# 8: Chemical Interactions

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Printed: April 8, 2016

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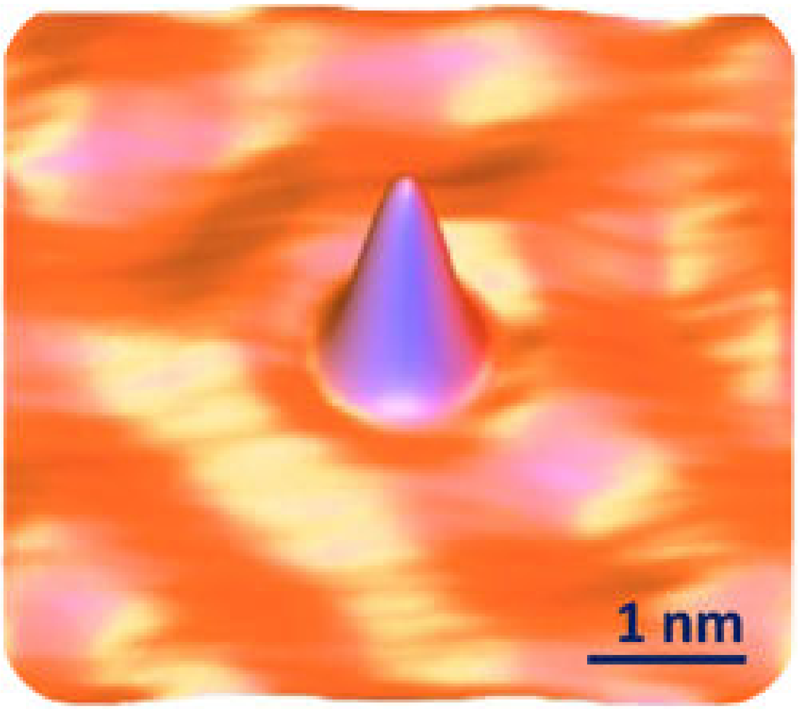
# CHAPTER 1 Atoms

## Chapter Outline






What could this hilly blue surface possibly be? Do you have any idea? The answer is a single atom of the element Cobalt. The picture was created using a scanning tunneling microscope. No other microscope can make images of things as small as atoms. How small are atoms? You will find out in this lesson.

# 1.1 Inside the Atom

## Lesson Objectives

- Compare and contrast protons, neutrons, and electrons.
- Describe the forces that hold the particles of atoms together.
- Define atomic number and mass number.
- Identify the particles called quarks.

## Vocabulary

- atomic mass unit (amu)
- atomic number
- electron
- mass number
- neutron
- nucleus
- proton
- quark

## Introduction

Atoms are the smallest particles of an element that still have the element's properties. They are the building blocks of all matter. Individual atoms are extremely small. In fact, they are so small that trillions of them would fit inside the period at the end of this sentence. Yet atoms, in turn, consist of even smaller particles.

## Parts of the Atom

**Figure** 1.1 represents a simple model of an atom. You will learn about more complex models in later lessons, but this model is a good place to start.

### The Nucleus

At the center of an atom is the **nucleus** (plural, nuclei). The nucleus contains most of the atom's mass. However, in size, it's just a tiny part of the atom. The model in **Figure** 1.1 is not to scale. If an atom were the size of a football stadium, the nucleus would be only about the size of a pea.

The nucleus, in turn, consists of two types of particles, called protons and neutrons. These particles are tightly packed inside the nucleus. Constantly moving about the nucleus are other particles called electrons. You can see a video about all three types of atomic particles at this URL: http://www.youtube.com/watch?v=lP57gEWcisY (1:57).

### Protons

A **proton** is a particle in the nucleus of an atom that has a positive electric charge. All protons are identical. It is the number of protons that gives atoms of different elements their unique properties. Atoms of each type of element

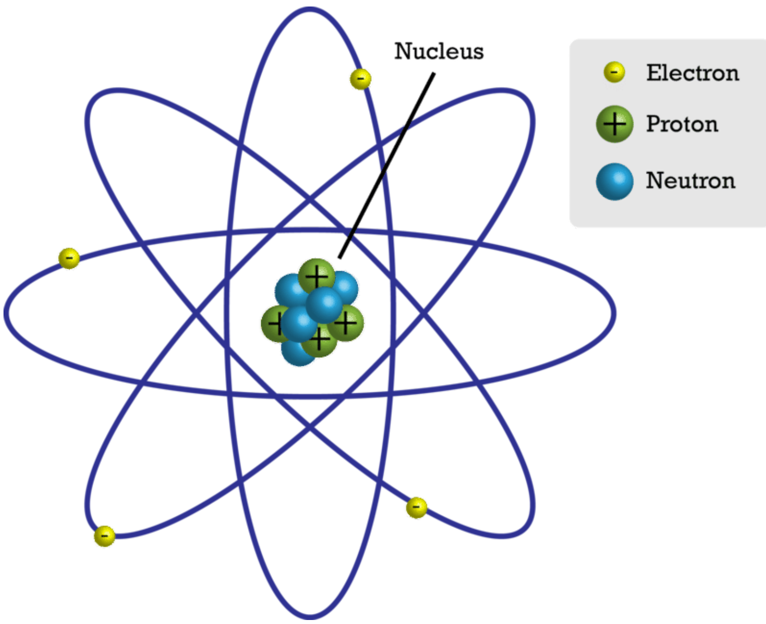



FIGURE 1.1
This simple atomic model shows the particles inside the atom.

have a characteristic number of protons. For example, each atom of carbon has six protons, as you can see in **Figure** 1.2. No two elements have atoms with the same number of protons.

### Neutrons

A **neutron** is a particle in the nucleus of an atom that has no electric charge. Atoms of an element often have the same number of neutrons as protons. For example, most carbon atoms have six neutrons as well as six protons. This is also shown in **Figure** 1.2.

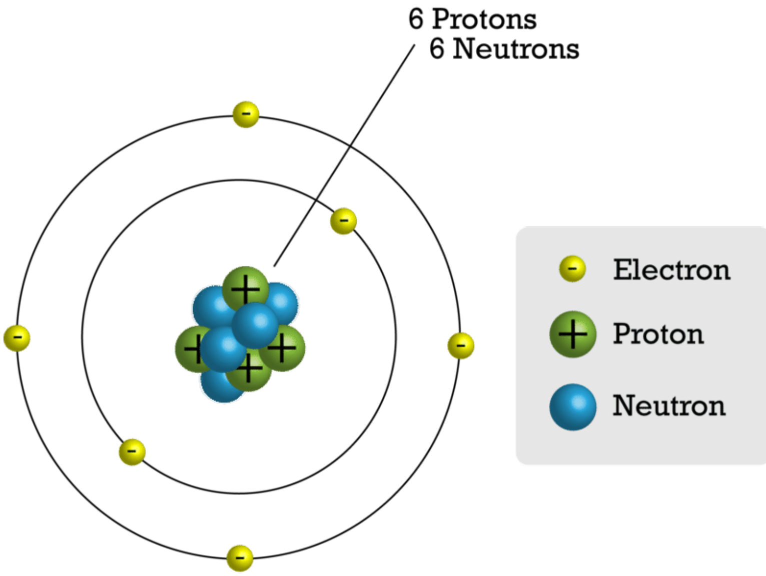



FIGURE 1.2
This model shows the particles that make up a carbon atom.

### Electrons

An **electron** is a particle outside the nucleus of an atom that has a negative electric charge. The charge of an electron is opposite but equal to the charge of a proton. Atoms have the same number of electrons as protons. As a result, the negative and positive charges "cancel out." This makes atoms electrically neutral. For example, a carbon atom has six electrons that "cancel out" its six protons.

## Atomic Number and Mass Number

Electrons have almost no mass. Instead, almost all the mass of an atom is in its protons and neutrons in the nucleus. The nucleus is very small, but it is densely packed with matter. The SI unit for the mass of an atom is the **atomic mass unit (amu)**. One atomic mass unit equals the mass of a proton, which is about $1.7 \times 10^{-24}$ g. Each neutron also has a mass of 1 amu. Therefore, the sum of the protons and neutrons in an atom is about equal to the atom's total mass in atomic mass units.

Two numbers are commonly used to distinguish atoms: atomic number and mass number. **Figure** 1.3 shows how these numbers are usually written.


FIGURE 1.3
The symbol He stands for the element helium. Can you infer how many electrons a helium atom has?

- The **atomic number** is the number of protons in an atom. This number is unique for atoms of each kind of element. For example, the atomic number of all helium atoms is 2.
- The **mass number** is the number of protons plus the number of neutrons in an atom. For example, most atoms of helium have 2 neutrons, so their mass number is 2 + 2 = 4. This mass number means that an atom of helium has a mass of about 4 amu.

**Problem Solving**

*Problem*: An atom has an atomic number of 12 and a mass number of 24. How many protons and neutrons does the atom have?

*Solution*: The number of protons is the same as the atomic number, or 12. The number of neutrons is equal to the mass number minus the atomic number, or 24 - 12 = 12.

**You Try It!**

*Problem*: An atom has an atomic number of 8 and a mass number of 16. How many neutrons does it have? What is the atom's mass in atomic mass units?

### Lesson Summary

- The nucleus is at the center of the atom. It contains positive protons and neutral neutrons. Negative electrons constantly move about the nucleus.

- Atomic number is the number of protons in an atom. It is unique for the atoms of each element. Mass number is the number of protons plus neutrons in an atom. It is about equal to the mass of the atom in atomic mass units (amu)

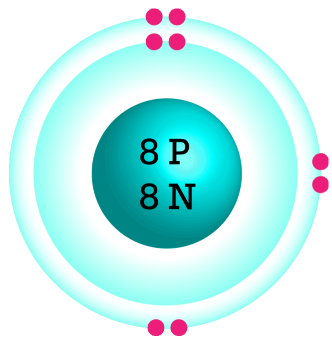
Lesson Review Questions

### Recall

1. Describe the nucleus of an atom.
2. What does the atomic number of an atom represent?

### Apply Concepts

6. What is the atomic mass (in atomic mass units) of the atom represented by the model below?




### Think Critically

8. Make a table comparing and contrasting protons, neutrons, and electrons. Include their location, mass, and electric charge.
9. Explain why atoms are neutral in electric charge.

## Points to Consider

In this lesson, you saw several simple models of atoms. Models are useful for representing things that are very small. Scientists have used models to represent atoms for more than 200 years. In the next lesson, "History of the Atom," you'll read about some of the earlier models.

- How might scientists have modeled atoms before the particles inside atoms were discovered?
- How do you think earlier models might have differed from the models in this lesson?

# 1.2 Modern Atomic Theory

### Lesson Objectives

- Define energy levels.
- Describe the electron cloud.

### Vocabulary

- electron cloud
- energy level
- Valence electrons

### Introduction

Rutherford's model of the atom was better than earlier models. But it wasn't the last word. Danish physicist Niels Bohr created a more accurate and useful model. Bohr's model was an important step in the development of modern atomic theory. The video at the URL below is a good introduction to modern atomic theory. It also reviews important concepts from the previous two lessons, "Inside the Atom" and "History of the Atom."

http://www.khanacademy.org/video/introduction-to-the-atom?playlist=Chemistry

### Bohr's Model of the Atom

Bohr's research focused on electrons. In 1913, he discovered evidence that the orbits of electrons are located at fixed distances from the nucleus. Remember, Rutherford thought that electrons orbit the nucleus at random. **Figure** 1.4 shows Bohr's model of the atom.

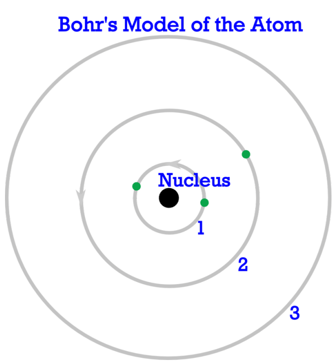



FIGURE 1.4
In Bohr's atomic model, electrons orbit at fixed distances from the nucleus. These distances are called energy levels.

### What Are Energy Levels?

**Energy levels** (also called electron shells) are fixed distances from the nucleus of an atom where electrons may be found. Electrons are tiny, negatively charged particles in an atom that move around the positive nucleus at the center.

Energy levels are a little like the steps of a staircase. You can stand on one step or another but not in between the steps. The same goes for electrons. They can occupy one energy level or another but not the space between energy levels.

The model in the **Figure** below shows the first four energy levels of an atom.Electrons in energy level I have the least amount of energy. As you go farther from the nucleus, electrons at higher levels have more energy, and their energy increases by a fixed, discrete amount. Electrons can jump from a lower to the next higher energy level if they absorb this amount of energy. Conversely, if electrons jump from a higher to a lower energy level, they give off energy, often in the form of light. This explains the fireworks pictured above. When the fireworks explode, electrons gain energy and jump to higher energy levels. When they jump back to their original energy levels, they release the energy as light. Different atoms have different arrangements of electrons, so they give off light of different colors. You can see an animation of electrons jumping from one energy level to another at this URL:http://cas.sdss.org/dr6/en/proj/advanced/spectraltypes/energylevels.asp .

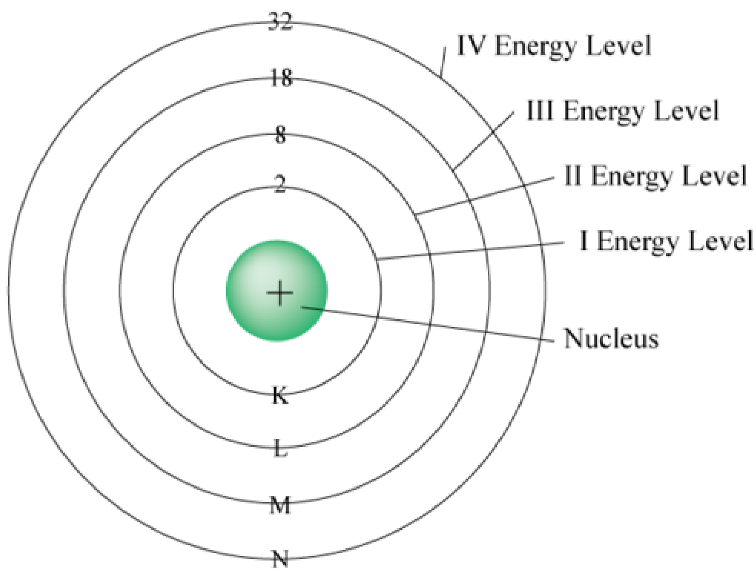



FIGURE 1.5
Energy levels in an atom

The smallest atoms are hydrogen atoms. They have just one electron orbiting the nucleus. That one electron is in the first energy level. Bigger atoms have more electrons. Electrons are always added to the lowest energy level first until it has the maximum number of electrons possible. Then electrons are added to the next higher energy level until that level is full, and so on.

How many electrons can a given energy level hold? The maximum numbers of electrons possible for the first four energy levels are shown in the **Figure** above . For example, energy level I can hold a maximum of two electrons, and energy level II can hold a maximum of eight electrons.

Bohr's idea of energy levels is still useful today. It helps explain how matter behaves. For example, when chemicals in fireworks explode, their atoms absorb energy. Some of their electrons jump to a higher energy level. When the electrons move back to their original energy level, they give off the energy as light. Different chemicals have different arrangements of electrons, so they give off light of different colors. This explains the blue- and purple-colored fireworks in **Figure** 1.6.

FIGURE 1.6
Atoms in fireworks give off light when their electrons jump back to a lower energy level.

### The Outermost Level

Electrons in the outermost energy level of an atom have a special significance. These electrons are called **valence electrons**, and they determine many of the properties of an atom. An atom is most stable if its outermost energy level contains as many electrons as it can hold. For example, helium has two electrons, both in the first energy level. This energy level can hold only two electrons, so helium's only energy level is full. This makes helium a very stable element. In other words, its atoms are unlikely to react with other atoms.

### Lesson Summary

- Bohr introduced the idea that electrons orbit the nucleus only at fixed distances, called energy levels. Electrons in energy levels farther from the nucleus have more energy.
- Today, electrons are represented by an electron cloud model. Orbitals in the cloud show where electrons are most likely to be.

### Lesson Review Questions

#### Recall

1. What are energy levels?
2. Which energy level has the smallest amount of energy?
3. Define valance electrons

### Apply Concepts

5. A change in energy caused an electron in an atom to jump from energy level 2 to energy level 3. Did the atom gain or lose energy? Explain.
6. Create a sketch to model the concept of the electron cloud.

### Think Critically

7. Explain how orbitals are related to energy levels.

## Points to Consider

In this chapter, you learned that atoms of each element have a unique number of protons. This is one way that each element differs from all other elements.

- How could the number of protons be used to organize elements?
- If one element has more protons than another element, how do their numbers of electrons compare?

# 1.3 References

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# CHAPTER 2 Periodic Table

## Chapter Outline



Imagine going to the library and finding all the books in big messy piles like the one above. It could take a very long time to find the book you wanted. You might give up without even trying. Of course, in most libraries, books are arranged in an orderly way, like the books shown below. For example, novels, like those pictured here, are arranged in alphabetical order by author's last name. Not only can you quickly find the book you want, you can also scan the books nearby to find others by the same author. It's clear that grouping books in an organized way is very useful.


The same is true of chemical elements. For many years, scientists looked for a good way to organize them. This became increasingly important as more and more elements were discovered. In this chapter, you'll read how elements were first organized and how they are organized today. You'll see why an orderly arrangement of elements, like the books in a library, is also very useful.

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# 2.1 How Elements Are Organized

## Lesson Objectives

- Describe Mendeleev's periodic table of the elements.
- Give an overview of the modern periodic table of the elements.

## Vocabulary

- group
- period
- periodic table

## Introduction

Scientists first started looking for a way to organize the elements in the 1700s. They were trying to find a method to group together elements with similar properties. No one could come up with a good solution. It wasn't until the 1860s that a successful method was devised. It was developed by a Russian chemist named Dmitri Mendeleev. He is pictured in **Figure** 2.1. You can learn more about him and his work at this URL: http://videos.howstuffworks.com/science-channel/27862-100-greatest-discoveries-the-periodic-table-video.htm .

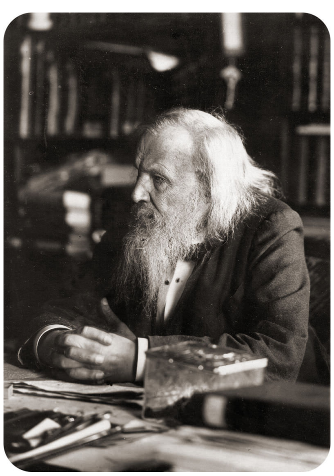


FIGURE 2.1

Dmitri Mendeleev developed the first periodic table of the elements in 1869.

## Mendeleev's Periodic Table of the Elements

Mendeleev was a teacher as well as a chemist. He was writing a chemistry textbook and needed a way to organize the elements so it would be easier for students to learn about them. He made a set of cards of the elements, similar to a deck of playing cards, with one element per card. On the card, he wrote the element's name, atomic mass, and known properties. He arranged and rearranged the cards in many different ways, looking for a pattern. He finally found it when he placed the elements in order by atomic mass.

### A Repeating Pattern

You can see how Mendeleev organized the elements in **Figure** 2.2. From left to right across each row, elements are arranged by increasing atomic mass. Mendeleev discovered that if he placed eight elements in each row and then continued on to the next row, the columns of the table would contain elements with similar properties. He called the columns **groups**. They are sometimes called families, because elements within a group are similar but not identical to one another, like people in a family.

| Reihen | Gruppo I. — $R^2O$ | Gruppo II. — $RO$ | Gruppo III. — $R^2O^3$ | Gruppo IV. $RH^4$ $RO^2$ | Gruppo V. $RH^3$ $R^2O^5$ | Gruppo VI. $RH^2$ $RO^3$ | Gruppo VII. $RH$ $R^2O^7$ | Gruppo VIII. — $RO^4$ |
|---|---|---|---|---|---|---|---|---|
| 1 | H=1 | | | | | | | |
| 2 | Li=7 | Be=9,4 | B=11 | C=12 | N=14 | O=16 | F=19 | |
| 3 | Na=23 | Mg=24 | Al=27,3 | Si=28 | P=31 | S=32 | Cl=35,5 | |
| 4 | K=39 | Ca=40 | —=44 | Ti=48 | V=51 | Cr=52 | Mn=55 | Fe=56, Co=59, Ni=59, Cu=63. |
| 5 | (Cu=63) | Zn=65 | —=68 | —=72 | As=75 | Se=78 | Br=80 | |
| 6 | Rb=85 | Sr=87 | ?Yt=88 | Zr=90 | Nb=94 | Mo=96 | —=100 | Ru=104, Rh=104, Pd=106, Ag=108. |
| 7 | (Ag=108) | Cd=112 | In=113 | Sn=118 | Sb=122 | Te=125 | J=127 | |
| 8 | Cs=133 | Ba=137 | ?Di=138 | ?Ce=140 | — | — | — | — — — — |
| 9 | (—) | — | — | — | — | — | — | |
| 10 | — | — | ?Er=178 | ?La=180 | Ta=182 | W=184 | — | Os=195, Ir=197, Pt=198, Au=199. |
| 11 | (Au=199) | Hg=200 | Tl=204 | Pb=207 | Bi=208 | — | — | |
| 12 | — | — | — | Th=231 | — | U=240 | — | — — — — |

FIGURE 2.2

Mendeleev's table of the elements organizes the elements by atomic mass. The table has a repeating pattern.

Mendeleev's table of the elements is called a **periodic table** because of its repeating pattern. Anything that keeps repeating is referred to as periodic. Other examples of things that are periodic include the monthly phases of the moon and the daily cycle of night and day. The term **period** refers to the interval between repetitions. In a periodic table, the periods are the rows of the table. In Mendeleev's table, each period contains eight elements, and then the pattern repeats in the next row.

### Predicting Missing Elements

Did you notice the blanks in Mendeleev's table (**Figure** 2.2)? They are spaces that Mendeleev left for elements that had not yet been discovered when he created his table. He predicted that these missing elements would eventually be discovered. Based on their position in the table, he could even predict their properties. For example, he predicted a missing element in row 5 of his group 3. He said it would have an atomic mass of about 68 and be a soft metal like other group 3 elements. Scientists searched for the missing element. They found it a few years later and named it gallium. Scientists searched for the other missing elements. Eventually, all of them were found.

An important measure of a good model is its ability to make accurate predictions. This makes it a useful model. Clearly, Mendeleev's periodic table was a useful model. It helped scientists discover new elements and make sense of those that were already known.

## The Modern Periodic Table of the Elements

A periodic table is still used today to classify the elements. **Figure** 2.3 shows the modern periodic table. You can see an interactive version at this URL: http://www.ptable.com/ .

### Basis of the Modern Periodic Table

In the modern periodic table, elements are organized by atomic number. The atomic number is the number of protons in an atom of an element. This number is unique for each element, so it seems like an obvious way to organize the elements. (Mendeleev used atomic mass instead of atomic number because protons had not yet been discovered when he made his table.) In the modern table, atomic number increases from left to right across each period. It also increases from top to bottom within each group. How is this like Mendeleev's table?

### Reading the Table

Besides atomic number, the periodic table includes each element's chemical symbol and class. Some tables include other information as well.

- The chemical symbol consists of one or two letters that come from the chemical's name in English or another language. The first letter is always written in upper case. The second letter, if there is one, is always written in lower case. For example, the symbol for lead is Pb. It comes from the Latin word *plumbum*, which means "lead." Find lead in **Figure** 2.3. What is its atomic number? You can access videos about lead and other elements in the modern periodic table at this URL: http://www.periodicvideos.com/index.htm .
- The classes of elements are metals, metalloids, and nonmetals. They are color-coded in the table. Blue stands for metals, orange for metalloids, and green for nonmetals. You can read about each of these three classes of elements later in the chapter, in the lesson "Classes of Elements."

### Periods

Rows of the modern table are called periods, as they are in Mendeleev's table. From left to right across a period, each element has one more proton than the element before it. In each period, elements change from metals on the left side of the table, to metalloids, and then to nonmetals on the right. **Figure** 2.4 shows this for period 4.

Some periods in the modern periodic table are longer than others. For example, period 1 contains only two elements. Periods 6 and 7, in contrast, are so long that some of their elements are placed below the main part of the table. They

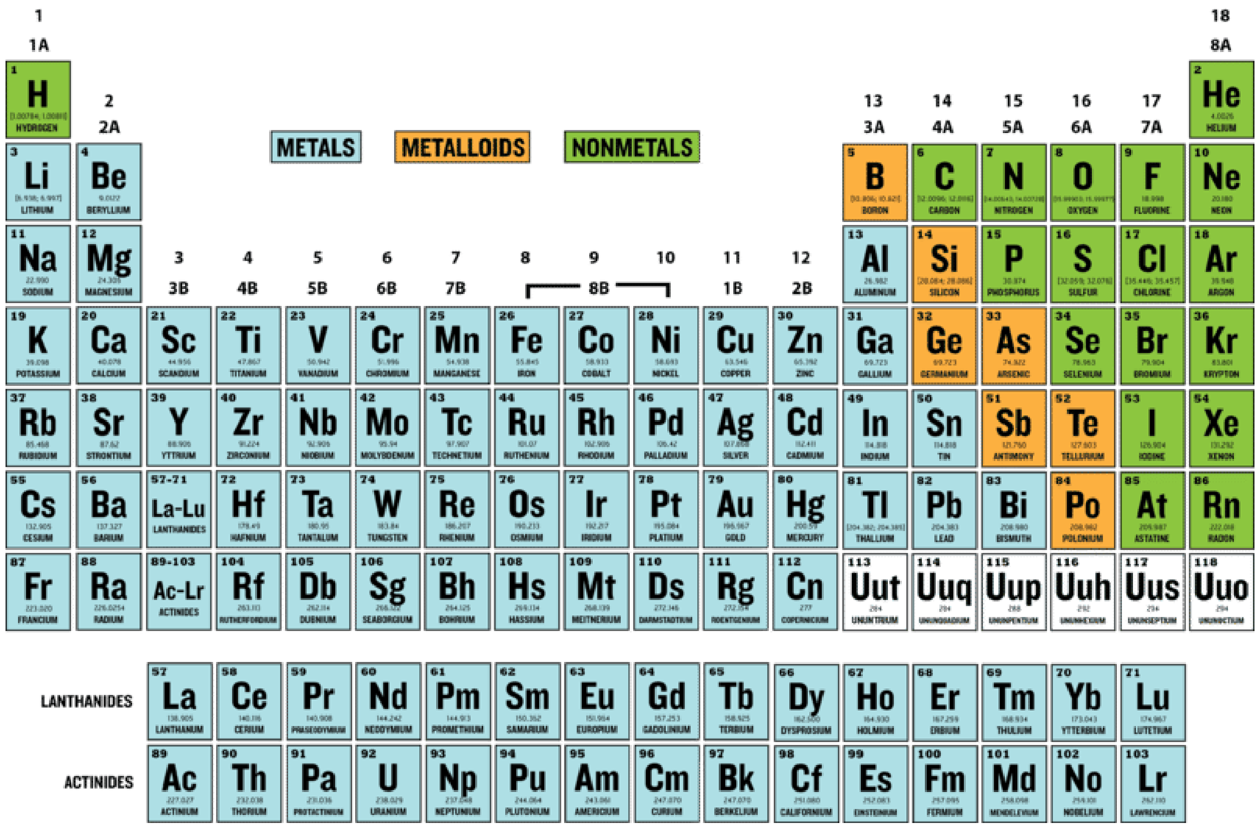



FIGURE 2.3

The modern periodic table of the elements is a lot like Mendeleev's table. But the modern table is based on atomic number instead of atomic mass. It also has more than 110 elements. Mendeleev's table only had about 65 elements.

are the elements starting with lanthanum (La) in period 6 and actinium (Ac) in period 7. Some elements in period 7 have not yet been named. They are represented by temporary symbols, such as Uub.

### Groups

Columns of the modern table are called groups, as they are in Mendeleev's table. However, the modern table has many more groups —18 to be exact. Elements in the same group have similar properties. For example, all elements in group 18 are colorless, odorless gases. You can read about the different groups of elements in this chapter's lesson on "Groups of Elements."

## Lesson Summary

- Mendeleev developed the first periodic table of the elements in 1869. He organized the elements by increasing atomic mass. He used his table to predict unknown elements. These were later discovered.
- The modern periodic table is based on atomic number. Elements in each period go from metals on the left to

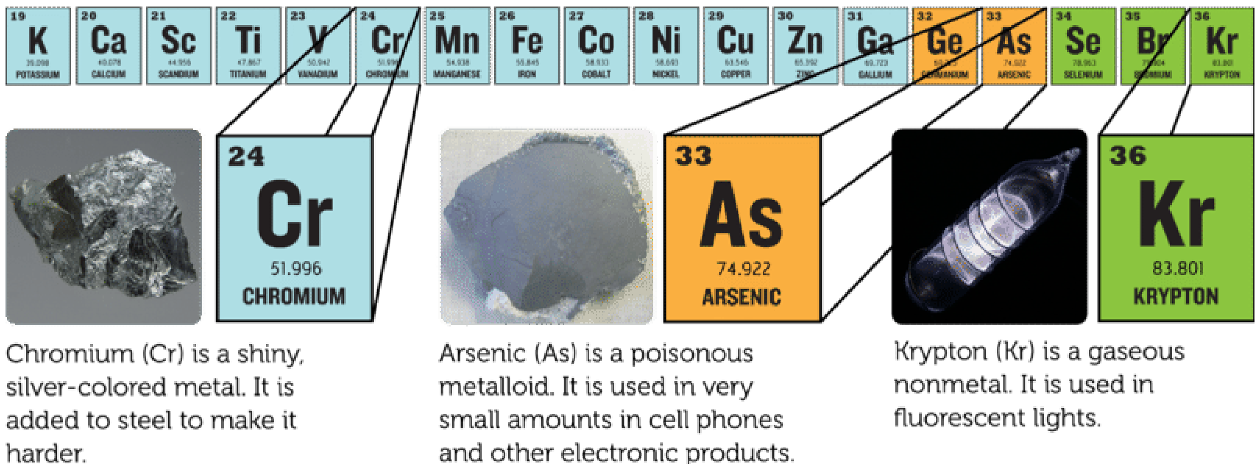



FIGURE 2.4

Like other periods, period 4 changes from metals on the left to metalloids and then nonmetals on the right.

metalloids and then nonmetals on the right. Within groups, elements have similar properties.

## Lesson Review Questions

### Recall

1. How did Mendeleev organize the elements?
2. How does the modern periodic table differ from Mendeleev's table?
3. What is a period in the periodic table?
4. What is a group in the periodic table?

### Apply Concepts

5. An unknown element has an atomic number of 44. Identify the element's symbol and the symbols of two other elements that have similar properties.

### Think Critically

6. Mendeleev's table and the modern periodic table organize the elements based on different information, yet most elements are in the same order in both tables. Explain why.

## Points to Consider

Elements are classified as metals, metalloids, or nonmetals.

- Do you know some examples of metals?
- How do you think metals might differ from the other two classes of elements?

# 2.2 Classes of Elements

## Lesson Objectives

- Identify properties of metals.
- List properties of nonmetals.
- Describe metalloids.
- Relate valence electrons to reactivity of elements by class.

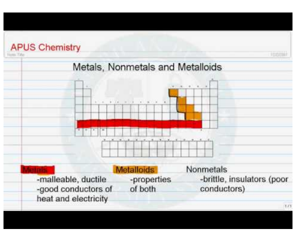
## Vocabulary

- metal
- metalloid
- nonmetal
- valence electron

## Introduction

Elements in different groups are lumped together in one of three classes, depending on their properties. The classes are metals, nonmetals, and metalloids. Knowing the class of an element lets you predict many of its properties. The video at the URL below is a good introduction to the classes.

http://www.youtube.com/watch?v=ZuQmionhkGU (2:04)



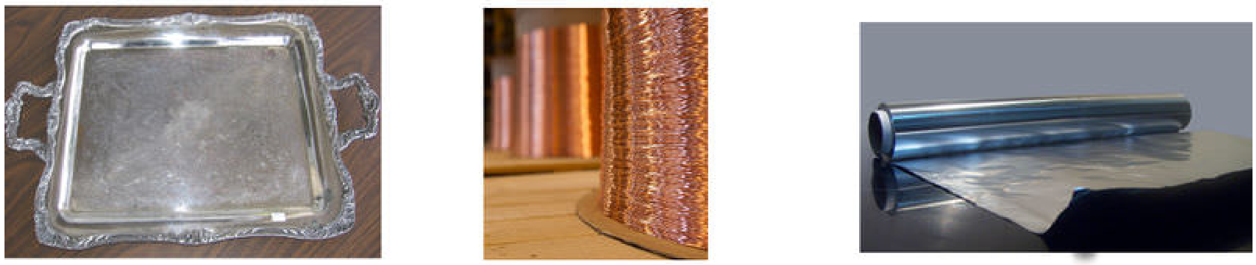

MEDIA

Click image to the left or use the URL below.

URL: https://www.ck12.org/flx/render/embeddedobject/5077

## Metals

**Metals** are elements that are good conductors of electricity. They are the largest of the three classes of elements. In fact, most elements are metals. Look back at the modern periodic table (**Figure** 2.3) in this chapter's lesson "How Elements Are Organized." Find the metals in the table. They are all the elements that are color-coded blue. Examples include sodium (Na), silver (Ag), and zinc (Zn).



Most metals are shiny. That's because they reflect a lot of light. This tray is made mainly of the metal silver (Ag).

Most metals are ductile. This means they can be pulled into long thin shapes, like these wires made of the metal copper (Cu).

Most metals are malleable. This means they can be formed into thin sheets without breaking, like this foil made of the metal aluminum (Al).

FIGURE 2.5

The three properties described here characterize most metals.

Metals have relatively high melting points, so almost all are solids at room temperature. The only exception is mercury (Hg), which is a liquid. Most metals are also good conductors of heat. That's why they are used for cooking pots and stovetops. Metals have other characteristic properties as well. Most are shiny, ductile, and malleable. These properties are illustrated in **Figure** 2.5. You can dig deeper into the properties of metals at this URL: http://www.bbc.co.uk/schools/gcsebitesize/science/add_gateway/periodictable/metalsrev1.shtml .

## Nonmetals

**Nonmetals** are elements that do not conduct electricity. They are the second largest class of elements. Find the nonmetals in **Figure** 2.3. They are all the elements on the right side of the table that are color-coded green. Examples of nonmetals include helium (He), carbon (C), and oxygen (O).

Nonmetals generally have properties that are the opposite of those of metals. They also tend to vary more in their properties than metals do. For example, nonmetals have relatively low boiling points, so many of them are gases at room temperature. But several nonmetals are solids, including carbon and phosphorus (P). One nonmetal, bromine (Br), is a liquid at room temperature.

Generally, nonmetals are also poor conductors of heat. In fact, they may be used for insulation. For example, the down filling in a down jacket is mostly air, which consists mainly of nitrogen (N) and oxygen (O). These nonmetal gases are poor conductors of heat, so they keep body heat in and cold air out. Solid nonmetals are dull rather than shiny. They are also brittle rather than ductile or malleable. You can see examples of solid nonmetals in **Figure** 2.6. You can learn more about specific nonmetals with the interactive table at this URL: http://library.thinkquest.org/3659/pertable/nonmetal.html .

These yellow piles of powder are sulfur (S), a nonmetal. Sulfur in rocks has been ground up to produce a powder. The powder has been heaped on a dock for shipment.

The "lead" in this pencil is actually graphite, a form of the nonmetal carbon (C). Graphite is brittle. It breaks easily if you put too much pressure on it.

These match heads are coated with the nonmetal phosphorus (P). Phosphorus is not malleable. If you tried to pound it flat, it would crumble into a powder.

FIGURE 2.6
Unlike metals, solid nonmetals are dull and brittle.

## Metalloids

Metalloids are elements that fall between metals and nonmetals in the periodic table. Just seven elements are metalloids, so they are the smallest class of elements. In **Figure** 2.3, they are color-coded orange. Examples of metalloids include boron (B), silicon (Si), and germanium (Ge).

**Metalloids** have some properties of metals and some properties of nonmetals. For example, many metalloids can conduct electricity but only at certain temperatures. These metalloids are called semiconductors. Silicon is an example. It is used in computer chips. It is also the most common metalloid on Earth. It is shiny like a metal but brittle like a nonmetal. You see a sample of silicon in **Figure** 2.7. The figure also shows other examples of metalloids. You can learn more about the properties of metalloids at this URL: http://library.thinkquest.org/3659/pertable/metaloid.html .


FIGURE 2.7
Metalloids share properties with both metals and nonmetals.

## Classes of Elements and Electrons

From left to right across the periodic table, each element has one more proton than the element to its left. Because atoms are always electrically neutral, for each added proton, one electron is also added. Electrons are added first to the lowest energy level possible until that level is full. Only then are electrons added to the next higher energy level.

### Electrons by Class

The increase in electrons across the periodic table explains why elements go from metals to metalloids and then to nonmetals from left to right across the table. Look at period 2 in **Figure** 2.8 as an example. Lithium (Li) is a metal, boron (B) a metalloid, and fluorine (F) and neon (Ne) are nonmetals. The inner energy level is full for all four elements. This level has just one orbital and can hold a maximum of two electrons. The outer energy level is a different story. This level has four orbitals and can hold a maximum of eight electrons. Lithium has just one electron in this level, boron has three, fluorine has seven, and neon has eight.

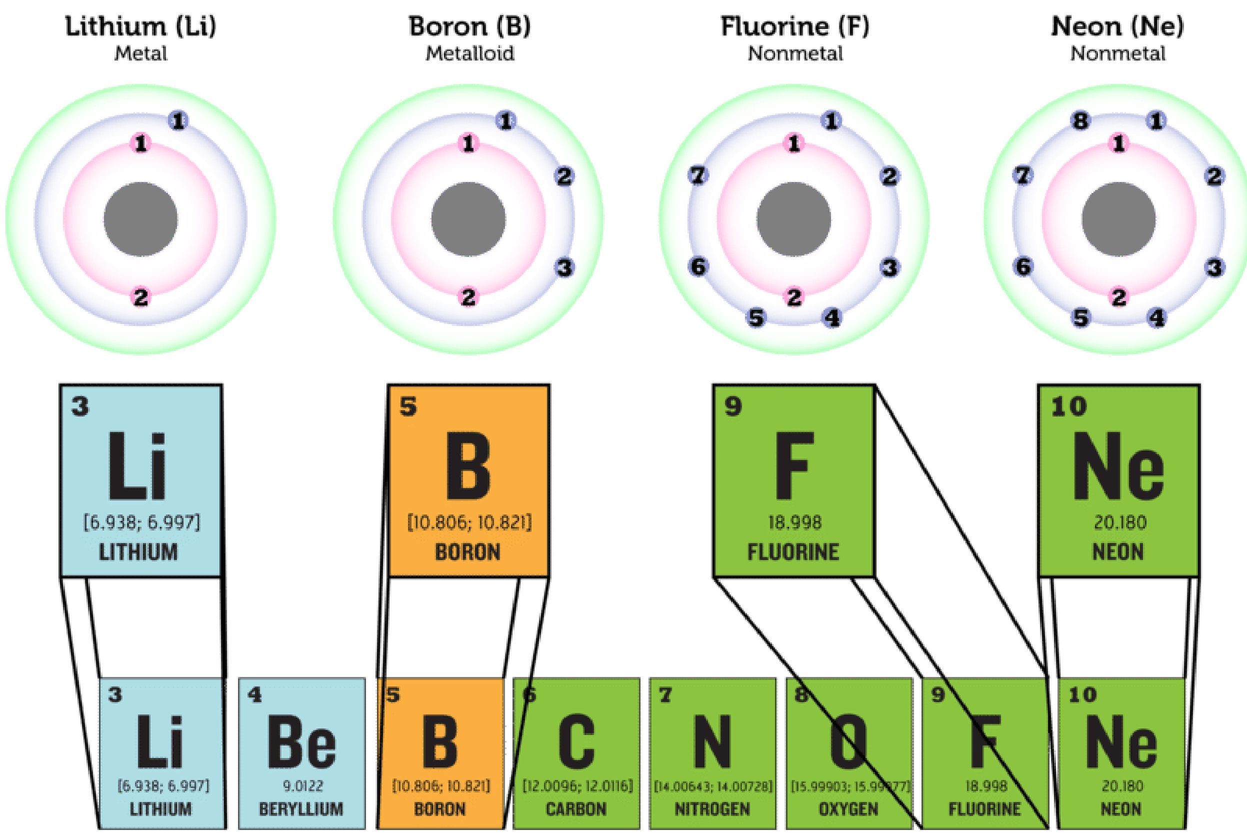



FIGURE 2.8

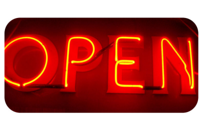
The number of electrons increases from left to right across each period in the periodic table. In period 2, lithium (Li) has the fewest electrons and neon (Ne) has the most. How do the numbers of electrons in their outer energy levels compare?

### Valence Electrons and Reactivity

The electrons in the outer energy level of an atom are called **valence electrons**. It is valence electrons that are potentially involved in chemical reactions. The number of valence electrons determines an element's reactivity, or how likely the element is to react with other elements. The number of valence electrons also determines whether the element can conduct electric current. That's because electric current is the flow of electrons. **Table** 2.1 shows how these properties vary in elements from each class.

- Metals such as lithium have an outer energy level that is almost empty. They "want" to give up their few valence electrons so they will have a full outer energy level. As a result, metals are very reactive and good conductors of electricity.
- Metalloids such as boron have an outer energy level that is about half full. These elements need to gain or lose too many electrons for a full outer energy level to come about easily. As a result, these elements are not very reactive. They may be able to conduct electricity but not very well.
- Some nonmetals, such as bromine, have an outer energy level that is almost full. They "want" to gain electrons so they will have a full outer energy level. As a result, these nonmetals are very reactive. Because they only accept electrons and do not give them up, they do not conduct electricity.
- Other nonmetals, such as neon, have a completely full outer energy level. Their electrons are already in the most stable arrangement possible. They are unreactive and do not conduct electricity.

**TABLE 2.1:** These examples show the relative reactivity of elements in the three classes.

| Element | Description |
| --- | --- |
| Lithium  | Lithium (Li) is a highly reactive metal. It has just one electron in its outer energy level. Lithium reacts explosively with water (see picture). It can react with moisture on skin and cause serious burns. |
| Boron   | Boron (B) is a metalloid. It has three valence electrons and is less reactive than lithium. Boron compounds dissolved in water form boric acid. Dilute boric acid is weak enough to use as eye wash. |
| Bromine | Bromine (Br) is an extremely reactive nonmetal. In fact, reactions with fluorine are often explosive, as you can see in the URL below. http://www.youtube.com/watch?v=vtWp45Eewtw |
| Neon  | Neon (Ne) is a nonmetal gas with a completely filled outer energy level. This makes it unreactive, so it doesn't combine with other elements. Neon is used for lighted signs like this one. You can learn why neon gives off light at this link: http://www.scientificamerican.com/article.cfm?id=how-do-neon-lights-work |

## Lesson Summary

- Metals are elements that are good conductors of electricity. They are the largest class of elements. Many metals are shiny, ductile, and malleable. They are also good conductors of heat. Almost all metals are solids are room temperature.
- Nonmetals are elements that do not conduct electricity. They are the second largest class of elements. Nonmetals are also poor conductors of heat. The majority of nonmetals are gases. Solid nonmetals are dull and brittle.
- Metalloids are elements that have properties of both metals and nonmetals. Some can conduct electricity but only at certain temperatures. They may be shiny but brittle. All metalloids are solids at room temperature.
- Atoms of elements in different classes vary in their number of valence electrons. This explains their differences in reactivity and conductivity.

## Lesson Review Questions

### Recall

1. What are metals? Name one example.
2. Define nonmetal, and give an example.
3. State one way that metalloids may be like metals and one way they may be like nonmetals.
4. What are valence electrons?

### Apply Concepts

5. A mystery element is a dull, gray solid. It is very reactive with other elements. Classify the mystery element as a metal, nonmetal, or metalloid. Explain your answer.

### Think Critically

6. Create a Venn diagram for metals, metalloids, and nonmetals. The diagram should show which properties are different and which, if any, are shared among the three groups of elements.
7. Relate number of valence electrons to reactivity of classes of elements.

## Points to Consider

The number of valence electrons increases from left to right across each period of the periodic table. By the end of the period, the outer energy level is full. Moving on to the next period of the table, electrons are added to the next higher energy level. This happens in each row of the periodic table.

- How do you think the number of valence electrons compares in elements within the same column (group) of the periodic table?
- How might this be reflected in the properties of elements within a group?

# 2.3 References


1. . http://commons.wikimedia.org/wiki/File:DIMendeleevCab.jpg . Public Domain
2. Dmitri Mendeleev. http://commons.wikimedia.org/wiki/File:Mendelejevs_periodiska_system_1871.png . Public Domain
3. Christopher Auyeung. CK-12 Foundation . CC BY-NC 3.0
4. Chromium and krypton: User:Jurii/Wikimedia Commons; Arsenic: Aram Dulyan (Wikimedia: Aramgutang). Chromium: http://commons.wikimedia.org/wiki/File:Chromium.jpg; Arsenic: http://commons.wikimedia.org/wiki/File:Native_arsenic.jpg; Krypton: http://commons.wikimedia.org/wiki/File:Krypton-glow.jpg . Chromium and krypton: CC-BY 3.0; Arsenic: Public Domain
5. Silver tray: The Living Room; Copper wire: Flickr:Flavio~; Aluminum foil: User:MdeVicente/Wikimedia Commons. Silver tray: http://www.flickr.com/photos/thelivingroominkenmore/4546677601/; Copper wire: http://www.flickr.com/photos/37873897@N06/6154424083/; Aluminum foil: http://commons.wikimedia.org/wiki/File:Aluminio.jpg . Silver tray: CC BY 2.0; Copper wire: CC BY 2.0; Aluminum foil: Public Domain
6. Sulfur: Ed Uthman; Pencil: Robert Lopez; Matches: Charles Knowles. Sulfur: http://www.flickr.com/photos/euthman/9183618186/; Pencil: CK-12 Foundation; Matches: http://www.flickr.com/photos/theknowlesgallery/4536616288/ . Sulfur: CC BY 2.0; Pencil: CC BY-NC 3.0; Matches: CC BY 2.0
7. User: Jurii/Wikimedia Commons. Silicon: http://commons.wikimedia.org/wiki/File:Silicon.jpg; Antimony:http://commons.wikimedia.org/wiki/File:Antimony-piece.jpg; Boron: http://commons.wikimedia.org/wiki/File:Boron.jpg . CC BY 3.0
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# CHAPTER 3 Chemical Bonding

## Chapter Outline



What do this lump of coal, diamond, and pencil "lead" all have in common? All three substances are forms of carbon. Are you surprised that one element can exist in forms that have such different properties? Do you know what explains it? The answer is chemical bonds. Carbon atoms chemically bond together in different ways to form these three substances. What are chemical bonds and how do they form? Read on to find out.

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# 3.1 Introduction to Chemical Bonds

## Lesson Objectives

- Define chemical bond.
- List general properties of compounds.

## Lesson Vocabulary

- chemical bond
- chemical formula

## Introduction

There is an amazing diversity of matter in the universe, but there are only about 100 elements. How can this relatively small number of pure substances make up all kinds of matter? Elements can combine in many different ways. When they do, they form new substances called compounds. For a video introduction to compounds, go this URL: http://www.youtube.com/watch?v=-HjMoTthEZ0 (3:53).


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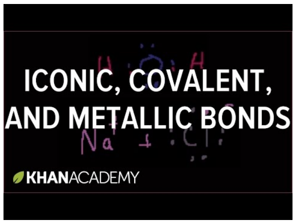
URL: https://www.ck12.org/flx/render/embeddedobject/195

## Chemical Bonding

Elements form compounds when they combine chemically. Their atoms join together to form molecules, crystals, or other structures. The atoms are held together by chemical bonds. A **chemical bond** is a force of attraction between atoms or ions. It occurs when atoms share or transfer valence electrons. Valence electrons are the electrons in the outer energy level of an atom. You can learn more about chemical bonds in this video: http://www.youtube.com/watch?v=CGA8sRwqIFg (13:21).

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Look at the example of water in **Figure** 3.1. A water molecule consists of two atoms of hydrogen and one atom of oxygen. Each hydrogen atom has just one electron. The oxygen atom has six valence electrons. In a water molecule, two hydrogen atoms share their two electrons with the six valence electrons of one oxygen atom. By sharing electrons, each atom has electrons available to fill its sole or outer energy level. This gives it a more stable arrangement of electrons that takes less energy to maintain.

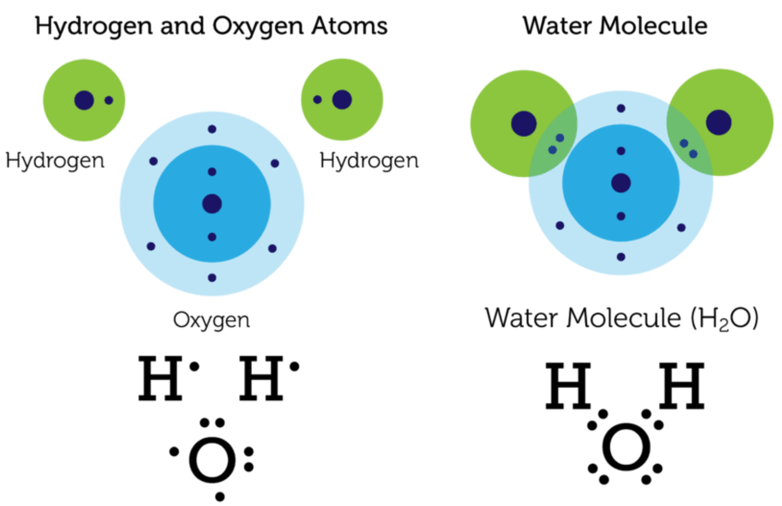



FIGURE 3.1

These diagrams show the valence electrons of hydrogen and water atoms and a water molecule. The diagrams represent electrons with dots, so they are called electron dot diagrams.

## Chemical Compounds

Water ($H_2O$) is an example of a chemical compound. Water molecules always consist of two atoms of hydrogen and one atom of oxygen. Like water, all other chemical compounds consist of a fixed ratio of elements. It doesn't matter how much or how little of a compound there is. It always has the same composition.

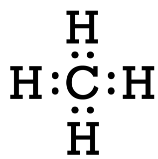
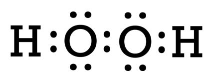
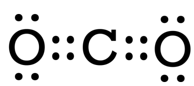
### Chemical Formulas

Elements are represented by chemical symbols. Examples are H for hydrogen and O for oxygen. Compounds are represented by **chemical formulas**. You've already seen the chemical formula for water. It's $H_2O$. The subscript 2 after the H shows that there are two atoms of hydrogen in a molecule of water. The O for oxygen has no subscript. When there is just one atom of an element in a molecule, no subscript is used. **Table** 3.1 shows some other examples of compounds and their chemical formulas.

**TABLE 3.1:** Examples of Chemical Compounds

| Name of Compound | Electron Dot Diagram | Numbers of Atoms | Chemical Formula |
| --- | --- | --- | --- |

**TABLE 3.1:** (continued)

| Name of Compound | Electron Dot Diagram | Numbers of Atoms | Chemical Formula |
|---|---|---|---|
| Hydrogen chloride |  | H = 1<br>Cl = 1 | HCl |
| Methane |  | C = 1<br>H = 4 | $CH_4$ |
| Hydrogen peroxide |  | H = 2<br>O = 2 | $H_2O_2$ |
| Carbon dioxide |  | C = 1<br>O = 2 | $CO_2$ |

**Problem Solving**

*Problem:* A molecule of ammonia consists of one atom of nitrogen (N) and three atoms of hydrogen (H). What is its chemical formula?

*Solution:* The chemical formula is $NH_3$.

**You Try It!**

*Problem:* A molecule of nitrogen dioxide consists of one atom of nitrogen (N) and two atoms of oxygen (O). What is its chemical formula?

### Same Elements, Different Compounds

The same elements may combine in different ratios. If they do, they form different compounds. **Figure** 3.2 shows some examples. Both water ($H_2O$) and hydrogen peroxide ($H_2O_2$) consist of hydrogen and oxygen. However, they have different ratios of the two elements. As a result, water and hydrogen peroxide are different compounds with different properties. If you've ever used hydrogen peroxide to disinfect a cut, then you know that it is very different from water! Both carbon dioxide ($CO_2$) and carbon monoxide (CO) consist of carbon and oxygen, but in different ratios. How do their properties differ?

### Types of Compounds

There are different types of compounds. They differ in the nature of the bonds that hold their atoms together. The type of bonds in a compound determines many of its properties. Three types of bonds are ionic, covalent, and metallic bonds. You will read about these three types in later lessons. You can also learn more about them by watching this video: http://www.youtube.com/watch?v=hEFeLYWTKX0 (7:18).

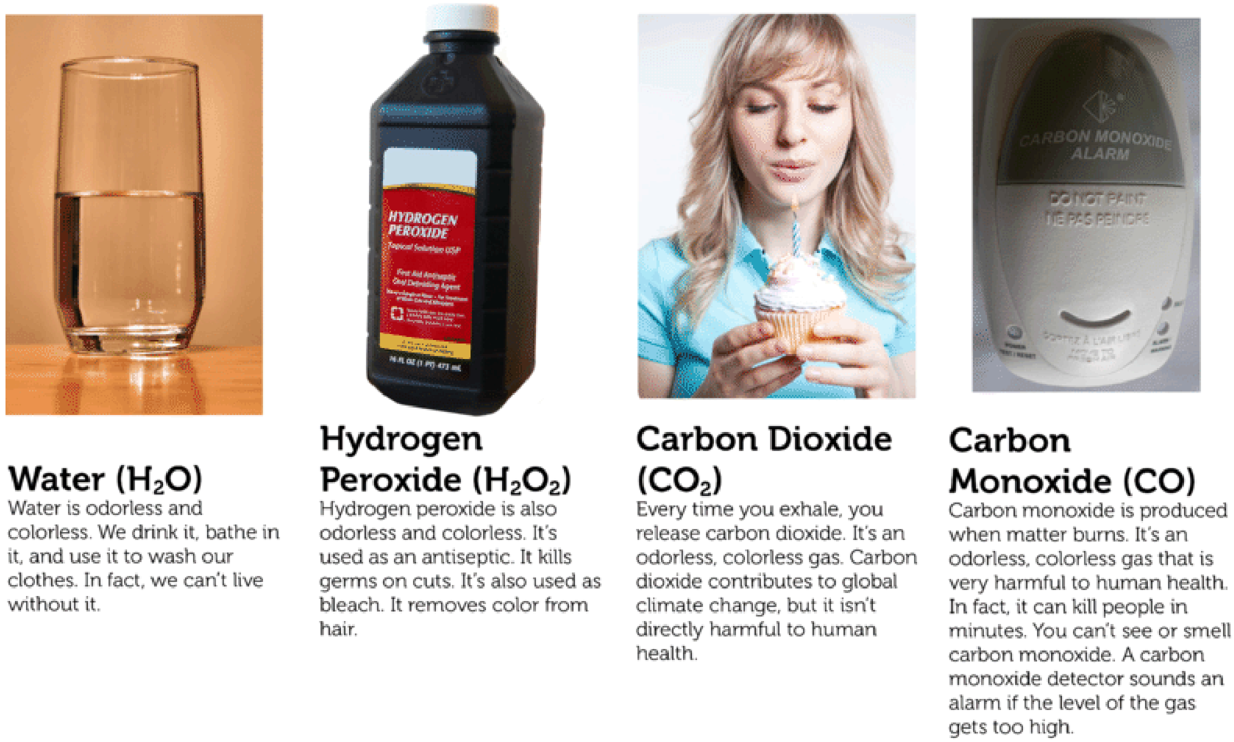
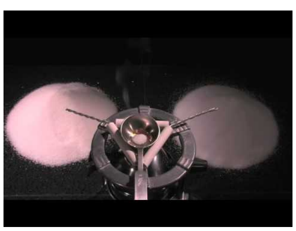



FIGURE 3.2

Different compounds may contain the same elements in different ratios. How does this affect their properties?



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## KQED: The Sweet Science of Chocolate

Chocolate: It's been revered for millennia by cultures throughout the world. But while it's easy to appreciate all of its delicious forms, creating this confection is a complex culinary feat. Local chocolate makers explain the elaborate engineering and chemistry behind this tasty treat. And learn why it's actually good for your health! For more information on the science of chocolate, see http://science.kqed.org/quest/video/the-sweet-science-of-chocolate/ .

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## Lesson Summary

- A chemical bond is a force of attraction between atoms. It occurs when atoms share or transfer electrons.
- A chemical compound is a new substance that forms when atoms of different elements form chemical bonds. A compound always consists of a fixed ratio of elements.

## Lesson Review Questions

### Recall

1. What is a chemical bond?
2. Define chemical compound.

### Apply Concepts

3. Which atoms and how many of each make up a molecule of sulfur dioxide? Write the chemical formula for this compound.

### Think Critically

4. Why does a molecule of water have a more stable arrangement of electrons than do individual hydrogen and oxygen atoms?
5. Explain how the ratio of elements in a compound is related to the compound's properties.

## Points to Consider

In this lesson, you learned about chemical bonds in a water molecule. The bonds form between atoms of hydrogen and oxygen when they share electrons. This type of bond is an example of a covalent bond.

- What might be other ways that atoms can bond together?
- How might ions form bonds?

# 3.2 Ionic Bonds

## Lesson Objectives

- Describe ions and isotopes
- Describe how ionic bonds form
- List properties of ionic compounds.

## Lesson Vocabulary

- ions
- isotope
- ionic bond
- ionic compound
- crystal lattice

## Introduction

All compounds form when atoms of different elements share or transfer electrons. In water, the atoms share electrons. In some other compounds, called **ionic compounds**, atoms transfer electrons. The electrons actually move from one atom to another. When atoms transfer electrons in this way, they become charged particles called ions. The ions are held together by ionic bonds.

The number of protons per atom is always the same for a given element. However, the number of neutrons may vary, and the number of electrons can change.

**Ions**

Sometimes atoms lose or gain electrons. Then they become **ions**. Ions have a positive or negative charge. That's because they do not have the same number of electrons as protons. If atoms lose electrons, they become positive ions, or **cation**s. If atoms gain electrons, they become negative ions, or **anions**.

Consider the example of fluorine in **Figure** below . A fluorine atom has nine protons and nine electrons, so it is electrically neutral. If a fluorine atom gains an electron, it becomes a fluoride ion with a negative charge of minus one.

**Isotopes of Atoms**

Some atoms of the same element may have different numbers of neutrons. For example, some carbon atoms have seven or eight neutrons instead of the usual six. Atoms of the same element that differ in number of neutrons are called **isotopes**. Many isotopes occur naturally. Usually one or two isotopes of an element are the most stable and common. Different isotopes of an element generally have the same chemical properties. That's because they have the same numbers of protons and electrons. For a video explanation of isotopes, go to this URL: http://www.youtube.com/watch?v=6w7raarHNA8 (5:23).

**An Example: Hydrogen Isotopes**

Hydrogen is a good example of isotopes because it has the simplest atoms. Three isotopes of hydrogen are modeled in **Figure** below . Most hydrogen atoms have just one proton and one electron and lack a neutron. They are just called hydrogen. Some hydrogen atoms have one neutron. These atoms are the isotope named deuterium. Other

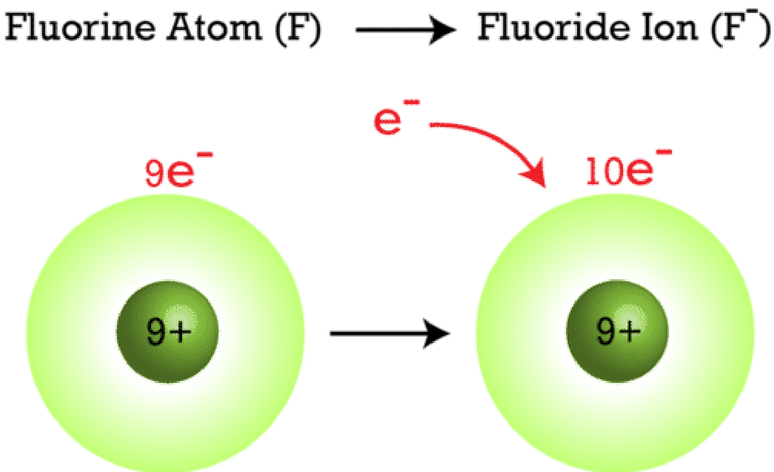



FIGURE 3.3

hydrogen atoms have two neutrons. These atoms are the isotope named tritium.

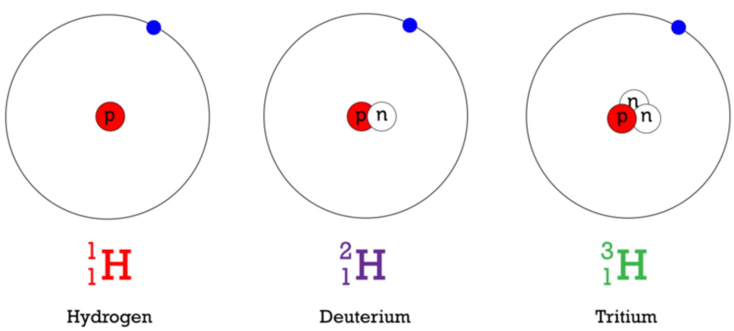



FIGURE 3.4

*All isotopes of a given element have the same number of protons (P), but they differ in the number of neutrons (N). What is the mass number of each isotope shown here?*

### Formation of Ionic Bonds

An **ionic bond** is the force of attraction that holds together positive and negative ions. It forms when atoms of a metallic element give up electrons to atoms of a nonmetallic element. **Figure** 3.5 shows how this happens.

In row 1 of **Figure** 3.5, an atom of sodium donates an electron to an atom of chlorine (Cl).

- By losing an electron, the sodium atom becomes a sodium ion. It now has one less electron than protons, giving it a charge of +1. Positive ions such as sodium are given the same name as the element. The chemical symbol has a plus sign to distinguish the ion from an atom of the element. The symbol for a sodium ion is $Na^+$.
- By gaining an electron, the chlorine atom becomes a chloride ion. It now has one more electron than protons, giving it a charge of -1. Negative ions are named by adding the suffix -*ide* to the first part of the element name. The symbol for chloride is $Cl^-$.

Sodium and chloride ions have equal but opposite charges. Opposites attract, so sodium and chloride ions attract each other. They cling together in a strong ionic bond. You can see this in row 2 of **Figure** 3.5. Brackets separate the

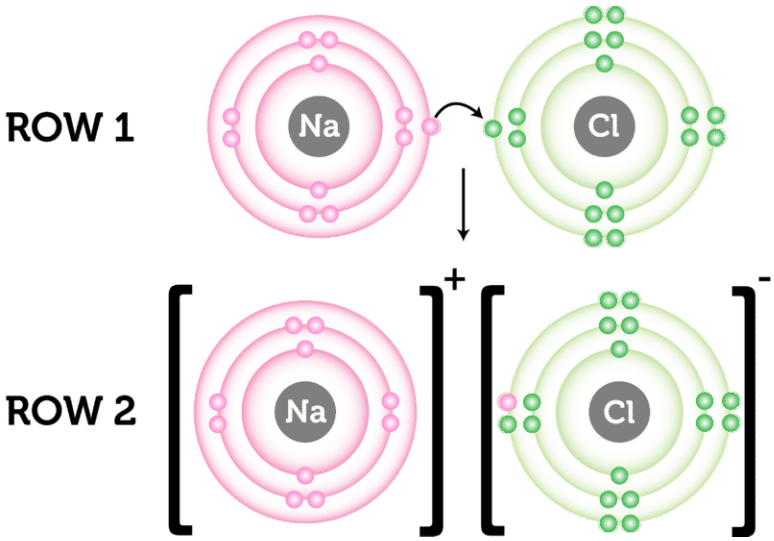



FIGURE 3.5
An ionic bond forms when the metal sodium gives up an electron to the non-metal chlorine.

ions in the diagram to show that the ions in the compound do not share electrons. You can see animations of sodium chloride forming at these URLs:

- http://web.jjay.cuny.edu/~acarpi/NSC/salt.htm
- http://www.visionlearning.com/library/module_viewer.php?mid=55

### Why Ionic Bonds Form

Ionic bonds form only between metals and nonmetals. Metals "want" to give up electrons, and nonmetals "want" to gain electrons. Find sodium (Na) in **Figure** 3.6. Sodium is an alkali metal in group 1. Like other group 1 elements, it has just one valence electron. If sodium loses that one electron, it will have a full outer energy level. Now find fluorine (F) in **Figure** 3.6. Fluorine is a halogen in group 17. It has seven valence electrons. If fluorine gains one electron, it will have a full outer energy level. After sodium gives up its valence electron to fluorine, both atoms have a more stable arrangement of electrons.

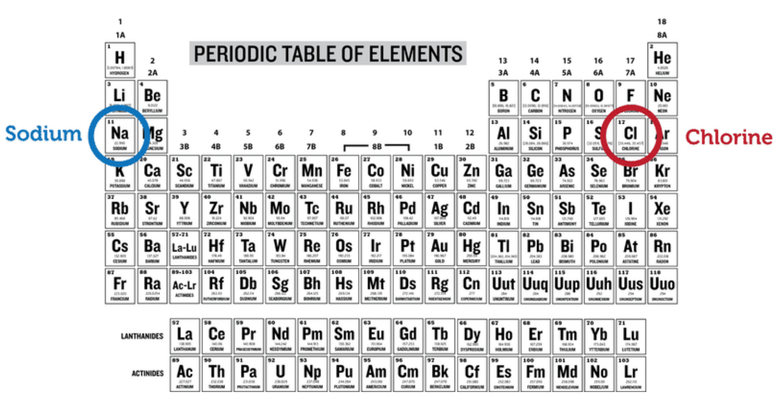



FIGURE 3.6
Sodium and chlorine are on opposite sides of the periodic table. How is this related to their numbers of valence electrons?

### Energy and Ionic Bonds

It takes energy to remove valence electrons from an atom. The force of attraction between the negative electrons and positive nucleus must be overcome. The amount of energy needed depends on the element. Less energy is needed to remove just one or a few electrons than many. This explains why sodium and other alkali metals form positive ions so easily. Less energy is also needed to remove electrons from larger atoms in the same group. For example, in group 1, it takes less energy to remove an electron from francium (Fr) at the bottom of the group than from lithium (Li) at the top of the group (see **Figure** 3.6). In bigger atoms, valence electrons are farther from the nucleus. As a result, the force of attraction between the electrons and nucleus is weaker.

What happens when an atom gains an electron and becomes a negative ion? Energy is released. Halogens release the most energy when they form ions. As a result, they are very reactive.

## Ionic Compounds

Ionic compounds contain ions of metals and nonmetals held together by ionic bonds. Ionic compounds do not form molecules. Instead, many positive and negative ions bond together to form a structure called a **crystal lattice**. You can see an example of a crystal in **Figure** 3.7. It shows the ionic compound sodium chloride. Positive sodium ions ($Na^+$) alternate with negative chloride ions ($Cl^-$). The oppositely charged ions are strongly attracted to each other.

**Helpful Hints**

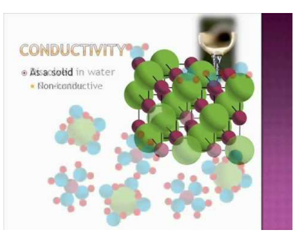
*Naming Ionic Compounds* Ionic compounds are named for their positive and negative ions. The name of the positive ion always comes first. For example, sodium and chloride ions form the compound named sodium chloride.

**You Try It!**

*Problem*: What is the name of the ionic compound composed of positive barium ions and negative iodide ions?

### Properties of Ionic Compounds

The crystal structure of ionic compounds is strong and rigid. It takes a lot of energy to break all those strong ionic bonds. As a result, ionic compounds are solids with high melting and boiling points (see **Table** 3.2). The rigid crystals are brittle and more likely to break than bend when struck. As a result, ionic crystals tend to shatter. You can learn more about the properties of ionic compounds by watching the video at this URL: http://www.youtube.com/watch?v=buWrSgs_ZHk (3:34).

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Compare the melting and boiling points of these ionic compounds with those of water (0°C and 100°C), which is *not* an ionic compound.

TABLE 3.2: Melting and Boiling Points of Select Ionic Compounds

| **Ionic Compound** | **Melting Point** (°C) | **Boiling Point** (°C) |
|---|---|---|
| Sodium chloride (NaCl) | 801 | 1413 |
| Calcium chloride ($CaCl_2$) | 772 | 1935 |

**TABLE 3.2:** (continued)

| **Ionic Compound** | **Melting Point** (°C) | **Boiling Point** (°C) |
|---|---|---|
| Barium oxide (BaO) | 1923 | 2000 |
| Iron bromide ($FeBr_3$) | 684 | 934 |

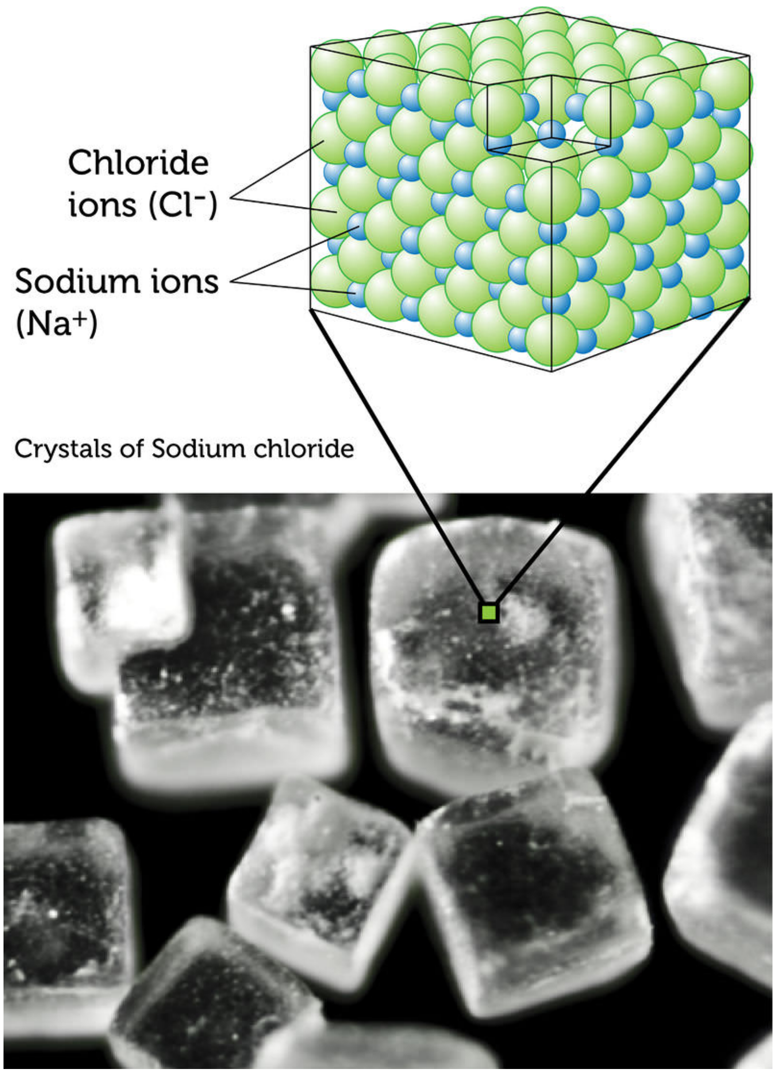


FIGURE 3.7
Sodium chloride crystals are cubic in shape. Other ionic compounds may have crystals with different shapes.

Solid ionic compounds are poor conductors of electricity. The strong bonds between ions lock them into place in the crystal. However, in the liquid state, ionic compounds are good conductors of electricity. Most ionic compounds dissolve easily in water. When they dissolve, they separate into individual ions. The ions can move freely, so they are good conductors of electricity. Dissolved ionic compounds are called electrolytes.

### Uses of Ionic Compounds

Ionic compounds have many uses. Some are shown in **Figure** 3.8. Many ionic compounds are used in industry. The human body also needs several ions for good health. Having low levels of the ions can endanger important functions such as heartbeat. Solutions of ionic compounds can be used to restore the ions.

## Lesson Summary

- If atoms lose or gain electrons, they become positive or negative ions. Atoms of the same element that have different numbers of neutrons are called isotopes.
- An ionic bond is the force of attraction that holds together oppositely charged ions. It forms when atoms of


FIGURE 3.8
Have you ever used any of these ionic compounds?

a metal transfer electrons to atoms of a nonmetal. When this happens, the atoms become oppositely charged ions.
- Ionic compounds form crystals instead of molecules. Ionic bonds are strong and the crystals are rigid. As a result, ionic compounds are brittle solids with high melting and boiling points. In the liquid state or dissolved in water, ionic compounds are good conductors of electricity.

## Lesson Review Questions

### Recall

1. Define isotope. Give an example.
2. What is an ionic bond?
3. List properties of ionic compounds.

### Apply Concepts

4. If an atom gains electrons, it becomes an ion. Is the ion positively or negatively charged? Explain your answer.
5. Create a model to represent the ionic bonds in a crystal of the salt lithium iodide (LiI).
6. A mystery compound is a liquid with a boiling point of 50°C. Is it likely to be an ionic compound? Why or why not?

### Think Critically

6. Explain why ionic bonds form only between atoms of metals and nonmetals.

## Points to Consider

Bonds form not only between atoms of metals and nonmetals. Nonmetals may also bond with nonmetals.

- How do you think bonds form between atoms of nonmetals?
- Can you think of examples of compounds that consist only of nonmetals?

# 3.3 Covalent Bonds

## Lesson Objectives

- Describe how covalent bonds form.

**Lesson Vocabulary**

- covalent bond
- covalent compound
- molecule

Introduction

Covalent bonds are bonds in which atoms share rather than transfer electrons. Compounds with covalent bonds are called covalent compounds.

## Formation of Covalent Bonds

A **covalent bond** is the force of attraction that holds together two atoms that share a pair of electrons. The shared electrons are attracted to the nuclei of both atoms. Covalent bonds form only between atoms of nonmetals. The two atoms may be the same or different elements. If the bonds form between atoms of different elements, a covalent compound forms. Covalent compounds are described in detail later in the lesson. To see a video about covalent bonding, go to this URL: http://www.youtube.com/watch?v=-Eh_0Dseg3E (6:20).

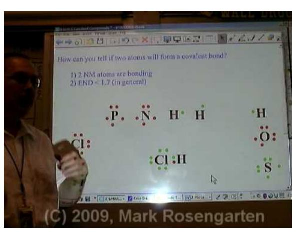


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**Figure** 3.9 shows an example of a covalent bond forming between two atoms of the same element, in this case two atoms of hydrogen. The two atoms share a pair of electrons. Hydrogen normally occurs in two-atom, or diatomic, molecules like this (*di-* means "two"). Several other elements also normally occur as diatomic molecules: nitrogen, oxygen, and all but one of the halogens (fluorine, chlorine, bromine, and iodine).

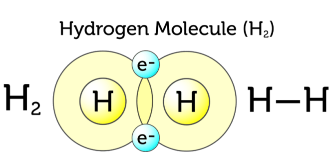



FIGURE 3.9

This figure shows three ways of representing a covalent bond. A dash (-) between two atoms represents one pair of shared electrons.

### Why Covalent Bonds Form

Covalent bonds form because they give atoms a more stable arrangement of electrons. Look at the hydrogen atoms in **Figure** 3.9. Alone, each hydrogen atom has just one electron. By sharing electrons with another hydrogen atom, it has two electrons: its own and the one in the other hydrogen atom. The shared electrons are attracted to both hydrogen nuclei. This force of attraction holds the two atoms together as a molecule of hydrogen.

Some atoms need to share more than one pair of electrons to have a full outer energy level. For example, an oxygen atom has six valence electrons. It needs two more electrons to fill its outer energy level. Therefore, it must form two covalent bonds. This can happen in many different ways. One way is shown in **Figure** 3.10. The oxygen atom in the figure has covalent bonds with two hydrogen atoms. This forms the covalent compound water.

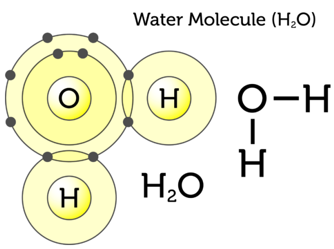



FIGURE 3.10
An oxygen atom has a more stable arrangement of electrons when it forms covalent bonds with two hydrogen atoms.

## Covalent Compounds

Covalent bonds between atoms of different elements form **covalent compounds**. The smallest, simplest covalent compounds have molecules with just two atoms. A **molecule** is two or more nonmetals. An example is hydrogen chloride (HCl). It consists of one hydrogen atom and one chlorine atom. The largest, most complex covalent molecules have thousands of atoms. Examples include proteins and carbohydrates. These are compounds in living things.

**Helpful Hints**

*Naming Covalent Compounds* Follow these rules in naming simple covalent compounds:

- The element closer to the left of the periodic table is named first.
- The second element gets the suffix *-ide*.
- Prefixes such as *di-* (2) and *tri-* (3) show the number of each atom in the compound. These are written with subscripts in the chemical formula.

*Example:* The gas that consists of one carbon atom and two oxygen atoms is named carbon dioxide. Its chemical formula is $CO_2$.

**You Try It!**

*Problem:* What is the name of the compound that contains three oxygen atoms and two nitrogen atoms? What is its chemical formula?

### Properties of Covalent Compounds

Covalent compounds have different properties than ionic compounds because of their bonds. Covalent compounds exist as individual molecules rather than crystals. It takes less energy for individual molecules than ions in a crystal to pull apart. As a result, covalent compounds have lower melting and boiling points than ionic compounds. Many are gases or liquids at room temperature. Covalent compounds have shared electrons. These are not free to move like the transferred electrons of ionic compounds. This makes covalent compounds poor conductors of electricity. Many covalent compounds also do not dissolve in water as all ionic compounds do.

The molecules of polar compounds are attracted to each other. You can see this in **Figure** below for water. A bond forms between the positive hydrogen end of one water molecule and the negative oxygen end of another water molecule. This type of bond is called a **hydrogen bond**. Hydrogen bonds are weak, but they still must be overcome when a polar substance changes from a solid to a liquid or from a liquid to a gas. As a result, polar covalent compounds may have higher melting and boiling points than nonpolar covalent compounds. To learn more about hydrogen bonding and when it occurs, see the video at this URL: http://www.youtube.com/watch?v=lkl5cbfqFRM (0:58).

## Lesson Summary

- A covalent bond is the force of attraction that holds together two atoms that share a pair of electrons. It forms between atoms of the same or different nonmetals. In polar covalent bonds, one atom attracts the shared electrons more strongly and becomes slightly negative. The other atom becomes slightly positive.
- Covalent compounds form individual molecules rather than crystals. Compared with ionic compounds, they have low melting and boiling points. They are also poor conductors of electricity.

## Lesson Review Questions

### Recall

**What is a covalent bond?**

**List general properties of covalent compounds.**

**Explain why covalent bonds form.**

## Points to Consider

You read in this lesson that covalent bonds may form between atoms of the same nonmetal element. For example, hydrogen atoms (H) commonly form covalent bonds to form hydrogen molecules ($H_2$).

1. Do you think bonds may also form between atoms of the same metallic element?
- Predict what these metallic bonds might be like.

# 3.4 Metallic Bonds

## Lesson Objectives

- Describe how metallic bonds form.
- Relate the nature of metallic bonds to the properties of metals.

## Lesson Vocabulary

- Metallic bond
- sea of electrons

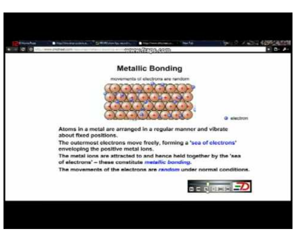
## Introduction

Special bonds form in metals that do not form in other classes of elements. They are called metallic bonds. The bonds explain some of the unique properties of elements in the metals class.

## Formation of Metallic Bonds

A **metallic bond** is the force of attraction between a positive metal ion and the valence electrons it shares with other ions of the metal. The positive ions form a lattice-like structure. You can see an example in **Figure** 3.11. (For an animated version, go to the URL below.) The ions are held together in the lattice by bonds with the valence electrons around them. These valence electrons include their own and those of other ions. Why do metallic bonds form? Recall that metals "want" to give up their valence electrons. This means that their valence electrons move freely. The electrons form a "sea" of negative charge surrounding the positive ions. This gives them the structure called **"sea of electrons."**

http://www.youtube.com/watch?v=c4udBSZfLHY (1:50)

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## Metallic Bonds and the Properties of Metals

Because of their freely moving electrons, metals are good conductors of electricity. Metals also can be shaped without breaking. They are ductile (can be shaped into wires) and malleable (can be shaped into thin sheets). Metals have these properties because of the nature of their metallic bonds.

A metallic lattice, like the one in **Figure** 3.11, may resemble a rigid ionic crystal. However, it is much more flexible. Look at **Figure** 3.12. It shows a blacksmith hammering a piece of red-hot iron in order to shape it. Why doesn't

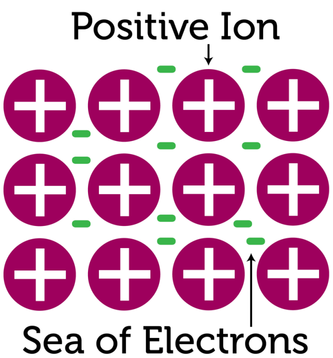


FIGURE 3.11
Positive metal ions and their shared electrons form metallic bonds.

the iron shatter, as an ionic crystal would? The ions of the metal can move within the "sea" of electrons without breaking the metallic bonds that hold them together. The ions can shift closer together or farther apart. In this way, the metal can change shape without breaking. You can learn more about metallic bonds and the properties of metals at this URL: http://www.youtube.com/watch?v=ap5pHBWwpu4 (6:12).

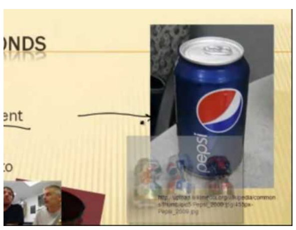

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FIGURE 3.12
A blacksmith shapes a piece of iron.

### Lesson Summary

- A metallic bond is the force of attraction between a positively charged metal ion and the valence electrons it shares with other ions of the metal. The electrons move freely around the positive ions, which form a lattice-like structure.

- With freely moving electrons, metals are good conductors of electricity. The positive ions of metals can also move within the "sea" of electrons without breaking the metallic bonds that hold them together. This allows metals to change shape without breaking.

## Lesson Review Questions

### Recall

1. What is a metallic bond?

### Apply Concepts

3. Create a model to represent the metallic bonds in solid iron (Fe).

### Think Critically

4. Relate metallic bonds to the properties of metals.
5. Compare and contrast metallic and ionic bonds.

## Points to Consider

Compounds form when atoms of different elements combine. This process is a chemical reaction.

- How would you define chemical reaction?
- How do you think chemical reactions are related to chemical changes in matter, such as wood burning and iron rusting?

# 3.5 References


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# CHAPTER 4 Chemical Reactions

## Chapter Outline



Does the term "chemical reaction" bring to mind an image like this one? In the picture, a chemist is mixing chemicals in a lab. Many chemical reactions take place in labs. However, most chemical reactions do not. Where do they occur? They happen in the world all around you. They even happen inside your own body. In fact, you would not be alive if it were not for the many chemical reactions that take place inside your cells. In this chapter, you will learn about chemical reactions, including two very important reactions that take place in living things.

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# 4.1 Introduction to Chemical Reactions

## Lesson Objectives

- Describe how chemical reactions occur.
- List signs that a chemical reaction has occurred.

## Lesson Vocabulary

- chemical reaction
- equilibrium
- product
- reactant

## Introduction

No doubt you've seen changes like those pictured in **Figure** 4.1. What do all these changes have in common? They are all chemical changes in matter. In a chemical change, matter changes into a different substance with different properties. Chemical changes occur because of chemical reactions. You can see more examples of chemical changes at this URL: http://www.youtube.com/watch?v=66kuhJkQCVM (2:05).

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## What Is a Chemical Reaction?

A **chemical reaction** is a process in which some substances change into different substances. Substances that start a chemical reaction are called **reactants**. Substances that are produced in the reaction are called **products**. Reactants and products can be elements or compounds. A chemical reaction can be represented by this general equation:

$$\text{Reactants} \rightarrow \text{Products}$$


FIGURE 4.1
Each of these pictures shows a chemical change taking place.

The arrow (→) shows the direction in which the reaction occurs. The reaction may occur quickly or slowly. For example, foam shoots out of a fire extinguisher as soon as the lever is pressed. But it might take years for metal to rust.

### Breaking and Reforming Chemical Bonds

In chemical reactions, bonds break in the reactants and new bonds form in the products. The reactants and products contain the same atoms, but they are rearranged during the reaction. As a result, the atoms are in different combinations in the products than they were in the reactants.

Look at the example in **Figure** 4.2. It shows how water forms. Bonds break in molecules of hydrogen and oxygen. Then new bonds form in molecules of water. In both reactants and products, there are four hydrogen atoms and two oxygen atoms. But the atoms are combined differently in water. You can see another example at this URL: http://www.avogadro.co.uk/h_and_s/bondenthalpy/bondenthalpy.htm .

### Reaction Direction and Equilibrium

The arrow in **Figure** 4.2 shows that the reaction goes from left to right, from hydrogen and oxygen to water. The reaction can also go in the reverse direction. If an electric current passes through water, water molecules break down into molecules of hydrogen and oxygen. This reaction would be represented by a right-to-left arrow (←) in **Figure** 4.2.

Many other reactions can also go in both forward and reverse directions. Often, a point is reached at which the forward and reverse reactions occur at the same rate. When this happens, there is no overall change in the amount of reactants and products. This point is called **equilibrium**, which refers to a balance between any opposing changes. You can see an animation of a chemical reaction reaching equilibrium at this URL: http://www.tutorvista.com/content/chemistry/chemistry-ii/chemical-equilibrium/chemical-equilibrium-animation.php .

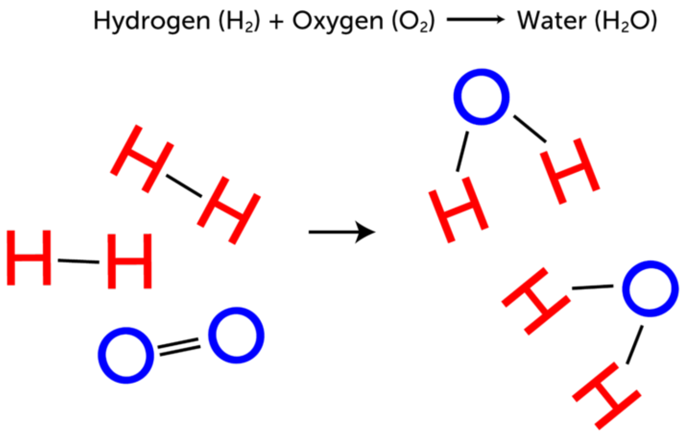



FIGURE 4.2
A chemical reaction changes hydrogen and oxygen to water.

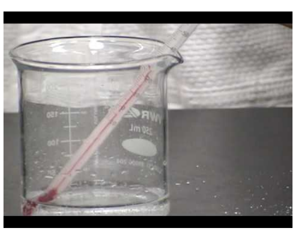
## Evidence of Chemical Reactions

Not all changes in matter involve chemical reactions. For example, there are no chemical reactions involved in changes of state. When liquid water freezes or evaporates, it is still water. No bonds are broken and no new products are formed.

How can you tell whether a change in matter involves a chemical reaction? Often, there is evidence. Four common signs that a chemical reaction has occurred are:

- Change in color: the products are a different color than the reactants.
- Change in temperature: heat is released or absorbed during the reaction.
- Production of a gas: gas bubbles are released during the reaction.
- Production of a solid: a solid settles out of a liquid solution. The solid is called a precipitate.

You can see examples of each type of evidence in **Figure** 4.3 and at this URL: http://www.youtube.com/watch?v=gs0j1EZJ1Uc (9:57).



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## Lesson Summary

- A chemical reaction is a process in which some substances change into different substances. In a chemical reaction, bonds break in reactants and new bonds form in products.
- Evidence that a chemical reaction has occurred include a change in color, a change in temperature, the production of a gas, or the formation of a precipitate.


FIGURE 4.3
Can you think of other examples of changes like these? If so, they probably indicate that a chemical reaction has occurred.

## Lesson Review Questions

### Recall

1. Define chemical reaction.
2. What are the reactants and products in a chemical reaction?
3. Describe what happens to the atoms involved in a chemical reaction.
4. List four common signs that a chemical reaction has occurred.

### Apply Concepts

5. Tina made a "volcano" by pouring vinegar over a "mountain" of baking soda. The wet baking soda bubbled and foamed. Did a chemical reaction occur? How do you know?

### Think Critically

6. Explain the meaning of the term "equilibrium" as it applies to a chemical reaction. How can you tell when a chemical reaction has reached equilibrium?

## Points to Consider

In **Figure** 4.2, you saw how hydrogen and oxygen combine chemically to form water.

- How could you use chemical symbols and formulas to represent this reaction?
- How many molecules of hydrogen and oxygen are involved in this reaction? How many molecules of water are produced? How could you include these numbers in your representation of the reaction?

# 4.2 Chemical Equations

## Lesson Objectives

- Describe how to write chemical equations.
- Demonstrate how to balance chemical equations.
- Relate the law of conservation of mass to balancing chemical equations.

## Lesson Vocabulary

- chemical equation

## Introduction

Chemists use a standard method to represent chemical reactions. It includes chemical symbols and formulas to stand for reactants and products. The symbols and formulas are used to write chemical equations.

## Writing Chemical Equations

A **chemical equation** is a symbolic representation of a chemical reaction. It is a shorthand way of showing how atoms are rearranged in the reaction. The general form of a chemical equation was introduced in this chapter's lesson "Introduction to Chemical Reactions." It is:

$$\text{Reactants} \rightarrow \text{Products}$$

Consider the simple example in **Figure** 4.4. When carbon (C) reacts with oxygen ($O_2$), it produces carbon dioxide ($CO_2$). The chemical equation for this reaction is:

$$C + O_2 \rightarrow CO_2$$

The reactants are one atom of carbon and one molecule of oxygen. When there is more than one reactant, they are separated by plus signs (+). The product is one molecule of carbon dioxide. If more than one product were produced, plus signs would be used between them as well.

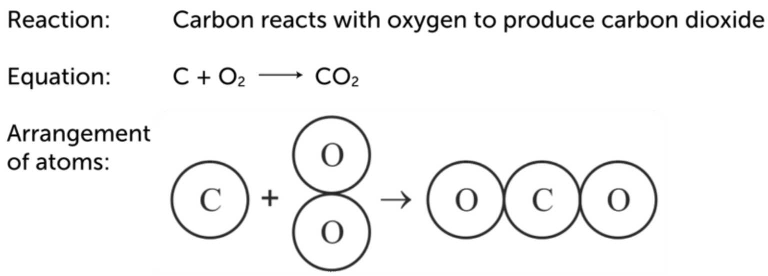



FIGURE 4.4
This figure shows a common chemical reaction. The drawing below the equation shows how the atoms are rearranged in the reaction. What chemical bonds are broken and what new chemical bonds are formed in this reaction?

## Balancing Chemical Equations

Some chemical equations are more challenging to write. Consider the reaction in which hydrogen ($H_2$) and oxygen ($O_2$) combine to form water ($H_2O$). Hydrogen and oxygen are the reactants, and water is the product. To write a chemical equation for this reaction, you would start by writing symbols for the reactants and products:

***Equation 1:*** $H_2 + O_2 \rightarrow H_2O$

Like equations in math, equations in chemistry must balance. There must be the same number of each type of atom in the products as there is in the reactants. In equation 1, count the number of hydrogen and oxygen atoms on each side of the arrow. There are two hydrogen atoms in both reactants and products. There are two oxygen atoms in the reactants but only one in the product. Therefore, equation 1 is not balanced.

### Using Coefficients

Coefficients are used to balance chemical equations. A coefficient is a number placed in front of a chemical symbol or formula. It shows how many atoms or molecules of the substance are involved in the reaction. For example, two molecules of hydrogen would be written as $2H_2$. A coefficient of 1 usually isn't written.

Coefficients can be used to balance equation 1 (above) as follows:

***Equation 2:*** $2H_2 + O_2 \rightarrow 2H_2O$

Equation 2 shows that two molecules of hydrogen react with one molecule of oxygen to produce two molecules of water. The two molecules of hydrogen each contain two hydrogen atoms. There are now four hydrogen atoms in both reactants and products. Is equation 2 balanced? Count the oxygen atoms to find out.

### Steps in Balancing a Chemical Equation

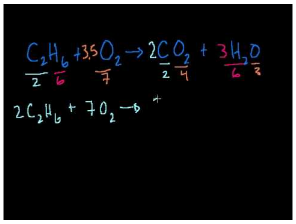
Balancing a chemical equation involves a certain amount of trial and error. In general, however, you should follow these steps:

1. Count the number of each type of atom in reactants and products. Does the same number of each atom appear on both sides of the arrow? If not, the equation is not balanced, and you need to go to step 2.
2. Add coefficients to increase the number of atoms or molecules of reactants or products. Use the smallest coefficients possible.
3. Repeat steps 1 and 2 until the equation is balanced.

**Helpful Hint**

When you balance chemical equations, never change the subscripts in chemical formulas. Changing subscripts changes the substances involved in the reaction. Change only the coefficients.

Work through the **Problem Solving** examples below. Then do the **You Try It!** problems to check your understanding. If you need more help, go to this URL: http://www.youtube.com/watch?v=RnGu3xO2h74 (14:28).



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**Problem Solving**

*Problem:* Balance this chemical equation: $N_2 + H_2 \rightarrow NH_3$

*Hints for balancing*

1. Two N are needed in the products to match the two N ($N_2$) in the reactants. Add the coefficient 2 in front of $NH_3$. Now N is balanced.
2. Six H are now needed in the reactants to match the six H in the products. Add the coefficient 3 in front of $H_2$. Now H is balanced.

*Solution:* $N_2 + 3H_2 \rightarrow 2NH_3$

*Problem:* Balance this chemical equation: $CH_4 + O_2 \rightarrow CO_2 + H_2O$

*Solution:* $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

**You Try It!**

*Problem:* Balance these chemical equations:

$$Zn + HCl \rightarrow ZnCl_2 + H_2$$
$$Cu + O_2 \rightarrow CuO$$

## Conserving Mass

Why must chemical equations be balanced? It's the law! Matter cannot be created or destroyed in chemical reactions. This is the law of conservation of mass. In every chemical reaction, the same mass of matter must end up in the products as started in the reactants. Balanced chemical equations show that mass is conserved in chemical reactions.

How do scientists know that mass is always conserved in chemical reactions? Careful experiments in the 1700s by a French chemist named Antoine Lavoisier led to this conclusion. For this and other contributions, Lavoisier has been called the father of modern chemistry.

Lavoisier carefully measured the mass of reactants and products in many different chemical reactions. He carried out the reactions inside a sealed jar, like the one in **Figure** 4.5. As a result, any gases involved in the reactions were captured and could be measured. In every case, the total mass of the jar and its contents was the same after the reaction as it was before the reaction took place. This showed that matter was neither created nor destroyed in the reactions. Another outcome of Lavoisier's research was his discovery of oxygen. You can learn more about Lavoisier and his important research at: http://www.youtube.com/watch?v=x9iZq3ZxbO8

FIGURE 4.5
Lavoisier carried out several experiments inside a sealed glass jar. Why was sealing the jar important for his results?

## Lesson Summary

- A chemical equation is a symbolic representation of a chemical reaction. It shows how atoms are rearranged in the reaction.
- Equations in chemistry must balance. There must be the same number of each type of atom in the products as there is in the reactants. Coefficients are used to balance chemical equations. They show how many atoms or molecules of a substance are involved in a reaction.
- Chemical equations must be balanced because matter cannot be created or destroyed. This is the law of conservation of mass. Experiments by Antoine Lavoisier led to this law.

## Lesson Review Questions

### Recall

1. What is a chemical equation? Give an example.
2. What is a coefficient? How are coefficients used in chemistry?
3. Describe how Antoine Lavoisier showed matter is conserved in chemical reactions.

### Apply Concepts

4. Draw a sketch that shows how atoms are rearranged in the chemical reaction represented by equation 2.
5. Balance this chemical equation: $Hg + O_2 \rightarrow HgO$.

### Think Critically

6. Explain why it is necessary to balance chemical equations.

## Points to Consider

In this lesson, you saw examples of chemical reactions in which two reactants combine to yield a single product. This is called a synthesis reaction. It is just one type of chemical reaction.

- What might be other types of chemical reactions?
- How might one reactant produce more than one product?

# 4.3 Types of Chemical Reactions

## Lesson Objectives

- Explain how synthesis reactions occur.
- Describe how decomposition reactions occur.
- Describe single and double replacement reactions.
- Explain how combustion reactions occur.

## Lesson Vocabulary

- combustion reaction
- decomposition reaction
- replacement reaction
- synthesis reaction

## Introduction

Most of the chemical reactions you have seen so far in this chapter are synthesis reactions. In this type of reaction, two or more reactants combine to synthesize a product. There are several other types of chemical reactions, including decomposition, replacement, and combustion reactions. You will read about all four types of reactions in this lesson. **Table** 4.1 summarizes the four types of chemical reactions you will read about in the rest of the lesson. You can see demonstrations of each type at this URL: http://www.youtube.com/watch?v=nVysOW0Lb8U .

TABLE 4.1: Four Types of Chemical Reactions

| Type of Reaction | General Equation | Example |
|---|---|---|
| Synthesis | $A+B \rightarrow C$ | $2Na+Cl_2 \rightarrow 2NaCl$ |
| Decomposition | $AB \rightarrow A+B$ | $2H_2O \rightarrow 2H_2+O_2$ |
| Replacement | $A+BC \rightarrow B+AC$ | $2K+2H_2O \rightarrow 2KOH+H_2$ |
| Single | $AB+CD \rightarrow AD+CB$ | $NaCl+AgF \rightarrow NaF+AgCl$ |
| Double | | |
| Combustion | fuel + oxygen $\rightarrow$ carbon dioxide + water | $CH_4+2O_2 \rightarrow CO_2+2H_2O$ |

## Synthesis Reactions

A **synthesis reaction** occurs when two or more reactants combine to form a single product. A synthesis reaction can be represented by the general equation:

$$A + B \rightarrow C$$

In this general equation (and others like it in this lesson), the letters $A, B, C$, and so on represent atoms or ions of elements. The arrow shows the direction of the reaction. The letters on the left side of the arrow are the reactants that begin the chemical reaction. The letters on the right side of the arrow are the product of the reaction. Two examples of synthesis reactions are described below. You can see more examples at this URL: http://www.youtube.com/watch?v=dxlWtsFinTM .

### Synthesis Example 1

An example of a synthesis reaction is the combination of sodium (Na) and chlorine (Cl) to produce sodium chloride (NaCl). This reaction is represented by the chemical equation:

$$2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$$

Sodium is a highly reactive metal, and chlorine is a poisonous gas (see **Figure** 4.6). The compound they synthesize has very different properties. It is table salt, which is neither reactive nor poisonous. In fact, salt is a necessary component of the human diet.


FIGURE 4.6
Sodium and chlorine combine to synthesize table salt.

### Synthesis Example 2

Another example of a synthesis reaction is illustrated in **Figure** 4.7. The brown haze in the air over the city of Los Angeles is smog. A major component of smog is nitrogen dioxide ($NO_2$). It forms when nitric oxide (NO), from sources such as car exhaust, combines with oxygen ($O_2$) in the air. The equation for this reaction is:

$$2NO + O_2 \rightarrow 2NO_2$$

Nitrogen dioxide is a toxic gas with a sharp odor. It can irritate the eyes and throat and trigger asthma attacks. It is a major air pollutant.

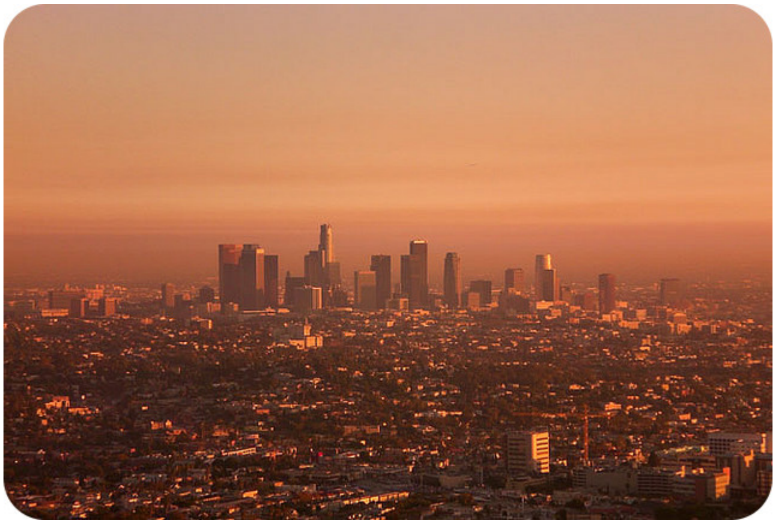

FIGURE 4.7
In this photo, the air over Los Angeles, California is brown with smog.

## Decomposition Reactions

A decomposition reaction is the reverse of a synthesis reaction. In a **decomposition reaction**, one reactant breaks down into two or more products. This can be represented by the general equation:

$$AB \rightarrow A + B$$

Two examples of decomposition reactions are described below. You can see other examples at this URL: http://www.youtube.com/watch?v=dxlWtsFinTM .

### Decomposition Example 1

An example of a decomposition reaction is the breakdown of carbonic acid ($H_2CO_3$) to produce water ($H_2O$) and carbon dioxide ($CO_2$). The equation for this reaction is:

$$H_2CO_3 \rightarrow H_2O + CO_2$$

Carbonic acid is synthesized in the reverse reaction. It forms when carbon dioxide dissolves in water. For example, some of the carbon dioxide in the atmosphere dissolves in the ocean and forms carbonic acid. The amount of carbon dioxide in the atmosphere has increased over recent decades (see **Figure** 4.8). As a result, the acidity of ocean water is also increasing. How do you think this might affect ocean life?

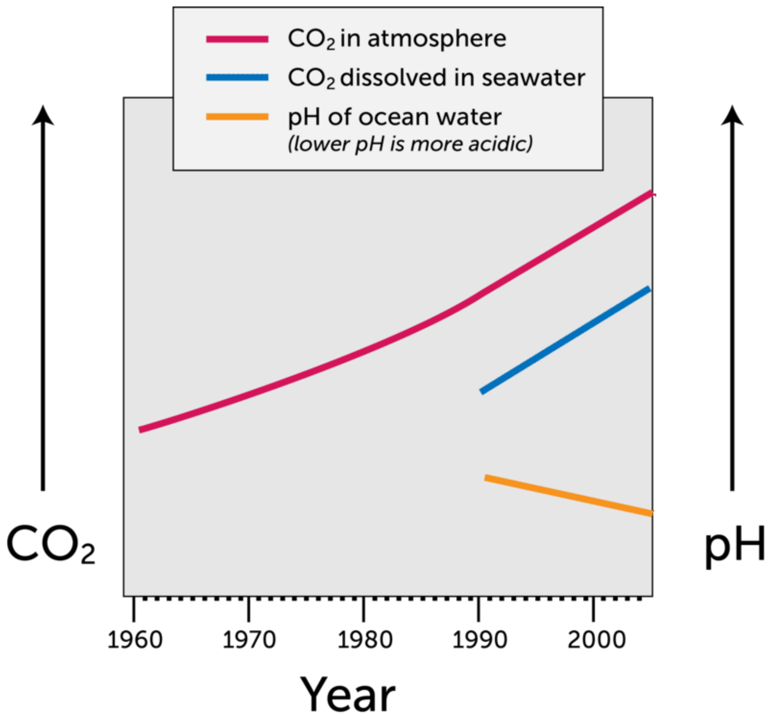


FIGURE 4.8
As carbon dioxide increases in the atmosphere, more carbon dioxide dissolves in ocean water.

### Decomposition Example 2

Another example of a decomposition reaction is illustrated in **Figure** 4.9. Water ($H_2O$) decomposes to hydrogen ($H_2$) and oxygen ($O_2$) when an electric current passes through it. This reaction is represented by the equation:

$$2H_2O \rightarrow 2H_2 + O_2$$

What is the reverse of this decomposition reaction? (*Hint*: How is water synthesized? You can look at this chapter's "Introduction to Chemical Reactions" lesson to find out.)

## Replacement Reactions

**Replacement reactions** involve ions. They occur when ions switch places in compounds. There are two types of replacement reactions: single and double. Both types are described below.

### Single Replacement Reactions

A single replacement reaction occurs when one ion takes the place of another in a single compound. This type of reaction has the general equation:

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

Do you see how $A$ has replaced $B$ in the compound? The compound $BC$ has become the compound $AC$.

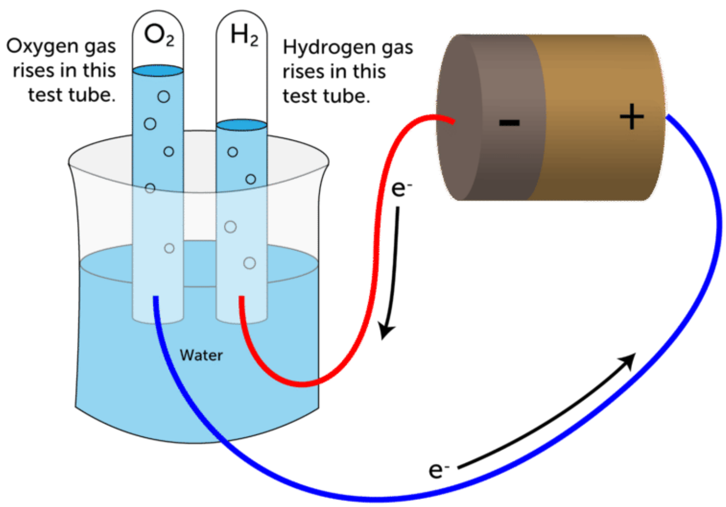


FIGURE 4.9
A decomposition reaction occurs when an electric current passes through water.

An example of a single replacement reaction occurs when potassium (K) reacts with water ($H_2O$). A colorless solid called potassium hydroxide (KOH) forms, and hydrogen gas ($H_2$) is released. The equation for the reaction is:

$$2K + 2H_2O \rightarrow 2KOH + H_2$$

Potassium is a highly reactive group 1 alkali metal, so its reaction with water is explosive. You can actually watch this reaction occurring at: http://commons.wikimedia.org/wiki/File:Potassium_water_20.theora.ogv .

### Double Replacement Reactions

A double replacement reaction occurs when two compounds exchange ions. This produces two new compounds. A double replacement reaction can be represented by the general equation:

$$AB + CD \rightarrow AD + CB$$

Do you see how *B* and *D* have changed places? Both reactant compounds have changed.

An example of a double replacement reaction is sodium chloride (NaCl) reacting with silver fluoride (AgF). This reaction is represented by the equation:

$$NaCl + AgF \rightarrow NaF + AgCl$$

Cl and F have changed places. Can you name the products of this reaction?

## Combustion Reactions

A **combustion reaction** occurs when a substance reacts quickly with oxygen ($O_2$). You can see an example of a combustion reaction in **Figure** 4.10. Combustion is commonly called burning. The substance that burns is usually referred to as fuel. The products of a combustion reaction include carbon dioxide ($CO_2$) and water ($H_2O$). The reaction typically gives off heat and light as well. The general equation for a combustion reaction can be represented by:

$$\text{Fuel} + O_2 \rightarrow CO_2 + H_2O$$

FIGURE 4.10
The burning of charcoal is an example of a combustion reaction.

### Combustion Example 1

The fuel that burns in a combustion reaction is often a substance called a hydrocarbon. A hydrocarbon is a compound that contains only carbon (C) and hydrogen (H). Fossil fuels, such as natural gas, consist of hydrocarbons. Natural gas is a fuel that is commonly used in home furnaces and gas stoves (see **Figure** 4.11). The main component of natural gas is the hydrocarbon called methane ($CH_4$). The combustion of methane is represented by the equation:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

### Combustion Example 2

Your own body cells burn fuel in combustion reactions. The fuel is glucose ($C_6H_{12}O_6$), a simple sugar. The process in which combustion of glucose occurs in body cells is called cellular respiration. This combustion reaction provides energy for life processes. Cellular respiration can be summed up by the equation:

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$$

Where does glucose come from? It is produced by plants during photosynthesis. In this process, carbon dioxide and water combine to form glucose. Which type of chemical reaction is photosynthesis?

FIGURE 4.11
The blue flame on this gas stove is produced when natural gas burns.

## Lesson Summary

- A synthesis reaction occurs when two or more reactants combine to form a single product.
- In a decomposition reaction, one reactant breaks down into two or more products. This is the reverse of a synthesis reaction.
- Replacement reactions occur when elements switch places in compounds. In a single replacement reaction, one element takes the place of another in a single compound. In a double replacement reaction, two compounds exchange elements.
- A combustion reaction occurs when a substance reacts quickly with oxygen. Combustion is commonly called burning. Carbon dioxide, water, and heat and light are products of combustion.

## Lesson Review Questions

### Recall

1. Write an equation for the chemical reaction in which hydrogen reacts with oxygen to form water. What type of reaction is this?
2. Write an equation for the reverse of the reaction in question 1. What type of reaction is this?
3. Name the type of reaction represented by this general equation: $AB + CD \rightarrow AD + CB$
4. In the general equation in question 3, what do the individual letters represent?
5. What are the reactants and products in a combustion reaction?

### Apply Concepts

6. Apply lesson concepts to classify the following chemical reactions:
   a. $Zn + 2HCl \rightarrow H_2 + ZnCl_2$
   b. $2KClO_3 \rightarrow 2KCl + 3O_2$
   c. $2KI + Cl_2 \rightarrow 2KCl + I_2$

d. $AgNO_3 + KCl \rightarrow AgCl + KNO_3$

### Think Critically

7. Compare and contrast the four types of reactions described in this lesson. Include an example of each type of reaction.

## Points to Consider

Combustion reactions release energy. Some other types of reactions absorb energy. They need a continuous supply of energy to occur.

- Can you think of any chemical changes that might absorb energy?
- What might be different about reactions that need energy to keep going?

# 4.4 Chemical Reactions and Energy

## Lesson Objectives

- Describe endothermic reactions.
- Describe exothermic reactions.
- Relate the law of conservation of energy to chemical reactions.
- Define activation energy.
- Identify factors that affect the rates of chemical reactions.

## Lesson Vocabulary

- activation energy
- catalyst
- concentration
- endothermic reaction
- exothermic reaction
- law of conservation of energy
- reaction rate

## Introduction

All chemical reactions involve energy. Energy is needed to break bonds in reactants. These bonds may be very strong. Energy is released when new bonds form in the products. That's because the atoms now have a more stable arrangement of electrons. Which energy is greater: that needed for breaking bonds in reactants or that released by bonds forming in products? It depends on the type of reaction. When it comes to energy, chemical reactions may be endothermic or exothermic.

## Endothermic Reactions

In an **endothermic reaction**, it takes more energy to break bonds in the reactants than is released when new bonds form in the products. The word "endothermic" literally means "taking in heat." A constant input of energy, often in the form of heat, is needed in an endothermic reaction. Not enough energy is released when products form to break more bonds in the reactants. Additional energy is needed to keep the reaction going. The general equation for an endothermic reaction is:

$$\text{Reactants} + \text{Energy} \rightarrow \text{Products}$$

In many endothermic reactions, heat is absorbed from the surroundings. As a result, the temperature drops. The drop in temperature may be great enough to cause liquid products to freeze. That's what happens in the endothermic reaction at this URL: http://www.bbc.co.uk/schools/gcsebitesize/science/add_aqa/chemreac/energychangesrev1.sht ml .

One of the most important endothermic reactions is photosynthesis. In this reaction, plants synthesize glucose ($C_6H_{12}O_6$) from carbon dioxide ($CO_2$) and water ($H_2O$). They also release oxygen ($O_2$). The energy for photosynthesis comes from light (see **Figure** 4.12). Without light energy, photosynthesis cannot occur. The chemical equation for photosynthesis is:

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

FIGURE 4.12
Plants can get the energy they need for photosynthesis from either sunlight or artificial light.

## Exothermic Reactions

In an **exothermic reaction**, it takes less energy to break bonds in the reactants than is released when new bonds form in the products. The word "exothermic" literally means "turning out heat." Energy, often in the form of heat, is released as an exothermic reaction occurs. The general equation for an exothermic reaction is:

$$\text{Reactants} \rightarrow \text{Products} + \text{Energy}$$

If the energy is released as heat, an exothermic reaction results in a rise in temperature. That's what happens in the exothermic reaction at the URL below.

http://www.bbc.co.uk/schools/gcsebitesize/science/add_aqa/chemreac/energychangesrev1.shtml

Combustion reactions are examples of exothermic reactions. When substances burn, they usually give off energy as heat and light. Look at the big bonfire in **Figure** 4.13. You can see the light energy it is giving off. If you were standing near the fire, you would also feel its heat.

## Conservation of Energy

Whether a reaction absorbs energy or releases energy, there is no overall change in the amount of energy. Energy cannot be created or destroyed. This is the **law of conservation of energy**. Energy can change form —for example, from electricity to light —but the same amount of energy always remains.

FIGURE 4.13
The combustion of wood is an exothermic reaction that releases energy as heat and light.

If energy cannot be destroyed, what happens to the energy that is absorbed in an endothermic reaction? The energy is stored in the chemical bonds of the products. This form of energy is called chemical energy. In an endothermic reaction, the products have more stored chemical energy than the reactants. In an exothermic reaction, the opposite is true. The products have less stored chemical energy than the reactants. The excess energy in the reactants is released to the surroundings when the reaction occurs. The graphs in **Figure** 4.14 show the chemical energy of reactants and products in each type of reaction.

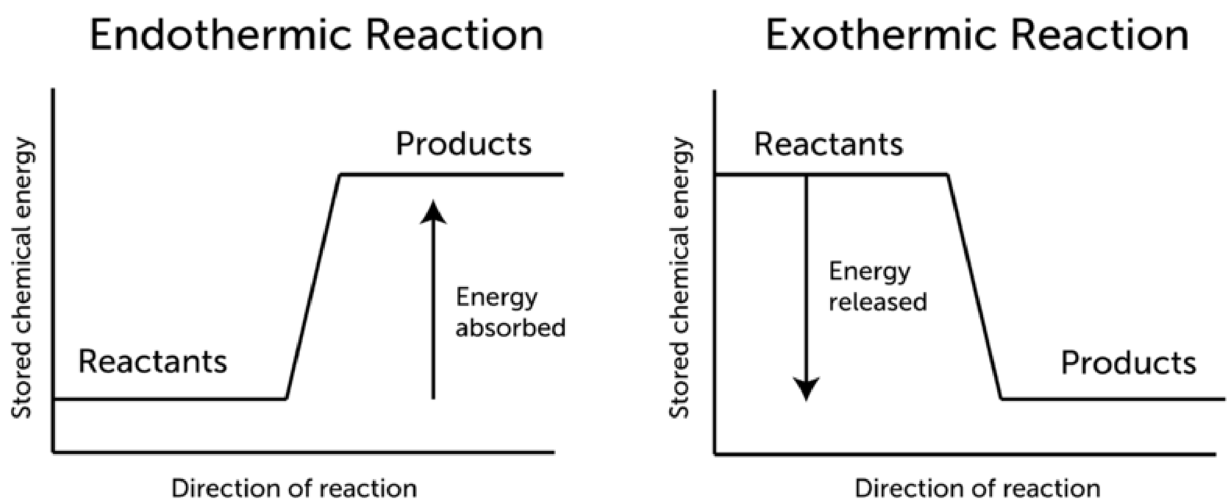


FIGURE 4.14
These graphs compare the energy changes in endothermic and exothermic reactions. What happens to the energy that is absorbed in an endothermic reaction?

## Activation Energy

All chemical reactions, even exothermic reactions, need a certain amount of energy to get started. This energy is called **activation energy**. For example, activation energy is needed to start a car. Turning the key causes a spark that activates the burning of gasoline in the engine. The combustion of gas won't occur without the spark of energy to begin the reaction.

Why is activation energy needed? A reaction won't occur unless atoms or molecules of reactants come together.

This happens only if the particles are moving, and movement takes energy. Often, reactants have to overcome forces that push them apart. This takes energy as well. Still more energy is needed to start breaking bonds in reactants. The graphs in **Figure** 4.15 show the changes in energy in endothermic and exothermic reactions. Both reactions need the same amount of activation energy in order to begin.

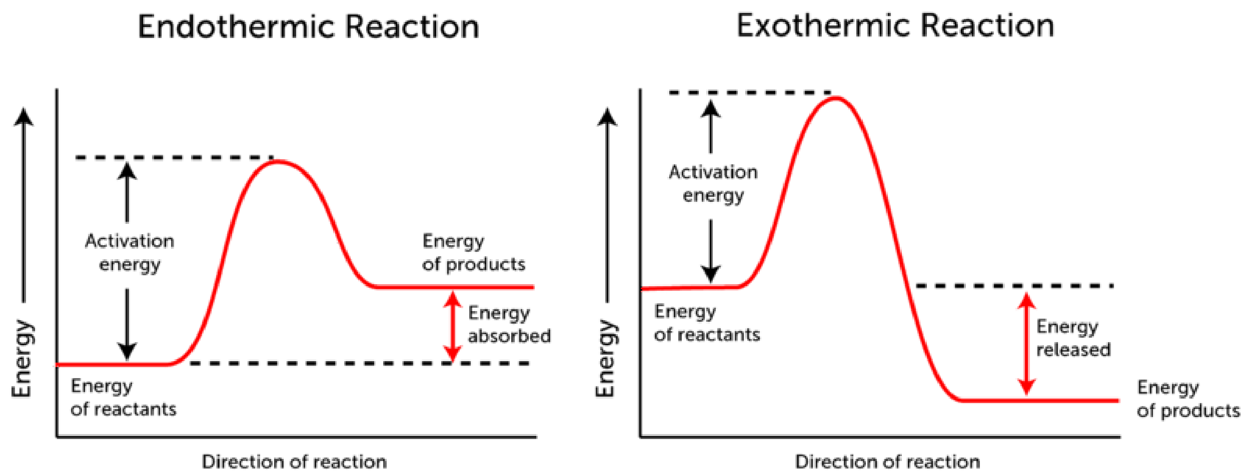


FIGURE 4.15
Even exothermic reactions need activation energy to get started.

You have probably used activation energy to start a chemical reaction. For example, if you've ever used a match to light a campfire, then you provided the activation energy needed to start a combustion reaction. Combustion is exothermic. Once a fire starts to burn, it releases enough energy to activate the next reaction, and the next, and so on. However, wood will not burst into flames on its own.

## Reaction Rate

Any factor that helps reactants come together so they can react lowers the amount of activation energy needed to start the reaction. If the activation energy is lowered, more reactant particles can react, and the reaction occurs more quickly. How fast a reaction occurs is called the **reaction rate**. Factors that affect the reaction rate include:

- temperature of reactants
- concentration of reactants
- surface area of reactants
- presence of catalysts

### Temperature of Reactants

When the temperature of reactants is higher, the rate of the reaction is faster. At higher temperatures, particles of reactants have more energy, so they move faster. They are more likely to bump into one another and to collide with greater force. For example, when you fry an egg, turning up the heat causes the egg to cook faster. The same principle explains why storing food in a cold refrigerator reduces the rate at which food spoils (see **Figure** 4.16). Both food frying and food spoiling are chemical reactions that happen faster at higher temperatures.

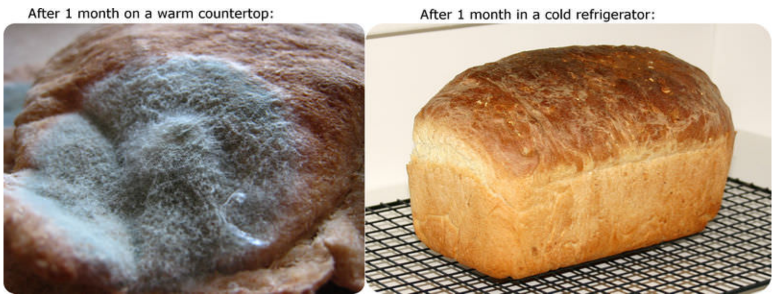


FIGURE 4.16
The chemical reactions that spoil food occur faster at higher temperatures.

### Concentration of Reactants

**Concentration** is the number of particles of a substance in a given volume. When the concentration of reactants is higher, the reaction rate is faster. At higher concentrations, particles of reactants are crowded closer together, so they are more likely to collide and react. Did you ever see a sign like the one in **Figure** 4.17? You might see it where someone is using a tank of pure oxygen for a breathing problem. The greater concentration of oxygen in the air makes combustion rapid if a fire starts burning.


FIGURE 4.17
It's dangerous to smoke or use open flames when oxygen is in use. Can you explain why?

### Surface Area of Reactants

When a solid substance is involved in a chemical reaction, only the matter at the surface of the solid is exposed to other reactants. If a solid has more surface area, more of it is exposed and able to react. Therefore, increasing the surface area of solid reactants increases the reaction rate. For example, crushing a solid into a powder exposes more of the substance to other reactants. This may greatly speed up the reaction. You can see another example in **Figure** 4.18. Iron rusts when it combines with oxygen in the air. The iron hammer head and iron nails will both rust eventually. Which will rust faster?

FIGURE 4.18
The nails have more surface area exposed to the air than the head of the hammer. How does this affect the rate at which they rust?

### Presence of a Catalyst

Some reactions need extra help to occur quickly. They need another substance, called a catalyst. A **catalyst** is a substance that increases the rate of a chemical reaction but is not changed or used up in the reaction. The catalyst can go on to catalyze many more reactions.

Catalysts are not reactants, but they help reactants come together so they can react. You can see one way this happens in the animation at the URL below. By helping reactants come together, a catalyst decreases the activation energy needed to start a chemical reaction. This speeds up the reaction.

http://www.saskschools.ca/curr_content/chem30/modules/module4/lesson5/explainingcatalysts.htm

Living things depend on catalysts to speed up many chemical reactions inside their cells. Catalysts in living things are called enzymes. Enzymes may be extremely effective. A reaction that takes a split second to occur with an enzyme might take billions of years without it!

## Lesson Summary

- In an endothermic reaction, it takes more energy to beak bonds in the reactants than is released when new bonds form in the products. Therefore, an endothermic reaction needs a constant input of energy to keep going.
- In an exothermic reaction, it takes less energy to break bonds in the reactants than is released when new bonds form in the products. Therefore, an exothermic reaction releases enough energy to keep going.
- In any chemical reaction, there is no overall change in the amount of energy. Energy cannot be created or destroyed. This is the law of conservation of energy.
- All chemical reactions, even exothermic reactions, need activation energy to get started. Activation energy is needed to bring reactants together so they can react.
- How fast a reaction occurs is called the reaction rate. Factors that affect the reaction rate include catalysts and the temperature, concentration, and surface area of reactants. A catalyst is a substance that increases the rate of a chemical reaction but is not changed or used up in the reaction.

## Lesson Review Questions

### Recall

1. What form of energy is needed for the endothermic reaction called photosynthesis?
2. What evidence shows that combustion reactions are exothermic?

3. What happens to the energy that is absorbed in an endothermic reaction?
4. In an exothermic reaction, which has more stored chemical energy: the reactants or the products?
5. Define activation energy.
6. List four factors that affect the rates of chemical reactions.

### Apply Concepts

7. Suppose you put a whole antacid tablet in one glass of water and a crushed antacid tablet in another glass containing the same amount of water. Both tablets would start reacting and producing bubbles of gas. Use lesson concepts to predict which tablet would stop producing bubbles first. Explain your prediction. Then, with the permission of an adult, do the activity. Do your results agree with your prediction?
8. Sketch a simple graph to show how energy changes in an exothermic reaction. Include activation energy in your graph.

### Think Critically

9. Compare and contrast endothermic and exothermic chemical reactions.

## Points to Consider

You read in this chapter that most fuels contain carbon. In the next chapter, "Chemistry of Carbon," you will learn much more about carbon.

- What do you already know about carbon?
- Based on carbon's position in the periodic table, predict how it reacts and the type of bonds it forms.

# 4.5 References


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