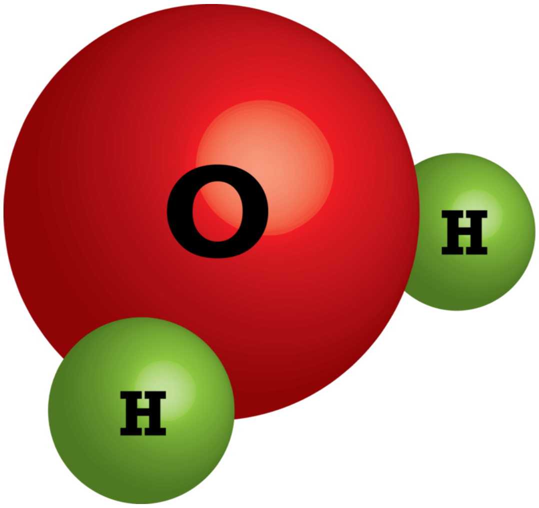
Capturing kinetic energy can be very effective, but you may already realize that there is a small problem: kinetic energy is only available when something is moving. When the wind stops blowing, there's no kinetic energy available. Imagine what it would be like trying to power your television set using the wind's kinetic energy. You could turn on the TV and watch your favorite program on a windy day, but every time the wind stopped blowing, your TV screen would flicker off because it would run out of energy.

You'd have noticed, however, that you can usually rely on your TV to stay on. This is largely because we don't rely on kinetic energy alone for power. Instead, we primarily use energy in its other form as potential energy. **Potential energy** is stored energy that remains available until we choose to use it. Think of a battery in a flashlight. If you leave a flashlight on, the battery will run out of energy within a couple of hours. If, instead, you only use the flashlight when you need it and turn it off when you don't, the battery will last for days or even months. Because the battery stores potential energy, you can choose to use the energy all at once, or you can save it and use a small amount at a time.

Any stored energy is potential energy and has the "potential" to be used at a later time. Unfortunately, there are a lot of different ways in which energy can be stored, making potential energy very difficult to recognize. Generally speaking, an object has potential energy due to its position relative to another object. For example, when you hold a rock above the earth, it has more potential energy than a rock on the ground. As long as you're holding the rock, the rock has potential energy stored. Once you drop the rock, though, the stored energy is released. This can confuse students because it doesn't seem like a falling rock is releasing energy. Remember, however, that energy is defined as the ability to do work or cause change.

For some examples of potential energy, though, it's harder to see how "position" is involved. In chemistry, we are often interested in what is called chemical potential energy. **Chemical potential energy** is energy stored in the atoms, molecules, and chemical bonds that make up matter. How does this depend on position? As you learned earlier, the world and all of the chemicals in it are made up of atoms. These atoms store potential energy that is dependent on their positions relative to one another. Although we cannot see atoms, scientists know a lot about the ways in which atoms interact. This allows them to figure out how much potential energy is stored in a specific quantity of a particular chemical. *Different chemicals have different amounts of potential energy* because they are made up of different atoms, and those atoms have different positions relative to one another.

The image below represents two hydrogen atoms chemically joined to an oxygen atom to form a water molecule. Scientists use their knowledge of what the atoms and molecules look like and how they interact to determine the potential energy that can be stored in any particular chemical substance.



Since different chemicals have different amounts of potential energy, scientists will sometimes say potential energy depends on not only position but also composition. Composition affects potential energy because it determines which molecules and atoms end up next to each other. For example, the total potential energy in a cup of pure water is different than the total potential energy in a cup of apple juice because the cup of water and the cup of apple juice are composed of different amounts of different chemicals.

The Law of Conservation of Matter and Energy

While it's important to understand the difference between kinetic energy and potential energy, the truth is energy is constantly changing. Kinetic energy is constantly being turned into potential energy, and potential energy is constantly being turned into kinetic energy. Even though energy can change form, it must still follow the fundamental law: energy cannot be created or destroyed**law of conservation of energy**. In a lot of ways, energy is like money. You can exchange quarters for dollar bills and dollar bills for quarters, but no matter how often you convert between the two, you won't end up with more or less money than you started with.

Think about what happens when you throw a ball into the air. When the ball leaves your hand, it has a lot of kinetic energy. At some point, the ball will stop momentarily in the air and then falls back down. When the ball stops, it no longer has any kinetic energy. According to the law of conservation of energy, the initial kinetic energy that the ball had does not just disappear. Instead, as the ball moves higher and higher into the sky, the kinetic energy is converted to potential energy. When the ball stops moving upward, all of the kinetic energy has been converted to potential energy. The ball then starts to fall back down, and the potential energy is once again changed into kinetic energy.

As it turns out, the law of conservation of energy isn't completely accurate. Energy and matter are actually interchangeable. In other words, energy can be created (made out of matter) and destroyed (turned into matter). As a result, the law of conservation of energy has been changed into the **law of conservation of matter and energy**. This law states that: *the total amount of mass and energy in the universe is conserved (does not change)*. This is one of the most important laws you will ever learn. Nevertheless, in chemistry we are rarely concerned with converting matter to energy or energy to matter. Instead, chemists deal primarily with converting one form of matter into another form of matter (through chemical reactions) and converting one form of energy into another form of energy.

Heat and Work

When we talk about using energy, we are really referring to transferring energy from one place to another. When you use energy to throw a ball, you transfer energy from your body to the ball, which causes the ball to fly through the air. When you use energy to warm your house, you transfer energy from the furnace to the air in your home, which causes the temperature in your house to rise. Although energy is used in many kinds of different situations, all of these uses rely on energy being transferred in one of two ways: as heat or as work. Unfortunately, both "heat" and "work" are used commonly in everyday speech, so you might think that you already know their meanings. In science, the words "heat" and "work" have very specific definitions that may be different from what you expect. Do not confuse the everyday meanings of the words "heat" and "work" with the scientific meanings.

When scientists speak of heat, they are referring to energy that is transferred from an object with a higher temperature to an object with a lower temperature as a result of the temperature difference. Heat will "flow" from the hot object to the cold object until both end up at the same temperature. When you cook with a metal pot, you witness energy being transferred in the form of heat. Initially, only the stove element is hot; the pot and the food inside the pot are cold. As a result, heat moves from the hot element to the cold pot, as illustrated in **Figure** 2.4. After a while, enough heat is transferred from the element to the pot, raising the temperature of the pot and all of its contents.

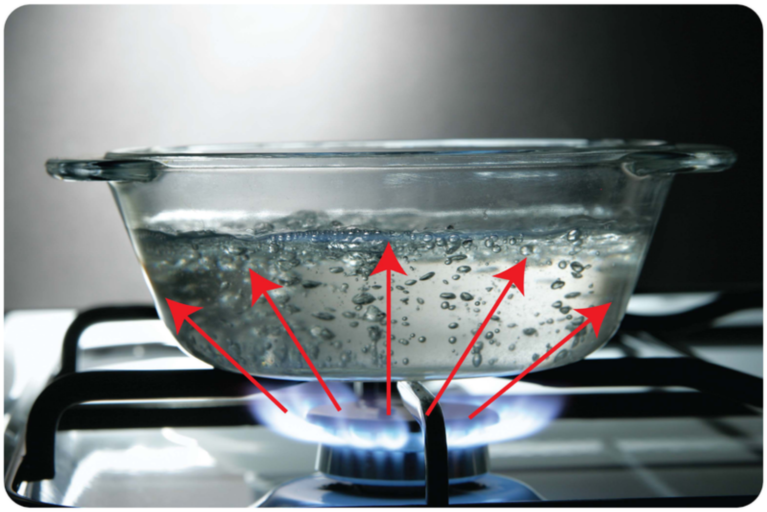
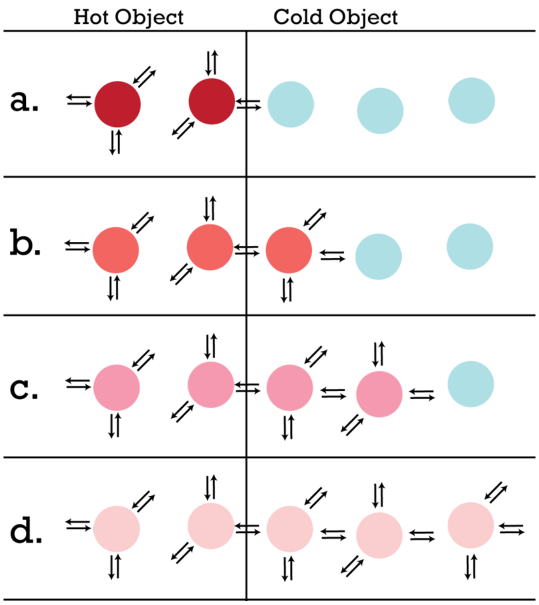


FIGURE 2.4

Energy is transferred as heat from the hot stove element to the cooler pot until the pot and its contents become just as hot as the element.

We've all observed heat moving from a hot object to a cold object, but you might wonder how the energy actually travels. Whenever an object is hot, the molecules within the object are shaking and vibrating vigorously. The hotter an object is, the more the molecules jiggle around. Anything that is moving has energy, and the more it's moving, the more energy it has. Hot objects have a lot of energy, and it's this energy that is transferred to the colder objects when the two come in contact.

The easiest way to visualize heat transfer is to imagine a domino effect. When the vibrating molecules of the hot object bump into the molecules of the colder object, they transfer some of their energy, causing the molecules in the colder object to start vibrating vigorously as well. In the image below, the red molecules are jiggling around and vibrating. As these molecules vibrate, they bump into their neighbors (the blue molecules) and transfer some of their energy. These colder molecules begin to heat up and begin to vibrate faster. Just like dominoes, the heat gets passed along the chain until the energy is spread equally between all of the molecules. At the end, all of the molecules will be at the same temperature.



Heat is only one way in which energy can be transferred. Energy can also be transferred as work. The scientific definition of **work** is force (any push or pull) applied over a distance. Whenever you push an object and cause it to move, you've done work and transferred some of your energy to the object. At this point, it is important to warn you of a common misconception. Sometimes we think that the amount of work done can be measured by the amount of effort put in. This may be true in everyday life, but this is not true in science. By definition, scientific work requires that force be applied over a distance. It doesn't matter how hard you push or pull. If you haven't moved the object, you haven't done any work. For example, no matter how much you sweat, if you cannot lift a heavy object off the ground, you have not done any work.

Lesson Summary

- Energy is the ability to do work or cause change.
- The two forms of energy are kinetic energy and potential energy.
- Kinetic energy is energy associated with motion.
- Potential energy is stored energy.
- Kinetic energy is constantly being turned into potential energy, and potential energy is constantly being turned into kinetic energy.
- Even though energy can change form, it must still follow the law of conservation of energy.
- The law of conservation of energy states that energy cannot be created or destroyed, it can only be changed from one form to another.
- When scientists speak of heat, they are referring to energy that is transferred from an object with a higher temperature to an object with a lower temperature as a result of the temperature difference.
- Heat will "flow" from the hot object to the cold object until both end up at the same temperature.

- Energy can also be transferred as work.
- Work is force (any push or pull) applied over a distance.

Further Reading / Supplemental Links

Summary of concepts of matter and energy and benchmark review.

- http://broncho2.uco.edu/funeral/Bill%20Lewis/BoardReview/ChemLessons/Lesson%201.pdf

Classroom videos about energy.

- http://www.energyclassroom.com/

Review Questions

1. Classify each of the following as energy primarily transferred as heat or energy primarily transferred as work.
    a. The energy transferred from your body to a shopping cart as you push the shopping cart down the aisle.
    b. The energy transferred from a wave to your board when you go surfing.
    c. The energy transferred from the flames to your hot dog when you cook your hot dog over a campfire.
2. Decide whether each of the following statements is true or false.
    a. When heat is transferred to an object, the object cools down.
    b. Any time you raise the temperature of an object, you have done work.
    c. Any time you move an object by applying force, you have done work.
    d. Any time you apply force to an object, you have done work.
3. Rank the following scenarios in order of increasing
    a. You apply 100 N of force to a boulder and successfully move it by 2 m.
    b. You apply 100 N of force to a boulder and successfully move it by 1 m.
    c. You apply 200 N of force to a boulder and successfully move it by 2 m.
    d. You apply 200 N of force to a boulder but cannot move the boulder.
4. In science, a vacuum is defined as space that contains absolutely no matter (no molecules, no atoms, etc.) Can energy be transferred as heat through a vacuum? Why or why not?
5. Classify each of the following energies as kinetic energy or potential energy:
    a. the energy in a chocolate bar.
    b. the energy of rushing water used to turn a turbine or a water wheel.
    c. the energy of a skater gliding on the ice.
    d. the energy in a stretched rubber band.
6. Decide which of the following objects has more kinetic energy.
    a. A 200 lb man running at 6 mph or a 200 lb man running at 3 mph.
    b. A 200 lb man running at 7 mph or a 150 lb man running at 7 mph.
    c. A 400 lb man running at 5 mph or a 150 lb man running at 3 mph.
7. A car and a truck are traveling along the highway at the same speed.
    a. If the car weighs 1500 kg and the truck weighs 2500 kg, which has more kinetic energy, the car or the truck?

b. Both the car and the truck convert the potential energy stored in gasoline into the kinetic energy of motion. Which do you think uses more gas to travel the same distance, the car or the truck?

2.4 References

1. Photo by Magnus Rosendahl. Melting lakeside. Creative Commons Public Domain License
2. PDPhoto.org. Fireworks. Public domain
3. Photograph by Stan Shebs. Tehachapi wind farm. CC-BY-SA-3.0
4. Image copyrighted by Roman Sigaev, modified by Christopher Auyeung. Glass Saucepan on the Gas Stove. Used under 2010 license from Shutterstock.com

CHAPTER 3 Electron Configuration and the Periodic Table

Chapter Outline

3.1 Mendeleev's Periodic Table

Lesson Objectives

The student will:

- identify the person credited for organizing the periodic table.
- state the basis for the organization of Mendeleev's periodic table.

Vocabulary

- periodic table

Introduction

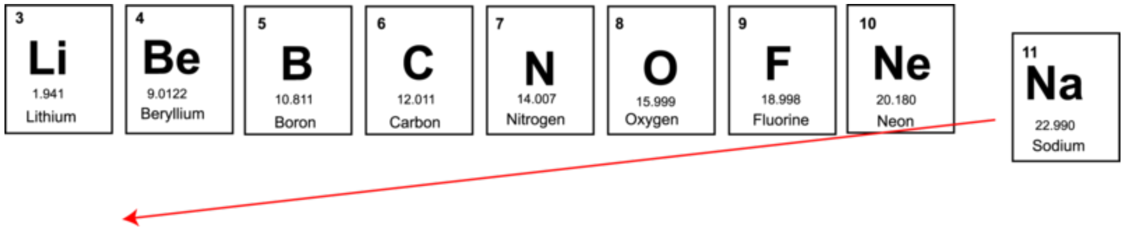
During the 1800s, when most of the elements were being discovered, many chemists tried to classify the elements according to their similarities. In 1829, Johann Dobringer noted chemical similarities in several groups of three elements and placed these elements into what he called triads. His groupings included the triads of 1) chlorine, bromine, and iodine; 2) sulfur, selenium, and tellurium; 3) calcium, strontium, and barium; and 4) lithium, sodium, and potassium.

In 1864, John Newlands saw a connection between the chemical properties of elements and their atomic masses. He stated that if the known elements, beginning with lithium, are arranged in order of increasing mass, the eighth element will have properties similar to the first element, the ninth similar to the second, the tenth similar to the third, and so on. Newlands called his relationship the law of octaves, comparing the elements to the notes in a musical scale. Newlands tried to force all the known elements to fit into his octaves, but many of the heavier elements, when discovered, did not fit into his patterns.

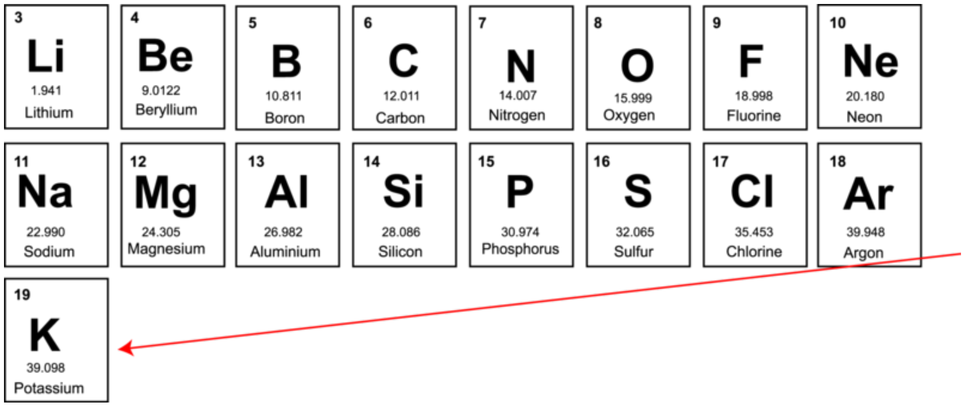
Mendeleev's Table Based on Chemical Behavior

The **periodic table** is a tabular arrangement of the chemical elements according to atomic number. In 1869, Dmitri Mendeleev, unaware of Newlands's work, also developed a table of the elements that showed a relationship between the chemical properties of elements and their atomic masses. His table would be similar to the one we use today. Mendeleev avoided Newlands's mistake of trying to force elements into groups where their chemistry did not match. Instead, he left empty spaces in his table for where no known element fit the properties and atomic mass for that space. He assumed that elements discovered in the future would fit in those spots. Mendeleev even predicted some of the chemical and physical properties the undiscovered elements would have.

Mendeleev's table listed the elements in order of increasing atomic mass and grouped elements with similar chemical behavior together. For example, lithium is a shiny metal that is soft enough to be cut with a spoon. It reacts readily with oxygen and violently with water. When lithium reacts with water, it produces hydrogen gas and lithium hydroxide. As we examine the next few elements in order of increasing mass, we will come to the element sodium. Sodium is also a shiny metal soft enough to be cut with a spoon. It reacts readily with oxygen and violently with water. When sodium reacts with water, it produces hydrogen gas and sodium hydroxide. You should note that the description of the chemical behavior of sodium is very similar to the chemical description of lithium. Whenever Mendeleev found an element whose chemistry was very similar to a previous element, he placed it below that similar element.

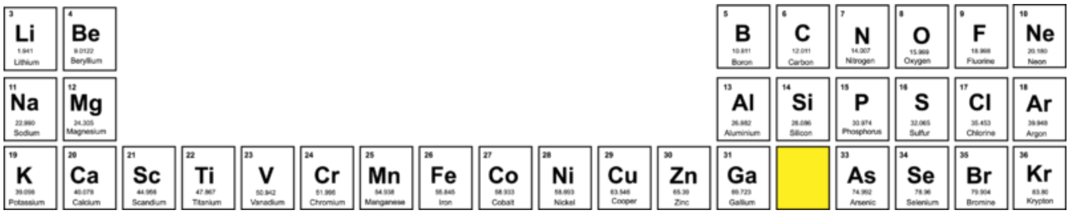


As each succeeding element was considered, Mendeleev continued to place an element underneath other elements with similar chemistry. For this reason, Mendeleev placed magnesium below beryllium, aluminum under boron, and so on.



When Mendeleev came to potassium, he found that it had very similar chemistry to the chemistry of lithium and sodium, so he placed potassium under sodium. You may notice that the atomic masses are no longer in increasing order with respect to potassium and argon. Although argon has a greater mass than potassium, the chemistry of argon was identical to that of neon, and the chemistry of potassium was almost identical to that of sodium. As a result, Mendeleev did not try to put argon beneath sodium. Instead, he allowed the chemical behavior determine where each element should be placed. As you go through the periodic table, you will find several elements whose atomic masses are not in increasing order.

Mendeleev's procedure was to place the elements in the column with matching chemical behavior regardless of atomic mass, so if no element with appropriate chemistry was known, the space would be left empty. For example, germanium had not been discovered when Mendeleev constructed his table. After gallium, the next known element in Mendeleev's time was arsenic. Arsenic did not match the chemical characteristics of carbon and silicon. Instead, arsenic matched the chemical characteristics of nitrogen and phosphorus. Mendeleev placed arsenic in the column which matched arsenic's chemistry and assumed that there was an undiscovered element that would fit chemically with the carbon column. As a result, Mendeleev left the space blank, as you can see in the table below. At a later date, when germanium was discovered, germanium was found to fit the space and was placed in the empty spot.



Lesson Summary

- The periodic table in its present form was organized by Dmitri Mendeleev.
- The basis for Mendeleev's organization was the chemical behavior of the elements.

Further Reading / Supplemental Links

The following websites explore the history behind the periodic table.

- http://www.wou.edu/las/physci/ch412/perhist.htm
- http://www.aip.org/history/curie/periodic.htm

Review Questions

1. On what basis did Mendeleev organize his periodic table?
    a. atomic number and atomic mass
    b. atomic mass and chemical behavior
    c. atomic number of number of valence electrons
    d. atomic mass and the Latin alphabet
    e. none of these

3.2 Families and Periods of the Periodic Table

Lesson Objectives

The student will:

- identify groups in the periodic table.
- state the number of valence electrons for each A group in the periodic table.
- explain the relationship between the chemical behavior of families in the periodic table and their electron configuration.
- identify periods in the periodic table.
- describe the similarities among elements in the same period in the periodic table.

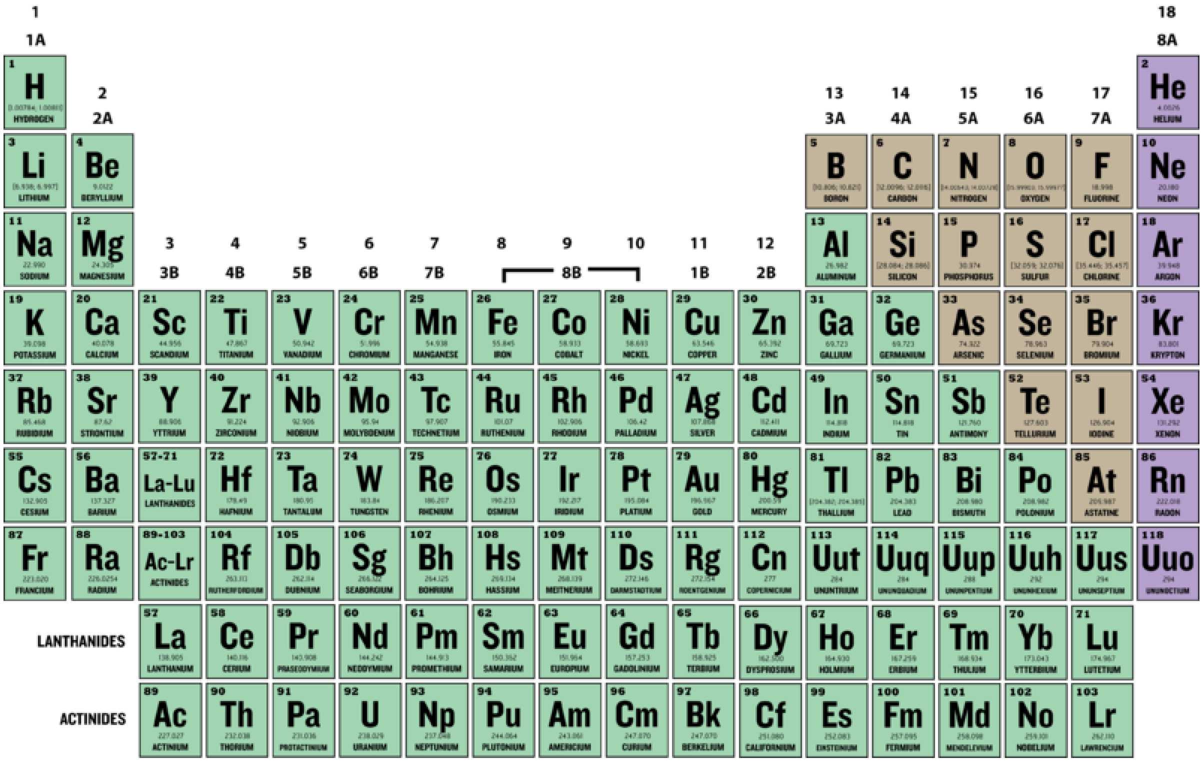
Vocabulary

- actinide series
- alkali metals
- alkaline earth metals
- group (family)
- halogens
- lanthanide series
- noble gases
- period
- transition elements

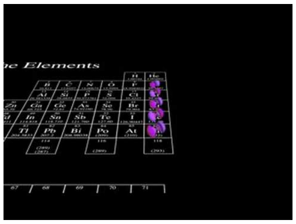
Introduction

When Mendeleev created his periodic table, he did not know why certain elements had similar chemistry. He placed the elements in their positions because they exhibited similar chemical behaviors. Thus, the vertical columns in Mendeleev's table were composed of elements with similar chemistry. These vertical columns are called **groups**, or families. In this section, you are going to see that the elements in the same groups are related to each other by their electron configurations. Since the families of elements were organized by their chemical behavior, it is predictable that the individual members of each chemical family will have similar electron configurations.

If you examine a periodic table, you will often find a number written above each group (column). These numbers serve as labels, and groups are often referred to by their labels. Depending on the source or age of your periodic table, you may see two different numbering systems for referring to the families on the periodic table. In the older system, the numbers 1 – 8 and the letters A and B were used to label the groups. The newer convention is to label each group from 1 – 18 in sequential order. However, the older labeling scheme helps to provide more insight into the electron configurations of each group. As a result, in this text we will use the older labeling scheme to present each group. The periodic table below shows both numbering systems.



For an introduction to the electronic organization of the periodic table **(1a, 1c, 1d, 1f)**, see http://www.youtube.com/watch?v=5MMWpeJ5dn4 (3:51).



MEDIA

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Group 1A

The electron configuration codes for the elements in Group 1A are:

lithium = $1s^22s^1$ or $[\mathrm{He}]2s^1$
sodium = $1s^22s^22p^63s^1$ or $[\mathrm{Ne}]3s^1$
potassium = $1s^22s^22p^63s^23p^64s^1$ or $[\mathrm{Ar}]4s^1$
rubidium = $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^1$ or $[\mathrm{Kr}]5s^1$
cesium = $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^1$ or $[\mathrm{Xe}]6s^1$
francium = $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^67s^1$ or $[\mathrm{Rn}]7s^1$

The fact that all 1A elements participate in similar chemistry despite having vastly different nuclear sizes further illustrates the fact that electrons, particularly valence electrons, are the primary contributors to chemical reactivity. The electron configuration for the outermost energy levels of the 1A elements is the same, with the only difference being the energy level involved. Each larger member of the family has its single s**alkali metals**) increases as the atoms become larger. Note that in the new labeling scheme, this group is labeled as Group 1.

All of the 1A elements have one valence electron and react in similar ways as the other members of the family – that's why they are a part of the same family. The 1A metals have similar electron configurations, which causes them to have similar chemistry. Although Mendeleev did not know about these elements had similar electron configurations, he recognized the similar chemistry and organized the periodic table based on this similar chemistry.

Group 2A

Here are the electron configuration codes for the first five elements in Group 2A:

beryllium = $1s^2 2s^2$ or $[\text{He}]2s^2$
magnesium = $1s^2 2s^2 2p^6 3s^2$ or $[\text{Ne}]3s^2$
calcium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ or $[\text{Ar}]4s^2$
strontium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$ or $[\text{Kr}]5s^2$
barium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$ or $[\text{Xe}]6s^2$

All of the elements in this family have two valence electrons and have very similar chemistry. This group of metals is called the **alkali earth metals**. As with the 1A family, the elements in this family also increase in reactivity as the elements become larger and the valence electrons are held more loosely. In the new labeling conventions, this group is labeled as Group 2.

Group 3A

The electron configurations of Group 3A (Group 3 according to new labeling conventions) show that all the members of this family have three valence electrons. The chemical behaviors of the elements in this family are similar to each other, but they are not as consistent from element to element as they are for other families of elements. Information in the following sections will explain why.

boron = $1s^2 2s^2 2p^1$ or $[\text{He}]2s^2 2p^1$
aluminum = $1s^2 2s^2 2p^6 3s^2 3p^1$ or $[\text{Ne}]3s^2 3p^1$
gallium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$ or $[\text{Ar}]4s^2 3d^{10} 4p^1$
indium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^1$ or $[\text{Kr}]5s^2 4d^{10} 5p^1$
thallium = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^1$ or $[\text{Xe}]6s^2 4f^{14} 5d^{10} 6p^1$

Groups 4A through 8A

Group 4A (also known as Group 14) members have four valence electrons.

carbon = $[\text{He}]2s^2 2p^2$
silicon = $[\text{Ne}]3s^2 3p^2$
germanium = $[\text{Ar}]4s^2 3d^{10} 4p^2$
tin = $[\text{Kr}]5s^2 4d^{10} 5p^2$
lead = $[\text{Xe}]6s^2 4f^{14} 5d^{10} 6p^2$

Group 5A (or Group 15) members have five valence electrons.

nitrogen = [He]$2s^22p^3$
phosphorus = [Ne]$3s^23p^3$
arsenic = [Ar]$4s^23d^{10}4p^3$
antimony = [Kr]$5s^24d^{10}5p^3$
bismuth = [Xe]$6s^24f^{14}5d^{10}6p^3$

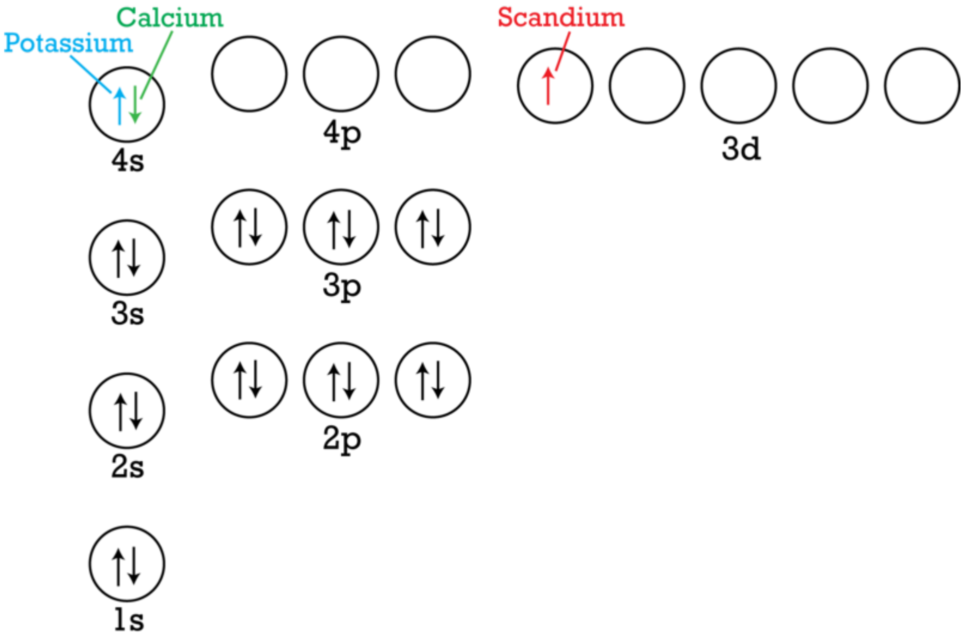
Instead of showing you the electron configurations for Groups 6A (or Group 16), 7A (or 17), and 8A (or 18), it would be good practice for you to practice writing the electron configurations for these families. When you finish writing them, check to make sure that the outer energy levels contain six valence electrons, seven valence electrons, and eight valence electrons for Groups 6A, 7A, and 8A, respectively. The group label provides a hint about the valence electrons: the number preceding the "A" is equal to the number of valence electrons for each atom in the group.

Groups 7A and 8A are also known by the names **halogens** and the **noble gases**, respectively. Groups 7A and 8A were two of the first families identified because the chemistry of their members are so similar to each other.

Transition Elements

If you have been checking where the different groups that have been introduced so far are located on the periodic table, you may have noticed that there are a series of elements that are not part of Groups 1A – 8A (sometimes refer to as the main group). In some periodic tables, these groups of elements are in families called B groups, and in newer periodic tables, these groups are numbered 3 – 12.

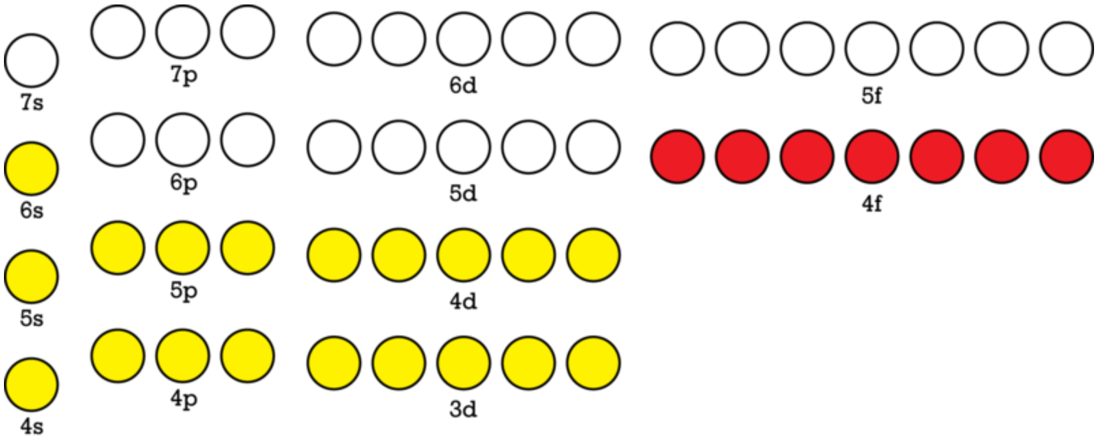
Consider the fourth horizontal row in the periodic table. For potassium, the first element in this row, one electron resides in the 4s*s* orbital. Beyond calcium, however, the pattern in which the electrons are added changes. Beginning with scandium, atomic number 21, the additional electrons do not enter the valence shell but instead enter the **d** sublevel of the third energy level ($n = 3$), as illustrated below. The electron configuration for scandium is [Ar]$4s^23d^1$.



The ten elements formed by filling in the 3d*d* orbitals, are called the **transition elements**. These elements, in general, differ from each other in the electron structure of the next-to-last energy level. For the most part, these elements have similar outer energy levels because they have two valence electrons. As a result, they have somewhat similar chemistry, even though they are not in the same vertical family.

The Lanthanide and Actinide Series

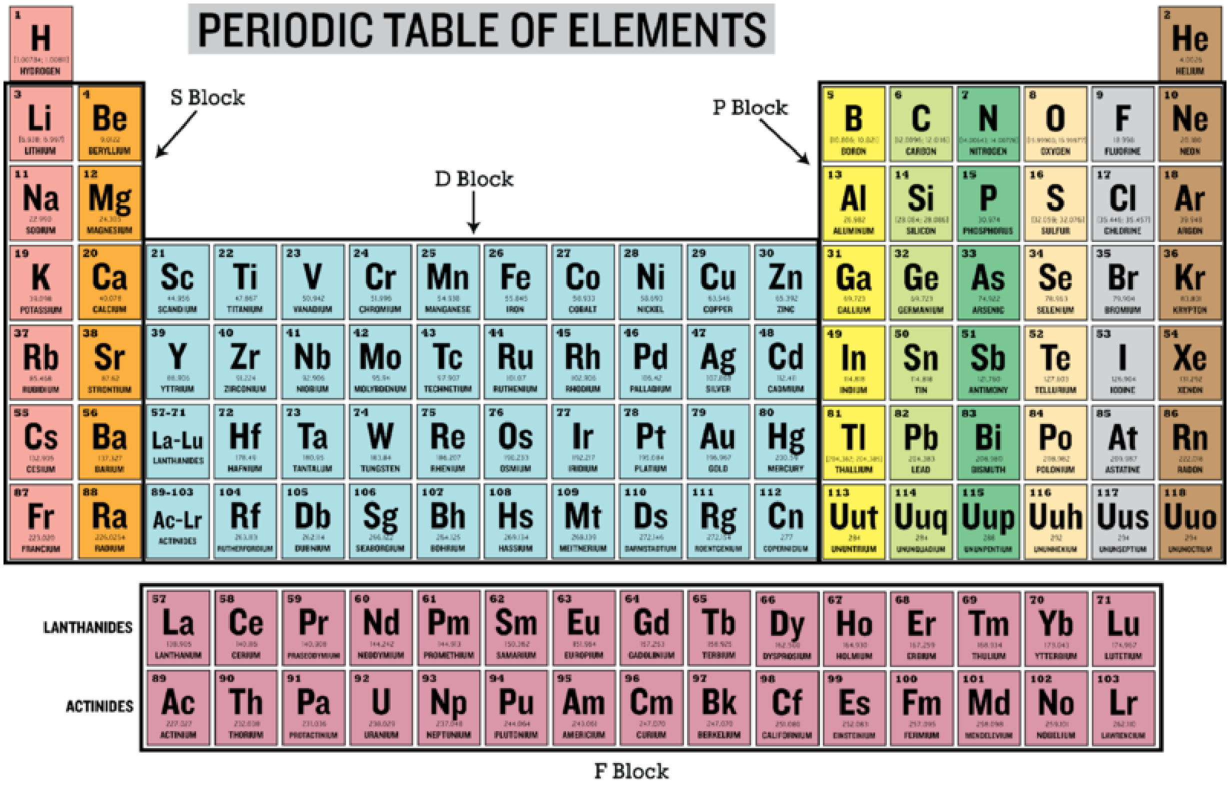
There is still one more block of elements on the periodic table that has not been introduced. This block is usually placed below the periodic table and represents elements with electrons in the f*s* orbital and produces the element cesium. The second electron fills the 6*s*, orbital producing the element barium. The next 14 electrons that are added then enter the 4*f* orbitals (marked in red in the figure below).



This group of elements, atomic numbers 57 – 70, is called the **lanthanide series**. Elements with atomic numbers 89 – 102 are called the **actinide series**. As in the case of the transition elements, these elements have the electrons added to an inner energy level, rather than the valence shell. The number of valence electrons in these elements remains at 2 while the electrons enter and eventually fill the *f* orbitals. For some of the elements with more complicated electron configuration, this is a somewhat simplified way to generate the electron configurations. Due to the fact that outer orbitals, such as 5*f* and 6*d*, have increasingly similar energies, other factors can come into play and change the exact configuration of the lanthanides, actinides, and some transition metals.

Sub-Level Filling Blocks on the Periodic Table

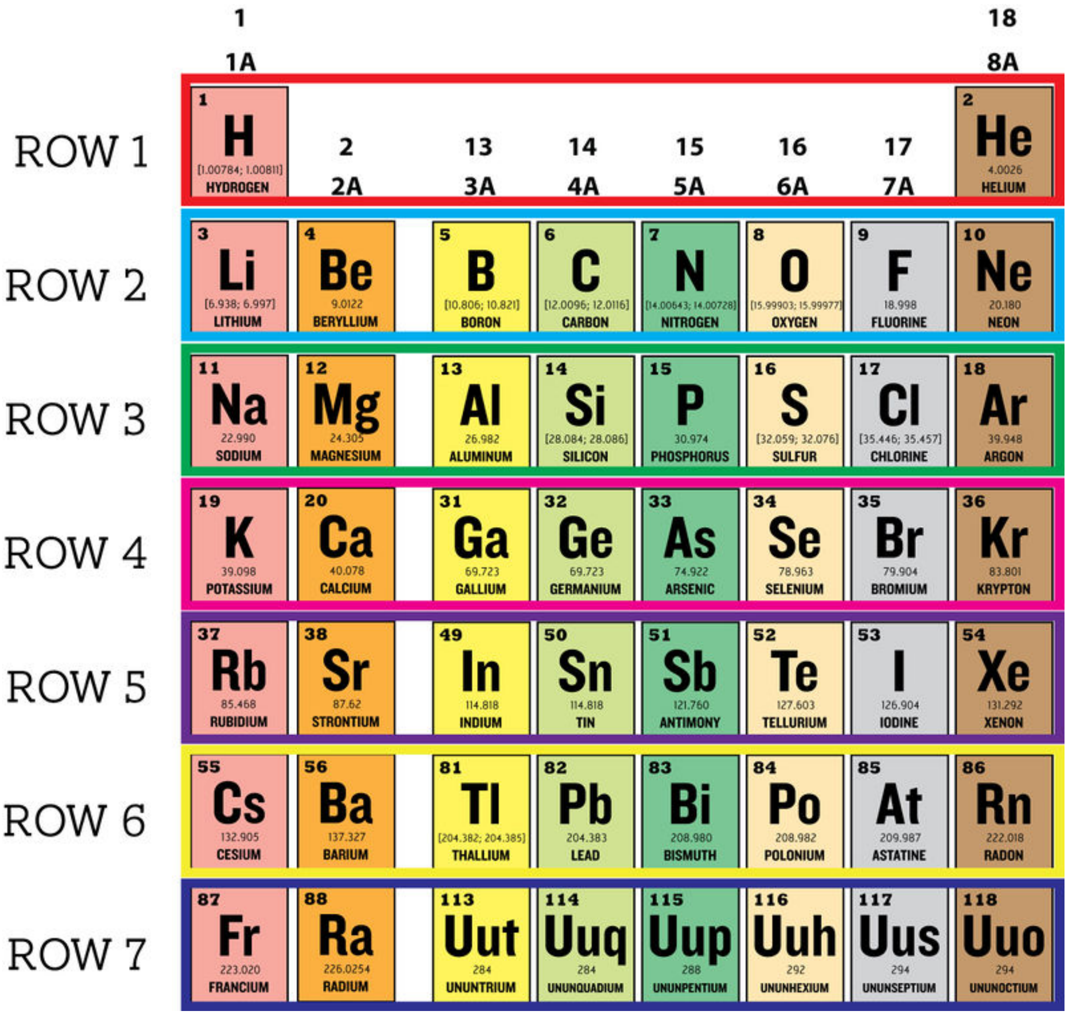
For Groups 1A and 2A, the last electron added enters an s*p* orbital. These six families of elements, therefore, make up an area called the "p block." The various sub-level blocks, including the "d block" and "f block," are indicated in the periodic table shown below.



Even though the periodic table was organized according to the chemical behavior of the elements, you can now see that the shape and design of the table is a perfect reflection of the electron configuration of the atoms. This is because the chemical behavior of the elements is also dictated by the electron configuration of the atoms.

Periods of the Periodic Table

In addition to providing information about the electron configuration, an element's position on the periodic table can also be used to figure out the energy level of the element's valence electrons. Let's try to figure out what we can learn from an element's period in the periodic table. A **period** is a horizontal row of elements on the periodic table. The figure below shows how the different rows in the periodic table are numbered.



	1 1A	2 2A	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
ROW 1	1 H [1.00784; 1.00811] HYDROGEN							2 He 4.0026 HELIUM
ROW 2	3 Li [6.938; 6.997] LITHIUM	4 Be 9.0122 BERYLLIUM	5 B [10.806; 10.821] BORON	6 C [12.0096; 12.0116] CARBON	7 N [14.00643; 14.00728] NITROGEN	8 O [15.99903; 15.99977] OXYGEN	9 F 18.998 FLUORINE	10 Ne 20.180 NEON
ROW 3	11 Na 22.990 SODIUM	12 Mg 24.305 MAGNESIUM	13 Al 26.982 ALUMINUM	14 Si [28.084; 28.086] SILICON	15 P 30.974 PHOSPHORUS	16 S [32.059; 32.076] SULFUR	17 Cl [35.446; 35.457] CHLORINE	18 Ar 39.948 ARGON
ROW 4	19 K 39.098 POTASSIUM	20 Ca 40.078 CALCIUM	31 Ga 69.723 GALLIUM	32 Ge 69.723 GERMANIUM	33 As 74.922 ARSENIC	34 Se 78.963 SELENIUM	35 Br 79.904 BROMIUM	36 Kr 83.801 KRYPTON
ROW 5	37 Rb 85.468 RUBIDIUM	38 Sr 87.62 STRONTIUM	49 In 114.818 INDIUM	50 Sn 114.818 TIN	51 Sb 121.760 ANTIMONY	52 Te 127.603 TELLURIUM	53 I 126.904 IODINE	54 Xe 131.292 XENON
ROW 6	55 Cs 132.905 CESIUM	56 Ba 137.327 BARIUM	81 Tl [204.382; 204.385] THALLIUM	82 Pb 204.383 LEAD	83 Bi 208.980 BISMUTH	84 Po 208.982 POLONIUM	85 At 209.987 ASTATINE	86 Rn 222.018 RADON
ROW 7	87 Fr 223.020 FRANCIUM	88 Ra 226.0254 RADIUM	113 Uut 284 UNUNTRIUM	114 Uuq 284 UNUNQUADIUM	115 Uup 288 UNUNPENTIUM	116 Uuh 292 UNUNHEXIUM	117 Uus 294 UNUNSEPTIUM	118 Uuo 294 UNUNOCTIUM

To understand what this means in terms of an element's electron configuration, let's consider the Group 1A metals. If we write the electron configuration for the Group 1A metal from each row of the Periodic Table, we have:

row	element	electron configuration
2	Li	$1s^2\mathbf{2s}^1$
3	Na	$1s^22s^22p^6\mathbf{3s}^1$
4	K	$1s^22s^22p^63s^23p^6\mathbf{4s}^1$
5	Rb	$1s^22s^22p^63s^23p^64s^23d^{10}4p^6\mathbf{5s}^1$
6	Cs	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6\mathbf{6s}^1$
7	Fr	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^6\mathbf{7s}^1$

Do you see any patterns? For Group 1A metals, it seems that element's row corresponds to the energy level of that element's valence electron. Lithium (Li), for instance, is found in row **2** of the periodic table, and its valence electron is found in the $n = \mathbf{2}$ energy level. Cesium (Cs) is found in row **6** of the periodic table, and its valence electron is found in the $n = \mathbf{6}$ energy level. Now consider the Group 2A metals:

row	element	electron configuration
2	Be	$1s^2\mathbf{2s}^2$
3	Mg	$1s^22s^22p^6\mathbf{3s}^2$
4	Ca	$1s^22s^22p^63s^23p^6\mathbf{4s}^2$
5	Sr	$1s^22s^22p^63s^23p^64s^23d^{10}4p^6\mathbf{5s}^2$
6	Ba	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6\mathbf{6s}^2$
7	Ra	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^6\mathbf{7s}^2$

For Group 2A metals, the same rule applies. Magnesium (Mg) is found in row **3** of the periodic table, and its valence electrons are found in the n = **3** energy level. Similarly, radium (Ra) is found in row **7** of the periodic table, and its valence electrons are found in the n = **7** energy level.

Consider the electron configurations for the Group 3A elements in each row:

row	element	electron configuration
2	B	$1s^2\mathbf{2}s^2\mathbf{2}p^1$
3	Al	$1s^22s^22p^6\mathbf{3}s^2\mathbf{3}p^1$
4	Ga	$1s^22s^22p^63s^23p^6\mathbf{4}s^23d^{10}\mathbf{4}p^1$
5	In	$1s^22s^22p^63s^23p^64s^23d^{10}4p^6\mathbf{5}s^24d^{10}\mathbf{5}p^1$
6	Tl	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6\mathbf{6}s^25d^{10}\mathbf{6}p^1$

Even though the valence electrons in Group 3A elements are found in both **s** and **p** orbitals, it turns out that an element's row still corresponds to the energy level of that element's valence electrons. For example, gallium (Ga) is found in row **4** of the periodic table, and its valence electrons are found in the n = **4** energy level. Likewise, thallium (Tl) is found in row **6** of the periodic table, and its valence electrons are found in the n = **6** energy level.

Let's consider one last example by examining the electron configurations for the Group 7A elements:

row	element	electron configuration
F	B	$1s^2\mathbf{2}s^2\mathbf{2}p^5$
Cl	Al	$1s^22s^22p^6\mathbf{3}s^2\mathbf{3}p^5$
Br	Ga	$1s^22s^22p^63s^23p^6\mathbf{4}s^23d^{10}\mathbf{4}p^5$
I	In	$1s^22s^22p^63s^23p^64s^23d^{10}4p^6\mathbf{5}s^24d^{10}\mathbf{5}p^5$
At	Tl	$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6\mathbf{6}s^25d^{10}\mathbf{6}p^5$

Once again, it turns out that an element's period can be used to determine the energy level of that element's valence electrons. Chlorine (Cl), for instance, is found in row **3** of the periodic table, and its valence electrons are found in the n = **3** energy level. Similarly, iodine (I) is found in row 5 of the periodic table, and its valence electrons are found in the n = **5** energy level.

Note, however, that determining the energy level of an element's valence electrons by using the element's row number works only for Groups 1A – 8A. This rule does not

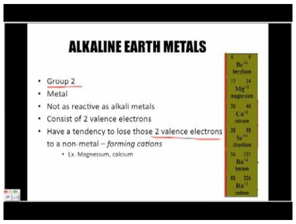
You may have noticed that the first row has not been mentioned at all. In all of the examples considered, only the elements starting from the second row were considered. The first row in the periodic table is a "special row" for several reasons. To begin with, the first row only contains hydrogen and helium. Can you figure out why there are only two elements in the first row?

Recall that an element's period number corresponds to the energy level of that element's valence electrons. Therefore, the first row must contain elements with valence electrons in the nn = 1 energy level, there is only one orbital: the 1s orbital. Hydrogen (electron configuration $1s^1$) has one valence electron in the 1s orbital, and helium (electron configuration $1s^2$) has two valence electrons in the 1s orbital. Since it is impossible to fit more than two electrons into the 1s orbital, atoms with a total of three or more electrons must have valence electrons in an energy level with n = 2 or greater. Atoms with a total of three or more electrons do not belong in the first row of the periodic table.

The first row is also special because its elements have special properties. Hydrogen, for example, is not a metal like the rest of the Group 1A elements. Instead, hydrogen atoms react with each other and form hydrogen gas, H_2. Hydrogen is different from the other Group 1A elements, so the chemical properties of hydrogen cannot be compared with the chemical properties of the other elements in Group 1A.

The second element in the first row, helium, is also a special atom. Helium is noticeably placed away from hydrogen at the far right-hand side of the periodic table. Helium's placement reflects its special chemical properties. While helium only has two valence electrons when the rest of the Group 8A elements have eight, helium is grouped with the 8A family because helium's chemical behavior is similar to the chemical behaviors of the other noble gases. This is because helium and the other noble gases all have a completed outer energy shell, even though the number of valence electrons is different.

This video describes the organization and family properties within the periodic table **(1b, 1c, 1f)**: http://www.youtube.com/watch?v=rD0maRGQhU8 (4:45).



ALKALINE EARTH METALS

- Group 2
- Metal
- Not as reactive as alkali metals
- Consist of 2 valence electrons
- Have a tendency to lose those 2 valence electrons to a non-metal – *forming cations*
    - Ex. Magnesium, calcium

MEDIA

Click image to the left for more content.

Lesson Summary

- The vertical columns on the periodic table are called groups or families because the elements in the column exhibit similar chemical behaviors.
- The groups on the periodic table also have similar electron configurations.
- The electron configuration in the outer energy level of the atoms in a family is the same, with the only difference in the electron configurations being the number of the outer energy level.
- All the members of a family of elements have the same number of valence electrons.
- The blocks of elements in the periodic table correspond to the sub-level in which the last electrons are being added.
- The horizontal rows on the periodic table are called periods.
- You can predict the energy level of an element's valence electrons using the element's period number in the periodic table.
- For the elements in Groups 1A, 2A, 3A, 4A, 5A, 6A, 7A, and 8A, the period in which an element is found corresponds to the energy level of that element's valence electrons.

Further Reading / Supplemental Links

This website provides a dynamic periodic table that also provides information about the orbitals and electron configurations of the elements.

- http://www.dayah.com/periodic
- A video on electron configuration and the periodic table.
    - http://video.yahoo.com/watch/193832/925988

This website reviews the different groups in the periodic table.

- http://www.chemtutor.com/perich.htm

This video is a ChemStudy film called "Transuranium Elements." The film is somewhat dated but the information is accurate.

- http://www.youtube.com/watch?v=SpQSEN1n4HE

This video is a ChemStudy film called "Inert Gas Compounds." The film is somewhat dated but the information is accurate.

- http://www.youtube.com/watch?v=QASDy4bLQxs

Review Questions

1. The horizontal rows in the periodic table are called _______________.
2. The vertical columns in the periodic table are called _______________.
3. Elements in a _______________ have similar chemical properties.
    a. period
    b. family
    c. both a and b
    d. neither a nor b
4. Which of the following elements is in the same family as fluorine?
    a. silicon
    b. antimony
    c. iodine
    d. arsenic
    e. None of these.
5. Which of the following is a noble gas?
    a. polonium
    b. astatine
    c. krypton
    d. hydrogen
    e. francium
6. Name the period (horizontal row) and the sub-level (s*p*, *d*, *f*) of the valence electrons for the following elements. Then indicate the number of valence electrons for that element.
    a. hydrogen
    b. carbon
    c. sulfur
    d. barium
    e. iodine
7. Of the s-block, p-block, d-block, and f-block of elements, which block contains most of the nonmetals?
8. Why are there no transition elements in periods 1 and 2?
9. Why do many of the transition elements have similar chemistry?
10. Use the periodic table to determine the energy level and sub-level of the highest energy electrons in each of the following elements:
    a. N
    b. Ca
    c. Rb
    d. P
    e. In
11. Decide whether each of the following statements is true or false.
    a. Li has valence electrons in the n
    b. Si has valence electrons in the n
    c. Ga has valence electrons in the n
    d. Xe has valence electrons in the n
    e. P has valence electrons in the n

CHAPTER 4 Chemical Periodicity

Chapter Outline

4.1 The Periodic Table

Lesson Objectives

The student will:

- explain the periodic law.
- describe the differences among metals, nonmetals, and metalloids.
- draw a rough sketch of the modern periodic table and indicate the portion of the table occupied by metals, nonmetals, and metalloids.
- identify the stair-step line that separates the metallic elements from the nonmetallic ones.

Vocabulary

- ductile
- malleable
- periodic law

Introduction

In the periodic table, the elements are arranged according to similarities in their properties. The elements are listed in order of increasing atomic number as you read from left to right across a period. In this chapter, you will learn the general behavior and trends within the periodic table that result from this arrangement in order to predict the properties of the elements.

The Periodic Law

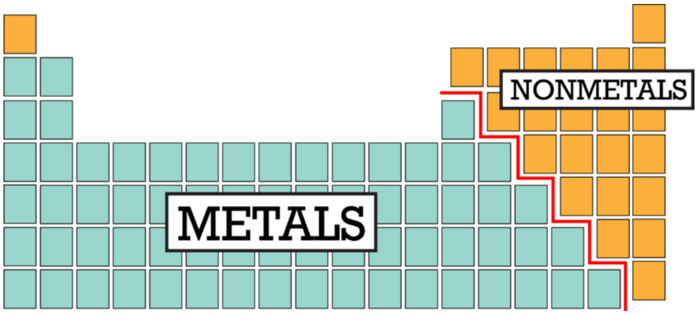
The periodic table is a powerful tool that provides a way for chemists to organize the chemical elements. The word "periodic" means happening or recurring at regular intervals. The **periodic law** states that the properties of the elements recur periodically as their atomic numbers increase. This is because the electron configurations of the atoms vary periodically with their atomic number. Since the physical and chemical properties of elements depend on their electron configurations, many of the physical and chemical properties of the elements also tend to repeat in a pattern.

You may recall that the periodic table was created by Russian scientist Dmitri Mendeleev. Mendeleev used similarities in properties to construct his periodic table of the elements. He was even able to predict the properties of several elements missing from the periodic table by using the properties of neighboring elements. Mendeleev arranged the elements in a table by increasing atomic weight, although he sometimes had to ignore the atomic weights in order to group elements with similar chemical behaviors together.

The work of Henry Moseley led to the arrangement of elements based on their properties and atomic numbers, not

Metals, Nonmetals, and Metalloids

There is a progression from metals to nonmetals across each period of elements in the periodic table. The diagonal line at the right side of the table shown below separates the elements into two groups: the metals and the nonmetals. The elements that are on the left of this line tend to be metals, while those to the right tend to be nonmetals. The elements that are directly on the diagonal line are metalloids. Metallic character generally increases from top to bottom down a group and right to left across a period. The noticeable exception is hydrogen, which is grouped with the metals but is actually a nonmetal.



Most of the chemical elements are metals. Most metals have the common properties of being shiny, being very dense, and having high melting points. Metals tend to be **ductile** (can be drawn out into thin wires) and **malleable** (can be hammered into thin sheets). Metals are good conductors of heat and electricity. All metals are solids at room temperature except for mercury. In chemical reactions, metals easily lose electrons to form positive ions. Examples of metals are silver, gold, and zinc.

Nonmetals are generally brittle, dull, and have low melting points. They are generally poor conductors of heat and electricity. In chemical reactions, they tend to gain electrons to form negative ions. Examples of nonmetals are hydrogen, carbon, and nitrogen.

Metalloids have properties of both metals and nonmetals. Metalloids can be shiny or dull. Electricity and heat can travel through metalloids, although not as easily as they can through metals. They are also called semi-metals. They are typically semiconductors, which means that they conduct electricity better than insulators, but not as well as conductors. Semiconductors are valuable in the computer chip industry. Examples of metalloids are silicon and boron.

Lesson Summary

- The word “periodic” means happening or recurring at regular intervals.
- The periodic law states that the properties of the elements recur periodically as their atomic numbers increase.
- There is a progression from metals to nonmetals across each period of elements in the periodic table.
- Metallic character generally increases from top to bottom down a group and right to left across a period.

Further Reading / Supplemental Links

This video is on the periodic table.

- http://www.youtube.com/watch?v=1geccHiylcU&feature=fvw

Review Questions

1. Why is the table of elements called "the periodic table"?
   a. It describes the periodic motion of celestial bodies.
   b. It describes the periodic recurrence of chemical properties.
   c. Because the rows are called periods.
   d. Because the elements are grouped as metals, metalloids, and nonmetals.
   e. None of these.
2. Which of the following elements is a nonmetal?
   a. oxygen
   b. lead
   c. iron
   d. zinc
   e. All of these are metals.
3. Which of the following metals is not
   a. gold
   b. silver
   c. copper
   d. bronze
   e. All of these are elements.
4. Which of the following elements is a metalloid?
   a. chlorine
   b. magnesium
   c. rhenium
   d. boron
   e. None of these.

4.2 Periodic Trends in Atomic Size

Lesson Objectives

The student will:

- define atomic radius.
- define the shielding effect.
- describe the factors that determine the trend in atomic size.
- describe the general trend in atomic size for groups and periods.
- use the general trends to predict the relative sizes of atoms.
- describe variations that occur in the general trend of atomic size in the transition metals.

Vocabulary

- atomic radius
- diatomic molecule
- nuclear charge
- shielding effect

Introduction

In the periodic table, there are a number of physical properties that are trend-like. This means that as you move down a group or across a period, you will see the properties changing in a general direction. The actual trends that are observed with atomic size have to do with three factors. These factors are:

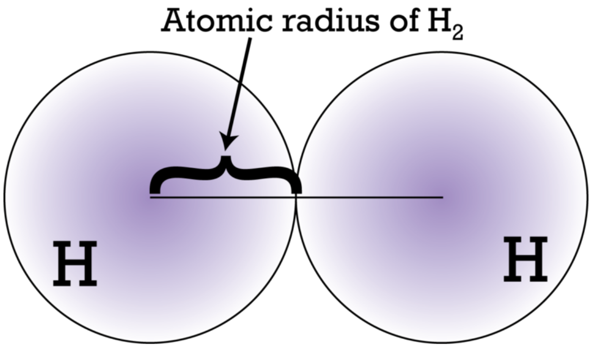
1. the number of protons in the nucleus (called the **nuclear charge**).
2. the number of energy levels holding electrons and the number of electrons in the outer energy level.
3. the number of electrons held between the nucleus and its outermost electrons (called the **shielding effect**).

Atomic Radius Defined

The gold foil experiment performed by Rutherford in 1911 (see the chapter "The Atomic Theory" for more details), was the first experiment that gave scientists an approximate measurement for the size of the atom. The atomic size is the distance from the nucleus to the valence shell, where the valence electrons are located. Using the technology available in the early part of the 1900s, Rutherford was able to determine quantitatively that the nucleus had a radius size smaller than 3×10^{-12} cm. The size of the atom is significantly larger, being approximately 2×10^{-8} cm in diameter.

The region in space occupied by the electron cloud of an atom is often thought of as a probability distribution of the electron positions. Consequently, there is no well-defined outer edge of the electron cloud. Because it is so difficult to measure atomic size from the nucleus to the outermost edge of the electron cloud, chemists use other approaches to get consistent measurements of atomic sizes.

Atomic size is defined in several different ways, which often produce some variations in the measurement of atomic sizes. One way that chemists define atomic size is by using the atomic radius. The **atomic radius** is one-half the distance between the centers of a homonuclear diatomic molecule, as illustrated below. A **diatomic molecule** is a molecule made of exactly two atoms, while homonuclear means both atoms are the same element.



Group Trends in Atomic Radii

Let's now look at how the atomic radii changes from the top of a family to the bottom. Take, for example, the Group 1A metals (see **Table** 4.1). Every atom in this family has the same number of electrons in the outer energy level (true for all main group families). Each period in the periodic table represents another added energy level. When we first learned about principal energy levels, we learned that each new energy level was larger than the one before. Therefore, as we move down the periodic table, each successive period represents the addition of a larger energy level, thus increasing the atomic radius.

TABLE 4.1: **Group 1A Data**

Element	**Number of Protons**	**Electron Configuration**
Lithium (Li)	3	[He]$2s^1$
Sodium (Na)	11	[Ne]$3s^1$
Potassium (K)	19	[Ar]$4s^1$
Rubidium (Rb)	37	[Kr]$5s^1$
Cesium (Cs)	55	[Xe]$6s^1$

One other contributing factor to atomic size is the shielding effect. The protons in the nucleus attract the valence electrons in the outer energy level, but the strength of this attraction depends on the size of the charges, the distance between the charges, and the number of electrons between the nucleus and the valence electrons. The presence of the core electrons weakens the attraction between the nucleus and the valence electrons. This weakening is called the shielding effect. Note that although valence electrons do participate in shielding, electrons in the same energy level do not shield each other as effectively as the core electrons do. As a result, the amount of shielding primarily depends on the number of electrons between the nucleus and the valence electrons. When the nucleus pulls strongly on the valence electrons, the valence shell can be pulled in tighter and closer to the nucleus. When the attraction is weakened by shielding, the valence shell cannot be pulled in as close. The more shielding that occurs, the further the valence shell can spread out.

For example, if you are looking at the element sodium, it has the electron configuration:

Na $1s^2 2s^2 2p^6 3s^1$

The outer energy level is $n = 3$. There is one valence electron, but the attraction between this lone valence electron and the nucleus, which has 11 protons, is shielded by the other 10 inner (or core) electrons.

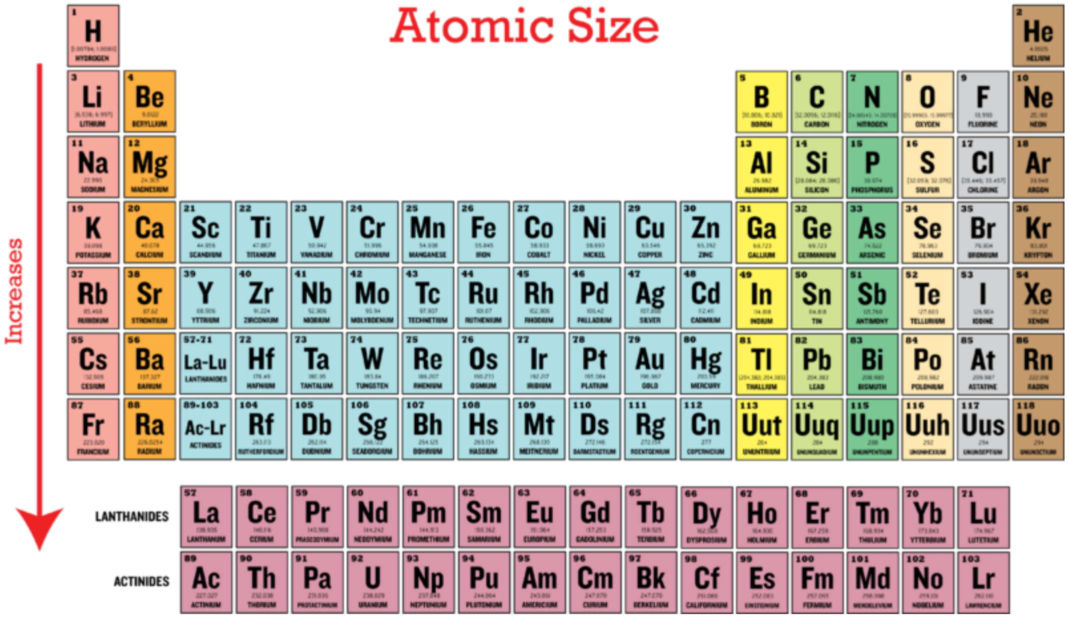
When we compare an atom of sodium to one of cesium, we notice that the number of protons increases, as well as the number of energy levels occupied by electrons. The increase in the number of protons, however, is also accompanied by the same increase in the number of shielding electrons.

Cs $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$

The result is that the valence electron in both atoms feels a similar pull from the nucleus, but the valence electron in the cesium atom is further from the nucleus because it is in a higher energy level. Compared to the shielding effect, the increase in the number of energy levels has a greater impact on the atom's size. Consequently, the size of a cesium atom is larger than that of a sodium atom.

This is true for not only Group 1A metals, but for all of the groups across the periodic table. For any given group, as you move downward in the periodic table, the size of the atoms increases. For instance, the largest atoms in the halogen family are bromine and iodine (astatine is radioactive and only exists for short periods of time, so we won't include it in the discussion). You can imagine that with the increase in the number of energy levels, the size of the atom must increase. The increase in the number of energy levels in the electron cloud takes up more space.

The periodic table below shows the trend of atomic size for groups, with the arrow indicating the direction of the increase.



Example:

Which of the following is larger? Explain.

1. As or Sb
2. Ca or Be
3. polonium or sulfur

Solution:

1. Sb, because it is below As in Group 15.
2. Ca, because it is below Be in Group 2.
3. Polonium, because it is below sulfur in Group 16.

Period Trends in Atomic Radii

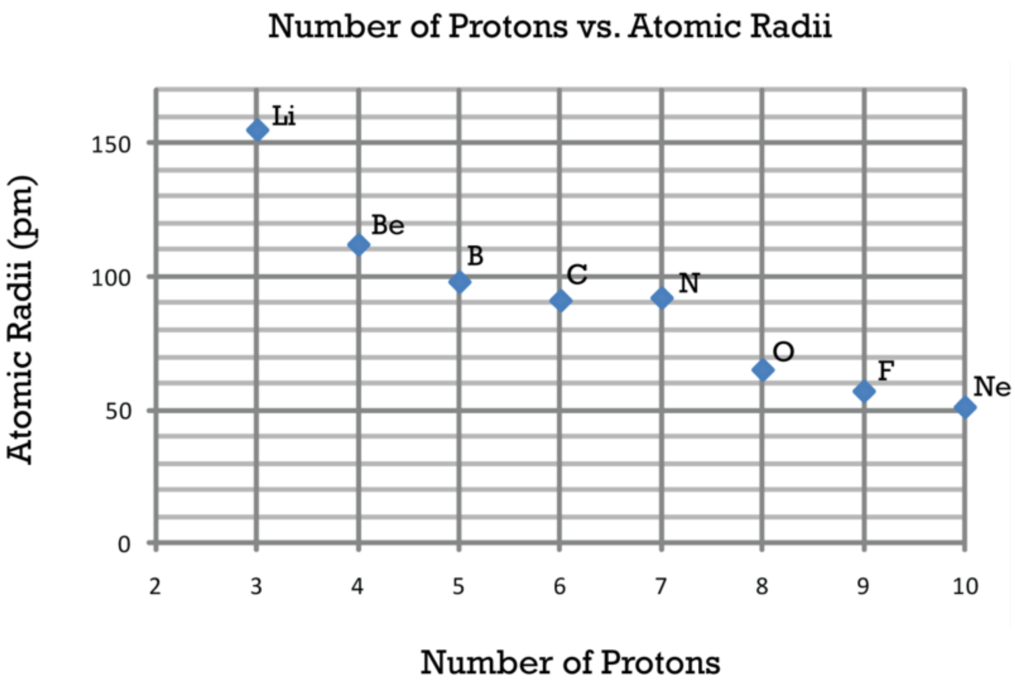
In order to determine the trend for the periods, we need to look at the number of protons (nuclear charge), the number of energy levels, and the shielding effect. For a row in the periodic table, the atomic number still increases (as it did for the groups), and thus the number of protons would increase. For a given period, however, we find that the outermost energy level does not change as the number of electrons increases. In period 2, for example, each additional electron goes into the second energy level, so the total number of energy levels does not go up. **Table** 4.2 shows the electron configuration for the elements in period 2.

TABLE 4.2: Electron Configurations for Elements in Period 2

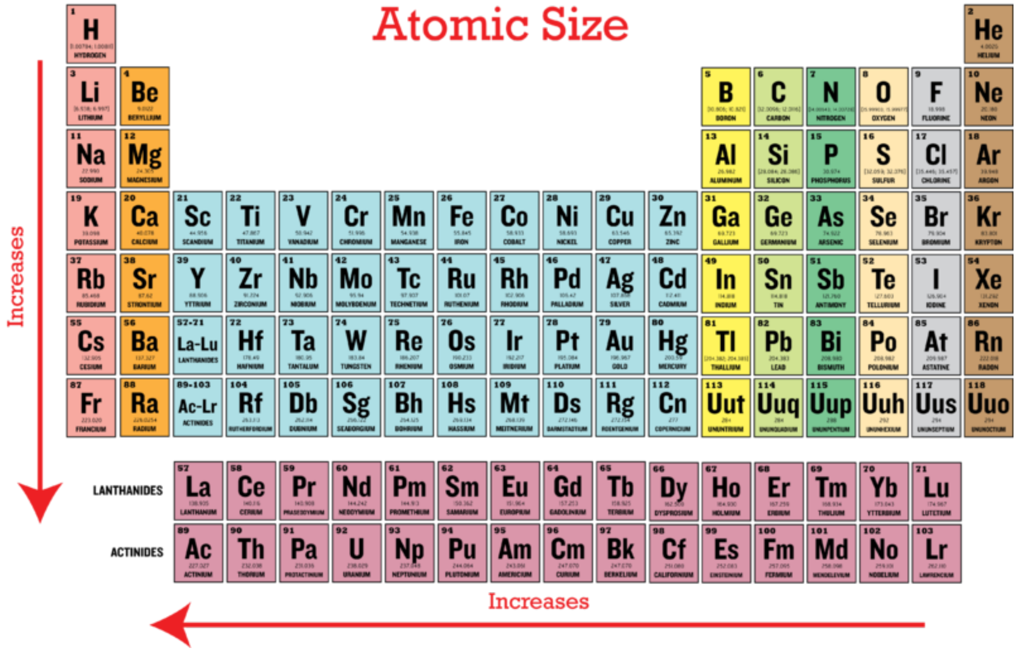
Element	**Number of Protons**	**Electron Configuration**
Lithium (Li)	3	$1s^2 2s^1$
Beryllium (Be)	4	$1s^2 2s^2$
Boron (B)	5	$1s^2 2s^2 2p^1$
Carbon (C)	6	$1s^2 2s^2 2p^2$
Nitrogen (N)	7	$1s^2 2s^2 2p^3$
Oxygen (O)	8	$1s^2 2s^2 2p^4$
Fluorine (F)	9	$1s^2 2s^2 2p^5$

Looking at the elements in period 2, the number of protons increases from three (for lithium) to nine (for fluorine). Therefore, the nuclear charge increases across a period. Meanwhile, the number of energy levels occupied by the electrons remains the same. How will this affect the radius? We know that every one of the elements in this period has two core electrons in the inner energy level (n

Consider the elements lithium, beryllium, and fluorine from period 2. With lithium, the two core electrons will shield the one valence electron from three protons. Beryllium has four protons being shielded by the $1s^2$ electrons. For fluorine, there are nine protons and nine electrons. All three of these elements have the same core electrons: the $1s^2$ electrons. As the number of protons increases, the nuclear charge increases. With an increase in nuclear charge, there is an increase in the pull between the protons and the outer level, pulling the outer electrons toward the nucleus. The amount of shielding from the nucleus does not increase because the number of core electrons remains the same. The net result is that the atomic size decreases going across the row. In the graph below, the values are shown for the atomic radii for period 2.

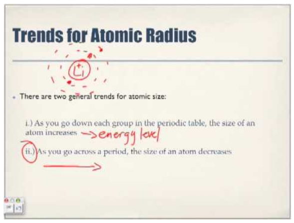


Let's add this new trend to the periodic table. In the diagram below, you will notice that the trend arrow for the period shows the atomic radii increase going from right to left, which is the equivalent to saying that the atomic radii decrease from left to right.



Considering these two trends, you will recognize that the largest atom, francium (atomic number 87), is at the bottom left-hand corner of the periodic table, while the smallest atom, helium (atomic number 2) is at the top right-hand corner of the table.

For an introduction to the electronic organization of the periodic table **(1c)**, see http://www.youtube.com/watch?v=35cWAxtHUGw (4:20).



Trends for Atomic Radius

- There are two general trends for atomic size:

i.) As you go down each group in the periodic table, the size of an atom increases → energy level

ii.) As you go across a period, the size of an atom decreases

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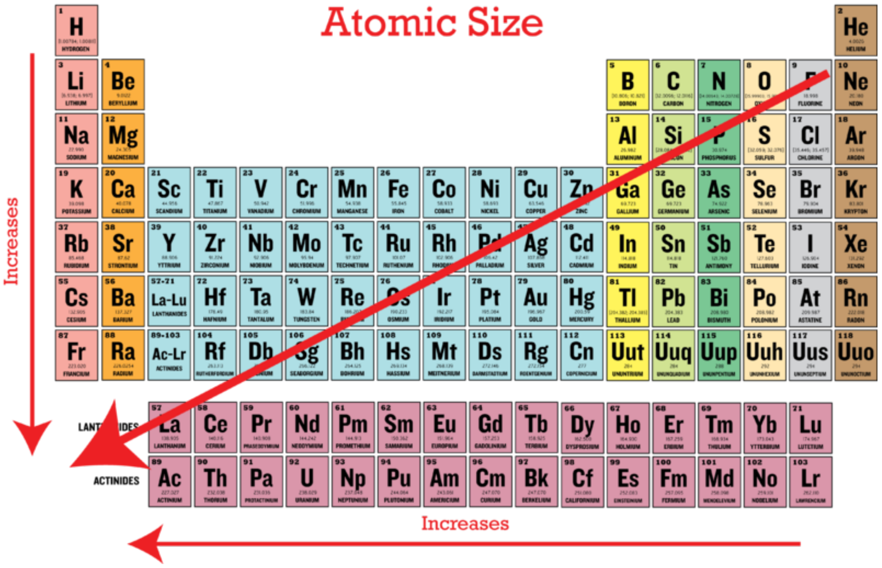
Atomic Radii of Transition Elements

The general trend for atomic radii in the periodic table would look similar to that illustrated in the diagram below. The elements with the smallest atomic radii are found in the upper right; those with the largest atomic radii are found in the lower left.



Until now, we have worked solely with the main group elements. Let's consider how three factors affecting atomic size affect transition metals. The first row of the transition metals all contain electrons in the 3d*d* metals. **Table** 4.3 shows the electron configuration for the ten elements in this row. The number of protons is increasing, so the nuclear charge is increasing.

TABLE 4.3: Electron Configuration for 3

Element	Number of Protons	Electron Configuration
Scandium (Sc)	21	$[Ar]3d^1 4s^2$
Titanium (Ti)	22	$[Ar]3d^2 4s^2$
Vanadium (V)	23	$[Ar]3d^3 4s^2$
Chromium (Cr)	24	$[Ar]3d^5 4s^1$
Manganese (Mn)	25	$[Ar]3d^5 4s^2$
Iron (Fe)	26	$[Ar]3d^6 4s^2$
Cobalt (Co)	27	$[Ar]3d^7 4s^2$
Nickel (Ni)	28	$[Ar]3d^8 4s^2$
Copper (Cu)	29	$[Ar]3d^{10} 4s^1$
Zinc (Zn)	30	$[Ar]3d^{10} 4s^2$

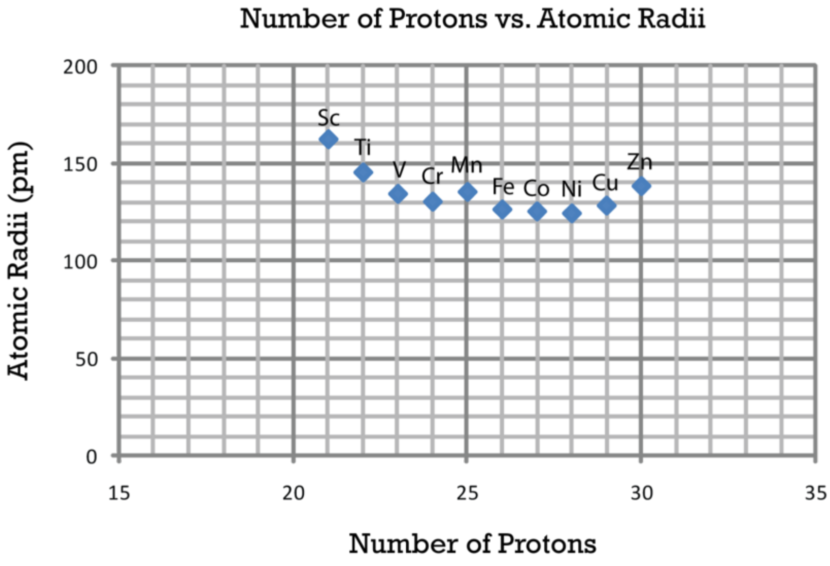
You may notice that some of these configurations are not what you would expect based on the information presented so far. Both chromium and copper have one of the 4s*d* orbital. A simplified explanation for these unusual electron configurations is that the *d* sub-level is particularly stable when it is half-full (5 electrons) or completely full (10 electrons). Since the 4*s* and 3*d* orbitals are close in energy, this added stabilization is enough to change the location of one valence electron.

TABLE 4.4: **Atomic Radii for 3d Metals**

Element	**Atomic Radii (pm)**
Scandium (Sc)	164
Titanium (Ti)	147
Vanadium (V)	135
Chromium (Cr)	129
Manganese (Mn)	137
Iron (Fe)	126
Cobalt (Co)	125
Nickel (Ni)	125
Copper (Cu)	128
Zinc (Zn)	137

Table 4.4 lists the atomic radii for the first row of the transition metals. It can be seen from this table that the period trend in atomic radii is not followed as closely by the transition metals. Since we are adding electrons to the 3*d* orbitals, we are actually adding to the core electrons and *not* to the valence orbitals. Although the nuclear charge is going up, the shielding is also increasing with each added electron. Because of this, there is less atomic contraction throughout the transition metals.

The graph of the number of protons versus the atomic radii for the 3d



Lesson Summary

- Atomic size is the distance from the nucleus to the valence shell.
- Atomic size is difficult to measure because it has no definite boundary.

- Atomic radius is a more definite and measurable way of defining atomic size. It is half the distance from the center of one atom to the center of another atom in a homonuclear diatomic molecule.
- There are three factors that help in the prediction of the trends in the periodic table: number of protons in the nucleus, number of energy levels, and the shielding effect.
- The atomic radii increase from top to the bottom in any group.
- The atomic radii decrease from left to right across a period.
- This trend is not as systematic for the transition metals because other factors come into play.

Review Questions

1. Why is the atomic size considered to have "no definite boundary"?
2. How is atomic size measured?
    a. using a spectrophotometer
    b. using a tiny ruler (called a nano ruler)
    c. indirectly
    d. directly
3. Which of the following would be smaller: indium or gallium?
4. Which of the following would be smaller: potassium or cesium?
5. Which of the following would be smaller: titanium or polonium?
6. Explain why iodine is larger than bromine.
7. What are three factors that affect atomic size?
8. Which of the following would have the largest atomic radius?
    a. Si
    b. C
    c. Sn
    d. Pb
9. Which of the following would have the smallest atomic radius?
    a. $1s^2 2s^2$
    b. $1s^2 2s^2 2p^6 3s^1$
    c. $1s^2$
    d. $1s^1$
10. Arrange the following in order of increasing atomic radius: Tl, B, Ga, Al, In.
11. Arrange the following in order of increasing atomic radius: Ga, Sn, C.
12. Which of the following would be larger: Rb or Sn?
13. Which of the following would be larger: Ca or As?
14. Describe the trend for the atomic size of elements in a row in the periodic table.
15. Which of the following would have the largest atomic radius?
    a. Sr
    b. Sn
    c. Rb
    d. In
16. Which of the following would have the smallest atomic radius?
    a. K
    b. Kr
    c. Ga
    d. Ge

17. Arrange the following in order of decreasing atomic radius: Ba, Tl, Se, Bi, Cs.

4.3 Periodic Trends in Ionic Size

Lesson Objectives

The student will:

- explain what an ion is.
- describe how cations and anions are formed.
- describe the factors that determine the trend in ionic size.
- describe the trend in ionic size for elements.
- use the general trends to predict the relative sizes of ions.

Vocabulary

- anion
- cation
- ion

Introduction

An atom is electrically neutral, which means that the number of protons is equal to the number of electrons. In chemical reactions, however, atoms can gain or lose electrons. This results in the formation of an ion. An **ion** is an atom with a positive or negative charge.

Atoms and Ions

Atoms of metallic elements tend to form positive ions by losing one or more electrons. A positive ion is called a **cation** (pronounced CAT-ion) and has fewer electrons than an electrically neutral atom. For example, an atom of sodium has eleven protons and eleven electrons. Its electron configuration is $[\text{Ne}]3s^1$. Sodium has one valence electron surrounding a stable core of ten electrons. In chemical reactions, a sodium atom tends to lose its one valence electron to become a sodium cation. Because this sodium ion has eleven protons and only ten electrons, it has a net charge of $+1$. An atom that loses two electrons will become an ion with a charge of $+2$, and an atom that loses three electrons will become an ion with a charge of $+3$.

Atoms of nonmetallic elements tend to form negative ions by gaining one or more electrons. A negative ion is called an **anion** (pronounced AN-ion). For example, an atom of fluorine has seventeen protons and seventeen electrons. Its electron configuration is $1s^2 2s^2 2p^5$, and it has seven valence electrons. In chemical reactions, a fluorine atom tends to gain one valence electron, becoming a fluoride anion. (Notice that the name of anions typically end in "-ide.") Because the fluoride ion has seventeen positive protons and eighteen negative electrons, it has a net charge of -1.

An atom that gains two electrons will become an ion with a charge of -2, and an atom that gains three electrons will become an ion with a charge of -3.

Group and Period Trends in Ionic Size

Cations are smaller than the atoms from which they are formed. The loss of outer shell electrons results in increased attraction between the nucleus and the remaining electrons. This results in less electron-electron repulsion and allows the nucleus and the electrons to come closer together. When compared to a neutral atom of sodium, Na, a sodium cation, symbolized by Na^+, has a smaller size.

Anions are larger than the atoms from which they are formed. The gain of outer shell electrons results in decreased attraction between the nucleus and the remaining electrons, and electron-electron repulsion forces them to spread apart. When compared to a neutral atom of fluorine, F, a fluoride anion, symbolized by F^-, has a larger size.

If we examine the ionic sizes of just the metals in the main group, we will find that the trends are the same as the trends in atomic sizes for the neutral elements. All the elements for each group of metals lose the same number of electrons, which means that the ionic sizes will be primarily affected by the number of energy levels in the electron cloud. Since the number of energy levels still increases from top to bottom, the ionic size also increases down a group of elements in the periodic table. For similar reasons, the trend across a period is the same for both ions and neutral atoms. All the metal elements in a given period will lose their outer shell electrons but still have the same number of core electrons. As a result, the nuclear charge increases from left to right, while the number of core electrons remains the same. This means that the ion size will decrease from left to right across a period.

Nonmetals also see the same trends in size as the neutral elements. The negative ions increase in size as you move down a group and decrease in size as you move from left to right across a period. In other words, as you go from top to bottom down a group or left to right across a period, the ionic size decreases as long as

Lesson Summary

- In chemical reactions, atoms can gain or lose electrons. This results in the formation of an ion. An ion is basically an atom with a positive or negative charge.
- Atoms of metallic elements tend to form positive ions (cations) by losing one or more electrons.
- Atoms of nonmetallic elements tend to form negative ions (anions) by gaining one or more electrons.
- Cations are smaller than the atoms from which they are formed.
- Anions are larger than the atoms from which they were formed.
- Ionic size increases from top to bottom down a group of elements in the periodic table.
- From left to right across a period, the ionic size decreases as long as you are comparing all metals or all nonmetals. Between the metals and nonmetals, the ionic size increases as you switch from cations to anions.

Review Questions

1. How is the size of a cation different from the size of the atom from which it was formed? Why?
2. How is the size of an anion different from the size of the atom from which it was formed? Why?
3. Mg^{2+} has the same number of electrons as F^- (they are said to be isoelectric). Which ion is larger and why?
4. Which of the following has the smallest ionic radius?
    a. O^{2-}

   b. S^{2-}
   c. Mg^{2+}
   d. Ca^{2+}

5. Which of the following has the largest ionic radius?
   a. Ba^{2+}
   b. Cs^{+}
   c. I^{-}
   d. Te^{2-}

4.4 Periodic Trends in Ionization Energy

Lesson Objectives

The student will:

- define ionization energy.
- describe the trends that exist in the periodic table for ionization energy.
- use the general trends to predict the relative ionization energies of atoms.

Introduction

Atoms are capable of forming ions by either losing or gaining electrons. Since the electrons are attracted to the positively charged nucleus, energy is needed to pull the electron away from the nucleus. In this lesson, we will gain an understanding of the energy required to remove an electron and recognize its trend on the periodic table.

Vocabulary

- effective nuclear charge
- ionization energy

Ionization Energy Defined

Consider lithium, which has an electron configuration of $1s^22s^1$ and has one electron in its outermost energy level. In order to remove this electron, energy must be added to the system. Look at the equation below:

$$\begin{array}{ccccccc} \text{energy} & + & \underset{1s^22s^1}{\text{Li}_{(g)}} & \rightarrow & \underset{1s^2}{\text{Li}^+_{(g)}} & + e^- \end{array}$$

With the addition of energy, a lithium atom can lose one electron and form a lithium ion. This energy is known as the ionization energy. The **ionization energy** is the energy required to remove the most loosely held electron from a gaseous atom or ion. The higher the value of the ionization energy, the harder it is to remove that electron. In the equation above, the subscript "g" indicates that the element is in the form of a gas. The definition for ionization energy specifies "in the gaseous phase" because when the atom or ion is in the liquid or solid phases, other factors are involved. The general equation for the ionization energy is as follows.

$$A_{(g)} + \text{energy} \rightarrow A^+_{(g)} + e^- \quad \text{IE}_1$$

If a second electron is to be removed from an atom, the general equation for the ionization energy is as follows:

$$A^{+}_{(g)} + \text{energy} \rightarrow A^{2+}_{(g)} + e^{-} \quad \text{IE}_2$$

After the first electron is removed, there are a greater number of protons than electrons. As a result, when a second electron is being removed, the energy required for the second ionization (IE_2) will be greater than the energy required for the first ionization (IE_1). In other words, $\text{IE}_1 < \text{IE}_2 < \text{IE}_3 < \text{IE}_4$.

Group and Period Trends in Ionization Energy

We can see a trend when we look at the ionization energies for the elements in period 2. **Table** 4.5 summarizes the electron configuration and the ionization energies for the elements in the second period.

TABLE 4.5: First Ionization Energies for Period 2 Main Group Elements

Element	**Electron Configuration**	**First Ionization Energy, IE_1**
Lithium (Li)	$[\text{He}]2s^1$	520 kJ/mol
Beryllium (Be)	$[\text{He}]2s^2$	899 kJ/mol
Boron (B)	$[\text{He}]2s^22p^1$	801 kJ/mol
Carbon (C)	$[\text{He}]2s^22p^2$	1086 kJ/mol
Nitrogen (N)	$[\text{He}]2s^22p^3$	1400 kJ/mol
Oxygen (O)	$[\text{He}]2s^22p^4$	1314 kJ/mol
Fluorine (F)	$[\text{He}]2s^22p^5$	1680 kJ/mol

We can see that as we move across the period from left to right, in general the ionization energy increases. At the beginning of the period with the alkali metals and the alkaline earth metals, losing one or two electrons allows these atoms to become ions.

$$\begin{array}{ccccc} \text{energy} & + & \underset{[\text{He}]2s^1}{\text{Li}_{(g)}} & \rightarrow & \underset{[\text{He}]}{\text{Li}^{+}_{(g)}} & + e^{-} \end{array}$$

$$\begin{array}{ccccc} \text{energy} & + & \underset{[\text{Ne}]3s^12}{\text{Mg}_{(g)}} & \rightarrow & \underset{[\text{Ne}]}{\text{Mg}^{2+}_{(g)}} & + 2e^{-} \end{array}$$

As we move across the period, the atoms become smaller, which causes the nucleus to have greater attraction for the valence electrons. Therefore, the electrons are more difficult to remove.

A similar trend can be seen for the elements within a family. **Table** 4.6 shows the electron configuration and the first ionization energies (IE_1) for some of the elements in the first group, the alkali metals.

TABLE 4.6: First Ionization Energies for Some Period 1 Elements

Element	**Electron Configuration**	**First Ionization Energy, IE_1**
Lithium (Li)	$[\text{He}]2s^1$	520 kJ/mol
Sodium (Na)	$[\text{Ne}]3s^1$	495.5 kJ/mol
Potassium (K)	$[\text{Ar}]4s^1$	418.7 kJ/mol

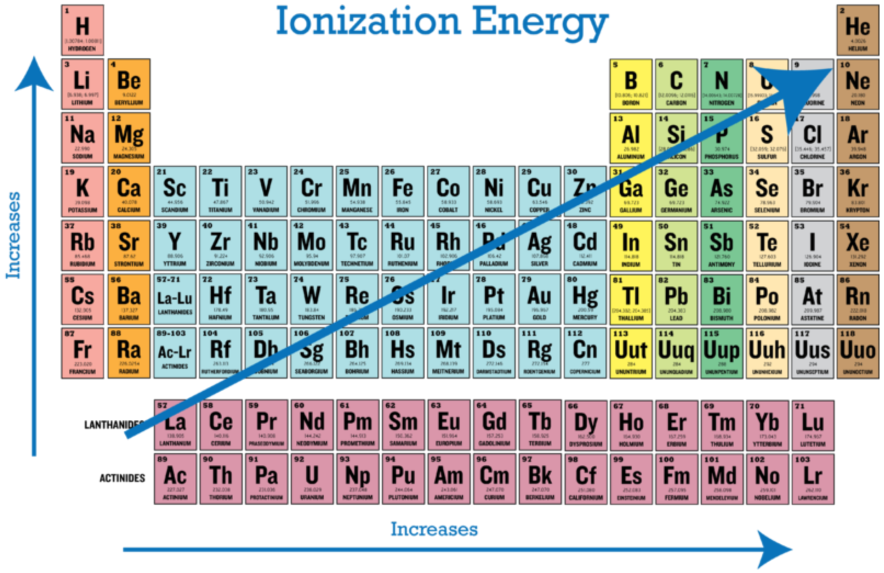
By comparing the electron configurations of lithium to potassium, we know that the valence electron is further away from the nucleus. We know this because the nIE_1 for potassium is less than IE_1 for lithium.

Why does the ionization energy increase going across a period? It has to do with two factors. One factor is that the atomic size decreases. The second factor is that the effective nuclear charge increases. The **effective nuclear charge** is the charge experienced by a specific electron within an atom. Recall that the nuclear charge was used to describe why the atomic size decreased going across a period. **Table** 4.7 shows the effective nuclear charge along with the ionization energy for the elements in period 2.

TABLE 4.7: **Effective Nuclear Charge for Period 2 Main Group Elements**

Element	**Electron Configuration**	**Number of Protons**	**Number of Core Electrons**	**Effective Nuclear Charge**	**Ionization Energy**
Lithium (Li)	[He]$2s^1$	3	2	1	520 kJ/mol
Beryllium (Be)	[He]$2s^2$	4	2	2	899 kJ/mol
Boron (B)	[He]$2s^22p^1$	5	2	3	801 kJ/mol
Carbon (C)	[He]$2s^22p^2$	6	2	4	1086 kJ/mol
Nitrogen (N)	[He]$2s^22p^3$	7	2	5	1400 kJ/mol
Oxygen (O)	[He]$2s^22p^4$	8	2	6	1314 kJ/mol
Fluorine (F)	[He]$2s^22p^5$	9	2	7	1680 kJ/mol

The electrons that are shielding the nuclear charge are the core electrons, which are the $1s^2$ electrons for period 2. The effective nuclear charge is approximately the difference between the total nuclear charge and the number of core electrons. Notice that as the effective nuclear charge increases, the ionization energy also increases. Overall, the general trend for ionization energy is summarized in the diagram below.



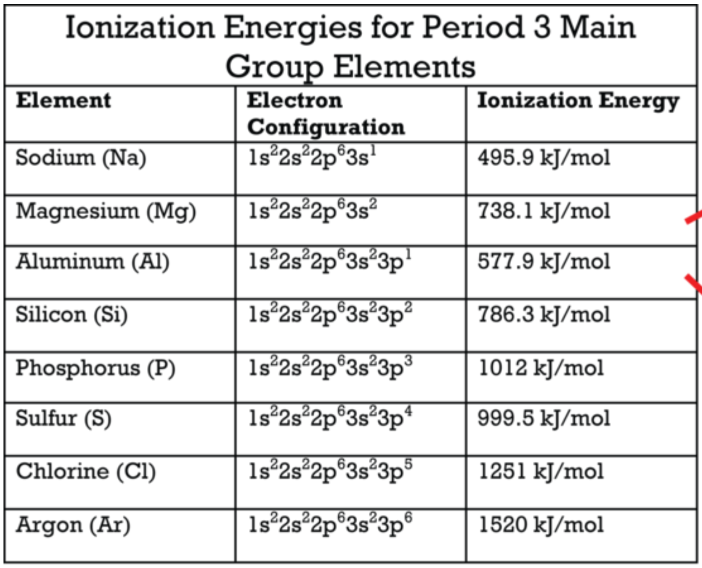
Example:

What would be the effective nuclear charge for chlorine? Would you predict the ionization energy to be higher or lower than the ionization energy for fluorine?

Solution:

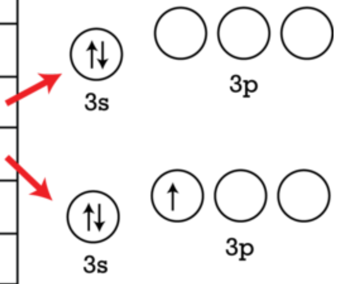
Chlorine has the electron configuration: $Cl = [Ne]3s^23p^5$. The effective nuclear charge is 7, which is the same as the nuclear charge for fluorine. Predicting the ionization energy with just this information would be difficult. The atomic size, however, is larger for chlorine than it is for fluorine because chlorine has three energy levels (chlorine is in period 3). Now we can conclude that the ionization energy for chlorine should be lower than that of fluorine because the electron would be easier to pull off when it is further away from the nucleus. (Indeed, the value for the first ionization energy of chlorine is 1251 kJ/mol, compared to 1680 kJ/mol for fluorine.)

A few anomalies exist with respect to the ionization energy trends. Going across a period, there are two ways in which the ionization energy may be affected by the electron configuration. When we look at period 3, we can see that there is an anomaly as we move from the 3*s* sublevels to the 3*p* sublevel. The table below shows the electron configurations and first ionization energy for the main group elements in period 3.



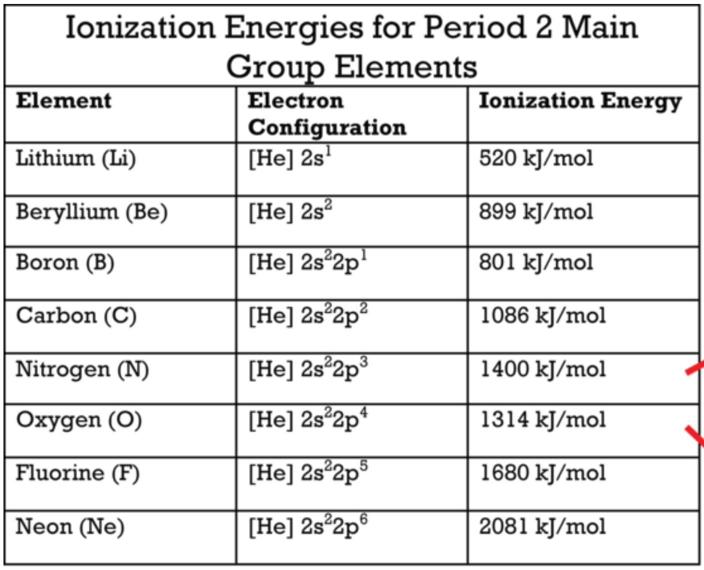
Ionization Energies for Period 3 Main Group Elements

Element	Electron Configuration	Ionization Energy
Sodium (Na)	$1s^22s^22p^63s^1$	495.9 kJ/mol
Magnesium (Mg)	$1s^22s^22p^63s^2$	738.1 kJ/mol
Aluminum (Al)	$1s^22s^22p^63s^23p^1$	577.9 kJ/mol
Silicon (Si)	$1s^22s^22p^63s^23p^2$	786.3 kJ/mol
Phosphorus (P)	$1s^22s^22p^63s^23p^3$	1012 kJ/mol
Sulfur (S)	$1s^22s^22p^63s^23p^4$	999.5 kJ/mol
Chlorine (Cl)	$1s^22s^22p^63s^23p^5$	1251 kJ/mol
Argon (Ar)	$1s^22s^22p^63s^23p^6$	1520 kJ/mol



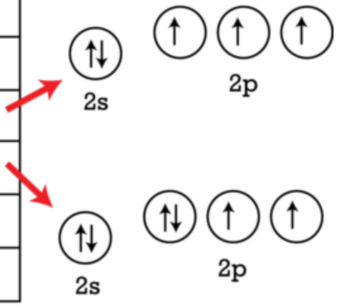
In the table, we see that when we compare magnesium to aluminum, the IE_1 decreases instead of increases. Why is this? Magnesium has its outermost electrons in the 3*s* sub-level. The aluminum atom has its outermost electron in the 3*p* sublevel. Since *p* electrons have just slightly more energy than *s* electrons, it takes a little less energy to remove that electron from aluminum. One other factor is that the electrons in $3s^2$ shield the electron in $3p^1$. These two factors allow the IE_1 for aluminum to be less than IE_1 for magnesium.

When we look again at the table, we can see that the ionization energy for nitrogen also does not follow the general trend.



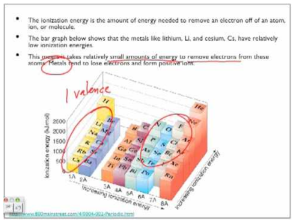
Ionization Energies for Period 2 Main Group Elements

Element	Electron Configuration	Ionization Energy
Lithium (Li)	[He] $2s^1$	520 kJ/mol
Beryllium (Be)	[He] $2s^2$	899 kJ/mol
Boron (B)	[He] $2s^22p^1$	801 kJ/mol
Carbon (C)	[He] $2s^22p^2$	1086 kJ/mol
Nitrogen (N)	[He] $2s^22p^3$	1400 kJ/mol
Oxygen (O)	[He] $2s^22p^4$	1314 kJ/mol
Fluorine (F)	[He] $2s^22p^5$	1680 kJ/mol
Neon (Ne)	[He] $2s^22p^6$	2081 kJ/mol



While nitrogen has one electron occupying each of the three p2*p* orbitals. The presence of two electrons in an orbital lead to greater electron-electron repulsion experienced by these $2p$ electrons, which lowers the amount of energy needed to remove one of these electrons. Therefore, IE_1 for oxygen is less than that for nitrogen.

This video discusses the ionization energy trends in the periodic table **(1c)**: http://www.youtube.com/watch?v=xE9YOBXdTSo (9:25).



- The ionization energy is the amount of energy needed to remove an electron off of an atom, ion, or molecule.
- The bar graph below shows that the metals like lithium, Li, and cesium, Cs, have relatively low ionization energies.
- This means it takes relatively small amounts of energy to remove electrons from these atoms. Metals tend to lose electrons and form positive ions.

http://www.800mainstreet.com/4/0004-002-Periodic.html

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Lesson Summary

- Ionization energy is the energy required to remove the most loosely held electron from a gaseous atom or ion. Ionization energy generally increases across a period and decreases down a group.
- Once one electron has been removed, a second electron can be removed, but $\text{IE}_1 < \text{IE}_2$. If a third electron is removed, $\text{IE}_1 < \text{IE}_2 < \text{IE}_3$, and so on.
- The effective nuclear charge is the charge of the nucleus felt by the valence electrons.
- The effective nuclear charge and the atomic size help explain the trend of ionization energy. Going down a group, the atomic size gets larger and the electrons can be more readily removed. Therefore, ionization energy decreases down a group. Going across a period, both the effective nuclear charge and the ionization energy increases, because the electrons are harder to remove.

Review Questions

1. Define ionization energy and write the general ionization equation.
2. Which of the following would have the largest ionization energy?
    a. Na
    b. Al
    c. H
    d. He
3. Which of the following would have the smallest ionization energy?
    a. K
    b. P
    c. S
    d. Ca
4. Place the following elements in order of increasing ionization energy: Na, O, Mg, Ne, K.
5. Place the following elements in order of decreasing ionization energy: N, Si, P, Mg, He.
6. Using experimental data, the first ionization energy for an element was found be 600 kJ/mol. The second ionization energy was found to be 1800 kJ/mol. The third, fourth, and fifth ionization energies were found to be, respectively, 2700 kJ/mol, $11,600$ kJ/mol, and $15,000$ kJ/mol. To which family of elements does this element belong? Explain.
7. Using electron configurations and your understanding of ionization energy, which would you predict to have a higher second ionization energy: Na or Mg?

8. Comparing the first ionization energies of Ca and Mg,
    a. calcium has a higher ionization energy because its radius is smaller.
    b. magnesium has a higher ionization energy because its radius is smaller.
    c. calcium has a higher ionization energy because it outermost sub-energy level is full.
    d. magnesium has a higher ionization energy because it outermost sub-energy level is full.
    e. they have the same ionization energy because they have the same number of valence electrons.
9. Comparing the first ionization energies of Be and B,
    a. beryllium has a higher ionization energy because its radius is smaller.
    b. boron has a higher ionization energy because its radius is smaller.
    c. beryllium has a higher ionization energy because it outermost sub-energy level is full.
    d. boron has a higher ionization energy because it outermost sub-energy level is full.
    e. they have the same ionization energy because boron only has one extra valence electron.

4.5 Periodic Trends in Electronegativity

Lesson Objectives

The student will:

- define electronegativity.
- describe the trends that exist in the periodic table for electronegativity.
- use the general trends to predict the relative electronegativities of atoms.

Vocabulary

- electronegativity

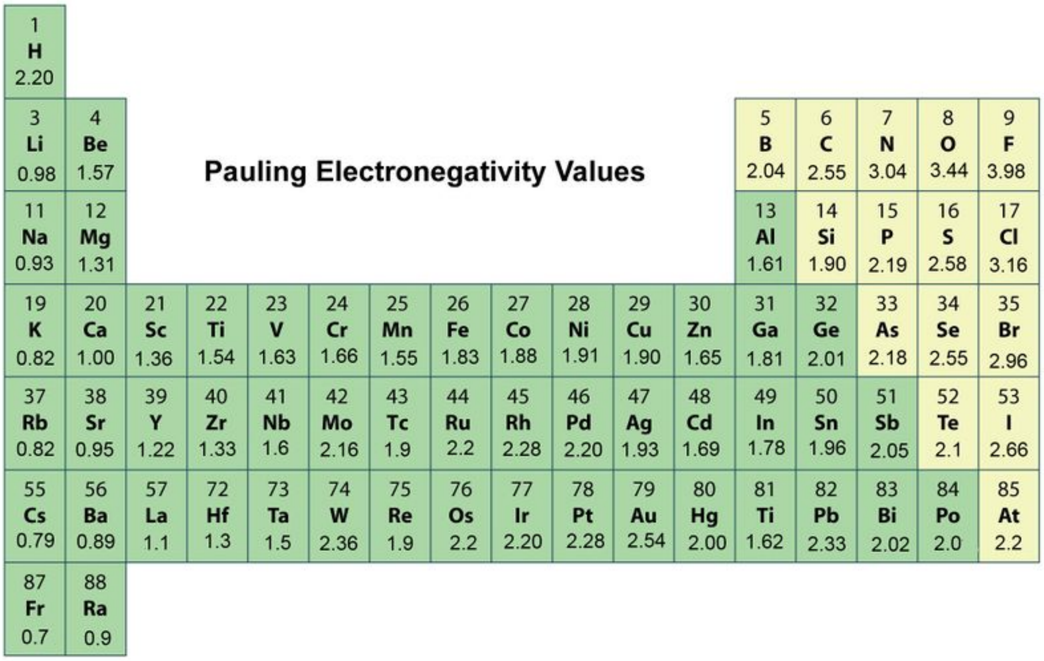
Introduction

Around 1935, the American chemist Linus Pauling developed a scale to describe the attraction an element has for electrons in a chemical bond. In this lesson, we will gain an understanding of this concept and recognize its trend on the periodic table.

Electronegativity Defined

In a molecule, some electrons are shared between the atoms making up the molecule. The ability of an atom in a molecule to attract shared electrons is called **electronegativity**. The higher the electronegativity of an atom, the greater its ability to attract shared electrons. The electronegativity of atoms has been defined in several ways. One method that is widely accepted is that developed by Linus Pauling.

On the Pauling scale, shown below, fluorine is the most electronegative element with an electronegativity of close to 4.0, and cesium and francium are the least electronegative with electronegativities of around 0.7.



Pauling Electronegativity Values

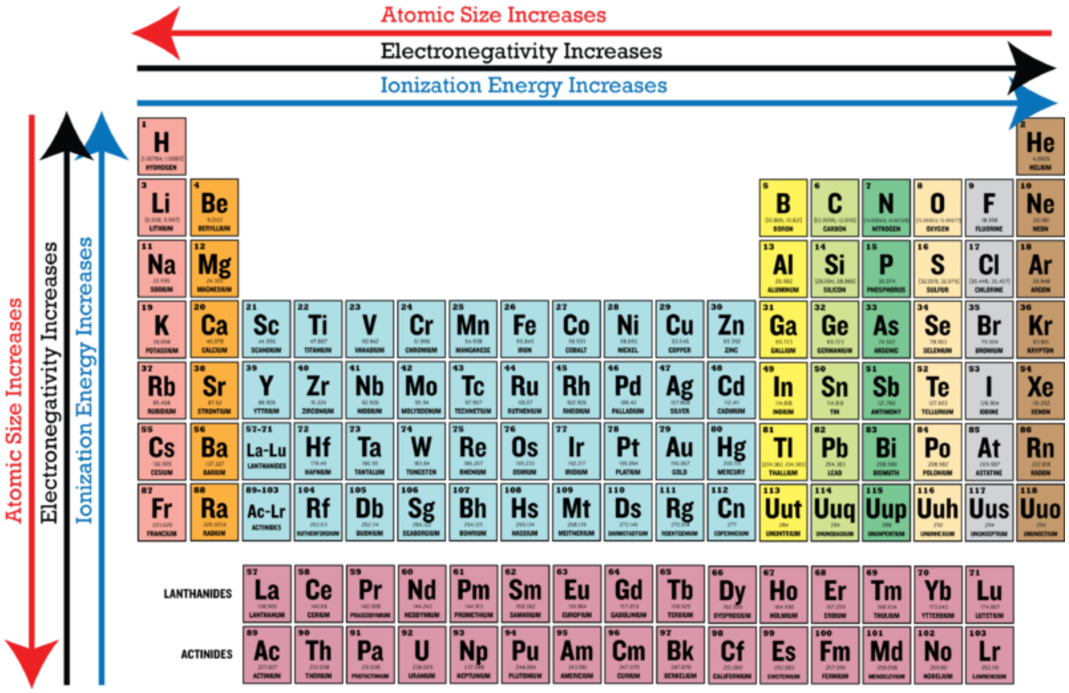
1 **H** 2.20																
3 **Li** 0.98	4 **Be** 1.57											5 **B** 2.04	6 **C** 2.55	7 **N** 3.04	8 **O** 3.44	9 **F** 3.98
11 **Na** 0.93	12 **Mg** 1.31											13 **Al** 1.61	14 **Si** 1.90	15 **P** 2.19	16 **S** 2.58	17 **Cl** 3.16
19 **K** 0.82	20 **Ca** 1.00	21 **Sc** 1.36	22 **Ti** 1.54	23 **V** 1.63	24 **Cr** 1.66	25 **Mn** 1.55	26 **Fe** 1.83	27 **Co** 1.88	28 **Ni** 1.91	29 **Cu** 1.90	30 **Zn** 1.65	31 **Ga** 1.81	32 **Ge** 2.01	33 **As** 2.18	34 **Se** 2.55	35 **Br** 2.96
37 **Rb** 0.82	38 **Sr** 0.95	39 **Y** 1.22	40 **Zr** 1.33	41 **Nb** 1.6	42 **Mo** 2.16	43 **Tc** 1.9	44 **Ru** 2.2	45 **Rh** 2.28	46 **Pd** 2.20	47 **Ag** 1.93	48 **Cd** 1.69	49 **In** 1.78	50 **Sn** 1.96	51 **Sb** 2.05	52 **Te** 2.1	53 **I** 2.66
55 **Cs** 0.79	56 **Ba** 0.89	57 **La** 1.1	72 **Hf** 1.3	73 **Ta** 1.5	74 **W** 2.36	75 **Re** 1.9	76 **Os** 2.2	77 **Ir** 2.20	78 **Pt** 2.28	79 **Au** 2.54	80 **Hg** 2.00	81 **Ti** 1.62	82 **Pb** 2.33	83 **Bi** 2.02	84 **Po** 2.0	85 **At** 2.2
87 **Fr** 0.7	88 **Ra** 0.9															

Group and Period Trends in Electronegativity

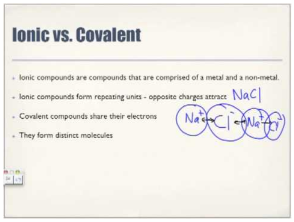
The electronegativity of atoms increases as you move from left to right across a period in the periodic table. This is because as you go from left to right across a period, the nuclear charge is increasing faster than the electron shielding, so the attraction that the atoms have for the valence electrons increases.

The electronegativity of atoms decreases as you move from top to bottom down a group in the periodic table. This is because as you go from top to bottom down a group, the atoms of each element have an increasing number of energy levels. The electrons in a bond are thus farther away from the nucleus and are held less tightly.

Atoms with low ionization energies have low electronegativities because their nuclei do not have a strong attraction for electrons. Atoms with high ionization energies have high electronegativities because the nucleus has a strong attraction for electrons.



Here is another video that describes ionization energy trends in the periodic table **(1c)**: http://www.youtube.com/watch?v=q3AiM1BYX-c (9:39).



Ionic vs. Covalent

- Ionic compounds are compounds that are comprised of a metal and a non-metal.
- Ionic compounds form repeating units - opposite charges attract NaCl
- Covalent compounds share their electrons
- They form distinct molecules

MEDIA

Click image to the left for more content.

Lesson Summary

- American chemist Linus Pauling developed the electronegativity scale to describe the attraction an element has for electrons in a chemical bond.
- The higher the electronegativity of an atom, the greater its ability to attract shared electrons.
- The electronegativity of atoms increases as you move from left to right across a period in the periodic table.
- The electronegativity of atoms decreases as you move from top to bottom down a group in the periodic table.

Further Reading / Supplemental Links

A series of selectable videos that show the properties and discuss the bonding of various elements.

- http://www.periodicvideos.com/#

Review Questions

1. Define electronegativity.
2. Choose the element in each pair that has the lower electronegativity.
    a. Li or N
    b. Cl or Na
    c. Ca or K
    d. Mg or F
3. Which of the following will have the largest electronegativity?
    a. Se
    b. F
    c. Ne
    d. Br
4. Which of the following will have the smallest electronegativity?
    a. Na
    b. Ne
    c. Al
    d. Rb
5. Describe the general trend for electronegativity in period 2.

4.6 Periodic Trends in Electron Affinity

Lesson Objectives

The student will:

- define electron affinity.
- describe the trends for electron affinity in the periodic table.

Vocabulary

- electron affinity

Introduction

We have talked about atomic structure, electronic configurations, size of atoms and ions, ionization energy, and electronegativity. The final periodic trend that we will examine is how atoms gain electrons.

Electron Affinity Defined

Atoms can gain or lose electrons. When an atom gains an electron, energy is given off and is known as the electron affinity. **Electron affinity** is defined as the energy released when an electron is added to a gaseous atom or ion.

$$T_{(g)} + e^- \rightarrow T^-_{(g)}$$

For most elements, the addition of an electron to a gaseous atom releases potential energy.

$$\mathrm{Br}_{(g)} + e^- \rightarrow \mathrm{Br}^-_{(g)} \qquad \Delta H = -325 \text{ kJ/mol}$$

Group and Period Trends in Electron Affinity

Let's look at the electron configurations of a few elements and the trend that develops within groups and periods. **Table** 4.8 shows the electron affinities for the halogen family.

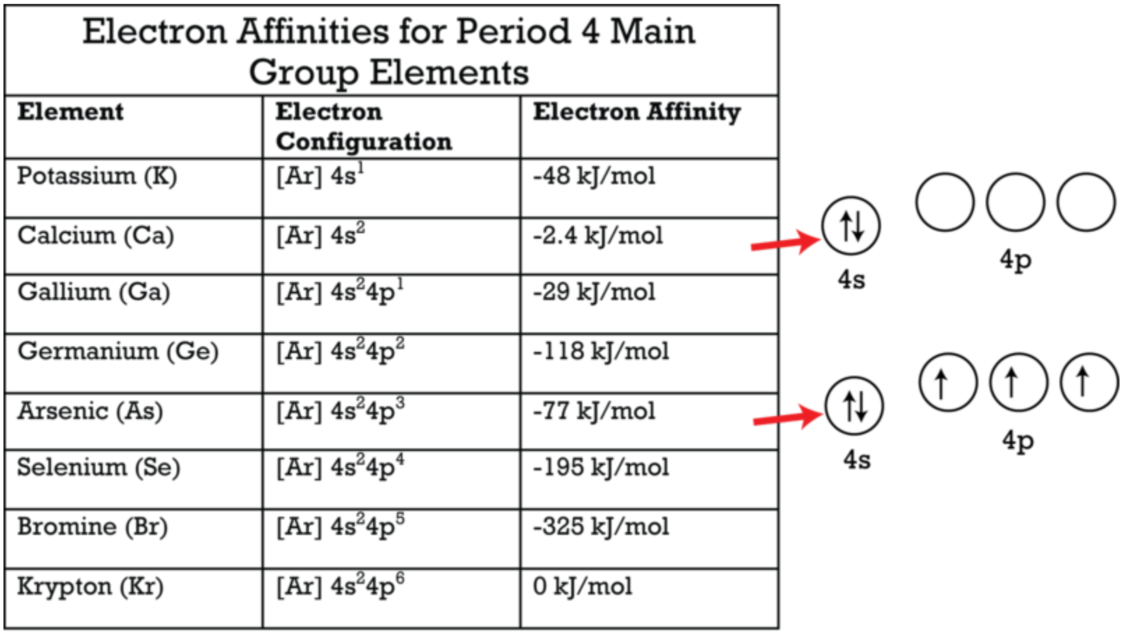
TABLE 4.8: Electron Affinities for Group 7A

Element	Electron Configuration	Electron Affinity, kJ/mol
Fluorine, F	[He]$2s^2 2p^5$	-328
Chlorine, Cl	[Ne]$3s^2 3p^5$	-349
Bromine, Br	[Ar]$4s^2 4p^5$	-325
Iodine, I	[Kr]$5s^2 5p^5$	-295

Going down a group, the electron affinity generally decreases because of the increase in size of the atoms. Remember that within a family, atoms located lower on the periodic table are larger because there are more filled energy levels. When an electron is added to a large atom, less energy is released because the electron cannot move as close to the nucleus as it can in a smaller atom. Therefore, as the atoms in a family get larger, the electron affinity gets smaller.

There are exceptions to this trend, especially when comparing the electron affinity of smaller atoms. In **Table** 4.8, the electron affinity for fluorine is less than that for chlorine. This phenomenon is observed in other families as well. The electron affinity of all the elements in the second period is less than the the electron affinity of the elements in the third period. For instance, the electron affinity for oxygen is less than the electron affinity for sulfur. This is most likely due to the fact that the elements in the second period have such small electron clouds (n = 2) that electron repulsion of these elements is greater than that of the rest of the family.

Overall, each row in the periodic table shows a general trend similar to the one below.



Electron Affinities for Period 4 Main Group Elements

Element	Electron Configuration	Electron Affinity
Potassium (K)	[Ar] $4s^1$	-48 kJ/mol
Calcium (Ca)	[Ar] $4s^2$	-2.4 kJ/mol
Gallium (Ga)	[Ar] $4s^2 4p^1$	-29 kJ/mol
Germanium (Ge)	[Ar] $4s^2 4p^2$	-118 kJ/mol
Arsenic (As)	[Ar] $4s^2 4p^3$	-77 kJ/mol
Selenium (Se)	[Ar] $4s^2 4p^4$	-195 kJ/mol
Bromine (Br)	[Ar] $4s^2 4p^5$	-325 kJ/mol
Krypton (Kr)	[Ar] $4s^2 4p^6$	0 kJ/mol

The general trend in the electron affinity for atoms is almost the same as the trend for ionization energy. This is because both electron affinity and ionization energy are highly related to atomic size. Large atoms have low ionization energy and low electron affinity. Therefore, they tend to lose electrons. In general, the opposite is true for small atoms. Since they are small, they have high ionization energies and high electron affinities. Therefore, the small atoms tend to gain electrons. The major exception to this rule is the noble gases. Noble gases follow the general trend for ionization energies, but do not follow the general trend for electron affinities. Even though the noble gases are small atoms, their outer energy levels are completely filled with electrons. Any added electron cannot enter their outer most energy level and would have to be the first electron in a new (larger) energy level. This causes the noble gases to have essentially zero electron affinity.

When atoms become ions, the process involves either releasing energy (through electron affinity) or absorbing energy (ionization energy). Therefore, the atoms that require a large amount of energy to release an electron will most likely be the atoms that give off the most energy while accepting an electron. In other words, nonmetals will gain electrons most easily since they have large electron affinities and large ionization energies. Metals will lose electrons since they have the low ionization energies and low electron affinities.

Lesson Summary

- Electron affinity is the energy released when an electron is added to a gaseous atom or ion.
- Electron affinity generally decreases going down a group and increases left to right across a period.
- Nonmetals tend to have the highest electron affinities.

Further Reading / Supplemental Links

This video shows the relationships between atomic size, ionization energy, and electron affinity.

- http://www.youtube.com/watch?v=iCwYjpl8eeY&feature=channel

This pdf document reviews the causes and relationships of the trends in atomic size, ionization energy, electronegativity, and electron affinity.

- http://www.oakland.k12.mi.us/Portals/0/Learning/PeriodicTable.pdf

Review Questions

1. Define electron affinity and write an example equation.
2. Choose the element in each pair that has the lower electron affinity.
    a. Li or N
    b. Cl or Na
    c. Ca or K
    d. Mg or F
3. Why is the electron affinity for calcium higher than that of potassium?
4. Which of the following will have the largest electron affinity?
    a. Se
    b. F
    c. Ne
    d. Br
5. Which of the following will have the smallest electron affinity?
    a. Na
    b. Ne
    c. Al
    d. Rb
6. Place the following elements in order of increasing electron affinity: Tl, Br, S, K, Al.

7. Describe the general trend for electron affinities in period 2.
8. Why does sulfur have a greater electron affinity than phosphorus does?

CHAPTER 5 Chemical Reactions

Chapter Outline

5.1 Chemical Reactions and Equations

Lesson Objectives

The student will:

- explain what happens during a chemical reaction.
- identify the reactants and products in any chemical reaction.
- convert verbal descriptions of chemical reactions into chemical equations, and vice versa.
- use the common symbols (s), (l), (g), (aq), and $\rightarrow$ appropriately.

Vocabulary

- chemical reaction
- products
- reactants

Introduction

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that make up the substances separate and re-arrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

In order to describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants**, and the substances present at the end are called **products**.

Writing Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. In the chemical equation shown below, sulfur dioxide and oxygen $(SO_2$ and $O_2)$ are reactants, and sulfur trioxide (SO_3) is the product.

$$\underbrace{2\,SO_{2(g)} + O_{2(g)}}_{\text{Reactants}} \longrightarrow \underbrace{2\,SO_{3(g)}}_{\text{Products}}$$

The general equation for a reaction is:

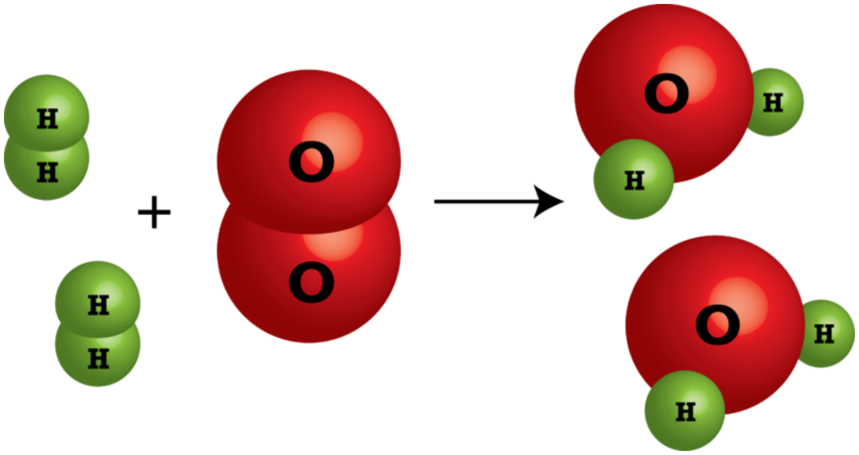
Reactants → Products

There are a few special symbols that we need to know in order to communicate in chemical shorthand. In **Table** 5.1 is a summary of the major symbols used in chemical equations. There are other symbols, but these are the main ones that we need to know.

TABLE 5.1: Common Symbols in Chemical Reactions

Symbol	Meaning	Example
$\rightarrow$	separates reactants from products; can be read as "to produce" or "to yield"	$2H_2 + O_2 \rightarrow 2H_2O$
+	separate reactants from each other or products from each other; can be read as "is added to"	$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
(s)	in the solid state	sodium in the solid state = $Na_{(s)}$
(l) or (L)	in the liquid state	water in the liquid state = $H_2O_{(l)}$
(g)	in the gaseous state	carbon dioxide in the gaseous state = $CO_{2(g)}$
(aq)	in the aqueous state, dissolved in water	sodium chloride solution = $NaCl_{(aq)}$

Chemists have a choice of methods for describing a chemical reaction. They could draw a picture of the chemical reaction, like in the image shown below.



Alternatively, they could describe the reaction in words. The image above can be described as two molecules of hydrogen gas reacting with one molecule of oxygen gas to produce two molecules of water vapor.

Chemists could also write the equation in chemical shorthand.

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

In the symbolic equation, chemical formulas are used instead of chemical names for reactants and products, and symbols are used to indicate the phase of each substance. It should be apparent that the chemical shorthand method is the quickest and clearest method for writing chemical equations. For example, we could write out that an aqueous solution of calcium nitrate is added to an aqueous solution of sodium hydroxide to produce solid calcium hydroxide and an aqueous solution of sodium nitrate. In shorthand, however, we could simply write:

$$Ca(NO_3)_{2(aq)} + 2NaOH_{(aq)} \rightarrow Ca(OH)_{2(s)} + 2NaNO_{3(aq)}$$

How much easier is that to read? Let's try it in reverse. Look at the following reaction in shorthand notation and describe the reaction in words.

$$Cu_{(s)} + AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + Ag_{(s)}$$

The description of this reaction might read something like "solid copper reacts with an aqueous solution of silver nitrate to produce a solution of copper(II) nitrate and solid silver."

Example:

Transfer the following symbolic equations into verbal descriptions or vice versa.

1. $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$
2. Gaseous propane, C_3H_8, burns in oxygen gas to produce gaseous carbon dioxide and liquid water.
3. Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

Solution:

1. An aqueous solution of hydrochloric acid reacts with an aqueous solution of sodium hydroxide to produce an aqueous solution of sodium chloride and liquid water.
2. $C_3H_{8(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$
3. $HF_{(g)} + K_2CO_{3(aq)} \rightarrow KF_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

Lesson Summary

- A chemical reaction is the process in which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that symbolizes "yields," and the products on the right.

Further Reading / Supplemental Links

This video shows ten amazing chemical reactions that are fun to watch but dangerous to carry out.

- http://listverse.com/2008/03/04/top-10-amazing-chemical-reactions/

Review Questions

1. Mothballs are commonly used to preserve clothing during off-season. We recognize mothballs by its smell because of a chemical compound known as naphthalene, $C_{10}H_8$. What are the different elements found in naphthalene, and how many atoms of each are found in the formula?
2. Give the verbal description of the following chemical equations.
    a. $H_2SO_{4(aq)} + NaCN_{(aq)} \rightarrow HCN_{(aq)} + Na_2SO_{4(aq)}$

  b. $Cu_{(s)} + AgNO_{3(aq)} \rightarrow Ag_{(s)} + Cu(NO_3)_{2(aq)}$
  c. $Fe_{(s)} + O_{2(g)} \rightarrow Fe_2O_{3(s)}$

3. Write the chemical equations for the following reactions.
   a. Solid calcium metal is placed in liquid water to produce aqueous calcium hydroxide and hydrogen gas.
   b. Aqueous sodium hydroxide is mixed with gaseous chlorine to produce aqueous solutions of sodium chloride and sodium hypochlorite plus liquid water.
   c. Solid xenon hexafluoride is mixed with liquid water to produce solid xenon trioxide and gaseous hydrogen fluoride.

5.2 Balancing Chemical Equations

Lesson Objectives

The student will:

- explain the roles of subscripts and coefficients in chemical equations.
- write a balanced chemical equation when given the unbalanced equation for any chemical reaction.
- explain the role of the law of conservation of mass in a chemical reaction.

Vocabulary

- balanced chemical equation
- coefficient
- subscript

Introduction

Even though chemical compounds are broken up to form new compounds during a chemical reaction, atoms in the reactants do not disappear, nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products. The atoms are merely re-organized into different arrangements. In a complete chemical equation, the two sides of the equation must be balanced. That is, in a **balanced chemical equation**, the same number of each atom must be present on the reactant and product sides of the equation.

Balancing Equations

The process of writing a balanced chemical equation involves three steps. As a beginning chemistry student, you will not know whether or not two given compounds will react or not. Even if you saw them react, you would not know what the products are without running any tests to identify them. Therefore, for the time being, you will be told both the reactants and products in any equation you are asked to balance.

Step 1: Know what the reactants and products are, and write a word equation for the reaction.

Step 2: Write the formulas for all the reactants and products.

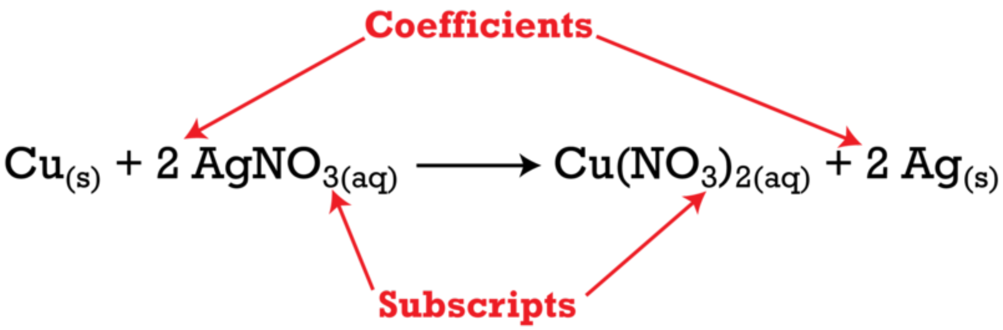
Step 3: Adjust the coefficients to balance the equation.

There are a number of elements shown in **Table** 5.2 that exist as diatomic molecules under normal conditions. When any of these elements appear in word equations, you must remember that the name refers to the diatomic molecule and insert the diatomic formula into the symbolic equation. If, under unusual circumstances, it was desired to refer to the individual atoms of these elements, the text would refer specifically to atomic hydrogen, atomic oxygen, and so on.

TABLE 5.2: Homonuclear Diatomic Molecules

Element	Formula for Diatomic Molecule	Phase at Room Temperature
Hydrogen	H_2	Gaseous
Oxygen	O_2	Gaseous
Nitrogen	N_2	Gaseous
Chlorine	Cl_2	Gaseous
Fluorine	F_2	Gaseous
Bromine	Br_2	Liquid
Iodine	I_2	Solid

There are two types of numbers that appear in chemical equations. There are **subscripts**, which are part of the chemical formulas of the reactants and products, and there are **coefficients** that are placed in front of the formulas to indicate how many molecules of that substance are used or produced. In the chemical formula shown below, the coefficients and subscripts are labeled.



The equation above indicates that one mole of solid copper is reacting with two moles of aqueous silver nitrate to produce one mole of aqueous copper(II) nitrate and two moles of solid silver. Recall that a subscript of 1 is not written - when no subscript appears for an atom in a formula, it is understood that only one atom is present. The same is true in writing balanced chemical equations. If only one atom or molecule is present, the coefficient of 1 is omitted.

The subscripts are part of the formulas, and once the formulas for the reactants and products are determined, the subscripts may not be changed. The coefficients indicate the mole ratios of each substance involved in the reaction and may be changed in order to balance the equation. Coefficients are inserted into the chemical equation in order to make the total number of each atom on both sides of the equation equal. Note that equation balancing is accomplished by changing coefficients, never

Example:

Write a balanced equation for the reaction that occurs between chlorine gas and aqueous sodium bromide to produce liquid bromine and aqueous sodium chloride.

Step 1: Write the word equation (keeping in mind that chlorine and bromine refer to the diatomic molecules).

chlorine + sodium bromide yields bromine + sodium chloride

Step 2: Substitute the correct formulas into the equation.

$Cl_2 + NaBr \rightarrow Br_2 + NaCl$

Step 3: Insert coefficients where necessary to balance the equation.

By placing a coefficient of 2 in front of NaBr, we can balance the bromine atoms. By placing a coefficient of 2 in front of the NaCl, we can balance the chlorine atoms.

$$\mathrm{Cl_2 + 2\ NaBr \rightarrow Br_2 + 2\ NaCl}$$

A final check (always do this) shows that we have the same number of each atom on the two sides of the equation. We have also used the smallest whole numbers possible as the coefficients, so this equation is properly balanced.

Example:

Write a balanced equation for the reaction between aluminum sulfate and calcium bromide to produce aluminum bromide and calcium sulfate. Recall that polyatomic ions usually remain together as a unit throughout a chemical reaction.

Step 1: Write the word equation.

aluminum sulfate + calcium bromide yields aluminum bromide + calcium sulfate

Step 2: Replace the names of the substances in the word equation with formulas.

$$\mathrm{Al_2(SO_4)_3 + CaBr_2 \rightarrow AlBr_3 + CaSO_4}$$

Step 3: Insert coefficients to balance the equation.

In order to balance the aluminum atoms, we must insert a coefficient of 2 in front of the aluminum compound in the products.

$$\mathrm{Al_2(SO_4)_3 + CaBr_2 \rightarrow 2\ AlBr_3 + CaSO_4}$$

In order to balance the sulfate ions, we must insert a coefficient of 3 in front of the product $\mathrm{CaSO_4}$.

$$\mathrm{Al_2(SO_4)_3 + CaBr_2 \rightarrow 2\ AlBr_3 + 3\ CaSO_4}$$

In order to balance the bromine atoms, we must insert a coefficient of 3 in front of the reactant $\mathrm{CaBr_2}$.

$$\mathrm{Al_2(SO_4)_3 + 3\ CaBr_2 \rightarrow 2\ AlBr_3 + 3\ CaSO_4}$$

The insertion of the 3 in front of the reactant $\mathrm{CaBr_2}$ also balances the calcium atoms in the product $\mathrm{CaSO_4}$. A final check shows that there are two aluminum atoms, three sulfur atoms, twelve oxygen atoms, three calcium atoms, and six bromine atoms on each side. This equation is balanced.

Note that this equation would still have the same number of atoms of each type on each side with the following set of coefficients:

$$\mathrm{2\ Al_2(SO_4)_3 + 6\ CaBr_2 \rightarrow 4\ AlBr_3 + 6\ CaSO_4}$$

Count the number of each type of atom on either side of the equation to confirm that this equation is "balanced." While this set of coefficients does "balanced" the equation, they are not the lowest set of coefficients possible. Chemical equations should be balanced with the simplest whole number coefficients. We could divide each of the coefficients in this equation by 2 to get another set of coefficients that still balance the equation and are whole numbers. Since it is required that an equation be balanced with the lowest whole number coefficients, the equation above is not

Example:

Balance the following skeletal equation. (The term "skeletal equation" refers to an equation that has the correct chemical formulas but does not include the proper coefficients.)

$Fe(NO_3)_3 + NaOH \rightarrow Fe(OH)_3 + NaNO_3$ (skeletal equation)

Solution:

We can balance the hydroxide ion by inserting a coefficient of 3 in front of the NaOH on the reactant side.

$Fe(NO_3)_3 + 3\ NaOH \rightarrow Fe(OH)_3 + NaNO_3$

We can then balance the nitrate ions by inserting a coefficient of 3 in front of the sodium nitrate on the product side.

$Fe(NO_3)_3 + 3\ NaOH \rightarrow Fe(OH)_3 + 3\ NaNO_3$

Counting the number of each type of atom on the two sides of the equation will now show that this equation is balanced.

Example:

Given the following skeletal (un-balanced) equations, balance them.

1. $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$
2. $H_2SO_{4(aq)} + Al(OH)_{3(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + H_2O_{(l)}$
3. $Ba(NO_3)_{2(aq)} + Na_2CO_{3(aq)} \rightarrow BaCO_{3(aq)} + NaNO_{3(aq)}$
4. $C_2H_{6(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

Solution:

1. $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$ (sometimes the skeletal equation is already balanced)
2. $3\ H_2SO_{4(aq)} + 2\ Al(OH)_{3(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + 6\ H_2O_{(l)}$
3. $Ba(NO_3)_{2(aq)} + Na_2CO_{3(aq)} \rightarrow BaCO_{3(aq)} + 2\ NaNO_{3(aq)}$
4. $2\ C_2H_{6(g)} + 7\ O_{2(g)} \rightarrow 4\ CO_{2(g)} + 6\ H_2O_{(l)}$

Conservation of Mass in Chemical Reactions

We already know from the law of conservation of mass that mass is conserved in chemical reactions. But what does this really mean? Consider the following reaction.

$Fe(NO_3)_3 + 3\ NaOH \rightarrow Fe(OH)_3 + 3\ NaNO_3$

Verify to yourself that this equation is balanced by counting the number of each type of atom on each side of the equation. We can demonstrate that mass is conserved by determining the total mass on both sides of the equation.

1 molecule of $Fe(NO_3)_3 \times$ molecular weight $= (1) \cdot (241.9$ daltons$)$ = 241.9 daltons
3 molecules of $NaOH \times$ molecular weight $= (3) \cdot (40.0$ daltons$)$ = 120. daltons
Total mass of reactants = 241.9 daltons + 120. daltons = 361.9 daltons

1 molecule of $Fe(OH)_3 \times$ molecular weight $= (1) \cdot (106.9$ daltons$)$ = 106.9 daltons
3 molecules of $NaNO_3 \times$ molecular weight $= (3) \cdot (85.0$ daltons$)$ = 255 daltons
Total mass of products = 106.9 daltons + 255 daltons = 361.9 daltons

As you can see, both the number of atoms and mass are conserved during chemical reactions. This is logically similar to saying that a group of 20 objects stacked in different ways will still have the same total mass no matter how you stack them.

Lesson Summary

- Chemical equations must always be balanced.
- Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio.
- Mass is always conserved in chemical reactions.

Review Questions

1. Explain in your own words why coefficients can change but subscripts must remain constant.
2. Which set of coefficients will properly balance the following equation: $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$?
   a. 1, 1, 1, 1
   b. 1, 3, 2, 2
   c. 1, 3.5, 2, 3
   d. 2, 7, 4, 6
3. When properly balanced, what is the sum of all the coefficients in the following chemical equation: $SF_4 + H_2O \rightarrow H_2SO_3 + HF$?
   a. 4
   b. 7
   c. 9
   d. None of the above
4. When the following equation is balanced, what is the coefficient found in front of the O_2: $P_4 + O_2 + H_2O \rightarrow H_3PO_4$?
   a. 1
   b. 3
   c. 5
   d. 7
5. Balance the following equations.
   a. $XeF_{6(s)} + H_2O_{(l)} \rightarrow XeO_{3(s)} + HF_{(g)}$
   b. $Cu_{(s)} + AgNO_{3(aq)} \rightarrow Ag_{(s)} + Cu(NO_3)_{2(aq)}$
   c. $Fe_{(s)} + O_{2(g)} \rightarrow Fe_2O_{3(s)}$
   d. $Al(OH)_3 + Mg_3(PO_4)_2 \rightarrow AlPO_4 + Mg(OH)_2$

5.3 Types of Reactions

Lesson Objectives

The student will:

- describe what is occurring in synthesis, decomposition, single replacement, double replacement, and combustion reactions.
- classify a chemical reaction as a synthesis, decomposition, single replacement, double replacement, or a combustion reaction.
- predict the products of simple reactions.

Vocabulary

- combustion reaction
- decomposition reaction
- double replacement reaction
- hydrocarbon
- single replacement reaction
- synthesis reaction

Introduction

Chemical reactions are classified into types to help us analyze them and to help us predict what the products of the reaction will be. The five major types of chemical reactions are synthesis, decomposition, single replacement, double replacement, and combustion.

Synthesis Reactions

A **synthesis reaction** is one in which two or more reactants combine to make one product. The general equation for a synthesis reaction is:

$$A + B \rightarrow AB$$

Synthesis reactions occur as a result of two or more simpler elements or molecules combining to form a more complex molecule. We can always identify a synthesis reaction because there is only one product. If you are given elemental reactants and told that the reaction is a synthesis reaction, you should be able to predict the products. For example, consider the equation below. Two elements (hydrogen and oxygen) combine to form one product (water).

$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(l)}$

You should also be able to write the chemical equation for a synthesis reaction if you are given a product by picking out its elements and writing the equation. As a result, we can write the synthesis reaction for sodium chloride just by knowing the elements that are present in the product.

$2\ Na_{(s)} + Cl_{2(g)} \rightarrow 2\ NaCl_{(s)}$

Example:

1. Write the chemical equation for the synthesis reaction of silver bromide, AgBr.
2. Predict the products for the following reaction: $CO_{2(g)} + H_2O_{(l)}$.
3. Predict the products for the following reaction: $Li_2O_{(s)} + CO_{2(g)}$.

Solution:

1. $2\ Ag_{(s)} + Br_{2(l)} \rightarrow 2\ AgBr_{(s)}$
2. $CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$
3. $Li_2O_{(s)} + CO_{2(g)} \rightarrow Li_2CO_{3(s)}$

Decomposition Reactions

When one type of reactant breaks down to form two or more products, we have a **decomposition reaction**. The best way to remember a decomposition reaction is that for all reactions of this type, there is only one reactant. The general equation for a decomposition reaction is:

$AB \rightarrow A + B$

Look at the equation below for an example of a decomposition reaction. In this reaction, ammonium nitrate breaks down to form dinitrogen oxide and water.

$NH_4NO_{3(s)} \rightarrow N_2O_{(g)} + 2H_2O_{(g)}$

Notice that there is only one reactant, NH_4NO_3, on the left of the arrow and that there is more than one on the right side of the arrow. This is the exact opposite of a synthesis reaction.

When studying decomposition reactions, we can predict the reactants in a similar manner as we did for synthesis reactions. Look at the formula for magnesium nitride, Mg_3N_2. What elements do you see in this formula? You see magnesium and nitrogen. Now we can write a decomposition reaction for magnesium nitride.

$Mg_3N_{2(s)} \rightarrow 3\ Mg_{(s)} + N_{2(g)}$

Notice there is only one reactant.

Example:

Write the chemical equation for the decomposition of the following compounds into their individual elements:

1. Al_2O_3
2. Ag_2S
3. MgO

Solution:

1. $2\ Al_2O_3 \rightarrow 4\ Al + 3\ O_2$
2. $Ag_2S \rightarrow 2\ Ag + S$
3. $2\ MgO \rightarrow 2\ Mg + O_2$

Single Replacement Reactions

A third type of reaction is the single replacement reaction. In **single replacement reactions**, one element reacts with one compound to form products. The single element is said to replace an element in the compound when the products form, hence the name single replacement.

There are actually three different types of single replacement reactions: 1) the single element is a metal and replaces the metal in the second reactant, 2) the single element is a metal and replaces the hydrogen in the second reactant, which is always an acid, and 3) the single element is a nonmetal and replaces the nonmetal in the second reactant.

Replacement of a Metal with a Metal

In this section, we will focus on single replacement reactions where an elemental metal reactant replaces the metal (or the cation) of a second compound. The general equation for this reaction is:

$$A + BC \rightarrow B + AC$$

Consider the following example. Notice there is only one reactant that is an element and one reactant that is a compound.

$$Zn_{(s)} + Cu(NO_3)_{2(aq)} \rightarrow Zn(NO_3)_{2(aq)} + Cu_{(s)}$$

When studying single replacement reactions, we can predict reactants in a similar manner as we did for synthesis and decomposition reactions. Suppose that we know a single replacement reaction will occur between solid aluminum and solid iron(III) oxide.

$$Al_{(s)} + Fe_2O_{3(s)}$$

In order to predict the products, we need to know that aluminum will replace iron and form aluminum oxide. Aluminum has a charge of $+3$ (it is in Group 3A), and oxygen has a charge of -2 (it is in Group 6A). The compound formed between aluminum and oxygen, therefore, will be $Al_2O_{3(s)}$. Since iron is replaced in the compound by aluminum, the iron product will now be a single element. The unbalanced equation will be:

$$Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + Fe_{(s)}.$$

The balanced equation will be:

$$2\ Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + 2\ Fe_{(s)}$$

Example:

1. Write the chemical equation for the single replacement reaction between zinc solid and lead(II) nitrate solution to produce zinc nitrate solution and solid lead.

2. Predict the products for the following reaction: $Fe + CuSO_4$.
3. Predict the products for the following reaction: $Al + CuCl_2$.

Solution:

1. $Zn_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow Pb_{(s)} + Zn(NO_3)_{2(aq)}$
2. $Fe_{(s)} + CuSO_{4(aq)} \rightarrow Cu_{(s)} + FeSO_{4(aq)}$
3. $2\ Al + 3\ CuCl_2 \rightarrow 3\ Cu + 2\ AlCl_3$

Replacement of Hydrogen with a Metal

These reactions are the same as those studied in the last section, except the compound in the reactant side of the equation is always an acid. Since you may not have studied acids yet, you should consider an acid to be a compound in which hydrogen is combined with an anion. Therefore, in this section, we will consider single replacement reactions where the element reactant replaces the hydrogen in the acid compound. The general reaction is:

$$A + 2\ HX \rightarrow AX_2 + H_2$$

The chemical equation below is an example of this type of reaction:

$$Zn_{(s)} + 2\ HBr_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

When studying these single replacement reactions, we can predict reactants in a similar manner as we did for the other types of single replacement reactions. Look at the reaction below. Since HCl is a compound that has hydrogen combined with an anion, it is an acid.

$$Mg_{(s)} + 2HCl_{(aq)}$$

In order to predict the products, we need to know that magnesium will replace hydrogen and form magnesium chloride. Magnesium has a charge of $+2$ (it is in Group 2A), and chlorine has a charge of -1 (it is in group 7A). Therefore, the compound formed will be $MgCl_2$. The replaced hydrogen forms the product elemental hydrogen, H_2. After the products are determined, all that remains is to balance the equation.

$$Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(s)} + H_{2(g)}$$

Notice that one reactant is the element Mg and one reactant is an acid compound. The Mg has replaced the hydrogen in HCl in the same manner as Zn replaced the hydrogen in the HBr in the example above.

Example:

1. Write the chemical equation for the single replacement reaction between iron solid and hydrochloric acid solution to produce iron(II) chloride solution and hydrogen gas.
2. Predict the products for the following reaction: $Zn_{(s)} + H_2SO_{4(aq)}$.
3. Predict the products for the following reaction: $Al_{(s)} + HNO_{3(aq)}$.

Solution:

1. $Fe_{(s)} + 2\ HCl_{(aq)} \rightarrow FeCl_{2(aq)} + H_{2(g)}$
2. $Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$
3. $2\ Al_{(s)} + 6\ HNO_{3(aq)} \rightarrow 2\ Al(NO_3)_{3(aq)} + 3\ H_{2(g)}$

Replacement of a Nonmetal with a Nonmetal

In this section, we will focus on the final type of single replacement reactions where the element reactant replaces the nonmetal (or the anion) in a compound. The general equation for this type of reaction is:

$$A+BC \rightarrow C+BA$$

Notice in the equation below that the chlorine replaced the iodine to produce solid iodine as a product. We can predict the products for these single replacement reactions in a similar manner as for all other single replacement reactions. The only difference here is that we have to remember that we are replacing the anion of the compound rather than the cation.

$$\mathrm{Cl}_{2(g)}+2\ \mathrm{KI}_{(aq)} \rightarrow 2\ \mathrm{KCl}_{(aq)}+\mathrm{I}_{2(s)}$$

Look at the reaction between chlorine gas and sodium bromide shown below. This is an actual method for extracting bromine from ocean water found to contain sodium bromide. Can you complete the reaction?

$$\mathrm{Cl}_{2(g)}+\mathrm{NaBr}_{(aq)} \rightarrow ?$$

In order to predict the products of this reaction, we need to know that chlorine will replace bromine and form sodium chloride. Sodium has a charge of $+1$ (it is in Group 1A) and chlorine has a charge of -1 (it is in group 7A). The compound formed will be NaCl.

$$\mathrm{Cl}_{2(g)}+2\ \mathrm{NaBr}_{(aq)} \rightarrow 2\ \mathrm{NaCl}_{(aq)}+\mathrm{Br}_{2(l)}$$

Notice, as with all of the other single replacement reactions, that the reactants include one element and one compound, and the products contain one element and one compound. This is the determining factor for identifying whether you have a single replacement reaction.

Example:

1. Write the chemical equation for the single replacement reaction between sodium iodide solution and liquid bromine to produce sodium bromide solution and solid iodine.
2. Predict the products for the following reaction: $\mathrm{Br}_{2(aq)}+\mathrm{KI}_{(aq)}$.
3. Predict the products for the following reaction: $\mathrm{MgI}_{2(aq)}+\mathrm{Cl}_{2(aq)}$.

Solution:

1. $2\ \mathrm{NaI}_{(aq)}+\mathrm{Br}_{2(l)} \rightarrow 2\ \mathrm{NaBr}_{(aq)}+\mathrm{I}_{2(s)}$
2. $\mathrm{Br}_{2(aq)}+2\ \mathrm{KI}_{(aq)} \rightarrow 2\ \mathrm{KBr}_{(aq)}+\mathrm{I}_{2(s)}$
3. $\mathrm{MgI}_{2(aq)}+\mathrm{Cl}_{2(aq)} \rightarrow \mathrm{MgCl}_{2(aq)}+\mathrm{I}_{2(s)}$

Double Replacement

For **double replacement reactions**, two reactants will react by having the cations exchange places. The key to identifying this type of reaction is to recognize that it has two compounds as reactants. This type of reaction is more common than any of the others, and there are many different types of double replacement reactions. Some double replacement reactions are more common than others. For example, precipitation and neutralization reactions are two of the most common double replacement reactions. Precipitation reactions are ones where two aqueous reactants combine to form products where one of the products is an insoluble solid. A neutralization reaction is one where the two reactant compounds are an acid and a base and the two products are a salt and water.

Example:

The following is a precipitation reaction because $AgCl_{(s)}$ is formed.

$$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$$

The following is a neutralization reaction because the acid, H_2SO_4, is neutralized by the base, NaOH.

$$2\ NaOH_{(aq)} + H_2SO_{4(aq)} \rightarrow Na_2SO_{4(aq)} + 2\ H_2O_{(l)}$$

In order to write the products for a double displacement reaction, you must be able to determine the correct formulas for the new compounds. Consider this common laboratory experiment that involves the reaction between lead(II) nitrate and sodium iodide, both clear solutions. Here is the start of the reaction:

$$Pb(NO_3)_{2(aq)} + NaI_{(aq)}$$

Now, predict the products based on what you know about charges. We know that the cations exchange anions. We now have to look at the charges of each of the cations and anions to see what the products will be.

We should presume the charge of the lead will remain $+2$, and since iodine forms ions with a charge of -1, one product will be PbI_2. The other product will form between the sodium ion, whose charge is $+1$, and the nitrate ion, whose charge is -1. Therefore, the second product will be $NaNO_3$. Once the products are written in, the equation can be balanced.

$$Pb(NO_3)_{2(aq)} + 2\ NaI_{(aq)} \rightarrow PbI_{2(s)} + 2\ NaNO_{3(aq)}$$

The experiment produces a brilliant yellow precipitate. If you have use of a solubility table, it is easy to determine that the precipitate will be the lead(II) iodide. Even without a solubility table, knowing that lead compounds tend to precipitate and sodium compounds are always soluble, we would still be able to determine that the PbI_2 is the brilliant yellow precipitate.

Look at the reaction between acetic acid and barium hydroxide below.

$$HC_2H_3O_{2(aq)} + Ba(OH)_{2(aq)} \rightarrow ?$$

Try to predict the products by having the cations exchange places and writing the correct formulas for the products formed.

$$HC_2H_3O_{2(aq)} + Ba(OH)_{2(aq)} \rightarrow Ba(C_2H_3O_2)_{2(aq)} + H_2O_{(l)} \text{ (not balanced)}$$

Therefore, the final balanced equation will be:

$$2\ HC_2H_3O_{2(aq)} + Ba(OH)_{2(aq)} \rightarrow Ba(C_2H_3O_2)_{2(aq)} + 2\ H_2O_{(l)} \text{ (balanced)}$$

This is an acid-base reaction yielding salt, barium acetate, and water. Notice that HOH and H_2O are the same.

Example:

1. Write a chemical equation for the double replacement reaction between calcium chloride solution and potassium hydroxide solution that produces potassium chloride solution and a precipitate of calcium hydroxide.
2. Predict the products for the following reaction: $AgNO_{3(aq)} + NaCl_{(aq)}$.

3. Predict the products for the following reaction: $FeCl_{3(aq)} + KOH_{(aq)}$.

Solution:

1. $CaCl_{2(aq)} + 2\ KOH_{(aq)} \rightarrow Ca(OH)_{2(s)} + 2\ KCl_{(aq)}$
2. $AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$
3. $FeCl_{3(aq)} + 3\ KOH_{(aq)} \rightarrow Fe(OH)_{3(s)} + 3\ KCl_{(aq)}$

Combustion

A special type of single replacement reaction deserves some attention. These reactions are **combustion reactions**. In a combustion reaction, oxygen reacts with another substance to produce carbon dioxide and water.

In a particular branch of chemistry, known as organic chemistry, we study compounds known as hydrocarbons. A **hydrocarbon** is an organic substance consisting of only hydrogen and carbon. Combustion reactions usually have a hydrocarbon reacting with oxygen to produce CO_2 and H_2O. In other words, the only part that changes from one combustion reaction to the next is the actual hydrocarbon involved in the reaction. The general equation for this reaction is given below. Notice oxygen, carbon dioxide, and water are listed in the general equation to show that these reactants and products remain the same from combustion reaction to combustion reaction.

$$\text{hydrocarbon} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$$

Look at the the combustion reaction of octane, C_8H_{18}, shown below. Octane has 8 carbon atoms, hence the prefix "oct-."

$$2\ C_8H_{18(l)} + 25\ O_{2(g)} \rightarrow 16\ CO_{2(g)} + 18\ H_2O_{(l)}$$

This reaction is referred to as complete combustion. Complete combustion reactions occur when there is enough oxygen to burn the entire hydrocarbon. This is why the only products are carbon dioxide and water.

Have you ever been in a lab and seen black soot appear on the bottom of a heated beaker? Or, have you ever seen the black puffs of smoke come out from the exhaust pipe of a car? If there is not enough oxygen, the result is an incomplete combustion reaction with $CO_{(g)}$ and $C_{(s)}$ (in the form of soot) also formed as products. Incomplete combustion reactions are actually quite dangerous because one of the products in the reaction is carbon monoxide, not carbon dioxide. Carbon monoxide is a gas that prevents oxygen from binding to the oxygen transport proteins in our blood cells. When the concentration of carbon monoxide in the blood becomes too high, not enough oxygen can be transported and the person can die.

Example:

Identify whether each of the following reactions are complete or incomplete combustions, and then balance the equation.

1. $C_7H_{16(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$
2. $C_3H_{8(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$
3. $CH_{4(g)} + O_{2(g)} \rightarrow CO_{(g)} + H_2O_{(l)}$
4. $C_5H_{12(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$
5. $C_2H_{6(g)} + O_{2(g)} \rightarrow CO_{(g)} + H_2O_{(l)}$

Solution:

1. Complete; $C_7H_{16(l)} + 11\ O_{2(g)} \rightarrow 7\ CO_{2(g)} + 8\ H_2O_{(l)}$

2. Complete; $C_3H_{8(g)} + 5\ O_{2(g)} \rightarrow 3\ CO_{2(g)} + 4\ H_2O_{(l)}$
3. Incomplete; $2\ CH_{4(g)} + 3\ O_{2(g)} \rightarrow 2\ CO_{(g)} + 4\ H_2O_{(l)}$
4. Complete; $C_5H_{12(l)} + 8\ O_{2(g)} \rightarrow 5\ CO_{2(g)} + 6\ H_2O_{(l)}$
5. Incomplete; $2\ C_2H_{6(g)} + 5\ O_{2(g)} \rightarrow 4\ CO_{(g)} + 6\ H_2O_{(l)}$

This video contains classroom demonstrations of several reaction types and then shows the balancing process for the reaction equations **(3a)** http://www.youtube.com/watch?v=4B8PFqbMNIw (8:40).

MEDIA

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Lesson Summary

TABLE 5.3: **The Five Types of Chemical Reactions**

Reaction Name	Reaction Description
synthesis	two or more reactants form one product.
decomposition	one type of reactant forms two or more products.
single replacement	one element reacts with one compound to form products.
double replacement	two compounds act as reactants.
combustion	a hydrocarbon reacts with oxygen gas.

Further Reading / Supplemental Links

The website below also reviews the different types of chemical reactions.

- http://library.thinkquest.org/2923/react.html

Review Questions

1. When balancing combustion reactions, did you notice a consistency relating to whether the number of carbons in the hydrocarbon was odd or even?
2. Distinguish between synthesis and decomposition reactions.
3. When dodecane, $C_{10}H_{22}$, burns in excess oxygen, the products would be:
   a. $CO_2 + 2\ H_2$
   b. $CO + H_2O$
   c. $CO_2 + H_2O$
   d. CH_4O_2
4. In the decomposition of antimony trichloride, which of the following products and quantities will be found?

   a. $An + Cl_2$
   b. $2\ An + 3\ Cl_2$
   c. $Sb + Cl_2$
   d. $2\ Sb + 3\ Cl_2$

5. Acetylsalicylic acid (aspirin), $C_9H_8O_{4(s)}$, is produced by reacting acetic anhydride, $C_4H_6O_{3(l)}$, with salicylic acid, $C_7H_6O_{3(s)}$. The other product in the reaction is acetic acid, $C_2H_4O_{2(l)}$. Write the balanced chemical equation.
6. When iron rods are placed in liquid water, a reaction occurs. Hydrogen gas evolves from the container, and iron(III) oxide forms onto the iron rod.
   a. Write a balanced chemical equation for the reaction.
   b. What type of reaction is this?
7. A specific fertilizer is being made at an industrial plant nearby. The fertilizer is called a triple superphosphate and has a formula $Ca(H_2PO_4)_2$. It is made by treating sand and clay that contains phosphate with a calcium phosphate solution and phosphoric acid. The simplified reaction is calcium phosphate reacting with phosphoric acid to yield the superphosphate. Write the balanced chemical equation and name the type of reaction.

CHAPTER 6

Acids and Bases

Chapter Outline

6.1 Properties of Acids and Bases

Lesson Objectives

The student will:

- list the properties of acids.
- list the properties of bases.
- name an acid or base when given the formula.
- write the formula for an acid or base given the name.

Vocabulary

- indicators

Introduction

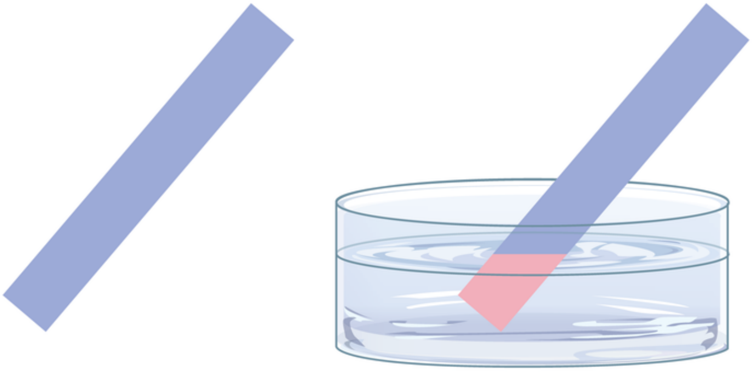
We interact with acids on a daily basis without even realizing it. For example, the chemical names for aspirin and vitamin C are acetylsalicylic acid and ascorbic acid; both will produce H^+ ions when dissolved in water. Acetic acid ($HC_2H_3O_2$) is the primary component in vinegar, and formic acid (HCO_2H) is what causes ant bites to sting. Hydrochloric acid (HCl) is stomach acid, phosphoric acid (H_3PO_4) is commonly found in dark soft drinks, and sulfuric acid (H_2SO_4) is used in car batteries. As you work your way through this chapter, try to notice how the properties of acids and bases manifest themselves in everyday situations.

Properties of Acids

One property that is common to all acids is a sour taste. You are probably most familiar with this in relation to citric acid, which is what makes lemons and other citrus fruits taste sour. In fact, sour taste buds are essentially just complicated H^+ sensors. The fact that one of our primary tastes is concerned solely with determining the acidity of what goes in our mouths further underscores the importance of acids in our lives.

However, testing whether something is acidic by taste is generally not a good idea. Another way to test for acidity is to use an indicator. **Indicators** are substances that can be used to determine the relative acidity or basicity of a solution, generally through a very distinct color change.

One common type of indicator is litmus paper. If a piece of blue litmus paper turns red when dipped into a solution, it means that the solution is acidic.



Another property common to many acids is that they can react with certain metals to form hydrogen gas. Examples of this type of reaction are shown below. Note that these are all single replacement reactions where a pure element reacts with a compound

$Zn_{(s)} + 2\ HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$
$Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
$Ba_{(s)} + 2\ HCl_{(aq)} \rightarrow BaCl_{2(aq)} + H_{2(g)}$

Names and Formulas of Acids

The chemical formula for an acid typically begins with one or more hydrogen atoms. This means that hydrogen is the cation. The name of an acid depends on the anion. When the anion ends in "–ide," such as chloride or sulfide, the prefix "hydro-" is used before the name of the nonmetal, and its "-ide" ending is replaced by "-ic acid." For example, the acid HCl is named hydrochloric acid. The anion is chloride, so the prefix "hydro-" is used, and the "-ide" ending is replaced by "-ic acid."

When the anion ends in "–ate," such as sulfate and phosphate, the "–ate" is replaced by "-ic acid." For example, the acid HNO_3 is named nitric acid. The anion is nitrate, so the "–ate" ending is replaced by "–ic" acid.

When the anion ends in "–ite," such as nitrite or sulfite, the "–ite" is replaced by "–ous acid." For example, the acid $HClO_2$ is named chlorous acid. The anion is chlorite, so the "–ite" is replaced by "–ous acid."

Example:

Write the names of each of the following acids:

1. HF
2. HNO_2
3. H_2SO_4

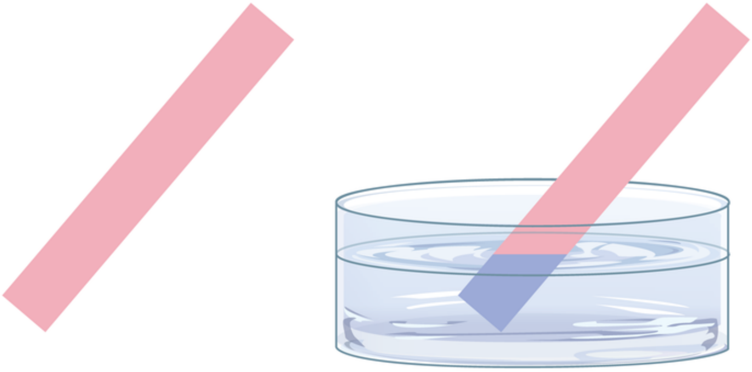
Solution:

1. hydrofluoric acid
2. nitrous acid
3. sulfuric acid

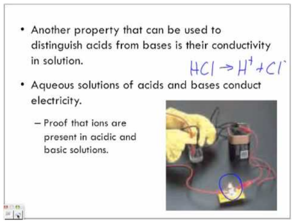
Properties of Bases

Bases also have a number of characteristic properties. Most bases are slippery and quite bitter (though not all bitter compounds are basic). Caffeine and milk of magnesia (chemical formula $Mg(OH)_2$) are two bases that you may have had the opportunity to taste, although the bitterness is generally masked by other flavors when these compounds are consumed. Other common bases are found in a number of cleaning products, including Drano (NaOH) and Windex (NH_4OH).

Like acids, bases can be identified by the use of an indicator. For example, if red litmus paper is dipped into a basic solution, it will turn blue.



This video discusses the properties of acids and bases **(5a)**: http://www.youtube.com/watch?v=mm7Hcff5b6g (8:02).



- Another property that can be used to distinguish acids from bases is their conductivity in solution.

$HCl \rightarrow H^+ + Cl^-$

- Aqueous solutions of acids and bases conduct electricity.
  - Proof that ions are present in acidic and basic solutions.

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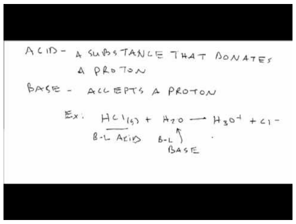
A discussion of the difference between strong and weak acids is available at **(5c)** http://www.youtube.com/watch?v=XTdkWGImtSc (9:47).



ACID - A SUBSTANCE THAT DONATES A PROTON

BASE - ACCEPTS A PROTON

Ex. $\underline{HCl}_{(g)} + H_2O \longrightarrow H_3O^+ + Cl^-$

B-L ACID B-L BASE

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Names and Formulas of Bases

The chemical formula for a strong base often ends with one or more hydroxide (OH^-) ions. This means that hydroxide is the anion. This type of base is named in the same way as ay other ionic compound.

Example:

Write the names of each of the following bases.

1. NaOH
2. KOH
3. $Ca(OH)_2$

Solution:

1. sodium hydroxide
2. potassium hydroxide
3. calcium hydroxide

Lesson Summary

- Acids turn blue litmus paper red, taste sour, and react with metals to produce hydrogen gases.
- Common acids include vinegar $(HC_2H_3O_2)$ and stomach acid (HCl).
- Bases turn red litmus paper blue, have a bitter taste, and are slippery to the touch.
- Common bases include Drano $(NaOH)$, soft soap (KOH), milk of magnesia $(Mg(OH)_2)$, and Windex (NH_4OH).

Further Reading / Supplemental Links

Visit this website to learn more about examples and properties of acids and bases.

- http://qldscienceteachers.tripod.com/junior/chem/acid.html

Review Questions

1. What are the properties of acids? Give a common example.
2. Which statement best describes a characteristic of acid solutions?
   a. They react with some metals to form hydrogen gas.
   b. They turn red litmus paper blue.
   c. They taste bitter.
   d. They are made from nonmetal oxides.
3. Write the reaction between:
   a. magnesium and sulfuric acid.
   b. calcium and acetic acid.

4. Which of the following will react with acids and produce hydrogen gas?
    a. chlorine
    b. ammonia
    c. carbon
    d. magnesium

6.2 Arrhenius Acids and Bases

Lesson Objectives

The student will:

- define an Arrhenius acid and list some substances that qualify as acids under this definition.
- define an Arrhenius base and list some substances that qualify as bases under this definition.

Vocabulary

- Arrhenius acid
- Arrhenius base

Introduction

Work done by Svante Arrhenius in the late 19th century set the groundwork for our current understanding of acid-base theory. Despite the relatively limited technology available to him, he was able to develop theories that adequately explained the properties and chemical reactivity of many acids and bases. Subsequent research led to the development of other commonly used definitions for acids and bases, but for now, we will focus on those proposed by Arrhenius.

Definitions

Take a look at the following dissociation reactions:

$$\begin{aligned}
&\mathrm{HCl}_{(aq)} \rightarrow \mathrm{H}^+_{(aq)} + \mathrm{Cl}^-_{(aq)}\\
&\mathrm{HNO}_{3(aq)} \rightarrow \mathrm{H}^+_{(aq)} + \mathrm{NO}^-_{3(aq)}\\
&\mathrm{HClO}_{4(aq)} \rightarrow \mathrm{H}^+_{(aq)} + \mathrm{ClO}^-_{4(aq)}\\
&\mathrm{NaOH}_{(s)} \rightarrow \mathrm{Na}^+_{(aq)} + \mathrm{OH}^-_{(aq)}\\
&\mathrm{Ba(OH)}_{2(s)} \rightarrow \mathrm{Ba}^{2+}_{(aq)} + 2\ \mathrm{OH}^-_{(aq)}
\end{aligned}$$

If you look at the products, you will notice some common features. All of the acids have H^+ as one of the resulting ions, and all the bases produce OH^-. It was on the basis of this observation that Arrhenius settled on his definitions for acids and bases. An **Arrhenius acid** is a substance that produces H^+ ions in solution, and an **Arrhenius base** is a substance that produces OH^- ions in solution.

Identifying Arrhenius Acids and Bases

How can we identify an acid or a base simply by looking at the chemical formula? Since we have defined acids and bases by the ions they release in solution, the first requirement is that they contain H or OH, respectively. However, there are plenty of compounds that contain oxygen and hydrogen atoms that are not acids and bases. What other clues can we look for?

Arrhenius bases are easy to spot because basically all of them are salts where the cation is a metal (or ammonium) and the anion is hydroxide. Due to the way we name ionic compounds, these substances will also have "hydroxide" in the name.

Acids can sometimes be a little trickier if all you have is the formula. For many of the acids that you will encounter, the formula will begin with an H: H_2SO_4, HCl, HNO_3, and $HClO_4$ are all acids. When a formula involves hydrogen plus an anion that you are familiar with, it is highly likely that the compound is an acid.

One common type of weak acid is generally written in the form R-COOH or R-CO_2H, where R is usually a hydrocarbon. An example is acetic acid, which is usually written as CH_3CO_2H. In solution, acetic acid dissociates to form acetate $(CH_3CO_2^-)$ and H^+. Other acids of this type have similar dissociation reactions. Note that when written in the form R-COOH, these compounds may appear to have an available hydroxide anion. Don't be fooled! When OH is attached to something covalently, it will not release when dissolved into solution.

Example:

Write the dissociation equation for the following compounds and state whether each is an acid, a base, or neither.

1. H_2SO_4
2. LiOH
3. $NaNO_3$
4. C_6H_5COOH
5. NaOH
6. KOH

Solution:

1. acid, $H_2SO_{4(aq)} \rightarrow 2\ H^+_{(aq)} + SO^{2-}_{4(aq)}$
2. base, $LiOH_{(s)} \rightarrow Li^+_{(aq)} + OH^-_{(aq)}$
3. neither, $NaNO_{3(s)} \rightarrow Na^+_{(aq)} + NO^-_{3(aq)}$
4. acid, $C_6H_5COOH_{(aq)} \rightarrow H^+_{(aq)} + C_6H_5COO^-_{(aq)}$
5. base, $NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$
6. base, $KOH_{(s)} \rightarrow K^+_{(aq)} + OH^-_{(aq)}$

Lesson Summary

- Arrhenius defined an acid as a substance that donates H^+ ions when dissociating in solution.
- An Arrhenius base is a substance that releases OH^- ions in solution.

Review Questions

1. Which of the following is the Arrhenius definition of an acid?
   a. An acid is a substance that donates protons.
   b. An acid is a substance that accepts protons.
   c. An acid is a substance that dissolves in water to form OH^- ions.
   d. An acid is a substance that reacts with water to form H^+ ions.

6.3 The pH Concept

Lesson Objectives

The student will:

- calculate $[H^+]$ and $[OH^-]$ for a solution of acid or base.
- define autoionization.
- state the $[H^+]$, $[OH^-]$, and K_w values for the autoionization of water.
- define pH and describe the pH scale.
- write the formulas for pH and pOH and express their values in a neutral solution at 25°C.
- explain the relationships among pH, pOH, and K_w.
- calculate $[H^+]$, $[OH^-]$, pH, and pOH given the value of any one of the other values.
- explain the relationship between the acidity or basicity of a solution and the hydronium ion concentration, $[H_3O^+]$, and the hydroxide ion concentration, $[OH^-]$, of the solution.
- predict whether an aqueous solution is acidic, basic, or neutral from $[H_3O^+]$, $[OH^-]$, or the pH.

Vocabulary

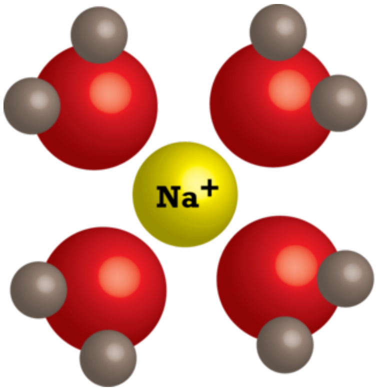
- autoionization
- hydronium ion
- ion product constant for water (K_w)
- pH
- pOH

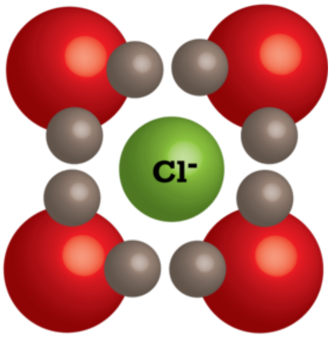
Introduction

We have learned many properties of water, such as pure water does not conduct electricity. The reason pure water does not conduct electricity is because the concentration of ions present when water ionizes is small. In this lesson, we will look a little closer at this property of water and how it relates to acids and bases.

The Hydronium Ion

Recall that ions in solution are hydrated. That is, water molecules are loosely bound to the ions by the attraction between the charge on the ion and the oppositely charged end of the polar water molecules, as illustrated in the figure below. When we write the formula for these ions in solution, we do not show the attached water molecules. It is simply recognized by chemists that ions in solution are always hydrated.





As with any other ion, a hydrogen ion dissolved in water will be closely associated with one or more water molecules. This fact is sometimes indicated explicitly by writing the **hydronium ion**, H_3O^+, in place of the hydrogen ion, H^+. Many chemists still use $H^+_{(aq)}$ to represent this situation, but it is understood that this is just an abbreviation for what is really occurring in solution. You are likely to come across both, and it is important for you to understand that they are actually describing the same entity. When using the hydronium ion in a chemical equation, you may need to add a molecule of water to the other side so that the equation will be balanced. This is illustrated in the equations below. Note that you are not really adding anything to the reaction. The (aq) symbol indicates that the various reaction components are dissolved in water, so writing one of these water molecules out explicitly in the equation does not change the reaction conditions.

$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$ (not showing hydronium)
$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$ (showing hydronium)

Relationship Between [H

Even totally pure water will contain a small amount of H^+ and OH^-. This is because water undergoes a process known as autoionization. **Autoionization** occurs when the same reactant acts as both the acid and the base. Look at the reaction below.

$$H_2O_{(aq)} + H_2O_{(aq)} \rightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$

The ionization of water is frequently written as:

$$H_2O_{(l)} \rightarrow H^+ + OH^-.$$

The equilibrium constant expression for this dissociation would be $K_w = [H^+][OH^-]$. From experimentation, chemists have determined that in pure water, $[H^+] = 1 \times 10^{-7}$ mol/L and $[OH^-] = 1 \times 10^{-7}$ mol/L.

Because this is a particularly important equilibrium, the equilibrium constant is given a subscript to differentiate it from other reactions. K_w, also known as the **ion product constant for water**, always refers to the autoionization of water. We can the calculate K_w because we know the value of $[H^+]$ and $[OH^-]$ for pure water at 25°C.

$$K_w = [H^+][OH^-]$$
$$K_w = (1 \times 10^{-7})(1 \times 10^{-7})$$
$$K_w = 1 \times 10^{-14}$$

A further definition of acids and bases can now be made:

When $[H_3O^+] = [OH^-]$ (as in pure water), the solution is *neutral*.
When $[H_3O^+] > [OH^-]$, the solution is an *acid*.
When $[H_3O^+] < [OH^-]$, the solution is a *base*.

Stated another way, an acid has a $[H_3O^+]$ that is greater than 1×10^{-7} and a $[OH^-]$ that is less than 1×10^{-7}. A base has a $[OH^-]$ that is greater than 1×10^{-7} and a $[H_3O^+]$ that is less than 1×10^{-7}.

The equilibrium between H^+, OH^-, and H_2O will exist in all water solutions, regardless of anything else that may be present in the solution. Some substances that are placed in water may become involved with either the hydrogen or hydroxide ions and alter the equilibrium state. However, as long as the temperature is kept constant at 25°C, the equilibrium will shift to maintain the equilibrium constant, K_w, at exactly 1×10^{-14}.

For example, a sample of pure water at 25°C has $[H^+]$ equal to 1×10^{-7} M and $[OH^-] = 1 \times 10^{-7}$ M. The K_w for this solution, of course, will be 1×10^{-14}. Suppose some HCl gas is added to this solution so that the H^+ concentration increases. This is a stress to the equilibrium system. Since the concentration of a product is increased, the reverse reaction rate will increase and the equilibrium will shift toward the reactants. The concentrations of both ions will be reduced until equilibrium is re-established. If the final $[H^+] = 1 \times 10^{-4}$ M, we can calculate the $[OH^-]$ because we know that the product of $[H^+]$ and $[OH^-]$ at equilibrium is always 1×10^{-14}.

$$K_w = [H^+][OH^-] = 1 \times 10^{-14}$$
$$[OH^-] = \frac{1 \times 10^{-14}}{[H^+]} = \frac{1 \times 10^{-14}}{1 \times 10^{-4}} = 1 \times 10^{-10}\text{ M}$$

Suppose, on the other hand, something is added to the solution that reduces the hydrogen ion concentration. As soon as the hydrogen ion concentration begins to decrease, the reverse rate decreases and the forward rate will shift the equilibrium toward the products. The concentrations of both ions will be increased until equilibrium is re-established. If the final hydrogen ion concentration is 1×10^{-12} M, we can calculate the final hydroxide ion concentration.

$$K_w = [H^+][OH^-] = 1 \times 10^{-14}$$
$$[OH^-] = \frac{1 \times 10^{-14}}{[H^+]} = \frac{1 \times 10^{-14}}{1 \times 10^{-12}} = 1 \times 10^{-2}\text{ M}$$

Using the K_w expression and our knowledge of the K_w value, as long as we know either the $[H^+]$ or the $[OH^-]$ in a water solution, we can always calculate the value for the other one.

Example:

What would be the $[H^+]$ for a grapefruit found to have a $[OH^-]$ of 1.26×10^{-11} mol/L? Is the solution acidic, basic, or neutral?

Solution:

$$K_w = [H^+][OH^-] = 1.00 \times 10^{-14}$$
$$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]} = \frac{1 \times 10^{-14}}{1.26 \times 10^{-11}} = 7.94 \times 10^{-4}\text{ M}$$

Since the $[H^+]$ in this solution is greater than 1×10^{-7} M, the solution is acidic.

pH and pOH

There are a few very concentrated acid and base solutions used in industrial chemistry and laboratory situations. For the most part, however, acid and base solutions that occur in nature, used in cleaning, and used in biochemistry applications are relatively dilute. Most of the acids and bases dealt with in laboratory situations have hydrogen ion concentrations between 1.0 M and 1.0×10^{-14} M. Expressing hydrogen ion concentrations in exponential numbers can become tedious, so a Danish chemist named Søren Sørensen developed a shorter method for expressing acid strength or hydrogen ion concentration with a non-exponential number. This value is referred to as **pH** and is defined by the following equation:

$$\text{pH} = -\log[\text{H}^+],$$

where p = log and H refers to the hydrogen ion concentration. The p from pH comes from the German word *potenz*, meaning power or the exponent of. Rearranging this equation to solve for $[\text{H}^+]$, we get $[\text{H}^+] = 10^{-\text{pH}}$. If the hydrogen ion concentration is between 1.0 M and 1.0×10^{-14} M, the value of the pH will be between 0 and 14.

Example:

Calculate the pH of a solution where $[\text{H}^+] = 0.01 \text{ mol/L}$.

Solution:

$$\begin{aligned}&\text{pH} = -\log(0.01)\\&\text{pH} = -\log(1 \times 10^{-2})\\&\text{pH} = 2\end{aligned}$$

Example:

Calculate the $[\text{H}^+]$ if the pH is 4.

Solution:

$$\begin{aligned}&[\text{H}^+] = 10^{-\text{pH}}\\&[\text{H}^+] = 10^{-4}\\&[\text{H}^+] = 1 \times 10^{-4} \text{ mol/L}\end{aligned}$$

Example:

Calculate the pH of saliva, where $[\text{H}^+] = 1.58 \times 10^{-6} \text{ mol/L}$.

Solution:

$$\begin{aligned}&\text{pH} = -\log[\text{H}^+] = -\log(1.58 \times 10^{-6})\\&\text{pH} = 5.8\end{aligned}$$

Example:

Fill in the rest of **Table** 6.1.

TABLE 6.1: Hydrogen ion concentration and corresponding pH.

$[\text{H}^+]$ in mol/L	$-\log[\text{H}^+]$	pH
0.1	1.00	1.00
0.2	0.70	0.70
1×10^{-5}	?	?
?	?	6.00
0.065	?	?
?	?	9.00

Solution:

The completed table is shown below (**Table** 6.2).

TABLE 6.2: Hydrogen ion concentration and corresponding pH.

$[H^+]$ in mol/L	$-\log[H^+]$	pH
0.1	1.00	1.00
0.2	0.70	0.70
1.00×10^{-5}	5	5
1.00×10^{-6}	6.00	6.00
0.065	1.19	1.19
1.00×10^{-9}	9.00	9.00

An acid with pH = 1, then, is stronger than an acid with pH = 2 by a factor of 10. Simply put, lower pH values correspond to higher H^+ concentrations and more acidic solutions, while higher pH values correspond to higher OH^- concentrations and more basic solutions. This is illustrated in the figure below. It should be pointed out that there are acids and bases that fall outside the pH range depicted. However, we will confine ourselves for now to those falling within the 0-14 range, which covers $[H^+]$ values from 1.0 M all the way down to 1×10^{-14} M.

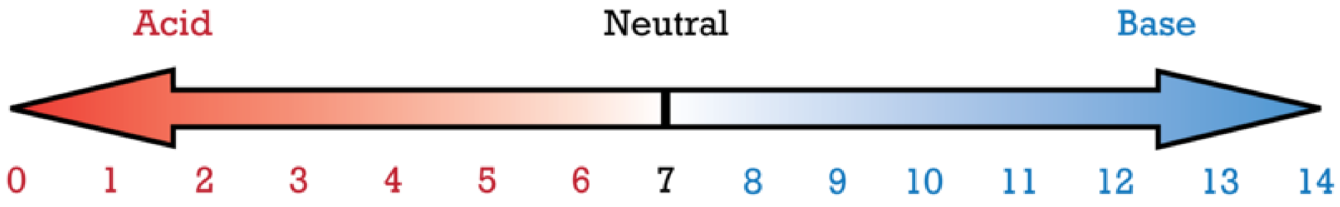
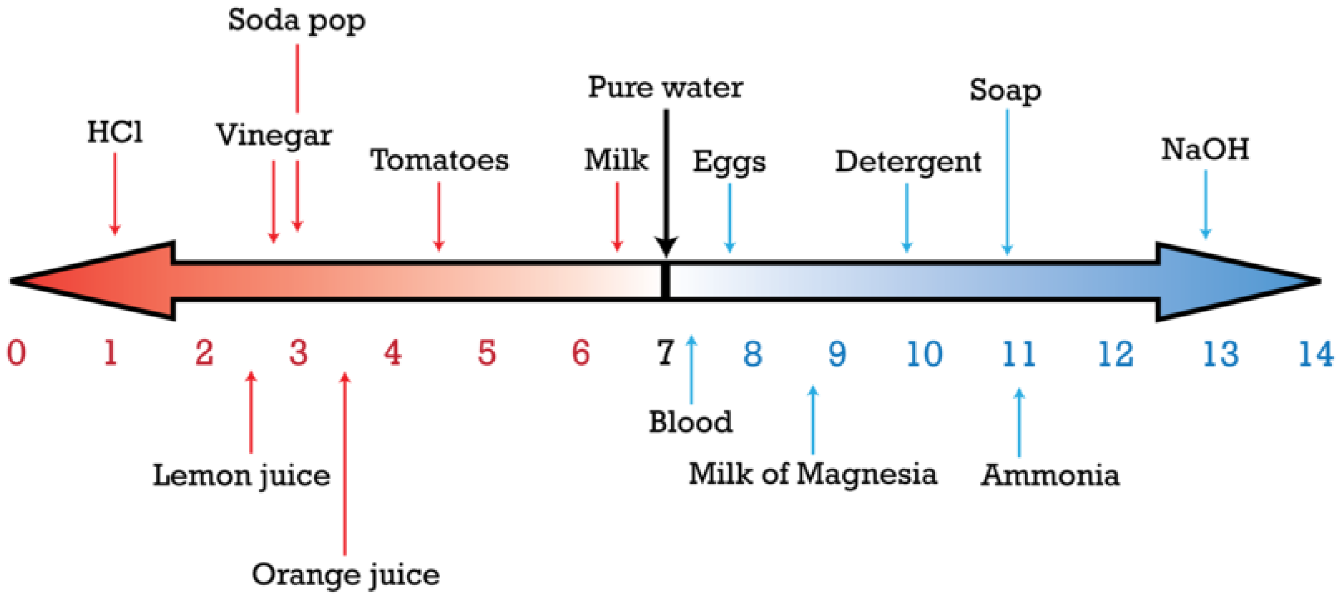


TABLE 6.3: pH versus Acidity

pH level	Solution
$pH < 7$	Acid
$pH = 7$	Neutral
$pH > 7$	Basic

Have you ever cut an onion and had your eyes water up? This is because of a compound with the formula C_3H_6OS that is found in onions. When you cut the onion, a variety of reactions occur that release a gas. This gas can diffuse into the air and eventfully mix with the water found in your eyes to produce a dilute solution of sulfuric acid. This is what irritates your eyes and causes them to water. There are many common examples of acids and bases in our everyday lives. Look at the pH scale below to see how these common examples relate in terms of their pH.



Even though both acidic and basic solutions can be expressed by pH, an equivalent set of expressions exists for the concentration of the hydroxide ion in water. This value, referred to as **pOH**, is defined as:

$$\text{pOH} = -\log[\text{OH}^-]$$

If the pOH is greater than 7, the solution is acidic. If the pOH is equal to 7, the solution is neutral. If the pOH is less than 7, the solution is basic.

If we take the negative log of the complete K_w expression, we obtain:

$$\begin{aligned}
&K_w = [\text{H}^+][\text{OH}^-] \\
&-\log K_w = (-\log[\text{H}^+]) + (-\log[\text{OH}^-]) \\
&-\log(1 \times 10^{-14}) = (-\log[\text{H}^+]) + (-\log[\text{OH}^-]) \\
&14 = \text{pH} + \text{pOH}
\end{aligned}$$

Therefore, the sum of the pH and the pOH is always equal to 14 (at 25°C). Remember that the pH scale is written with values from 0 to 14 because many useful acid and base solutions fall within this range. Now let's go through a few examples to see how this calculation works for problem-solving in solutions with an added acid or base.

Example:

What is the $[\text{H}^+]$ for a solution of NH_3 whose $[\text{OH}^-] = 8.23 \times 10^{-6}$ mol/L?

Solution:

$$\begin{aligned}
&[\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14} \\
&[\text{H}_3\text{O}^+] = \frac{1.00\times10^{-14}}{[\text{OH}^-]} = \frac{1.00\times10^{-14}}{8.23\times10^{-6}} = 1.26 \times 10^{-9}\ \text{M}
\end{aligned}$$

Example:

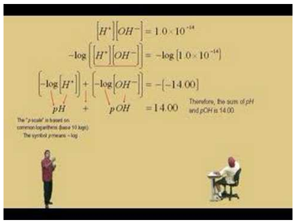
Black coffee has a $[\text{H}_3\text{O}^+] = 1.26 \times 10^{-5}$ mol/L. What is the pOH?

Solution:

$$\text{pH} = -\log[\text{H}^+] = -\log 1.26 \times 10^{-5} = 4.90$$

$\text{pH} + \text{pOH} = 14$
$\text{pOH} = 14 - \text{pH} = 14 - 4.90 = 9.10$

For a classroom demonstration of pH calculations **(5d, 5f; 1e IE Stand.)**, see http://www.youtube.com/watch?v=lca_puB1R8k (9:45).



MEDIA

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Lesson Summary

- Autoionization is the process where the same molecule acts as both an acid and a base.
- Water ionizes to a very slight degree according to the equation $H_2O_{(l)} \leftrightharpoons [H^+] + [OH^-]$.
- In pure water at 25°C, $[H^+] = [OH^-] = 1.00 \times 10^{-7}$ M.
- The equilibrium constant for the dissociation of water, K_w, is equal to 1.00×10^{-14} at 25°C.
- $\text{pH} = -\log[H^+]$
- $\text{pOH} = -\log[OH^-]$
- $\text{p}K_w = -\log K_w$
- $\text{pH} + \text{pOH} = \text{p}K_w = 14.0$

Further Reading / Supplemental Links

The websites below have more information about pH.

- http://www.johnkyrk.com/pH.html
- http://purchon.com/chemistry/ph.htm

Review Questions

1. What is the $[H^+]$ ion concentration in a solution of 0.350 mol/L H_2SO_4?
   a. 0.175 mol/L
   b. 0.350 mol/L
   c. 0.700 mol/L
   d. 1.42×10^{-14} mol/L
2. A solution has a pH of 6.54. What is the concentration of hydronium ions in the solution?
   a. 2.88×10^{-7} mol/L
   b. 3.46×10^{-8} mol/L
   c. 6.54 mol/L
   d. 7.46 mol/L

3. A solution has a pH of 3.34. What is the concentration of hydroxide ions in the solution?
    a. 4.57×10^{-4} mol/L
    b. 2.19×10^{-11} mol/L
    c. 3.34 mol/L
    d. 10.66 mol/L
4. A solution contains 4.33×10^{-8} M hydroxide ions. What is the pH of the solution?
    a. 4.33
    b. 6.64
    c. 7.36
    d. 9.67
5. Fill in **Table** 6.4 and rank the solutions in terms of increasing acidity.

TABLE 6.4: Table for Problem 5

Solutions	$[H^+]$ (mol/L)	$-\log\,[H^+]$	pH
A	0.25	0.60	0.60
B	?	2.90	?
C	1.25×10^{-8}	?	?
D	0.45×10^{-3}	?	?
E	?	1.26	?

6. It has long been advocated that red wine is good for the heart. Wine is considered to be an acidic solution. Determine the concentration of hydronium ions in wine with pH 3.81.
7. What does the value of K_w tell you about the autoionization of water?
8. If the pH of an unknown solution is 4.25, what is the pOH?
    a. $10^{-4.25}$
    b. $10^{-9.75}$
    c. 9.75
    d. $14.0 - 10^{-9.75}$
9. A solution contains a hydronium ion concentration of 3.36×10^{-4} mol/L. What is the pH of the solution?
    a. 3.36
    b. 3.47
    c. 10.53
    d. none of the above
10. A solution contains a hydroxide ion concentration of 6.43×10^{-9} mol/L. What is the pH of the solution?
    a. 5.80
    b. 6.48
    c. 7.52
    d. 8.19
11. An unknown solution was found in the lab. The pH of the solution was tested and found to be 3.98. What is the concentration of hydroxide ions in this solution?
    a. 3.98 mol/L
    b. 0.67 mol/L
    c. 1.05×10^{-4} mol/L
    d. 9.55×10^{-11} mol/L

6.4 Strength of Acids and Bases

Lesson Objectives

The student will:

- describe the difference between a strong and weak acid.
- identify specific acids as strong or weak.
- define weak acids and bases.
- use K_a or K_b to find $[\mathrm{H}^+]$ and vice versa.
- use K_a or K_b to find pH.

Vocabulary

- K_a
- K_b
- strong acid
- strong base
- weak acid
- weak base

Introduction

A great number of people associate strong acids with their ability to react with skin, essentially "melting" it away from bone. On a popular crime show, this very chemistry was used as a method for a crime. The crime show used sulfuric acid. Why sulfuric acid and not acetic acid? What makes the difference? How can we tell if an acid is strong or weak? The answers to these questions will be the focus of the lesson that follows.

Strong Acids and Bases

Strong acids and **strong bases** are those that completely dissociate when dissolved in water. HCl is an example of a strong acid, as seen in the equation below:

$$\mathrm{HCl}_{(g)} + \mathrm{H_2O}_{(l)} \rightarrow \mathrm{H_3O}^+_{(aq)} + \mathrm{Cl}^-_{(aq)}$$

Notice that this equation only has a forward arrow instead of the equilibrium arrows seen in many dissociation reactions. This indicates that the reaction goes to completion, so when the reaction is over, there are essentially no HCl molecules left in solution, only H^+ and Cl^- ions. Since strong acids fully dissociate, many ions are produced, making the solution a good conductor of electricity.

There are only six common strong acids. These acids are shown in **Table** 6.5. Each of the acids found in this table will completely dissociate in water.

TABLE 6.5: **Strong Acids**

Name	Formula
Hydrochloric acid	HCl
Hydrobromic acid	HBr
Hydroiodic	HI
Nitric acid	HNO_3
Perchloric acid	$HClO_4$
Sulfuric acid	H_2SO_4

Weak Acids and Bases

Weak acids and **weak bases** do *not* completely dissociate when dissolved in water. The less dissociation that takes place, the weaker the acid since there will be fewer H^+ ions in solution. For example, when acetic acid is placed in water, only about 5% of the acetic acid molecules separate into H^+ ions and $C_2H_3O_2^-$ ions:

$$HC_2H_3O_{2(aq)} \rightleftharpoons H^+_{(aq)} + C_2H_3O^-_{2(aq)}$$

Notice that we now use the equilibrium arrows in the chemical equation. At equilibrium, this solution will contain more acetic acid molecules than hydronium and acetate ions.

Let's look at citric acid, a weak acid. Citric acid, $C_6H_8O_7$, is commonly found in everyday products like lemons, limes, and soft drinks. It is the substance responsible for making these foods taste sour. If we were to write an ionization equation for citric acid, it would appear as written below:

$$C_6H_8O_{7(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + C_6H_7O^-_{7(aq)}$$

All acids that you will encounter at this level are weak acids if they are not one of the six listed in **Table** 6.5 are weak. Even though these weak acids are very soluble in water, they dissolve as molecules, and only a few of the molecules break into ions in the solution. Since weak acids do not produce many ions, a weak acid solution will only conduct a small electric current.

Example:

Write dissociation equations for only those acids that are weak.

1. Sulfuric acid (H_2SO_4)
2. Hydrofluoric acid (HF)
3. Trichloroacetic acid (CCl_3COOH)

Solution:

1. H_2SO_4 is a strong acid (one of the six).
2. HF is a weak acid (not one of the six); $HF_{(aq)} \rightleftharpoons H^+_{(aq)} + F^-_{(aq)}$
3. CCl_3COOH is a weak acid (not one of the six). $CCl_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CCl_3COO^-_{(aq)}$

Weak Acids and Weak Bases as Equilibrium Systems

The dissociation of a weak acid is another special type of reaction that is given its own equilibrium constant. The equilibrium constant for the dissociation of a weak acid is designated **K**. The dissociation reaction for acetic acid, one of the primary components of vinegar, is shown below.

$$HC_2H_3O_{2(aq)} \rightleftharpoons H^+_{(aq)} + C_2H_3O^-_{2(aq)}$$

The equilibrium constant, K_a, for acetic acid is 1.8×10^{-5}. This value is small, indicating that the equilibrium position lies more to the left than to the right. In other words, there are more acetic acid molecules at equilibrium than there are acetate ions or hydronium ions. Since K_a is less than 1, we know that $[HC_2H_3O_2] > [C_2H_3O_2^-][H^+]$ at equilibrium.

Similarly, the equilibrium constant for the reaction between a weak base and water is designated **K**. An example of a weak base reaction is shown below.

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH^+_{4(aq)} + OH^-_{(aq)}$$

The equilibrium constant for this reaction is, coincidentally, also 1.8×10^{-5}. Again, a value much less than one indicates that the equilibrium favors the reactants. In this solution, very few ammonia molecules are able to remove a proton from water to create ammonium and hydroxide ions.

Example:

Put the following acids in order of decreasing acid strength. Write an equilibrium expression for each. Remember dissociation increases with increasing acid strength (or decreasing pH).

Formic acid (HCOOH) $K_a = 6.3 \times 10^{-4}$
Phosphoric acid (H_3PO_4) $K_a = 7.2 \times 10^{-3}$
Oxalic Acid (HO_2CCO_2H) $K_a = 5.6 \times 10^{-2}$
Arsenic acid (H_3AsO_4) $K_a = 6.0 \times 10^{-3}$

Solution:

Order of decreasing acid strength: Oxalic acid > Phosphoric acid > Arsenic acid > Formic acid

Equilibrium Equations:

1. Oxalic acid: $HO_2CCO_2H_{(aq)} \rightleftharpoons H^+_{(aq)} + HO_2CCO^-_{2(aq)}$
2. Phosphoric acid: $H_3PO_{4(aq)} \rightleftharpoons H^+_{(aq)} + H_2PO^-_{4(aq)}$
3. Arsenic acid: $H_3AsO_{4(aq)} \rightleftharpoons H^+_{(aq)} + H_2AsO^-_{4(aq)}$
4. Formic acid: $HCOOH_{(aq)} \rightleftharpoons H^+_{(aq)} + HCOO^-_{(aq)}$

Equilibrium Constants for Acid and Base Dissociation

The pH for solutions of strong acids and strong bases can be calculated simply by knowing the original concentration of acid or base. Consider a solution that is 0.010M HCl. HCl is a strong acid, so the acid molecules dissociate completely. At equilibrium, this solution of HCl will be 0.010 M in H^+ ions and 0.010 M in Cl^- ions. Plugging the value of the hydrogen ion concentration into the pH formula, we can determine that this solution has a pH of 2.

Consider a 0.0010 M NaOH solution. NaOH is a strong base, so this solution will be 0.0010 M in sodium ions and also in hydroxide ions. Since the solution is 1.0×10^{-3}M in hydroxide ions, it will be 1.0×10^{-11} M in hydrogen ions. Therefore, this solution will have a pH = 11.

A strong acid such as H_2SO_4 is only slightly more complicated. Suppose we wish to determine the pH of a 0.00010 M solution of H_2SO_4.

$$\mathrm{H_2SO_{4(\mathit{aq})}} \rightarrow 2\ \mathrm{H^+} + \mathrm{SO_4^{2-}}$$

Since the original solution was 0.00010 M and sulfuric acid is a strong acid, then complete dissociation will produce a solution that contains $[\mathrm{H^+}] = 0.00020$ M. Substituting this hydrogen ion concentration in the pH formula yields:

$$\mathrm{pH} = -\log(2.0 \times 10^{-4}) = -(0.30 - 4) = -(-3.7) = 3.7$$

Let's now consider the process for finding the pH of weak acids and bases. In these cases, you need more information than you need for strong acids and bases. Not only do you need to know the concentration of the original acid or base solution, but you also must know the K_a or K_b. Suppose we wish to know the pH of a 1.0 M solution of ascorbic acid, $\mathrm{H_2C_6H_6O_{6(\mathit{aq})}}$, whose $K_a = 7.9 \times 10^{-5}$.

$$\mathrm{H_2C_6H_6O_{6(\mathit{aq})}} \rightleftharpoons \mathrm{H^+_{(\mathit{aq})}} + \mathrm{HC_6H_6O^-_{6(\mathit{aq})}} \qquad K_a = 7.9 \times 10^{-5}$$

The K_a expression for this reaction would be written

$$K_a = \frac{[\mathrm{H^+}][\mathrm{HC_6H_6O_6^-}]}{[\mathrm{H_2C_6H_6O_6}]}$$

To find the hydrogen ion concentration from this K_a expression and the original concentration of the acid, we need a little algebra. Let the molarity of the acid that has already dissociated be represented by x. In other words, the molarity of the hydrogen ions and ascorbate ions in solution will also be represented by x. The molarity of the undissociated acid must therefore be $1.0 - x$. We can now substitute these variables into the K_a expression and set it equal to the given K_a value.

$$K_a = \frac{(x)(x)}{(1.0-x)} = 7.9 \times 10^{-5}$$

When this equation is simplified, we find that it is a quadratic equation, which, of course, can be solved by the quadratic formula.

$$x^2 + (7.9 \times 10^{-5})x - (7.9 \times 10^{-5}) = 0$$
$$x = 8.9 \times 10^{-3}\ \mathrm{M}$$

However, there is a shortcut available to solve this problem that simplifies the math greatly. It involves significant figures and adding or subtracting extremely small numbers from large numbers. If you are working to 3 significant figures and you are required to subtract 0.00005 from 1.00, when you carry out the subtraction and round to 3 significant figures, you discover that you get the original number before you subtracted.

$1.00 - 0.00005 = 0.99995$, which to 3 significant figures is 1.00.

In the problem we solved above about ascorbic acid, the K_a value is very small, 1.8×10^{-5}. This indicates that the amount of ascorbic acid that dissociates, represented by x, is tiny. When we assigned the variables in that problem, we see that the molarity of ascorbic acid remaining after dissociation is represented by $1.0 - x$. Since this x is very tiny, the result of this subtraction will still be 1.0 M. Therefore, the K_a expression from above,

$$K_a = \frac{(x)(x)}{(1.0-x)} = 7.9 \times 10^{-5}$$

can quite safely be written as

$$K_a = \frac{(x)(x)}{(1.0)} = 7.9 \times 10^{-5}$$

We are assuming x is so small that it will not alter the value of 1.0 when we subtract and then round to proper significant figures. This is a safe assumption when the value of the K_a is very small (less than 1×10^{-3}).

When we solve this second expression,

$$\frac{(x)(x)}{(1.0)} = 7.9 \times 10^{-5}$$
$$x^2 = 7.9 \times 10^{-5}$$
$$x = 8.9 \times 10^{-3}\text{ M}$$

We no longer need to use the quadratic formula, and note that the answer (when rounded to the proper number of significant figures) is exactly the same as when the expression was solved with the quadratic formula.

Let's go through another example using a hypothetical weak acid, 0.10 M HA, whose $K_a = 4.0 \times 10^{-7}$.

$$HA_{(aq)} \rightleftharpoons H^+ + A^- \quad K_a = 4.0 \times 10^{-7}$$
$$K_a = \frac{[H^+][A^-]}{[HA]} = 4.0 \times 10^{-7}$$

Let x represent the molarity of HA that dissociates, so $[H^+] = [A^-] = x$, and $[HA] = 0.10 - x$.

$$K_a = \frac{[x][x]}{[0.10-x]} = 4.0 \times 10^{-7}$$

Once again, since the K_a value is very small, the x subtracted in the denominator can be neglected. and the equation becomes

$$K_a = \frac{[x][x]}{[0.10]} = 4.0 \times 10^{-7}$$
$$x^2 = 4.0 \times 10^{-8}$$
$$x = 2.0 \times 10^{-4}\text{ M}$$

Therefore, the hydrogen ion concentration in this solution is 2.0×10^{-4} M. Substituting this value into the pH formula yields:

$$\text{pH} = -\log(2.0 \times 10^{-4}) = -(0.3 - 4) = -(-3.7) = 3.7$$

The same process is used for weak bases. There is one additional step when working with weak bases: once the hydroxide ion concentration is determined, you must then find the hydrogen ion concentration before substituting the value into the pH formula. K_b represents the equilibrium constant for the reduction between a weak base and water. Consider dimethylamine (a weak base used in making detergents). We will calculate the pH of a 1.0 M solution of this weak base.

$$(\text{CH}_3)_2\text{NH}_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons (\text{CH}_3)_2\text{NH}^+_{2(aq)} + \text{OH}^-_{(aq)} \quad K_b = 5.9 \times 10^{-7}$$
$$K_b = \frac{[(\text{CH}_3)_2\text{NH}_2^+][\text{OH}^-}{[(\text{CH}_3)_2\text{NH}} = 5.9 \times 10^{-7}$$

Allowing x to represent the molarity of $(\text{CH}_3)_2\text{NH}$ that dissociates results in $[(\text{CH}_3)_2\text{NH}_2^+] = x$ and $[\text{OH}^-] = x$. The molarity of undissociated $(\text{CH}_3)_2\text{NH}$ will be $1.0 - x$.

Substituting the variables into the K_b expression yields

$$K_b = \frac{[x][x]}{[1.0-x]} = 5.9 \times 10^{-7}$$

and neglecting the x in the denominator because it is beyond the significant figures of the problem yields

$$K_b = \frac{[x][x]}{[1.0]} = 5.9 \times 10^{-7}$$

Therefore,

$$x^2 = 5.9 \times 10^{-7}$$
$$x = 7.7 \times 10^{-4} \text{ M}$$

Now that we know the hydroxide ion concentration in the solution, we calculate the hydrogen ion concentration by dividing the $[\text{OH}^-]$ into the K_w. This will yield $[\text{H}^+] = 1.3 \times 10^{-11}$ M. The final step is to plug the hydrogen ion concentration into the pH formula.

$$\text{pH} = -\log(1.3 \times 10^{-11}) = 10.9$$

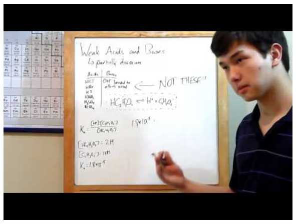
Example:

Acetic acid is mixed with water to form a 0.10 mol/L $HC_2H_3O_{2(aq)}$ solution at 25°C. If the equilibrium concentrations of $H_3O^+_{(aq)}$ and $C_2H_3O^-_{2(aq)}$ are both 1.34×10^{-3} mol/L and the equilibrium concentration of $HC_2H_3O_{2(aq)}$ is 0.0999 mol/L, determine the K_a and the pH of the solution at equilibrium.

Solution:

$$HC_2H_3O_{2(aq)} \rightleftharpoons H^+_{(aq)} + C_2H_3O^-_{2(aq)}$$
$$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = \frac{(1.34\times10^{-3})(1.34\times10^{-3})}{0.0999} = 1.80 \times 10^{-5}$$
$$\text{pH} = -\log[H^+] = -\log(1.34 \times 10^{-3}) = 2.87$$

This video shows an example of how to plug values into the Ka for a weak acid to find [H+] and pH **(5c)**: http://www.youtube.com/watch?v=MOv7Z16FMK0 (3:26).



MEDIA

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Lesson Summary

- Strong acids and bases undergo 100% dissociation in water.
- Weak acids undergo less than 100% dissociation in water.
- It is safe to assume that an acid is weak unless it is one of the six strong acids listed in the chapter.
- K_a represents the equilibrium constant for the dissociation of a weak acid.
- K_b represents the equilibrium constant for the dissociation of a weak base.
- Equilibrium calculations are the same for weak acids and bases as they were for all other equilibrium reactions.

Further Reading / Supplemental Links

Visit the link below for more information about the strengths of acids and bases.

- http://chemistry.about.com/od/acidsbases/a/strengthacids.htm

Review Questions

1. What makes weak acids and bases a special case for equilibrium reactions?
2. What do the constants K_a and K_b represent?
3. Oxalic acid is a weak acid. Its ionization reaction is represented as: $H_2C_2O_{4(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + HC_2O_4{}^-{}_{(aq)}$. Which of the following best represents the acid ionization constant expression, K_a?
   a. $K_a = \frac{[H_3O^+][HC_2O_4^-]}{[H_2C_2O_4][H_2O]}$
   b. $K_a = \frac{[H_2C_2O_4][H_2O]}{[H_3O^+][HC_2O_4^-]}$
   c. $K_a = \frac{[H_3O^+][HC_2O_4^-]}{[H_2C_2O_4]}$
   d. $K_a = \frac{[H_2C_2O_4]}{[H_3O^+][HC_2O_4^-]}$
4. Choose the weakest acid from the list below.
   a. $HNO_{2(aq)}; K_a = 5.6 \times 10^{-3}$
   b. $HF_{(aq)}; K_a = 6.6 \times 10^{-4}$
   c. $H_3PO_{4(aq)}; K_a = 6.9 \times 10^{-3}$
   d. $HCOOH_{(aq)}; K_a = 1.8 \times 10^{-4}$
5. Choose the reaction that would have an equilibrium constant represented by K_b.
   a. $H_2PO_4{}^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HPO_4{}^{-2}{}_{(aq)} + H_3O^+{}_{(aq)}$
   b. $NH_4{}^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^+{}_{(aq)}$
   c. $NH_4{}^-{}_{(aq)} + OH^-{}_{(aq)} \rightleftharpoons NH_{3(aq)} + H_2O_{(l)}$
   d. $F^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HF_{(aq)} + OH^-{}_{(aq)}$
6. A 0.150 mol/L solution of a weak acid having the general formula HA is 15.0% ionized in aqueous solution. Which expression best represents the calculation of the acid ionization constant K_a for this acid?
   a. $K_a = \frac{(0.150)(0.150)}{(0.150)}$
   b. $K_a = \frac{(0.0225)(0.0225)}{(0.128)}$
   c. Not enough information is given.
7. Put the following bases in order of increasing base strength. Write equilibrium reactions for each.
   a. ethanolamine $(HOCH_2CH_2NH_2)$, $K_b = 3.2 \times 10^{-5}$
   b. piperidine $(C_5H_{10}NH)$, $K_b = 1.3 \times 10^{-3}$
   c. triethylamine $((CH_3CH_2)_3N)$, $K_b = 5.2 \times 10^{-4}$
   d. ethylenediamine $(H_2NCH_2CH_2NH_2)$, $K_b = 8.5 \times 10^{-5}$

6.5 Brønsted-Lowry Acids and Bases

Lesson Objectives

The student will:

- define Brønsted-Lowry acids and bases.
- identify Brønsted-Lowry acids and bases in chemical equations.
- define conjugate acids and bases.
- write the formula for the conjugate acid of any base and for the conjugate base of any acid.
- identify conjugate acids and bases in equations.
- given the strengths of acids and bases, identify the strength the conjugate acids and bases.

Vocabulary

- amphoteric
- Brønsted-Lowry acid
- Brønsted-Lowry base
- conjugate acid
- conjugate base

Introduction

Arrhenius provided chemistry with the first definition of acids and bases, but like a lot of scientific ideas, these definitions have been refined over time. Two chemists, named Brønsted and Lowry, derived a more generalized definition for acids and bases that we use in conjunction with the Arrhenius theory. The Brønsted-Lowry theory is the focus of this lesson. As the Brønsted-Lowry definition unfolded, the number of acids and bases that were able to fit into each category increased.

Brønsted-Lowry Acids and Bases

Arrhenius made great in-roads into the understanding of acids and bases and how they behaved in chemical reactions. Brønsted and Lowry slightly altered the Arrhenius definition and greatly increased the number of compounds that qualify as bases. A **Brønsted-Lowry acid** is a substance that is a proton donor, and a **Brønsted-Lowry base** is a proton acceptor. Look at the equation below in which hydrochloric acid is reacting with water:

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

The Brønsted-Lowry concept of acids and bases states that the acid donates a proton and the base accepts a proton. Therefore, HCl acts as the acid and donates an H^+ ion to H_2O to form Cl^-. The H_2O acts as the base and accepts an H^+ ion from HCl to form H_3O^+. Look at the equation below for another example of Brønsted-Lowry acid and base:

$$H_2PO^-_{4(aq)} + OH^-_{(aq)} \rightleftharpoons HPO^{2-}_{4(aq)} + H_2O_{(l)}$$

The equation shows $H_2PO_4^-$ donating a proton to OH^- and forming HPO_4^{2-}, while OH^- is accepting the proton to form H_2O. Thus, $H_2PO_4^-$ is acting as the acid and OH^- is acting as the base.

Example:

Identify the Brønsted-Lowry acids and bases from each of the following equations:

1. $HC_2H_3O_{2(aq)} \rightleftharpoons C_2H_3O^-_{2(aq)} + H_3O^+_{(aq)}$
2. $HCN_{(aq)} + H_2O_{(l)} \rightleftharpoons CN^-_{(aq)} + H_2O^+_{(aq)}$

Solution:

1. $HC_2H_3O_2$ is acting as the acid and H_2O is acting as the base.
2. HCN is acting as the acid and H_2O is acting as the base.

If you think about the definition of an Arrhenius acid, it includes substances such as HCl, HNO_3, $HC_2H_3O_2$ - in essence all substances that contain H^+ ions. This is because according to Arrhenius, the acid dissociates in water to produce H^+ ions. This definition limits what can fit under the umbrella of the definition of acid. The Brønsted-Lowry definition of the acid is broader in that it defines the acid as a proton donor. With this broader definition there is the ability to include more compounds in the category of acid.

It needs to be pointed out that if a substance is an acid in the Arrhenius definition, it will be an acid in the Brønsted–Lowry definition. The same relationship holds for bases. The reverse, however, is not true. Consider the equation below. The hydroxide ion, OH^-, is both an Arrhenius base and a Brønsted–Lowry base. In other words, the Brønsted-Lowry definition can be viewed as an extension to the Arrhenius definition rather than a replacement of it.

$$HSO^-_{4(aq)} + OH^-_{(aq)} \rightleftharpoons H_2O_{(l)} + SO^{2-}_{4(aq)}$$

With the Arrhenius theory, water was a necessary part of the equation. Arrhenius said that an acid must produce H^+ ions in a water solution. Therefore, the following equation would be representative of an Arrhenius acid:

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

Now consider the following reaction:

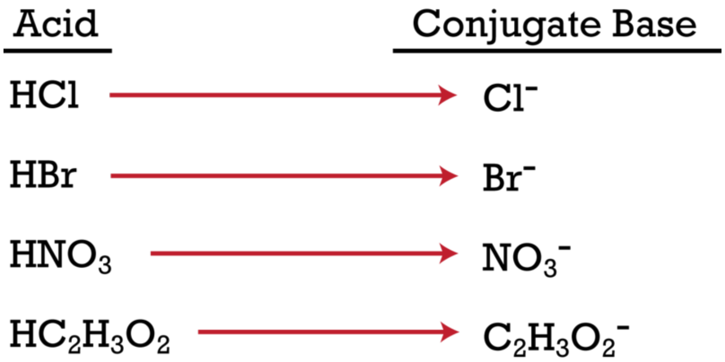
$$NH_{3(aq)} + NH_{3(aq)} \rightleftharpoons NH^+_{4(aq)} + NH^-_{2(aq)}$$

The first NH_3 molecule is accepting a proton to form NH_4^+ and is therefore a Brønsted-Lowry base, the second NH_3 molecule is donating a proton to form NH_2^- and is therefore a Brønsted-Lowry acid. Ammonia molecules, however, do not donate hydrogen ions in water, so they do not qualify as Arrhenius acids. The Brønsted-Lowry theory has provided a broader theory for acid-base chemistry.

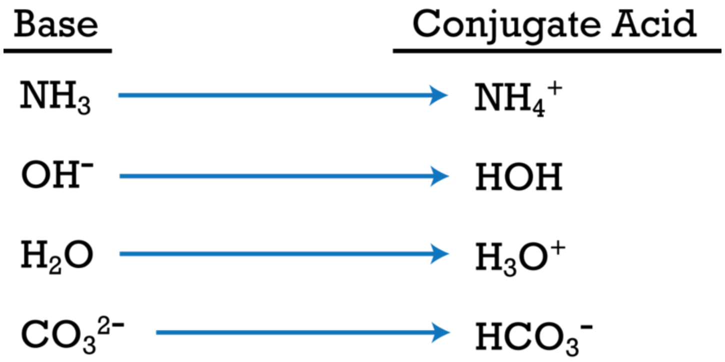
It should be noted that NH_3 is an example of an amphoteric species. **Amphoteric** species are those that can act as either an acid or a base, depending on the situation. That is, in some circumstances, they donate a proton, and in other circumstances, they accept a proton. Water is another example of an amphoteric species.

Acid-Base Conjugate Pairs

There is one more aspect of the Brønsted-Lowry theory that was a significant breakthrough to acid-base chemistry. Brønsted and Lowry said that in acid-base reactions, there are actually pairs of acids and bases in the reaction itself. According to Brønsted-Lowry, for every acid there is a conjugate base associated with that acid. The **conjugate base** is the result of the acid losing (or donating) a proton. Therefore, if you look at the figure below, you can see the acid on the left and the conjugate base on the right.



Notice that the difference between the acid and its conjugate base is simply a proton. Similarly, for every base in the acid-base reaction, there must be a corresponding conjugate acid. The **conjugate acid** is the result of the base gaining (or accepting) the proton. Look at the figure below to see the difference between the base and the corresponding conjugate acid.



Now that we know what conjugate acids and bases are, let's try to identify them in acid-base reactions. Consider the reaction between acetic acid and water shown below:

$$HC_2H_3O_{2(aq)} + H_2O_{(l)} \rightleftharpoons C_2H_3O^-_{(aq)} + H_3O^+_{(aq)}$$

Step 1: Identify the acid and base on the reactants side.

$HC_2H_3O_2$ is the acid and H_2O is the base.

Step 2: Identify the conjugate acid and base on the products side.

Look at the product side to see what product has gained a proton (this is the conjugate acid) and which product has lost a proton (this is the conjugate base).

$C_2H_3O_2^-$ is the conjugate base of $HC_2H_3O_2$, and H_3O^+ is the conjugate acid of H_2O.

As a result, the conjugate acid/base pairs are $HC_2H_3O_2$ / $C_2H_3O_2^-$ and H_3O^+ / H_2O.

Let's consider another example. Identify the conjugate acid-base pairs in the following equation.

$$CH_3NH_{2(aq)} + HClO_{(aq)} \rightleftharpoons ClO^-_{(aq)} + CH_3NH^+_{3(aq)}$$

Step 1: Identify the acid and base on the reactants side.

CH_3NH_2 is the base and HClO is the acid.

Step 2: Identify the conjugate acid and base on the products side.

ClO^- is the conjugate base and $CH_3NH_3^+$ is the conjugate acid.

Hence, the conjugate acid/base pairs are CH_3NH_2 / $CH_3NH_3^+$ and HClO / ClO^-.

Example:

Identify the conjugate acid-base conjugate pairs in each of the following equations:

1. $NH_{3(aq)} + HCN_{(aq)} \rightleftharpoons NH^+_{4(aq)} + CN^-_{(aq)}$
2. $CO^{2-}_{3(aq)} + H_2O_{(l)} \rightleftharpoons HCO^-_{3(aq)} + OH^-_{(aq)}$

Solution:

1. NH_3 / NH_4^+ and HCN / CN^-
2. CO_3^{2-} / HCO_3^- and H_2O / OH^-

The Strength of Conjugate Acids and Bases

As a general rule, the stronger an acid is, the weaker its conjugate base will be. The same is true for the relationship between bases and their conjugate acids. Overall:

Strong acid → Very weak conjugate base
Weak acid → Weak conjugate base
Weak base → Weak conjugate acid
Strong base → Very weak conjugate acid

Since acid strength and conjugate base strength are inversely related, you might think that strong acids would have weak conjugate bases, while weak acids would have strong conjugate bases. Although this is true in a relative sense, it does not match up to the way we have been defining "strong" acids and bases so far. This can often be confusing for students just beginning their study of chemistry.

Let's re-examine how we have defined strong and weak acids and bases. This can be illustrated most clearly by looking at some representative examples. Consider the following equilibria:

$$HCl_{(aq)} + H_2O_{(l)} \rightleftharpoons Cl^-_{(aq)} + H_3O^+_{(aq)} \qquad K_a = \text{Very large}$$
$$CH^-_{3(aq)} + H_2O_{(l)} \rightleftharpoons CH_{4(g)} + OH^-_{(aq)} \qquad K_b = \text{Very large}$$
$$CH_3CO_2H_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3CO^-_{2(aq)} + H_3O^+_{(aq)} \qquad K_a = 1.74 \times 10^{-5}$$

$$CH_3CO^-_{2(aq)} + H_2O_{(l)} \rightleftharpoons CH_3CO_2H_{(aq)} + OH^-_{(aq)} \qquad K_b = 5.6 \times 10^{-10}$$

HCl is an example of a strong acid because it completely dissociates in water. Another way to say this is that HCl is a very strong acid *compared to* H_3O^+, because when Cl^- (its conjugate base) and water are competing for a proton, water wins nearly every time. Cl^- is such a weak base that, when considered outside the context of an acid/base reaction, you would not even consider it a base. For example, a solution of NaCl contains Cl^- ions, but you would never say that the solution is basic.

On the opposite side of things, we can consider the reaction of CH_3^- with water. (The strong bases you have learned about so far all contain OH-, but using one of those bases as an example makes the following explanation a little confusing.) CH_3^- is one of the strongest bases known, and it is at least 10^{25} times stronger than OH^-. In the competition for control of a proton, CH_3^- beats OH^- by a landslide. On the other side of the equation, we see that its conjugate "acid" is methane, CH_4. Again, outside this context, you would never consider methane to be an acid, because almost nothing is strong enough to remove one of its protons. Thus, CH_3^- is an extremely strong base *compared to* OH^-. Notice that the species we are using for comparison changes when we are talking about acids versus bases.

Now let's look at the dissociation of acetic acid. You can see by the equilibrium constant that the reactants should be favored. This means that when the base and the conjugate base (in this case, acetate ions and water molecules) are competing for a proton, acetate wins most of the time. Compared toH_3O^+, acetic acid is a much weaker acid. This is equivalent to the statement that acetic acid will not dissociate much in water, which is how we have previously defined weak acids.

However, acetate ion is not*Compared to water*, acetate is a good base. However, we generally define "strong" and "weak" bases by comparing them to OH−, not H_2O. *Compared to* OH^-, acetate is a weak base.

Thus, we can see why a weak acid can still produce a weak conjugate base, and vice versa. It is important to remember that "acidic" and "basic" are relative terms. It is very useful to have standard benchmarks (H_3O^+ and OH^-) to use for comparison, but figuring out what is "strong" and what is "weak" can be confusing if you don't realize that the reference point we use when talking about acids is different than the one we use when talking about bases.

Lesson Summary

- The Brønsted-Lowry concept of acids and bases states that the acid donates a proton and the base accepts a proton.
- A conjugate acid is a substance that results when a base gains (or accepts) a proton.
- A conjugate base is a substance that results when an acid loses (or donates) a proton.
- Strong acids result in very weak conjugate bases when they lose a proton, and weak acids result in very strong conjugate bases when they lose a proton.
- Strong bases result in very weak conjugate acids when they gain a proton, and weak bases result in very strong conjugate acids when they gain a proton.

Further Reading / Supplemental Links

The learner.org

- http://www.learner.org/vod/vod_window.html?pid=808

Review Questions

1. What improvements did Brønsted-Lowry make over the Arrhenius definition for acids-bases?
2. What is the Brønsted-Lowry definition of an acid?
   a. A substance that donates protons.
   b. A substance that accepts protons.
   c. A substance that dissolves in water to form OH^- ions.
   d. A substance that dissolves in water to form H^+ ions.
3. If H_3O^+ is an acid according to the Brønsted-Lowry theory, what is the conjugate base of this acid?
   a. ${H_4O^{2+}}_{(aq)}$
   b. ${H^+}_{(aq)}$
   c. $H_2O_{(l)}$
   d. $OH^-_{(aq)}$
4. What is the conjugate base of ${H_2PO_4}^-$?
   a. ${H_3O^+}_{(aq)}$
   b. $H_3PO4_{(aq)}$
   c. ${HPO_4^{2-}}_{(aq)}$
   d. ${PO_4^{3-}}_{(aq)}$
5. In the following reactions, which are the Brønsted-Lowry acids? $H_3PO_{4(aq)} + H_2O_{(l)} \rightleftharpoons {H_3O^+}_{(aq)} + {H_2PO_4^-}_{(aq)}$, ${H_2PO_4^-}_{(aq)} + H_2O_{(l)} \rightleftharpoons {H_3O^+}_{(aq)} + {HPO_4^{2-}}_{(aq)}$
   a. $H_2PO_4^-, H_2O, HPO_4^{2-}$
   b. $H_3PO^4, H_2O, H_2PO_4^-$
   c. H_3O^+, H_2O, HPO_4^{2-}
   d. $H_3PO4, H_3O^+, H_2PO_4^-$
6. Label the conjugate acid-base pairs in each reaction.
   a. $HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$
   b. $H_2PO_4^- + H_2O \rightleftharpoons H_3O^+ + HPO_4^{2-}$
   c. $CN^- + H_2O \rightleftharpoons HCN + OH^-$
   d. $HF_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + F^-_{(aq)}$
7. Complete the following reactions. When done, label the conjugate acid/base pairs.
   a. $BrO_3^- + H_2O \rightleftharpoons$
   b. $HNO_3 + H_2O \rightarrow$
   c. $HSO_4^- + C_2O_4^{2-} \rightleftharpoons$
8. For the reactions in question 7, which are the weak conjugate bases and which are the strong conjugate bases?

6.6 Lewis Acids and Bases

Lesson Objectives

The student will:

- define a Lewis acid and a Lewis base.
- identify Lewis acids and bases in equations.

Vocabulary

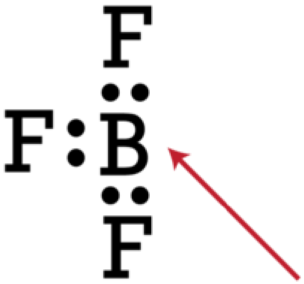
- Lewis acid
- Lewis base

Introduction

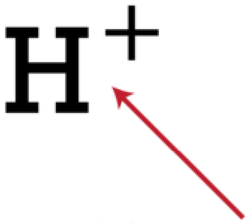
In the early 1930s, Gilbert Lewis saw the need for an even more general definition for acids and bases. Using the work of Brønsted and Lowry, he saw that in some cases, the acids may not have protons to donate but may still be able to donate an atom to the electron pair of a base. Most chemists today use the Brønsted-Lowry definitions of acids-bases, except for specialized situations that require the Lewis definitions.

Lewis Acids

Lewis defined an acid as a substance that accepts a pair of electrons from another substance. Therefore, **Lewis acids** must have room in their structure to accept a pair of electrons. Remember that each central atom can hold eight valence electrons. What this means is that if the atom has six or less valence electrons, then it can accept one more pair. Look at the structures below. Notice how, in each case, there is room to accept a pair of electrons.



The BF_3 molecule can hold 8 electrons surrounding the boron atom, but in this case has only 6 electrons around the boron. Therefore, this molecule can take on two electrons and that makes it a Lewis acid.

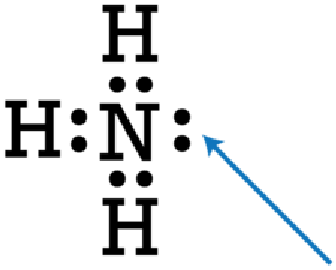


This hydrogen ion can hold 2 electrons around the hydrogen atom, but in this case has zero electrons around the hydrogen. Therefore, this ion can take on two electrons and that makes it a Lewis acid.

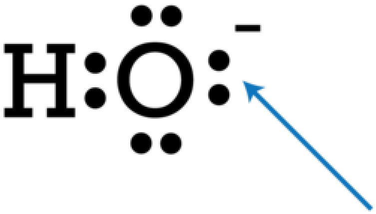
The Lewis acid will accept the electron pair in order to form a bond. The bond that forms between the two atoms will be covalent. Remember that covalent bonds are formed when electrons are shared between two atoms.

Lewis Bases

A **Lewis base** is an electron-pair donor. These electrons are not transferred, but rather shared covalently to complete the acid/base reaction. If we look at the example of bases in the figure below, we can see that each of them has a lone pair of electrons available to donate.



This ammonia molecule has four pairs of electrons around its central atom and only three of the pairs are shared. Therefore, this molecule has another pair of electrons that it can share and that makes it a Lewis base.



This hydroxide ion has four pairs of electrons around its central atom and only one of the pairs is shared. Therefore, this ion has unshared pairs of electrons available to share and that makes it a Lewis base.

This video contains textbook definitions of Arrhenius, Bronsted, and Lewis acids and bases **(5b, 5e)**: http://www.youtube.com/watch?v=jLcoIRUlz20 (3:14).

MEDIA

Click image to the left for more content.

Lesson Summary

- Lewis defined an acid as a substance that accepts a pair of electrons from a substance to form a bond.
- Lewis defined a base as a substance that donates a pair of electrons to a substance to form a bond.

Review Questions

1. How do the Lewis definitions of acids and bases compare to the Brønsted-Lowry definitions of acids and bases?
2. In the following reversible reaction, which of the reactants is acting as a Lewis base: $Cd^{2+}{}_{(aq)} + 4\ I^{-}{}_{(aq)} \rightleftharpoons CdI_4{}^{2-}{}_{(aq)}$?
   a. Cd^{2+}
   b. I^{-}
   c. $CdI_4{}^{2-}$
   d. None of the above, this is not an acid-base reaction.
3. Which of the following statements is false?
   a. NH_3 is a Lewis base.
   b. $B(OH)_3$ is a Lewis acid.
   c. CO_2 is a Lewis base.
   d. Ag^{+} Is a Lewis acid.
4. Which of the following statements is true?
   a. NH_3 is a Lewis base.
   b. $B(OH)_3$ is a Lewis acid.
   c. CO_2 is a Lewis base.
   d. Ag^{+} Is a Lewis acid.
5. Classify each of the following as a Lewis acid or base.
   a. H_2O
   b. BF_3
   c. S^{2-}
   d. Cu^{2+}
   e. O^{2-}
6. Write the balanced chemical equation between $SO_3{}^{2-}$ and H_2O and label the Lewis acids and bases.
7. Identify the Lewis acid and Lewis base in the following reaction: $Cu^{2+}{}_{(aq)} + 6\ H_2O_{(l)} \rightarrow Cu(H_2O)_6{}^{2+}{}_{(aq)}$.

CHAPTER 7 Chemistry Glossary

Chapter Outline

7.1 A

absolute zero
the temperature at which molecules stop moving and therefore have zero kinetic energy

accuracy
reflects how close the measured value is to the actual value

acidic salt
a salt formed in a neutralization reaction between a strong acid and a weak base

actinide series
elements with atomic numbers 89 – 102

activation energy
the minimum amount of energy that needs to be supplied to the system so that a reaction can occur

activated complex
a high energy transitional structure that exists just after the bonds in the reactants are broken and before the products are formed

group of un-bounded particles that exist right after the bonds in the reactant have been broken and before the product bonds have formed

actual yield
the actual amount of product that is formed in the laboratory experiment

alcohol
an organic compound that contains the hydroxyl (-OH) group

aldehyde
an organic compound that contains a terminal carbonyl group

alkali metals
Group 1A of the periodic table

alkaline earth metals
Group 2A of the periodic table

alkane
a hydrocarbon containing only single bonds

alkene
a double bond between two carbon atoms, or a hydrocarbon containing a carbon-carbon double bond

alkyne
a triple bond between two carbon atoms, or a hydrocarbon containing a carbon-carbon triple bond

allotropes
different forms of a pure element caused by different bonding arrangements

alloy
a substance with metallic properties composed of a mixture of two or more elements

alpha decay
a common mode of radioactive decay in which a nucleus emits an alpha particle (a helium-4 nucleus)

alpha particle
a helium-4 nucleus

amino acid
a molecule that contains an amine group $(-NH_2)$ and a carboxyl group $(-COOH)$

amphoteric
a substance that acts as both acids and bases in reactions

amplitude (of a wave)
the "height" of a wave

angular momentum quantum number
a number that describes the sub-shell in which an electron can be found
anion
a negative ion
anode
the electrode at which oxidation occurs
Arrhenius acid
a substance that produces H^+ ions in solution
Arrhenius base
a substance that produces OH^- ions in a solution
artificial radioactivity
induced radioactivity that is produced by bombarding an element with high-velocity particles
atmosphere
unit of pressure equivalent to the pressure under standard conditions at sea level and to 760 mm of Hg
atom
the basic building block of all matter
atomic mass
the weighted average of the masses of the naturally occurring isotopes of an element
atomic mass unit (amu)
used to express atomic and molecular masses, it is the approximate mass of a hydrogen atom, a proton, or a neutron and is one-twelfth of the mass of an unbound carbon-12 atom at rest; it is equivalent to the dalton
atomic number
the number of protons in the nucleus of an atom
atomic radius
one-half the distance between the centers of a homonuclear diatomic molecule
atomos
Democritus's word for the tiny, indivisible, solid objects that he believed made up all matter in the universe
Aufbau principle
states that as electrons are added to "build up" the elements, each electron is placed in the lowest energy orbital available
autoionization
when the same reactant acts as both the acid and the base
Avogadro's law
describes the relationship between the volume and the number of moles of gas
Avogadro's number
the number of objects in a mole; equal to 6.02×10^{23}

7.2 B

background radiation
radiation that comes from environment sources including the earth's crust, the atmosphere, cosmic rays, and radioisotopes

balanced chemical equation
a chemical equation in which the number of each type of atom is equal on the two sides of the equation

barometer
an instrument used to measure atmospheric pressure

base unit
a unit that cannot be expressed in terms of other units, such as the gram (base unit of mass), the meter (base unit of length), and the liter (base unit of volume)

basic salt
a salt formed in a neutralization reaction between a weak acid and a strong base

benzene ring
equivalent resonance structures representing a 6-carbon ring with alternating C-C double bonds

beta decay
a common mode of radioactive decay in which a nucleus emits beta particles

beta particle
a high speed electron, specifically an electron of nuclear origin

binding energy
the amount of energy that holds a nucleus together and also the amount of energy required to decompose a nucleus into its component nucleons

biochemistry
the study of the structure and properties of molecules in living organisms

black body radiation
the energy that would be emitted from an ideal black body.

boiling point elevation
the difference in the boiling points of the pure solvent and the solution

bond energy
the energy required to break a given chemical bond

bond length
the distance between the nuclei of the two atoms connected by a bond

Boyle's law
describes the relationship between the pressure and volume of a gas

Brønsted-Lowry acid
a substance that donates a proton (H+)

Brønsted-Lowry base
a substance that accepts a proton (H+)

buffer
a solution of a weak acid and its conjugate base or a weak base and its conjugate acid that resists changes in pH when an acid or base is added to it

7.3 C

carbohydrate
molecule that contains carbon, hydrogen, and oxygen and has the general formula $C_x(H_2O)_y$

carbonyl group
a carbon atom double bonded to an oxygen atom

catalyst
a substance that speeds up the rate of the reaction without itself being consumed by the reaction

cathode
the electrode at which reduction occurs

cathode ray tube
a small glass tube with a cathode (a negatively charged metal plate) and an anode (a positively charged metal plate) at opposite ends

cation
a positive ion

chain reaction
a multi-stage nuclear reaction that sustains itself in a series of fissions in which the release of neutrons from the splitting of one atom leads to the splitting of others

Charles's law
describes the relationship between the volume and temperature of a gas

chemical change
change that occurs when one substance is turned into another substance

chemical kinetics
the study of rates of chemical reactions and how factors affect rates of reactions

chemical nomenclature
the systematic procedure for naming chemical compounds

chemical potential energy
energy stored in the atoms, molecules, and chemical bonds that make up matter

chemical property
property that can be observed only when a substance is changed into a new substance

chemical reaction
the process in which one or more substances are changed into one or more new substances

chemistry
the scientific study of matter and the changes that it undergoes

chromatography
any of various techniques for the separation of complex mixtures that rely on the differential affinities of substances for a mobile solvent and a stationary medium through which they pass

coefficient
a whole number that appears in front of a formula in a balanced chemical equation

colligative property
a property that is due only to the number of particles in solution and not related to the chemical properties of the solute

collision frequency
the total number of collisions per second

collision theory
explains why reactions occur at this particle level between atoms, ions, and/or molecules and enables prediction about what conditions are necessary for a successful reaction to take place

colloid
type of mixture in which the size of the particles is between 1×10^3 pm and 1×10^6 pm

column chromatography
a method of chromatography that uses a vertical column filled with absorbent material as the medium through which the mixture is separated

combined gas law
combines Boyle's, Charles's, and Gay-Lussac's laws and describes the relationship between the temperature, pressure, and volume of a gas

combustion
a group of chemical reactants in which the reactants are fuel and oxygen gas

combustion reaction
a reaction in which oxygen reacts with another substance to produce carbon dioxide and water

compound
a pure substance that is made up of more than one type of atom

concentrated
a solution in which there is a large amount of solute in a given amount of solvent

concentration
the measure of how much of a given substance is mixed with another substance

condensation
the process whereby a gas or vapor is changed to a liquid

conductor
a substance that can transmit heat and/or electricity

conjugate acid
the substance that results when a base gains (or accepts) a proton

conjugate base
the substance that results when an acid loses (or donates) a proton

control rod
made of chemical elements capable of absorbing many neutrons, it is used to control the rate of a fission chain reaction in a nuclear reactor

controlled experiment
an experiment that compares the results of an experimental sample to a control sample, where the control sample is identical to the experimental sample in all ways except for the one variable being tested

conversion factor
a factor used to convert one unit of measurement into another unit

coordinate covalent bonds
a type of covalent bond when the two shared electrons of a covalent bond are both donated by the same atom

core electrons
the electrons that are not in the outermost energy level of an atom

covalent bond
a type of chemical bond where two atoms are connected to each other by the sharing of two or more electrons in overlapped orbitals

crest
highest point in a wave pattern (peak of a hill)

critical mass
the smallest mass of a fissionable material that will sustain a nuclear chain reaction at a constant level

critical pressure
the pressure required to liquefy a gas at its critical temperature

critical temperature
the highest temperature at which it is possible to liquefy the substance by increasing pressure

cubic meter
the SI unit of volume

7.4 D

dalton

used to express atomic and molecular masses, it is the approximate mass of a hydrogen atom, a proton, or a neutron and is one-twelfth of the mass of an unbound carbon-12 atom at rest; it is equivalent to the atomic mass unit

Dalton's atomic theory

the first scientific theory to relate chemical changes to the structure, properties, and behavior of the atom

Dalton's law of partial pressures

states that for a mixture of gases in a container, the total pressure exerted is the sum of the pressures that each gas would exert if it were alone

decomposition reaction

a reaction in which one reactant breaks down to form two or more products

delocalized electrons

electrons that are free to move between the bond (in multiple bonding)

derived units

units that are defined in terms of other SI base units

diatomic molecule

a molecule containing exactly two atoms

diffraction

the bending of waves around a barrier

diffusion

the movement of particles from areas of higher concentration to areas of lower concentration

dilute

a solution in which there is a small amount of solute in a given amount of solvent

dimensional analysis

a technique that involves the study of the dimensions (units) of physical quantities

dipeptide

two amino acids joined together

distillation

the evaporation and subsequent collection of a liquid by condensation as a means of purification

DNA (deoxyribonucleic acid)

DNA is a polynucleotide that carries our genetic coding; its function is to direct the body in the synthesis of proteins

double bond

a bond in which two pairs of electrons are shared

double replacement reaction

a reaction in which two reactants form products by having the cations exchange places

ductile

can be drawn out into thin wires

dynamic equilibrium

a state of equilibrium where change is still occurring on a molecular level even though the macroscopic properties remain constant; occurs when the rate of the forward reaction is equal to the rate of the reverse reaction

7.5 E

effective collision
a collision that results in a reaction

effective nuclear charge
the net charge experienced by a specific electron within an atom

electric current
flow of electric charge

electrochemical cell
an arrangement of electrodes and ionic solutions in which a spontaneous redox reaction is used to produce a flow of electrons in an external circuit

electrolysis
a chemical reaction brought about by an electric current

electrolyte
a solution that contains ions and is capable of conducting an electric current

electromagnetic spectrum
a list of all the possible types of light in order of decreasing frequency, increasing wavelength, or decreasing energy; spectrum includes gamma rays, X-rays, UV rays, visible light, IR radiation, microwaves, and radio waves

electron
a negatively charged subatomic particle

electron affinity
the energy released when an electron is added to a gaseous atom or ion

electron configuration code
a code that represents the arrangement of electrons of an atom

electron promotion
moving an electron from a lower sub-energy level to a higher sub-energy level within an energy level

electronegativity
the ability of an atom in a molecule to attract shared electrons

electronic geometry
the geometric arrangement of orbitals containing the shared and unshared electron pairs surrounding the central atom of a molecule or polyatomic ion

electroplating
a process in which electrolysis is used as a means of coating an object with a layer of metal

electrostatic attraction
the force of attraction between opposite electric charges

element
a pure substance that is made up of only one type of atom

elementary step
a single, simple step in a multi-step reaction involving one or two particles

emission spectrum
the unique pattern of electromagnetic radiation frequencies obtained when an element is subjected to specific excitation

empirical formula
a formula that gives the simplest whole number ratio of atoms of each element present in the compound

endothermic reaction
reaction that has a potential energy difference between the products and reactants that is positive

endpoint

the point in the titration where the indicator changes color

energy

is the ability to do work or cause change

energy levels

possible orbits an electron can have in the electron cloud of an atom

enthalpy

amount of energy a system or substance contains

enthalpy of formation

the heat required to form one mole of a substance from its elements at standard temperature and pressure

entropy

a measure of the disorder of a system

enzyme

a subset of proteins that function to speed up a chemical reaction

equilibrium

state where the macroscopic properties of a system do not change over time

equilibrium constant (K)

a mathematical ratio that shows the concentrations of the products divided by concentration of the reactants

equilibrium vapor pressure

the pressure that is exerted, at a given temperature, by the vapor of a solid or liquid in equilibrium with the vapor

equivalence point

the point in the titration where the number of moles of acid equals the number of moles of base

ester

an organic compound with the general formula RCO_2R', where R and R' are alkyl groups

ether

an organic compound that contains an oxygen atom bonded to two alkyl groups

evaporation

the escape of molecules from a liquid into the gaseous state at a temperature below the boiling point

excess reactant

any reactant present in an amount that is more than enough to react with the limiting reactant

exothermic reactions

reaction that has a potential energy difference between the products and reactants that is negative

experiment

a controlled method of testing a hypothesis

extrapolation

the process of creating data points beyond the end of a line graph, using the basic shape of the curve as a guide

7.6 F

fatty acid

a carboxylic acid having anywhere from four to thirty-six carbon atoms

fissile

a substance capable of sustaining a chain reaction of nuclear fission

fission

a nuclear reaction in which a heavy nucleus splits into two or more smaller fragments, releasing large amounts of energy

fissionable

a material capable of undergoing fission

formula mass

the sum of the atomic masses of the atoms in a formula

formula unit

one unit of the empirical formula for an ionic compound

fractional distillation

a special type of distillation used to separate a mixture of liquids using their differences in boiling points

freezing

the phase change from liquid to solid

freezing point

the temperature at which a liquid changes to a solid

freezing point depression

the difference in the freezing points of the solution from the pure solvent

frequency (f)

for a wave, the frequency refers the number of waves passing a specific reference point per unit time; in this text, frequency is symbolized by $f\lambda$

functional group

an atom or a group of atoms that replaces hydrogen in an organic compound and is responsible for the characteristic properties of the compound

fusion

depending on the context, fusion could mean the change of a liquid to a solid or a nuclear reaction in which two or more smaller nuclei combine to form a single nucleus

7.7 G

gamma ray
a very high energy form of electromagnetic radiation, emitted from the nucleus

gas chromatography
method, a mixture of liquids are vaporized and passed through a long tube of solid absorbent material

Gay-Lussac's law
describes the relationship between temperature and pressure of a gas

Geiger counter
an instrument used to detect radiation, usually alpha and beta radiation, although some models can also detect gamma radiation

Gibbs free energy
the maximum energy available to do useful work

Graham's law
states that under the same conditions of temperature and pressure, gases diffuse at a rate inversely proportional to the square root of the molecular masses

graph
a pictorial representation of the relationship between variables on a coordinate system

gravimetric analysis
the process for separating and identifying ions by selective precipitation and filtration

group (family)
a vertical column in the periodic table

7.8 H

half-life
the time interval required for a quantity of material to decay to half its original value

halogens
Group 7A of the periodic table

heat
the flow of thermal energy from a warmer object to a cooler object

heat content
the total of all forms of energy for a substance

heat of condensation
the quantity of heat released when a unit mass of a vapor condenses to liquid at constant temperature

heat of fusion
the quantity of heat released when a unit mass of liquid freezes to a solid at a constant temperature

heat of vaporization
the quantity of heat required to vaporize a unit mass of liquid at constant temperature

Heisenberg uncertainty principle
it is impossible to know both the precise location and the precise velocity of an electron at the same time

Henry's law
states that at a given temperature, the solubility of a gas in a liquid is proportional to the partial pressure of the gas above the liquid

hertz (Hz)
the SI unit used to measure frequency; one hertz is equivalent to 1 cycle per second

Hess's Law
if multiple reactions are combined, the enthalpy $(\triangle H)$ of the combined reaction is equal to the sum of all the individual enthalpies

heterogeneous mixture
a mixture that consists of visibly different substances

homogeneous mixture
a mixture that is uniform throughout

Hund's rule
a rule that states that no electrons are paired in a given orbital until all the orbitals of the same sub-level have received at least one electron

hydration
the process of water molecules attaching to ions

hydrocarbon
an organic substance consisting of only hydrogen and carbon

hydrogen bond
a type of strong polar attraction between a hydrogen atom in one molecule and a highly electronegative atom, such as N, O, or F, in another molecule

hydronium ion
a positively charged ion consisting of three hydrogen atoms and one oxygen atom (H_3O^+)

hydroxyl group
an oxygen atom bound covalently with a hydrogen atom

hypothesis
a tentative explanation that can be tested by further investigation

7.9 I

ideal gas

a gas that follows the universal gas law and satisfies the two assumptions that molecules are points that do not take up any space and that there are no attractions between molecules; this is generally true for gases at low pressure and high temperature

immiscible

liquids that do not have the ability to dissolve in each other

incompressible

describes the inability of molecules in a substance to be compacted (pushed closer together)

indicator

substance used to determine the relative acidity or basicity of a solution, generally through a very distinct color change

ionizing power

ability of radiation to damage molecules

insoluble

unable to dissolve

instantaneous rate

the rate of change at a particular time interval

internal energy

total energy contained in a thermodynamic system; often called enthalpy

interference

the addition of two or more waves that result in a new wave pattern

International System of Units (Le Système International d' Unites)

the internationally agreed upon standard metric system, also abbreviated as the SI system (derived from the French name)

interpolation

the process of estimating values between measured values

ion

an atom with a positive or negative charge

ion product constant for water, Kw

the product of the hydronium ion and the hydroxide ion concentrations in the autoionization of water

ionic bond

the electrostatic attraction between ions resulting from the transfer of electrons from one of the bonding atoms to the other

ionic conduction

movement of ions through solution

ionic equation

chemical equation written such that the actual free ions are shown for each species in aqueous form

ionization energy

the energy required to remove the most loosely held electron from a gaseous atom or ion

irreversible reaction

a reaction that continues to form products until reactants are fully consumed

isotopes

atoms of the same element that have the same number of protons but different numbers of neutrons

7.10 J

joule

the SI unit for energy, equivalent to 1 $\frac{\text{kg}\cdot\text{m}^2}{\text{s}^2}$

7.11 K

K$_a$
the equilibrium constant for the dissociation of a weak acid

K$_b$
the equilibrium constant for the dissociation of a weak base

Kelvin temperature scale
a temperature scale which has its zero at absolute zero

ketone
an organic compound that has a carbonyl group linked to two alkyl groups

kilogram
the SI unit of mass

kinetic energy
energy associated with motion

kinetic molecular theory
the foundation for the collision theory on the atomic level, the theory considers the collisions between particles to be elastic in nature

7.12 L

lanthanide series
elements with atomic numbers 57 – 70

lattice structure
a systematic, symmetrical network of atoms forming an ionic solid.

law of conservation of energy
states that energy cannot be created or destroyed, it can only be changed from one form to another

law of conservation of matter and energy
states that the total amount of mass and energy in the universe is conserved (does not change)

law of constant composition
law that states that the ratio by mass of the elements in a chemical compound is always the same, regardless of the source of the compound

law of definite proportions
states that in a given chemical substance, the elements are always combined in the same proportions by mass

law of multiple proportions
states that when two elements react to form more than one substance and the same amount of one element is used in each substance, then the ratio of the masses used of the other element will be in small whole numbers

Le Châtelier's Principle
applying a stress to a system at equilibrium causes a shift in the equilibrium position that partially counteracts the stress

length
the measurement of anything from end to end

Lewis acid
a substance that accepts a pair of electrons from another substance

Lewis base
a substance that donates a pair of electrons to a substance

Lewis dot diagram
a diagram that shows valence electrons as dots around the symbol of an element

limiting reactant
the reactant that determines the maximum amount of product that can be formed in a chemical reaction

lipids
fats and oils (triglycerides) produced for the purpose of storing energy

London dispersion forces
the electrostatic attractions of molecule or atoms for nearby molecules or atoms caused by the temporary unsymmetrical distribution of electrons in electron clouds

7.13 M

macroscopic properties
properties that describe the system as a whole and can be observed and measured without determining the properties of individual molecules
magnetic quantum number
a number that describes the orientation in space of a particular orbital
malleable
property of being able to be hammered or rolled into sheets
manometer
a liquid column pressure-measuring device
mass
a measure of the amount of matter in an object
mass defect
the difference between the sum of the masses of the nuclear components and the mass of the corresponding nucleus
mass number
the total number of protons and neutrons in the nucleus of an atom
mass percent
the number of grams of the solute in the number of grams of solution
matter
anything that has mass and volume
melting
the phase change from solid to liquid
melting point
the temperature at which a substance changes from the solid phase to the liquid phase
metallic bond
the attractive force that binds metal atoms together
metallic conduction
movement of electrons through a piece of metal
meter
the SI unit of length
metric system
an international decimal-based system of measurement
miscible
liquids that have the ability to dissolve in each other
model
a descriptive, graphic, or three-dimensional representation of a hypothesis or theory used to help enhance understanding
moderator
a medium that reduces the velocity of fast neutrons
molality
the number of moles of solute per kilograms of solvent
molar mass
the mass in grams of one mole of a substance
molar volume
the volume occupied by one mole of a substance in the form of a solid, liquid, or gas
molarity

the number of moles of solute per liter of solution

mole

an Avogadro's number of objects; 1.00 mole of carbon-12 atoms has a mass of 12.0 grams and contains $6.02 \times 10^{2}3$ atoms

mole ratio

a quantitative relationship between two components of a chemical reaction, determined by the coefficients in the balanced equation

molecular formula

a formula indicating the actual number of each kind of atom contained in a molecule

molecular geometry

the specific three-dimensional arrangement of atoms in a molecule

molecular mass

the mass of a molecule found by adding the atomic masses of the atoms comprising the molecule

molecule

the smallest particle of a compound

multi-step reaction

a reaction that occurs in more than one step

7.14 N

natural indicator
an indicator that is a naturally occurring substance

natural radioactivity
the radioactivity that occurs naturally, as opposed to induced radioactivity; also known as spontaneous fission

net ionic equation
the overall equation that results when spectator ions are removed from the ionic equation

nuclear symbol
contains the symbol for the element and the numbers that relate to the number of protons and neutrons in that particular nucleus

nucleon
a collective name for neutrons and protons

neutral salt
a salt formed in a neutralization reaction between a strong acid and a strong base or a weak acid and a weak base

neutralization
a reaction between an acid and a base that produces water and a salt

neutron
a subatomic particle with no charge

noble gases
Group 8A of the periodic table

non-electrolyte
a solution that does not contain ions and is not capable of conducting an electric current

non-spontaneous event (or reaction)
an unfavorable reaction that requires an external energy source in order for the reaction to take place

nuclear charge
the number of protons in the nucleus

nucleus
the center of the atom

7.15 O

observation

using the senses to gather information about the natural world

octet rule

an expression of the tendency for atoms to gain or lose the appropriate number of electrons so that the resulting ion has either completely filled or completely empty outer energy levels

orbital hybridization

orbitals involved in covalent bonding in an atom are hybridized so that they are identical in properties, and those properties are intermediate to the properties of the original orbitals

orbital representation

a method that uses circles or lines to represent the orbitals where electrons in an atom are located

organic acid

an organic compound containing one or more carboxyl groups

oxidation

a loss of electrons in an atom or an increase in the oxidation state of an atom

oxidation number

in ionic compounds, it is equal to the ionic charge; in covalent compounds, it is the charge assigned to the atom in accordance with a set of rules

oxidizing agent

a substance that gains electrons in a chemical reaction or undergoes an increase in its oxidation state

7.16 P

paper chromatography
a method of chromatography that uses paper as the medium upon which the mixture to be separated is applied

partial pressure
the pressure that one component of a mixture of gases would exert if it were alone in a container

parts per million (ppm)
the mass of solute per mass of solution multiplied by 1 million

pascal
the SI unit of pressure, equivalent to 1 Newtons/meter^2

Pauli exclusion principle
no two subatomic particles can be in states characterized by the same set of quantum numbers

penetration power
ability of radiation to pass through matter

percent composition
the proportion of an element present in a compound, found by dividing the mass of the element by the mass of the whole compound and multiplying by 100%

percent error
a measurement of how far an experimental value is from the accepted value

percent yield
the ratio of the actual yield to the theoretical yield expressed as a percentage

period
a horizontal row in the periodic table

periodic law
states that the properties of the elements recur periodically as their atomic numbers increase

periodic table
a tabular arrangement of the chemical elements according to atomic number

pH
the negative logarithm of the hydrogen ion concentration

phase
any of the forms or states (solid, liquid, gas, or plasma) in which matter can exist

phlogiston
the “fire substance” from a former theory of combustion

phospholipid
a combination of fatty acids, glycerol and a phosphate group joined together

photoelectric effect
a phenomenon in which electrons are emitted from the surface of a material after the absorption of energy

photon
a particle of light

physical change
change that does not alter the identity of a substance

physical property
property that can be observed without changing the identity of the substance

pi bond
a covalent bond in which **p** orbitals share an electron pair occupying the space above and below the line joining the atoms

pOH

the negative logarithm of the hydroxide ion concentration

polar covalent bonds

covalent bonds in which the shared bonding electrons are not shared equally, resulting in a dipole on the molecule

polyatomic ion

a group of covalently bound atoms that carry an overall charge, which interacts with other molecules as a single unit

polymer

a large organic molecule that contains hundreds or even thousands of atoms

polypeptide

many amino acids combined together

potential energy

stored energy

potential energy diagram

potential energy diagram in the study of kinetics shows how the potential energy changes during reactions from reactants to products

precipitate

the solid that is formed as a result of a precipitation reaction

precipitation reaction

a reaction in which soluble ions in separate solutions are mixed together to form an insoluble compound that settles out of the solution

precision

reflects how close the values in a set of measurements are to each other

principal quantum number

a number that indicates the main energy level of an electron in an atom

problem

the purpose for a scientific investigation

products

materials present at the end of a reaction

protein

polymer that is an amino acid

proton

a positively charged subatomic particle

7.17 Q

qualitative observation
observation that yields descriptive, nonnumerical results

quantitative observation
observation that yields meaningful, numerical results

quantum (singular form of quanta)
small unit of energy

quantum mechanics
the branch of physics that deals with the behavior of matter at the atomic and subatomic level

quantum numbers
four special numbers that completely describe the state of an electron in an atom

quantum theory
the theory that energy can only exist in discrete amounts (quanta)

quark
physical particles that form one of the two basic constituents of matter; various species of quarks combine in specific ways to form protons and neutrons, in each case taking exactly three quarks to make the composite particle

7.18 R

rate-determining step

the slowest step in a multi-step reaction

rate of reaction

the measure at which the products are formed over a time interval or the rate at which the reactants are consumed over a time interval

reactants

the starting materials in a reaction

reaction mechanism

the series of elementary reactions describing what occurs in a multi-step reaction

real gas

a gas that does not follow the ideal gas laws, which is generally the case for gases at high pressures or low temperatures

redox reaction

short for oxidation-reduction reaction, a chemical reaction that involves electrons being transferred from one substance to another

reducing agent

the substance in a redox reaction that loses electrons or increases its oxidation state

reduction

the gain of electrons or decrease in oxidation state in a chemical reaction

resonance

a condition occurring when more than one valid Lewis structure can be written for a particular molecule; the actual electronic structure is not represented by any one of the Lewis structures but by the average of all of them

retention factor

the ratio of the distance a substance moves up the stationary phase to the distance the solvent have moved

reversible reaction

a reaction that can also proceed in the reverse direction

7.19 S

salt bridge
a U-shaped tube containing an electrolyte that connects two half-cells in an electrochemical cell
saturated compound
an organic molecule containing only single bonds
saturated solution
a solution in which a given amount of solvent has dissolved the absolute maximum amount solute at that temperature
scientific law
a statement that summarizes the results of many observations and experiments
scientific method
a method of investigation involving experimentation and observation to acquire new knowledge, solve problems, and answer questions
scientific notation
a shorthand method of writing very large and very small numbers by expressing them as a product of a decimal number between 1 and 10 multiplied by an integral power of 10
second
the SI unit for time
shielding effect
the effect where the inner electrons help "shield" the outer electrons and the nucleus from each other
sigma bond
a covalent bond in which the electron pair is shared in an area centered on a line running between the atoms
significant figure
includes all of the digits that can be known with certainty in a measurement plus an estimated last digit
single replacement reaction
a reaction in which an element reacts with a compound to form products
slope
the ratio of the change in one variable with respect to the other variable
solubility
the maximum amount of a substance that will dissolve in a given amount of solvent at a specific temperature
solubility product constant, K_{sp}
equilibrium constant for a slightly soluble salt
soluble
able to dissolve
solute
the substance in a solution present in the least amount
solution
a homogeneous mixture of substances
solvent
the substance in a solution present in the greatest amount
specific heat
the amount of energy necessary to raise 1.00 gram of a substance by 1.00°C
spectator ions
the ions in the total ionic equation that appear in the same form on both sides of the equation indicating they do not participate in the overall reaction
spin quantum number

a number that indicates the orientation of the angular momentum of an electron in an atom

spontaneous event (or reaction)

a change that occurs without outside inference

standard solution

a solution whose concentration is known exactly and is used to find the exact concentration of the titrant

steroid

compound where four carbon rings are bounded together with branches and functional groups bounded to the rings

stoichiometry

the calculation of quantitative relationships of the reactants and products in a balanced chemical equation

strong acid

acid that undergoes 100% dissociation in water

strong base

base that undergoes 100% dissociation in water

strong nuclear force

the strong interaction that is responsible for binding the protons and neutrons in the atomic nuclei together

structural isomers

molecules with the same formula but different structures

subatomic particles

particles that are smaller than the atom

subscript

part of the chemical formula that indicates the number of atoms of the preceding element

surroundings

everything but the reactants and products in the reaction

suspension

type of mixture in which the particles settle to the bottom of the container and can be separated by filtration

synthetic indicator

an indicator that is synthesized in the laboratory

synthesis reaction

a reaction in which two or more reactants combine to make one product

system

the reactants and products in the reaction

7.20 T

temperature
the average kinetic energy of the particles that make up a material

theoretical yield
the amount of product that could be formed from a chemical reaction based on the balanced chemical equation

theory
an explanation that summarizes a hypothesis or a set of hypotheses and has been supported with repeated testing

thin-layer chromatography
a method of chromatography that uses silica gel or a similar inert material on a glass microscope slide or plastic sheet to which the mixture is applied

threshold energy
the minimum amount of energy necessary for a reaction to take place

titrant
the solution of known concentration used in a titration

titration
the process in which a known concentration of base (or acid) is added to a solution of acid (or base) of unknown concentration

titration curve
a graph of the pH versus the volume of titrant added

torr
unit of pressure, equivalent to 1 mm of Hg

transition elements
Groups 3 – 12 of the periodic table

triple bond
a bond in which three pairs of electrons are shared

trough
lowest point in a wave pattern (low point of a valley)

Tyndall effect
the scattering of light by particles

7.21 U

universal gas law

relates the volume of a gas with pressure, temperature, and number of moles of gas; written mathematically as $PV = nRT$

universal gas law constant (R)

constant equal to $\frac{PV}{nT}$, where the pressure, volume, moles, and temperature of the gas are represented by P, V, n, and T, respectively; the value and units of R depend on the units of P and V, and two commonly used values of R are 0.08206 L atm K^{-1} mol^{-1} and 8.314 J K^{-1} mol^{-1}

unsaturated compound

an organic molecule containing double or triple bonds

unsaturated solution

a solution in which a given amount of solvent has dissolved less than the absolute maximum amount solute at that temperature

unshared electron pair

an unshared electron pair, also known as a non-bonding pair of electrons or as a lone pair of electrons, is found in a filled valence orbital that is not used in the formation of a covalent bond

7.22 V

valence electrons
the electrons in the outermost principal quantum level of an atom
valence shell
the outermost energy level of an atom
van’t Hoff factor
the number of particles that the solute will dissociate into upon mixing with the solvent
vapor
the gaseous phase of a substance that exists even though the temperature is below the boiling point of the substance
vapor pressure
pressure of the vapor in the space above the liquid
velocity
distance traveled in one second.
voltage
the potential difference between two points in an electric circuit
volume
the amount of space an object occupies
VSEPR (valence shell electron pair repulsion) theory
a theory whose main postulate is that the structure around a given atom in a molecule is determined by minimizing electron-pair repulsion

7.23 W

wavelength (λ)	the length of a single wave from peak to peak (crest to crest or trough to trough)
wave-particle duality	the concept that all matter and energy exhibit both wave-like and particle-like properties
weak acid	acid that does not completely dissociate in water
weak base	base that does not completely dissociate in water
weight	the force of attraction between the object and the earth (or whatever large, gravity-producing body the object is located on)
work	force (any push or pull) applied over a distance

7.24 X

7.25 Y

7.26 Z