CHAPTER 3

MATTER AND ENERGY, ATOMS AND MOLECULES

We all learn the scientific method when we take a chemistry course, yet few of us think that we'll ever have to use it in our personal lives. This is the story of Joan Penn, who used the training and knowledge she received in a chemistry course to help her 11-year-old son Mike over a serious illness.

Our story begins on a Sunday afternoon in late September. Mike returned home from playing soccer complaining that he had a headache. By morning he felt better, so Mike went off to school. When he came home, he was coughing and running a low-grade fever. Two days later his mother took him to the doctor who diagnosed a viral infection. The doctor recommended that Mike rest until the virus passed.

After two days, his conditioned worsened. He was coughing and running a high fever. In addition, a rash appeared on his arms. A strep test proved positive and Mike was placed on an antibiotic. After two days on the medication, his condition continued to worsen, and he was now coughing severely and having bronchial spasms.

Joan was quite concerned. After all, the antibiotic Mike was on should have certainly wiped out any strep infection, yet Mike was still getting worse. She thought that perhaps her son was actually not suffering from a virus, but rather that he was having an allergic reaction to something environmental—for instance, something in the park, where he played every day. She decided to retrace Mike's schedule over the past two weeks to look for something—anything—that stood out from his regular schedule.

Joan checked Mike's school for any unusual incidents. Nothing. She checked the parks where he had played; she called the township to see if pesticides were being sprayed in the parks. All results proved negative. Joan decided to ask Mike if there was anything he had done over the past two weeks that he usually didn't do. One event stood out as slightly unusual.

On the day his initial symptoms appeared, Mike was at a friend's house. He and his friend had decided to build a tree house in the backyard, and they worked several hours cutting branches and putting the tree house together. Upon its completion, his friend's Mom bought them pizza to celebrate their fine work, which the boys gobbled down without having washed up.

Joan decided to find out more about the tree house and discovered that the boys built it using juniper branches. Joan called the county health department to find out if junipers could cause allergic reactions. The health officer stated that sticky sap from the pines *could* cause allergic reactions in humans, and he recited a list of symptoms that were identical to Mike's. The puzzle was solved! Mike was allergic to the juniper sap. When Joan later told her son about her findings, Mike recalled getting the sticky sap on his arms, where his rash had formed. He also said that he may have ingested some of the sap when he ate the pizza, which would explain his breathing problems.

Joan called Mike's doctor and told her about the information she had uncovered. Mike was immediately placed on Ventolin to relieve the bronchial spasms. He was also given a nebulizer to help him breathe more easily. Two weeks later Mike's fever was down, the rash had disappeared, and his breathing had returned to normal. Thanks to some good detective work by his mother, Mike was once again healthy.

LEARNING GOALS

After you've studied this chapter, you should be able to:

1. Explain what is meant by the scientific method.
2. Explain the Law of Conservation of Mass and Energy.
3. Explain the difference between physical and chemical properties.
4. Describe the difference between homogeneous and heterogeneous matter, between mixtures and compounds, and between compounds and elements.
5. Describe the difference between an atom and a molecule.
6. Explain the Law of Definite Composition (or Definite Proportions).
7. Explain the terms *atomic mass, formula mass*, and *molecular mass*.
8. Determine the formula or molecular mass of a compound when you are given the formula for the compound.

INTRODUCTION

This chapter is really the beginning of your study of chemistry. We start with discussions of the most elementary concepts, those of matter and energy. Then, in this chapter and later chapters, we build on and extend these concepts. As we do this, we shall be discussing the results of centuries of scientific research—the theories and laws of modern chemistry. These theories and laws are sometimes presented to students as though each resulted from a quick flash of insight on the part of some scientist. Actually they are the fruit of years—and sometimes decades or centuries—of hard work by many people.

3.1 The Scientific Method

Chemistry is an experimental science that is concerned with the behavior of matter. Much of the body of chemical knowledge consists of abstract concepts and ideas. Without application, these concepts and ideas would have little impact on society. Chemical principles are applied for the benefit of society through technology. Useful products are developed by the union of basic science and applied technology.

Over the past 200 years, science and technology have moved forward at a rapid pace. Ideas and applications of these ideas are developed through carefully planned experimentation, in which researchers adhere to what is called the **scientific method**. The scientific method is composed of a series of logical steps that allow researchers to approach a problem and try to come up with solutions in the most effective way possible. It is generally thought of as having four parts:

1. *Observation and classification*. Scientists begin their research by carefully observing natural phenomena. They carry out experiments, which are observations of natural events in a controlled setting. This allows results to be duplicated and rational conclusions to be reached. The data the scientists collect are analyzed, and the facts that emerge are classified.
2. *Generalization*. Once observations are made and experiments carried out, the researcher seeks regularities or patterns in the results that can lead to a generalization. If this generalization is basic and can be communicated in a concise statement or a mathematical equation, the statement or equation is called a *law*.
3. *Hypothesis*. Researchers try to find reasons and explanations for the generalizations, patterns, and regularities they discover. A hypothesis expresses a tentative explanation of a generalization that has been stated. Further experiments then test the validity of the hypothesis.

4. *Theory*. The new experiments are carried out to test the hypothesis. If they support it without exception, the hypothesis becomes a theory. A theory is a tested model that explains some basic phenomenon of nature. It cannot be proven to be absolutely correct. As further research is performed to test the theory, it may be modified or a better theory may be developed.

The scientific method represents a systematic means of doing research. There are times when discoveries are made by accident, but most knowledge has been gained via careful, planned experimentation. In your study of chemistry you will examine the knowledge and understanding that researchers using the scientific method have uncovered.

3.2 Matter and Energy

We begin with the two things that describe the entire universe: *matter* and *energy*. **Matter** is *anything that occupies space and has mass*. That includes trees, clothing, water, air, people, minerals, and many other things. Matter shows up in a wide variety of forms.

Energy is the *ability to perform work*. Like matter, energy is found in a number of forms. Heat is one form of energy, and light is another. There are also chemical, electrical, and mechanical forms of energy. And energy can change from one form to another. In fact, matter can also change form or change into energy, and energy can change into matter, but not easily.

3.3 Law of Conservation of Mass and Energy

The **Law of Conservation of Mass** tells us that when a chemical change takes place, no detectable difference in the mass of the substances is observed. In other words, mass is neither created nor destroyed in an ordinary chemical reaction. This law has been tested by extensive experimentation in the laboratory, and the work of the brilliant French chemist-physicist Antoine Lavoisier provides evidence for this conclusion. Lavoisier performed many experiments involving matter. In one instance he heated a measured amount of tin and found that part of it changed to a powder. He also found that the *product* (powder plus tin) weighed *more* than the original piece of tin. To find out more about the added weight, he heated metals in sealed jars, which, of course, contained air. He measured the mass of his starting materials (*reactants*), and when the reaction concluded and the metal no longer changed to powder, he measured the mass of the products. In every such reaction, the mass of the reactants (oxygen from the air in the jar plus the original metal) equaled the mass of the products (the remaining metal plus the powder). Today we know that the reaction actually stopped when all of the oxygen in the sealed jar combined with the metal to form the powder. Lavoisier concluded that when a chemical change occurs, *matter is neither created nor destroyed, it just changes from one form to another* (Figure 3.1), which is a statement of the Law of Conservation of Mass.

An experimenter puts a test tube containing a lead nitrate solution into a flask containing a potassium chromate solution. The experimenter weighs the flask and contents, then turns the flask upside down to mix the two solutions. A chemical reaction takes place, producing a yellow solid. The experimenter weighs the flask and contents again and finds no change in mass.

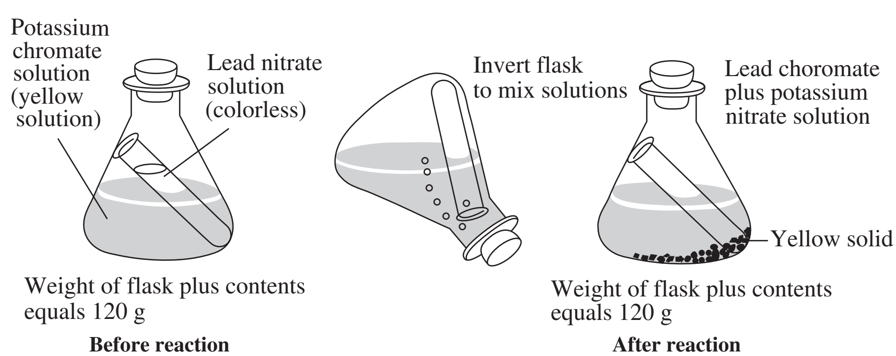


FIGURE 3.1 An experiment like Lavoisier's.

Whenever a chemical change occurs, it is accompanied by an energy transformation. In the 1840s, more than half a century after Lavoisier, three scientists—the Englishman James Joule and the Germans Julius von Mayer and Hermann von Helmholtz—performed a number of experiments in which energy transformations were studied. They provided experimental evidence that led to the discovery of the **Law of Conservation of Energy**. The law tells us that *in any chemical or physical change, energy is neither created nor destroyed, it is simply converted from one form to another*.

An auto engine provides a good example of how one form of energy is converted to a different form. *Electrical* energy from the battery generates a spark that contains *heat* energy. The heat ignites the gasoline-air mixture, which explodes, transforming chemical energy into heat and *mechanical* energy. The mechanical energy causes the pistons to rise and fall, rotating the engine crankshaft and moving the car.

At the same time, in the same engine, matter is changing from one form to another. When the gasoline explodes and burns, it combines with oxygen in the cylinders to form carbon dioxide and water vapor. (Unfortunately, carbon monoxide and other dangerous gases may also be formed. This is one of the major causes of air pollution.

To appreciate the significance of these facts, think of the universe as a giant chemical reactor or system. At any given time there are certain amounts of matter and energy present, and the matter has a certain mass. Matter is always changing from one form to another, and so is energy. Besides that, matter is changing to energy and energy to matter. But *the sum of all the matter (or mass) and energy in the universe always remains the same*. This repeated observation is called the **Law of Conservation of Mass and Energy**.

3.4 Potential Energy and Kinetic Energy

Which do you think has more energy, a metal cylinder held 1 foot above the ground or an identical cylinder held 5 feet above the ground? If you dropped them on your foot, you would know immediately that the cylinder with more energy was the one that was 5 feet above the ground. But where does this energy come from?

Work had to be done to raise the two cylinders to their respective heights—to draw them up against the pull of gravity. And energy was needed to do that work. The energy used to lift each cylinder was "stored" in each cylinder. The higher the cylinder was lifted, the more energy was stored in it—due to its position. *Energy that is stored in an object by virtue of its position* is called **potential energy**.

If we drop the cylinders, they fall toward the ground. As they do so, they lose potential energy because they lose height. But now they are moving; their potential energy is converted to "energy of motion." The more potential energy they lose, the more energy of motion they acquire. *The energy that an object possesses by virtue of its motion* is called **kinetic energy**. The conversion of potential energy to kinetic energy is a very common phenomenon. It is observed in a wide variety of processes, from downhill skiing to the generation of hydroelectric power.

3.5 The States of Matter

Matter may exist in any of the three physical states: solid, liquid, and gas.

A **solid** has a definite shape and volume that it tends to maintain under normal conditions. The particles composing a solid stick rigidly to one another. Solids most commonly occur in the **crystalline** form, which means they have a fixed, regularly repeating, symmetrical internal structure. Diamonds, salt, and quartz are examples of crystalline solids. A few solids, such as glass and paraffin, do not have a well-defined crystalline structure, although they do have a definite shape and volume. Such solids are called **amorphous solids**, which means they have no definite internal structure or form.

A **liquid** has a definite volume but does not have its own shape since it takes the shape of the container in which it is placed. Its particles cohere firmly, but not rigidly, so the particles of a liquid have a great deal of mobility while maintaining close contact with one another.

A **gas** has no fixed shape or volume and eventually spreads out to fill its container. As the gas particles move about they collide with the walls of their container causing *pressure*, which is a force exerted over an area. Gas particles move independently of one another. Compared with those of a liquid or solid, gas particles are quite far apart. Unlike solids and liquids, which cannot be compressed very much at all, gases can be both compressed and expanded.

Often referred to as the fourth state of matter, **plasma** is *a form of matter composed of electrically charged atomic particles*. Many objects found in the earth's outer atmosphere, as well as many celestial bodies found in space (such as the sun and stars), consist of plasma. A plasma can be created by heating a gas to extremely high temperatures or by passing a current through it. A plasma responds to a magnetic field and conducts electricity well.

3.6 Physical and Chemical Properties

Matter—whether it is solid, liquid, or gas—possesses two kinds of properties: physical and chemical. These unique properties separate one substance from another and ensure that no two substances are alike in every way. The **physical properties** are those that can be observed or measured without changing the chemical composition of the substance. These properties include state, color, odor, taste, hardness, boiling point, and melting point. A **physical change** is one that alters at least one of the physical properties of the substance without changing its chemical composition. Some examples of physical change are (1) altering the physical state of matter, such as what occurs when an ice cube is melted; (2) dissolving or mixing substances together, such as what happens when we make coffee or hot cocoa; and (3) altering the size or shape of matter, such as what happens when we grind or chop something.

Chemical properties stem from the ability of a substance to react or change to a new substance that has different properties. This often occurs in the presence of another substance. For example, iron reacts with oxygen to produce iron (III) oxide (rust). This is an example of a **chemical change**. The chemical properties can be observed or measured when a substance undergoes chemical change. The rusting of iron is an example of a chemical property of iron. When we pass an electric current through water, it decomposes to form hydrogen gas and oxygen gas. This reaction is an example of a chemical property of water.

Sometimes it is difficult to differentiate a chemical change from a physical change. In fact, physical changes almost always accompany chemical changes. Some of the signs of physical change that tell us that a chemical change has occurred include the presence of a large amount of heat or light, the presence of a flame, the formation of gas bubbles, a change in color or odor, or the formation of a solid material that settles out of a solution.

3.7 Mixtures and Pure Substances

Since matter consists of all the material things that compose the universe, many distinctly different types of matter are known. Matter that has a definite and *fixed composition* is called a **pure substance**, which is a substance that cannot be separated into any other form of matter by physical change. Some of the pure substances that you are familiar with are helium, oxygen, table salt, water, gold, and silver. Two or more pure substances can be combined to form a **mixture**, whose *composition can be varied*. The substances in a mixture can be separated by physical means; we can separate the substances without chemical change.

Matter can also be classified as heterogeneous or homogeneous (Figure 3.2). **Homogeneous matter** has the same parts with the same properties throughout, and **heterogeneous matter** is made up of different parts

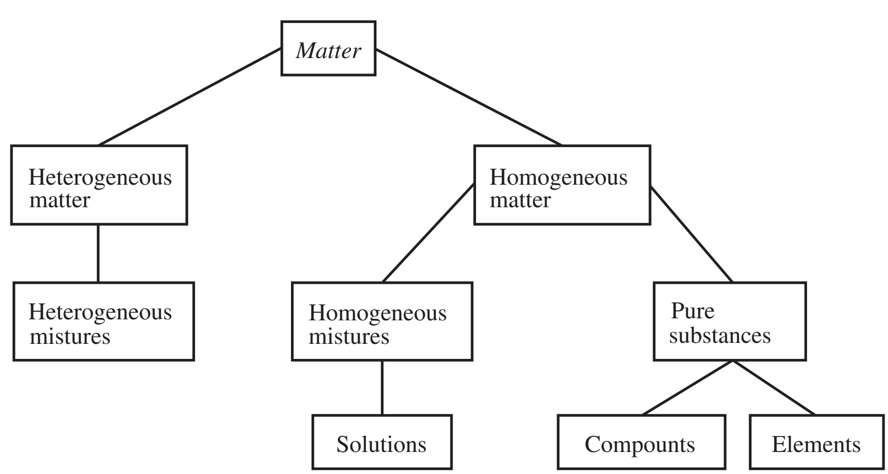


FIGURE 3.2 Classification of matter.

FIGURE 3.3 Separating sand from saltwater.

with different properties. A combination of salt and pepper is an example of heterogeneous matter, whereas a teaspoonful of sugar is an example of homogeneous matter. Another example of homogeneous matter is a teaspoonful of salt dissolved in a glass of water. We call this a **homogeneous mixture** because it is a uniform blend of two or more substances, and its proportion can be varied. In a homogeneous mixture every part is exactly like every other part. The salt can be separated from the water by physical means. Seawater and air are also examples of homogeneous mixtures.

We know that there are two types of homogeneous matter: pure substances and homogeneous mixtures. According to this classification scheme, matter can be broken down even further. Let's look at both homogeneous mixtures and heterogeneous mixtures. In a **heterogeneous mixture** different parts have different properties. A salt-sand mixture is heterogeneous because it is composed of two substances, *each of which retains its own unique properties*. It does not have the same composition or properties throughout, and its composition can be varied. A salt-sand mixture can be separated by physical means. If the mixture is placed in water, the salt will dissolve. The sand can be filtered out, and the salt can be recovered by heating the saltwater until the water evaporates (Figure 3.3).

We call any part of a system with uniform composition and properties a **phase**. A **system** is the body of matter being studied. A heterogeneous mixture is composed of two or more phases separated by physical boundaries. Additional examples of heterogeneous mixtures are oil and water (two liquids) and a tossed salad (several solids). It is important to note that although a pure substance is always homogeneous in composition, it may actually exist in more than one phase in a heterogeneous system. Think of a glass of ice water. This is a two-phase system composed of water in the solid phase and water in the liquid phase. In each phase the water is homogeneous in composition, but since two phases are present, the system is heterogeneous.

3.8 Solutions

Solutions are homogeneous mixtures. That is, they are uniform in composition. Every part of a solution is exactly like every other part. The salt-and-water mixture described in Section 3.7 is a solution. Even if we add more water to the solution, it will still be homogeneous, because the salt particles will continue to be distributed evenly throughout the solution. We would get the same result if we added more salt to the solution. However, we couldn't do this indefinitely. Homogeneity would end when the solution reached *saturation* (the point at which no more salt could dissolve in the limited amount of water). We will discuss this further in Chapter 12.

3.9 Elements

An **element** is *a pure substance that cannot be broken down into simpler substances, with different properties, by physical or chemical means*. The elements are the basic building blocks of all matter. There are now 118 known elements. Each has its own unique set of physical and chemical properties. (The elements are tabulated on the inside front cover of this book, along with their *chemical symbols*—a shorthand notation for their names. Some common elements are listed in Table 3.1.) The elements can be classified into three types: **metals**, **nonmetals**, and **metalloids**.

Examples of metallic elements are sodium (which has the symbol Na), calcium (Ca), iron (Fe), cobalt (Co), and silver (Ag). These elements are all classified as metals because they have certain properties in common. They have luster (in other words, they are shiny), they conduct electricity well, they conduct heat well, and they are malleable (can be pounded into sheets) and ductile (can be drawn into wires).

Some examples of nonmetals are chlorine, which has the symbol Cl (note that the second letter of this symbol is a lower-case "el" and not the numeral *one*), oxygen (O), carbon (C), and sulfur (S). These elements are classified as nonmetals because they have certain properties in common. They don't shine, they don't conduct electricity well, they don't conduct heat well, and they are neither malleable nor ductile.

The metalloids have some properties like those of metals and other properties like those of nonmetals. Some examples are arsenic (As), germanium (Ge), and silicon (Si). These particular metalloids are used in manufacturing transistors and other semiconductor devices (Table 3.1).

3.10 Atoms

Suppose we had a chunk of some element, say gold, and were able to divide it again and again, into smaller and smaller chunks. Eventually we could get a particle that could not be divided any further without losing its identity. This particle would be an atom of gold. An **atom** is *the smallest particle of an element that enters into chemical reactions*.

The atom is the ultimate particle that makes up the elements. Gold is composed of gold atoms, iron of iron atoms, and neon of neon atoms. These atoms are so small that billions of them are needed to make a speck large enough to be seen with a microscope. In 1970, Albert Crewe and his staff at the University of Chicago's Enrico Fermi Institute took the first black-and-white pictures of single atoms, using a special type of electron microscope. In late 1978, Crewe and his staff took the first time-lapse moving pictures of individual uranium atoms.

3.11 Compounds

A **compound** is *a pure substance that is made up of two or more elements chemically combined in a definite proportion by mass*. Unlike mixtures, compounds have a definite composition. Water, for instance, is made up of hydrogen and oxygen in the ratio of 11.1% hydrogen to 88.9% oxygen by mass. No matter what the source

TABLE 3.1 Names and Symbols of Some Common Elements

ELEMENT	SYMBOL	ELEMENT	SYMBOL
Aluminum	Al	Iodine	I
Bromine	Br	Magnesium	Mg
Calcium	Ca	Nickel	Ni
Carbon	C	Nitrogen	N
Chlorine	Cl	Oxygen	O
Chromium	Cr	Phosphorus	P
Fluorine	F	Silicon	Si
Helium	He	Sulfur	S
Hydrogen	H	Zinc	Zn

of the water, it is always composed of hydrogen and oxygen in this ratio. This idea, that *every compound is composed of elements in a certain fixed proportion*, is called the **Law of Definite Composition** (or the Law of Definite Proportions). It was first proposed by the French chemist Joseph Proust in about 1800.

The properties of a compound need not be similar to the properties of the elements that compose it. For example, water is a liquid, whereas hydrogen and oxygen are both gases. When two or more elements form a compound, they truly form a new substance.

Compounds can be broken apart into elements only by chemical means—unlike mixtures, which can be separated by physical means. More than thirty million compounds have been reported to date, and millions more may be discovered. Some compounds we are all familiar with are sodium chloride (table salt), which is composed of the elements sodium and chlorine, and sucrose (cane sugar), which is composed of the elements carbon, hydrogen, and oxygen.

3.12 Molecules

We have discussed what happens when a chunk of an element is continually divided: We eventually get down to a single atom. What happens when we keep dividing a chunk of a compound? Suppose we do so with the compound sugar. As we continue to divide a sugar grain, we eventually reach a small particle that can't be divided any further without losing the physical and chemical properties of sugar. This ultimate particle of a compound, *the smallest particle that retains the properties of the compound*, is called a **molecule**. Like atoms, molecules are extremely small, but with the aid of an electron microscope we can observe some of the very large and more complex molecules. Molecules are uncharged particles. That is, they carry neither a positive nor a negative electrical charge.

Molecules, as you might have guessed, are made up of two or more atoms. They may be composed of different kinds of atoms (for instance, water contains hydrogen and oxygen) or the same kind of atoms (for instance, a molecule of chlorine gas contains two atoms of chlorine). (See Figure 3.4.) *The Law of Definite Composition states that the atoms in a compound are combined in definite proportions by mass*. We can see in Figure 3.4 that they are also combined in definite proportions by number. For example, water molecules always contain two hydrogen atoms for every oxygen atom.

3.13 Molecular Versus Ionic Compounds

As we just learned, a molecule is the smallest *uncharged* part of a compound formed by the chemical combination of two or more atoms. Such compounds are known as *molecular compounds*, and we usually say that such compounds are composed of molecules. Water is a molecular compound composed of water molecules. Many molecular compounds are composed of atoms of nonmetallic elements that are chemically combined.

However, there are many compounds composed of oppositely charged *ions*. An ion is *a positively or negatively charged atom or group of atoms*. Compounds composed of ions are known as *ionic compounds*. These compounds are held together by attractive forces between the positive and negative ions that compose the compound. Ordinary table salt, sodium chloride, is such a compound. For these compounds, it is more proper to talk about a **formula unit** of the compound, rather than a molecule, as being *the smallest part of an ionic compound that retains the properties of the compound*. Ionic compounds are composed of metallic and nonmetallic elements. (We'll have much more to say about ions and ionic compounds when we discuss chemical bonding in Chapter 7.)

FIGURE 3.4 Molecules of water and chlorine.

3.14 Symbols and Formulas of Elements and Compounds

We have already mentioned the chemical symbols—the shorthand for the names of the elements. In some cases, as in the symbol O (capital "oh") for oxygen, the chemical symbol is the first letter of the element's name, capitalized. Often, though, the symbol for an element contains two letters. In these cases only the first letter is capitalized; the second letter is never capitalized. For instance, the symbol for neon is Ne and the symbol for cobalt is Co. (Be careful of this: CO does not represent cobalt, but a combination of the elements carbon and oxygen, which is a compound.) Some symbols come from the Latin names of the elements: iron is Fe, from the Latin *ferrum*, and lead is Pb, from the Latin *plumbum*. (See Table 3.2.)

Because compounds are composed of elements, we can use the chemical symbols as a shorthand for compounds too. We use the symbols to write the **chemical formula**, which shows the elements that compose the compound. For example, sodium chloride (table salt), an ionic compound, contains one atom of sodium (Na) and one atom of chlorine (Cl). The formula unit for sodium chloride is NaCl. Water, a molecular compound, is another example. The water molecule contains two atoms of hydrogen (H) and one atom of oxygen, so we write it as H_2O. The number 2 in the formula indicates that there are two atoms of hydrogen in the molecule; note that it is written as a *subscript*. A molecule of ethyl alcohol contains two atoms of carbon, six of hydrogen, and one of oxygen. It is written as follows:

$$C_2H_6O \longleftarrow \text{One oxygen atom}$$

Two carbon atoms   Six hydrogen atoms

EXAMPLE 3.1

State the number of atoms of each element in a molecule or formula unit of the following compounds: (a) $C_6H_{12}O_6$ (b) $Ca(OH)_2$ (c) $C_3H_6O_2$ (d) $Al_2(SO_4)_3$

Solution

UNDERSTAND THE PROBLEM

We ask, "What does the subscript mean?" Subscripts tell us the number of atoms in a molecule or formula unit of the substance.

DEVISE A PLAN

Our plan will be to look at the subscripts noted in each case. We can then determine the number of atoms of each element in a molecule or formula unit of the compound.

TABLE 3.2 Elements with Symbols Derived from Latin Names

ELEMENT	SYMBOL	LATIN NAME
Copper	Cu	Cuprum
Gold	Au	Aurum
Iron	Fe	Ferrum
Lead	Pb	Plumbum
Mercury	Hg	Hydrargentum
Potassium	K	Kalium
Silver	Ag	Argentum
Sodium	Na	Natrium
Tin	Sn	Stannum

CARRY OUT THE PLAN

(a) There are 6 atoms of C, 12 atoms of H, and 6 atoms of O.

(b) There is 1 atom of Ca, 2 atoms of O, and 2 atoms of H. (In this case, the subscript 2 means multiply everything inside the parentheses by 2.)

(c) There are 3 atoms of C, 6 atoms of H, and 2 atoms of O.

(d) There are 2 atoms of Al, 3 atoms of S, and 12 atoms of O. (In this case, the subscript 3 means multiply everything inside the parentheses by 3.)

LOOK BACK

Recheck to be sure that you have followed the steps correctly.

Practice Exercise 3.1

State the number of atoms of each element in a molecule or formula unit of the following compounds: (a) C_2H_7N (b) $(NH_4)2SO_4$

3.15 Atomic Mass

Suppose we want to find the relative masses of the atoms of the various elements. Suppose also that we have a double-pan balance that can weigh a single atom. We begin by assigning an arbitrary mass to one of the elements. Let's say we decide to assign a mass of 1 unit to hydrogen. (This is what chemists did originally, because hydrogen was known to be the lightest element even before the relative atomic masses were determined.)

Then, to find the relative mass of, say, carbon, we place an atom of carbon on one pan. We place hydrogen atoms on the other pan, one by one, until the pans are exactly balanced. We find that it takes 12 hydrogen atoms to balance 1 carbon atom, so we assign a relative atomic mass of 12 to carbon. In the same way for an oxygen atom, we find that it takes 16 hydrogen atoms to balance 1 oxygen atom. So we assign a relative mass of 16 to oxygen. By doing this experiment for all the other elements, we can determine all the masses relative to hydrogen (which is assigned mass 1).

Unfortunately this kind of balance has never existed, and more elaborate means had to be developed to find the relative atomic masses of the elements. But the logic we used here is the same as that used by the many scientists who determined the relative masses. (The atomic mass scale has been revised. The present scale is based on a particular type of carbon atom, called carbon-12. This carbon is assigned the value of 12 atomic mass units, amu.) The periodic table (inside the front cover) lists the relative atomic masses of the elements below their symbols. From now on, instead of referring to "relative atomic mass," we will use the simpler term **atomic mass**.

EXAMPLE 3.2

Using the periodic table (inside the front cover), look up the atomic masses of the following elements. (For this exercise, round all atomic masses to one decimal place.) (a) I (b) Ba (c) As (d) S

Solution

Look up the atomic mass of each element in the periodic table. Remember, the atomic mass is the number below the symbol of the element.

(a) I is 126.9

(b) Ba is 137.3

(c) As is 74.9

(d) S is 32.1

Practice Exercise 3.2

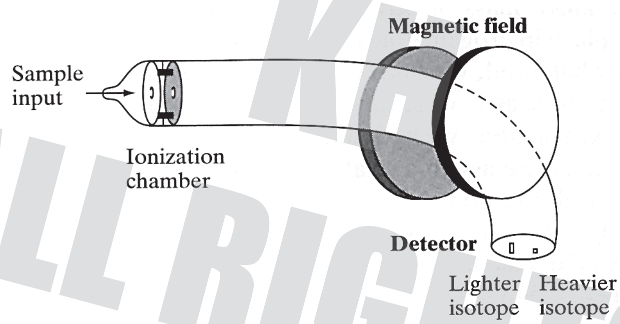
Using the periodic table (inside the front cover), look up the atomic masses of the following elements. (For this exercise, round all atomic masses to one decimal place.) (a) La (b) Fe (c) Ar (d) Sn

3.16 Formula Mass and Molecular Mass

The **molecular mass** of a compound is *the sum of the atomic masses of all the atoms that make up a molecule of the compound*. The term *molecular mass* is applied to compounds that exist as molecules. For example, the molecule P_2O_5 has two phosphorus atoms and five oxygen atoms. The atomic mass of each P (to one decimal place) is 31.0, and the atomic mass of each O (to one decimal place) is 16.0. Therefore the molecular mass of the compound is

$$(2 \times 31.0) + (5 \times 16.0) = 62.0 + 80.0 = 142.0$$

(Check this calculation and see whether you get the same answer. If not, you may have forgotten that in an equation like this, multiplication and division are done *before* addition and subtraction.)



The masses of individual atoms are determined with a mass spectrometer. Electrons are removed from atoms (or molecules), and the resultant ions are accelerated through a magnetic field. The amount of bending in the path of the ions is related to the mass and the charge of the ions.

The **formula mass** of a compound is *the sum of the atomic masses of all the ions that make up a formula unit of the compound*. The term *formula mass* is applied to compounds that are written as formula units and exist mostly as ions (charged atoms or groups of atoms). For example, the formula unit of aluminum oxide is $A1_2O_3$. This means that a formula unit of aluminum oxide has two aluminum atoms (actually aluminum ions) and three oxygen atoms (actually oxide ions). The atomic mass of aluminum (to one decimal place) is 27.0, and the atomic mass of oxygen (to one decimal place) is 16.0. Therefore the formula mass of $A1_2O_3$ is

$$(2 \times 27.0) + (3 \times 16.0) = 54.0 + 48.0 = 102.0$$

Now let's try to determine the formula and molecular masses of some additional compounds.

EXAMPLE 3.3

Find the molecular or formula masses (to one decimal place) of the following compounds: (a) H_2O (b) NaCl (c) $Ca(OH)_2$ (d) $Zn_3(PO_4)_2$

Note: In a chemical formula, parentheses followed by a subscript mean that everything inside the parentheses is multiplied by the subscript. For example, in one formula unit of $Ca(OH)_2$, there is one Ca atom plus two O atoms and two H atoms.

Solution

We must find the atomic mass of each element in the periodic table and then add the masses of all the atoms in each compound.

(a) The atomic mass of H is 1.0, and the atomic mass of O is 16.0.

$$\begin{aligned}\text{Molecular mass of } H_2O &= (2 \times 1.0) + (1 \times 16.0)\\ &= 2.0 + 16.0 = 18.0\end{aligned}$$

(b) The atomic mass of Na is 23.0, and the atomic mass of Cl is 35.5.

$$\begin{aligned}\text{Formula mass of } NaCl &= (1 \times 23.0) + (1 \times 35.5)\\ &= 23.0 + 35.5 = 58.5\end{aligned}$$

(c) The atomic mass of Ca is 40.1, the atomic mass of O is 16.0, and the atomic mass of H is 1.0.

$$\begin{aligned}\text{Formula mass of } Ca(OH)_2 &= (1 \times 40.1) + (2 \times 16.0) + (2 \times 1.0)\\ &= 40.1 + 32.0 + 2.0 = 74.1\end{aligned}$$

(d) The atomic mass of Zn is 65.4, the atomic mass of P is 31.0, and the atomic mass of O is 16.0.

$$\begin{aligned}\text{Formula mass of } Zn_3(PO_4)_2 &= (3 \times 65.4) + (2 \times 31.0) + (8 \times 16.0)\\ &= 196.2 + 62.0 + 128.0 = 386.2\end{aligned}$$

Practice Exercise 3.3

Find the molecular or formula masses (to one decimal place) of the following compounds: (a) Na_2CO_3 (b) $CoCl_2$ (c) $C1_2O$ (d) N_2O_4

SUMMARY

The body of knowledge called science has been developed through the scientific method: observation and classification of data, generalization of observations, and testing of generalizations. Of importance in all the sciences is the idea that the universe is made up of only matter and energy. Matter and energy can be neither created nor destroyed, but they can be changed to other forms. And matter can be transformed into energy, and vice versa. Potential energy is energy that is stored in a body because of its position. Kinetic energy is energy that is due to motion.

Matter may exist in any of three states—solid, liquid, or gas—and may be either heterogeneous (nonuniform) or homogeneous (uniform). Mixtures are combinations of two or more kinds of matter, each retaining its own chemical and physical properties. Mixtures too may be either homogeneous or heterogeneous, and solutions are homogeneous mixtures. Elements are the basic building blocks of matter. The 109 known elements are, in turn, made up of atoms. A compound is a substance that is made up of two or more elements chemically combined in definite proportions by mass. Molecules are the smallest particles that retain the properties of a compound, and atoms are the smallest particles that enter into chemical reactions.

Each element has its own symbol, and every compound has its own chemical formula. Each element has a unique atomic mass. The atomic mass of an element is found in the periodic table. The formula mass or molecular mass of a compound is the sum of the atomic masses in a formula unit or molecule of the compound.

KEY TERMS

amorphous solid **(3.5)**
atom **(3.10)**
atomic mass **(3.15)**
chemical change **(3.6)**
chemical formula **(3.14)**
chemical property **(3.6)**
chemical symbols **(3.9)**
compound **(3.11)**
crystalline **(3.5)**
element **(3.9)**
energy **(3.2)**
formula mass **(3.16)**
formula unit **(3.13)**
gas **(3.5)**
heterogeneous matter **(3.7)**
heterogeneous mixture **(3.7)**
homogeneous matter **(3.7)**
homogeneous mixture **(3.7)**
kinetic energy **(3.4)**
Law of Conservation of Energy **(3.3)**
Law of Conservation of Mass **(3.3)**
Law of Conservation of Mass and Energy **(3.3)**

Law of Definite Composition **(3.11)**
liquid **(3.5)**
matter **(3.2)**
metal **(3.9)**
metalloid **(3.9)**
mixture **(3.7)**
molecular mass **(3.16)**
molecule **(3.12)**
nonmetal **(3.9)**
phase **(3.7)**
physical change **(3.6)**
physical property **(3.6)**
plasma **(3.5)**
potential energy **(3.4)**
pure substance **(3.7)**
scientific method **(3.1)**
solid **(3.5)**
solution **(3.8)**
system **(3.7)**

SELF-TEST EXERCISES

LEARNING GOAL 1

The Scientific Method

1. (a) Explain the difference between a theory and a hypothesis.
   (b) Explain the difference between a theory and a scientific law.
2. Suppose that you are a researcher and you believe that you have found a vaccine that can prevent AIDS. How would you use the scientific method to determine whether the vaccine is effective?

LEARNING GOAL 2

Law of Conservation of Mass and Energy

◀ 3. According to the Law of Conservation of Energy, in any chemical or physical change, energy is neither created nor destroyed. Then why are we always worried about running out of energy in the future?
4. What is the significance of the Law of Conservation of Mass and Energy in terms of your study of chemistry?

LEARNING GOALS 3 & 4

Physical and Chemical Changes/Types of Matter

5. Match each word on the left with its definition on the right.

(a) Homogeneous	1. The basic building block of matter
(b) Heterogeneous	2. The word used to describe matter that is uniform throughout
(c) Mixture	3. A type of matter in which each part retains its own properties
(d) Compound	4. A chemical combination of two or more elements
(e) Element	5. The word used to describe matter that is not uniform throughout

◀ 6. (a) Distinguish among an element, a compound, and a mixture.
   (b) What type of matter is uniform throughout?
   (c) What type of matter is not uniform throughout?

◀ 7. State whether each of the following processes involves physical or chemical changes:

   (a) Shredding paper (b) Burning paper (c) Cooking an egg (d) Mixing egg whites with egg yolk (e) Digesting food (f) Toasting bread

8. Determine whether each of the following processes involves chemical or physical processes:

   (a) Ice melts.
   (b) Sugar dissolves in water.

(c) Milk sours.
(d) Eggs become rotten.
(e) Water boils.
(f) An egg is hard-cooked.

LEARNING GOAL 5

Difference Between Atom and Molecule

9. Describe the difference between an atom and a molecule.

◀ 10. (a) What is the smallest particle of matter that can enter into a chemical combination?
(b) What is the smallest uncharged individual unit of a compound that is composed of two or more atoms?

11. Classify each of the following elements as metal, metalloid, or nonmetal:

(a) Ba (b) Si (c) O (d) Hg (e) Ge (f) In (g) U

12. Classify each of the following elements as metal, metalloid, or nonmetal:

(a) Mn (b) Nd (c) Al (d) At (e) Pt (f) Cl (g) Ra

13. Explain the difference between Co and CO.
14. Explain the difference between Si and SI.

LEARNING GOAL 6

Law of Definite Composition

◀ *15*. How does the following information obtained from several experiments confirm the Law of Definite Composition?

Experiment 1: $1\overline{00}$ g of water are decomposed by electrolysis into its elements: 88.9 g of oxygen gas and 11.1 g of hydrogen gas are obtained.
Experiment 2: 25.0 g of water are decomposed by electrolysis into its elements: 22.2 g of oxygen gas and 2.8 g of hydrogen are obtained.
Experiment 3: $5\overline{00}$ g of water are decomposed by electrolysis into its elements: 444.5 g of oxygen gas and 55.5 g of hydrogen are obtained.

16. Give an example of the Law of Definite Composition (Proportions).
17. State the number of atoms of each element in a molecule or formula unit of the following compounds:

(a) $C_{12}H_{22}O_{11}$ (b) K_2CrO_4 (c) $H_8N_2O_3S_2$ (d) $Zn(NO_3)_2$

◀ 18. State the number of atoms of each element in a molecule or formula unit of the following compounds:

(a) H_2SeO_4 (b) $C_{21}H_{27}FO_6$ (c) $(NH_4)_3PO_4$ (d) $Fe_3(AsO_4)_2$

LEARNING GOAL 7

Atomic Mass, Formula, Mass, and Molecular Mass

19. Using examples, explain the difference among atomic mass, formula mass, and molecular mass.
20. State which term—atomic mass, formula mass, or molecular mass—is best suited to describe each of the following substances:

(a) $C_6H_{12}O_6$ (b) NaCl (c) Fe (d) CO_2 (e) H_2O (f) $A1_2O_3$ (g) Ca (h) $Ca(OH)_2$

◀**21*. If in the periodic table oxygen were assigned an atomic mass of 1, what would be the atomic mass of sulfur?

◀*22. If in the periodic table neon were assigned an atomic mass of 1, what would be the atomic mass of bromine?

23. Using the periodic table (inside front cover), look up the atomic masses of the following elements. (For this exercise, round all atomic masses to one decimal place.)

(a) Rb (b) Cr (c) U (d) Se (e) As

24. Using the periodic table (inside front cover), look up the atomic masses of the following elements. (For this exercise, round all atomic masses to one decimal place.)

(a) S (b) N (c) Li (d) Cs (e) Au

LEARNING GOAL 8

Formula or Molecular Mass of a Compound from the Formula

◀ ***25***. Determine the molecular or formula mass of each of the following compounds. (For this exercise, round all atomic masses to one decimal place.)

(a) FeO (b) Fe_2O_3 (c) CuI_2 (d) Na_3PO_4 (e) $Mg(OH)_2$ (f) $NiBr_2$ (g) $Hg_3(PO_4)2$
(h) $(NH_4)_2CO_3$

26. Determine the molecular or formula mass of each of the following compounds. (For this exercise, round all atomic masses to one decimal place.)

(a) H_2O (b) H_2SO_4 (c) $NaCl$ (d) $Ca_3(PO_4)_2$ (e) P_2O_5 (f) $SrSO_4$ (g) C_2H_6O (h) SO_2

27. Determine the molecular or formula mass of each of the following compounds. (For this exercise, round all atomic masses to one decimal place.)

(a) SiO_2 (b) H_2SO_3 (c) $Sr(OH)_2$ (d) RbF (e) $Cu(NO_3)_2$ (f) $CoBr_2$ (g) $(NH_4)_3PO_4$
(h) $HC_2H_3O_2$

◀ 28. Determine the molecular or formula mass of each of the following compounds. (For this exercise, round all atomic masses to one decimal place.)

(a) $LiOH$ (b) Na_2CO_3 (c) $CoCl_2$ (d) $NaBr$ (e) SO_3 (f) C_2H_6 (g) OF_2 (h) $(NH_4)_2SO_3$

EXTRA EXERCISES

29. State what each of the symbols and subscripts mean in the chemical formulas for (a) H_2O (b) $C_6H_{12}O_6$ (c) $Ca(OH)_2$ (d) H_2

◀ 30. Make a list of heterogeneous mixtures and homogeneous mixtures that you encounter in everyday life. Do the same for elements and compounds.

31. Write the names and symbols for the fourteen elements that have a one-letter symbol.

32. Write the names and symbols of all the metalloids.

◀ 33. How many metals are there in the periodic table? How many nonmetals? How many metalloids?

34. Write the names of the eleven elements whose symbols are not derived from their English names.

35. Name the elements present in each of the following compounds:

(a) $MgCl_2$ (b) N_2O (c) $(NH_4)_2SO_4$ (d) H_3PO_4

36. Write the chemical formula of each of the following, given the number of atoms in a molecule or formula unit of the compound:

(a) one nitrogen atom, two oxygen atoms (nitrogen dioxide) (b) two sodium atoms, one sulfur atom (sodium sulfide) (c) three potassium atoms, one arsenic atom, four oxygen atoms (potassium arsenate) (d) two phosphorus atoms, five oxygen atoms (diphosphorus pentoxide)

37. Classify each of the following as an element, a compound, or a mixture:

(a) gold
(b) air
(c) carbon dioxide
(d) wine
(e) table salt

38. State whether each of the following involves a physical or chemical change:

(a) toasting bread
(b) water freezing
(c) tearing paper
(d) burning wood

39. Determine the molecular or formula mass of each of the following compounds. (For this exercise, round all atomic masses to one decimal place.)

(a) OsO_4 (b) HNO_3 (c) $Fe(OH)_2$ (d) $Ba_3(PO_4)_2$

40. Explain the difference between Hf and HF.

CUMULATIVE REVIEW Chapters 1–3*

Indicate whether each of the following statements is true or false.

1. Chemistry is the science that deals with matter and the changes it undergoes.
2. Organic chemicals are those chemicals that have been derived from living or once-living organisms.
3. Sugar and salt are examples of inorganic chemicals.
4. The art of *khemia* flourished until A.D. 1600.
5. Alchemists were able to transmute lead and iron into gold.
6. The idea that every theory must be proved by experiment was advanced by Robert Boyle.
7. In 420 B.C., Greek philosophers were able to prove that atoms existed.
8. Some carbon-containing compounds are classified as inorganic chemicals.
9. The two main branches of chemistry that existed during the 1700s and 1800s still exist today.
10. Nuclear chemistry, biochemistry, and analytical chemistry are three subdivisions into which chemistry can be divided.

Answer the following questions to help sharpen your test-taking skills.

◀ 11. Find the area of a rectangular room that measures 12.5 m by 10.2 m. Use the proper number of significant figures in reporting your answer.
12. Calculate the volume of a cardboard box that measures 15.25 cm long, 12.00 cm wide, and 24.85 cm high.
13. Determine the volume of a cylindrical solid whose radius is 14.50 cm and whose height is 25.05 cm.
14. Convert 0.28 kg to (a) grams, (b) decigrams, (c) milligrams.
15. Convert 6.8 m to (a) decimeters, (b) centimeters, (c) millimeters.
16. Convert 125 mm to (a) centimeters, (b) meters, (c) kilometers.
17. Convert 25,595 mL to (a) liters, (b) deciliters.
18. Express 50.0 m in (a) feet, (b) inches.
19. Express 25.5 g in (a) pounds, (b) ounces.
20. Determine the number of cubic centimeters in a cubic inch without using a conversion table.
21. Express 165 g in pounds.
22. Express 12.0 gallons in (a) liters, (b) ounces.
23. Determine the density of a cube that has a mass of 500.0 g and measures 12.0 cm on each side.
24. A spherical balloon with a volume of 113.0 mL is filled with an unknown gas weighing 20.0 g. Will the balloon float on air?

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*Answers to cummulative-review questions are given in the back of the book.

◀ ***25***. A small metal sphere weighs 90.0 g. The sphere is placed in a graduated cylinder containing 15.0 mL of water. Once the sphere is submerged, the water in the cylinder measures 30.0 mL. What is the density of the sphere?

26. A small metal sphere with a density of 3.50 g/cm^3 has a mass of 937.83 g. Calculate the radius of the sphere.

27. Solve the following problems using the proper number of significant figures:

(a) $28.64 + 3.2$
(b) $125.4 \div 13.5$
(c) 6.55×12.1
(d) $98.4 - 0.12$

◀ ***28***. Express each of the following numbers in scientific notation:

(a) 5,000 (b) 0.0005 (c) 602,300,000 (d) 35,$\overline{000}$,000

29. How many significant figures are there in each of the following numbers?

(a) 5,500.0 (b) 0.5123 (c) 12.000 (d) 3,5$\overline{00}$

30. Convert each of the following temperatures from °F to °C:

(a) 45.0°F (b) −10.0°F (c) 450°F (d) −100.0°F

31. Convert each of the following temperatures from °C to °F:

(a) 88.0°C (b) −12.5°C (c) 65.6°C (d) 4$\overline{00}$°C

32. Distinguish between an amorphous solid and a crystalline solid.

33. Determine whether each of the following processes involves chemical or physical changes:

(a) A match burns.
(b) Glucose dissolves in water.
(c) Bread becomes moldy.
(d) A piece of wood is sawed.

34. Distinguish between heterogeneous and homogeneous matter.

35. State the number of atoms of each element in a molecule or formula unit of the following compounds:

(a) $Zn(C_2H_3O_2)_2$ (b) $(NH_4)_2CrO_4$

36. Why do we use the term *molecular mass* for some compounds *and formula mass* for other compounds?

37. If in the periodic table calcium were assigned an atomic mass of 1, what would be the atomic mass of mercury?

38. Using the periodic table (inside front cover), look up the atomic masses of the following elements. (For this exercise, round all atomic masses to one decimal place.)

(a) Yb (b) At (c) P (d) Ag

39. Determine the molecular or formula mass of each of the following compounds. (For this exercise, round all atomic masses to one decimal place.

(a) $Fe(C_2H_3O_2)_3$ (b) $(NH_4)_2SO_4$ (c) $C_9H_8O_4$ (d) $CO(SO_4)_2$

40. Explain in detail what the chemical formula $A1_2(SO_4)_3$ means.

41. Determine whether each of the following is an element, a compound, or a mixture:

(a) air
(b) arsenic
(c) carbon dioxide
(d) water
(e) gold
(f) root beer soda
(g) gasoline

42. Determine whether each of the following is an example of a heterogeneous or homogeneous mixture:

(a) beach sand
(b) ethyl alcohol and water
(c) tossed salad
(d) soda water (club soda)

43. Using the periodic table (inside front cover), write the names and symbols for all elements whose symbol begins with the letter A. (*Hint:* There are eight.)

44. Write the names and symbols of all the nonmetallic elements.

45. Name the elements present in each of the following compounds:

(a) K_2S (b) Ag_2CrO_4 (c) $KMnO_4$ (d) $Hg_3(PO_4)_2$

46. Write the chemical formula of each of the following, given the number of atoms in a molecule or formula unit of the compound:

(a) two nitrogen atoms, one oxygen atom (dinitrogen monoxide)
(b) two potassium atoms, one chromium atom, four oxygen atoms (potassium chromate)
(c) one nitrogen atom, three hydrogen atoms (ammonia)
(d) one strontium atom, one sulfur atom, four oxygen atoms (strontium sulfate)

47. State whether each of the following involves a physical or chemical change:

(a) a candle burning
(b) table sugar dissolving in water
(c) tooth decaying
(d) snow melting

48. Explain the difference between No and NO.

◀ ***49***. A piece of paper is weighed and is then burned. The resulting ash weighs less than the paper. Is this a violation of the Law of Conservation of Mass? Explain.

50. You are given a mixture of sodium chloride (table salt) and sand. Explain how you would separate this mixture.