



WELCOME

When you look around, do you ever wonder where everything came from and how it was made? Have you ever contemplated why a tree is hard, a sponge is soft, and a breeze is invisible?

By faith we understand that the universe was formed at God's command, so that what is seen was not made out of what was visible.

—Hebrews 11:3, NIV

Welcome to the world of chemistry! This year you are going to take a journey that allows you to explore God's creation with the eyes of a scientist.

Studying God's creation at the molecular and atomic level can enable you to understand how wonderfully everything fits together in this world God has created for us.

How many are your works, Lord! In wisdom you made them all; the earth is full of your creatures.

—Psalm 104:24, NIV

Science is the endeavor of explaining the truth of the world around us, and God is the source of both creation and truth. You will discover that proper application of scientific principles will help you uncover how the world around you operates. Since science and faith both search for truth, they complement each other. The more you know about your world, the more you will wonder at the complex beauty of God's creation.

He has made everything beautiful in its time. He has also set eternity in the human heart; yet no one can fathom what God has done from beginning to end.

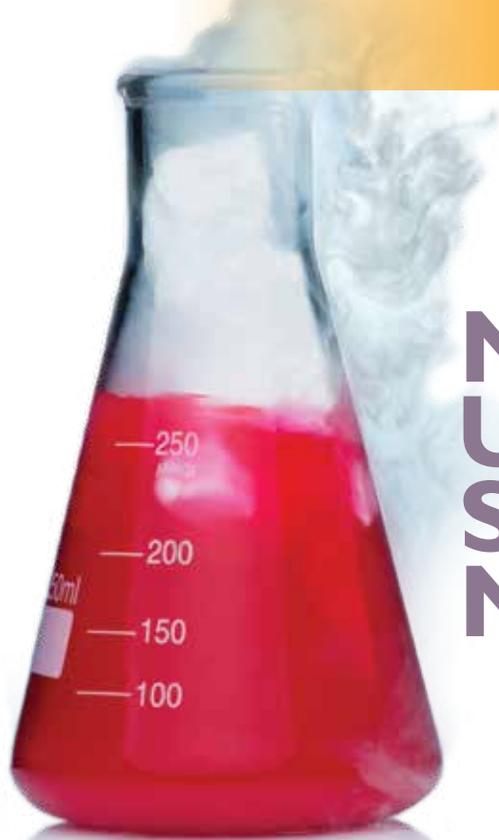
—Ecclesiastes 3:11, NIV

God gave humans dominion over the earth, so we can understand many things about it. This textbook is not just a compilation of facts and figures for you to memorize.

This textbook is designed to take you on a remarkable journey that involves facts about chemistry, figures to help you understand the facts, and truth from your Creator. We at Apologia pray that this text will enable you to say:

“How great are your works, LORD, how profound your thoughts!”

—Psalm 92:5, NIV



MEASUREMENT, UNITS, AND THE SCIENTIFIC METHOD

Our understanding of life has changed more in the past 2 centuries than all the previously recorded span of human history. The earth's population has increased more than 5-fold since 1800, and our life expectancy has nearly doubled because of our ability to synthesize medicines, control diseases, and increase crop yields. Many goods are now made of polymers (plastics) and ceramics instead of wood and metal because of our ability to manipulate and manufacture materials with properties unlike any found in nature. In one way or another, all of these changes involve **chemistry**.

What is chemistry? Quite simply, chemistry is the study of **matter**. Of course, this definition doesn't do us much good unless we know what matter is. So, to understand what chemistry is, we first need to define matter.

Matter—Anything that has mass and takes up space.

If matter is defined in this way, almost *everything* around us is matter. Your family car has a lot of mass. That's why it's so heavy. It also takes up a lot of space in the driveway or the garage. Your car must be made of matter. The food you eat isn't as heavy as a car, but it still has some mass. It also takes up space. So food must be made up of matter as well. Indeed, almost everything you see around you is made up of matter because nearly everything has mass and takes up space. There is one thing that has no mass and takes up no space. It's all around you right now. Can you think of what it might be?

You might think that the answer is air. However, that's not the right answer. Perform experiments 1.1 and 1.2 to see what we mean.

first reactions

When God told Noah to build an ark 300 cubits long, Noah had to know how long a cubit was in order to succeed. Imagine if Noah had built an ark 300 inches long! As you begin to learn the foundational language of chemistry, keep in mind that it is only the very beginning of true understanding.

EXPERIMENT 1.1

PURPOSE: To determine if air has mass.

MATERIALS:

- Meterstick (A yardstick will work as well, but a 12-inch ruler is not long enough.)
- Two 8-inch or larger balloons
- 2 pieces of string long enough to tie the balloons to the meterstick
- Tape
- Safety goggles

QUESTION: Does air have mass?

HYPOTHESIS: Pick one: Either air has mass or air does not have mass.

PROCEDURE:

1. Without blowing them up, tie the balloons to the strings. Be sure to make the knots loose so that you can untie one of the balloons later in the experiment.
2. Tie the other end of each string to an end of the meterstick. Try to attach the strings as close to the ends of the meterstick as possible.
3. Once the strings have been tied to the meterstick, tape them down so that they cannot move.
4. Go into your bathroom and pull back the shower curtain so that a large portion of the curtain rod is bare. Balance the meterstick (with the balloons attached) on the bare part of the shower curtain rod. You should be able to balance it very well. If you don't have a shower curtain rod or you are having trouble using yours, you can use any surface that is adequate for delicate balancing like the upper part of a chair.
5. Once you have the meterstick balanced, stand back and look at it. The meterstick balances now because the total mass on one side equals the total mass on the other side. To knock it off balance, you would need to move the meterstick or add more mass to one side. You will do the latter.
6. Have someone else hold the meterstick so that it does not move. For this experiment to work properly, the meterstick must stay stationary.
7. While the meterstick is held stationary, remove one of the balloons from its string (do not untie the string from the meterstick), and blow up the balloon.
8. Tie the balloon closed so that the air does not escape, then reattach it to its string.
9. Have the person holding the meterstick let go. If the meterstick was not moved while you were blowing up the balloon, it will tilt toward the side with the inflated balloon as soon as the person lets it go. This is because you added air to the balloon. Since air has mass, it knocks the meterstick off balance. So air does have mass!
10. Clean up and return everything to the proper place.

CONCLUSION: What did you think? Write something about what you observed related to the fact that air has mass.

EXPERIMENT 1.2

PURPOSE: To determine if air takes up space.

MATERIALS:

- Tall glass
- Paper towel
- Sink full of water
- Safety goggles

QUESTION: Does air take up space?

HYPOTHESIS: Pick one: Either air takes up space or air does not take up space.

PROCEDURE:

1. Fill the sink with water until the water level is high enough to submerge the entire glass.
2. Make sure the inside of the glass is dry.
3. Wad up the paper towel and shove it down into the bottom of the glass.
4. Turn the glass upside down and be sure that the paper towel does not fall out of the glass.
5. Submerge the glass upside down in the water, being careful not to tip the glass at any time.
6. Wait a few seconds and remove the glass, still being careful not to tilt it.
7. Pull the paper towel out of the glass. You will find that the paper towel is completely dry. Even though the glass was submerged in water, the paper towel never got wet. Why? When you tipped the glass upside down, there was air inside the glass. When you submerged it in the water, the air could not escape the glass, so the glass was still full of air. Since air takes up space, there was no room for water to enter the glass, so the paper towel stayed dry.
8. Repeat the experiment, but this time be sure to tip the glass while it is underwater. You will see large bubbles rise to the surface of the water. When you pull the glass out, you will find that it has water in it and that the paper towel is wet. This is because tilting the glass allowed the air trapped inside it to escape. Once the air escaped, there was room for the water to come into the glass.
9. Clean up and return everything to the proper place.

CONCLUSION: What did you think? Write something about what you observed related to the fact that air takes up space.

think about this

Air is typically used as a metaphor for nothingness. It is, however, very complex. Aristotle (384–322 BC) is generally given credit for being the first to state that air has weight, although many did not believe him. Would it surprise you to know that some 1,400 years *before* Aristotle, it was known that air had weight? It's true. The Bible tells us that God “gave to the wind its weight and apportioned the waters by measure” (Job 28:25). Think about this verse in the context of the timeline of scientific knowledge. Job may have lived anywhere from 2300 to 1700 BC. God created everything visible and invisible, including air. The Bible is His word, and we can trust it to be true. We explore science to understand what God already knows about His creation.

Now that you see that air does have mass and does take up space, have you figured out the correct answer to our original question? What very common thing that is surrounding you right now has no mass and takes up no space? The answer is light. As far as scientists can tell, light does not have any mass and takes up no space. Light is not considered matter. Instead, it is pure energy. Everything else that you see around you is considered matter. Chemistry, then, is the study of nearly everything! As you can imagine, studying nearly everything can be a very daunting task. However, chemists have found that even though there are many forms of matter, they all behave according to a few fundamental laws. If we can clearly understand these laws, then we can clearly understand the nature of the matter that exists in God's creation.

Before we start trying to understand these laws, we must step back and ask a more fundamental question: *How* do we study matter? The first thing we have to be able to do in order to study matter is to measure it. If we want to study an object, we first must learn things like how big it is, how heavy it is, and how old it is. To learn these things, we have to make some measurements. The rest of this module explains how scientists measure things and what those measurements mean.

UNITS OF MEASUREMENT

Let's suppose you are making curtains for a friend's windows. You ask him to measure the window and give you the dimensions so that you can make the curtains the right size. Your friend tells you that his windows are 50 by 60, so that's how big you make the curtains. When you go over to his house, it turns out that your curtains are more than twice as big as his windows! Your friend tells you that he's certain he measured the windows correctly, and you tell your friend that you are certain you measured the curtains correctly.

How can this be? The answer is quite simple. Your friend measured the windows in centimeters. You, on the other hand, measured the curtains in inches. The problem was not caused by measuring incorrectly. Instead, the problem was the result of measuring with different **units**.

think about this

“Who has measured the waters in the hollow of his hand and marked off the heavens with a span, enclosed the dust of the earth in a measure and weighed the mountains in scales and the hills in a balance?” Isaiah 40:12

Can you determine what the author is trying to do here? He is trying to describe the greatness of God by using measurements that would have been understood by anyone living at this time. How would scientists of this day try to describe the greatness of God?

When we are making measurements, the units we use are just as important as the numbers that we get. If your friend had told you that his windows were 50 centimeters by 60 centimeters, there would have been no problem. You would have known exactly how big to make the curtains. Since he failed to do this, the numbers that he gave you (50 by 60) were useless. A failure to indicate the units involved in measurements can lead to serious problems. For example, the Mars Climate Orbiter, a NASA (National Aeronautics and Space Administration) spacecraft built for the exploration of Mars, vanished during an attempt to put it into orbit around the planet. In an investigation that followed, NASA determined that a mix-up in units had caused the disaster. One team of engineers had used metric units in its calculations, while another team had used English units in executing an engine burn. The teams did not indicate the units they were using, and as a result, we lost a spacecraft worth several billion dollars.

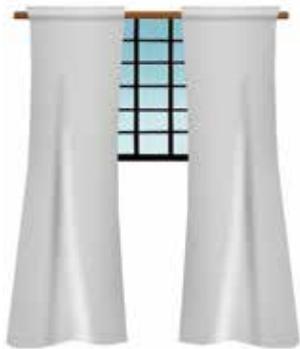
Scientists should never simply report numbers. They must always include units so that everyone knows exactly what the numbers mean. That will be the rule in this chemistry course.

If you answer a question or solve a problem and do not list units with the numbers, your answer will be considered incorrect.

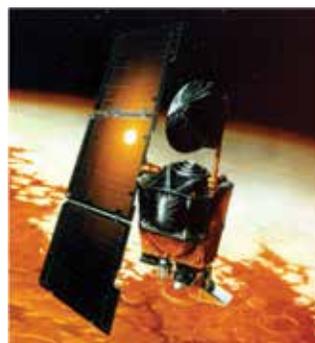
FIGURE 1.1

Two Consequences of Not Using Units Properly

Window illustration by David Weiss. Mars Climate Orbiter image courtesy of NASA/JPL/Caltech



These curtains are too long for this window because the window was measured in centimeters, but the curtains were made assuming the measurements were in inches.



The Mars Climate Orbiter did not successfully make it into orbit because 2 engineering teams involved used different units in their designs.

Since scientists use units in all of their measurements, it is convenient to define a standard set of units that will be used by everyone. This system of standard units is called the **metric system**. If you do not fully understand the metric system, don't worry. By the end of this module, you will be an expert at using it. If you do fully understand the metric system, you can skim this section as a review.

THE METRIC SYSTEM

We need to measure many different things when studying matter. First, we must determine how much matter exists in the object that we want to study. We know that there is a lot

more matter in a car than there is in a feather because a car is heavier. To study an object precisely, we need to know exactly how much matter is in it. To accomplish this, we need to measure mass and the amount of space the object takes up. So how do we measure the object's mass? In the metric system, the unit for mass is the **gram**. If an object has a mass of 10 grams, we know that it has 10 times the matter in an object with a mass of 1 gram. To give you an idea of the size of a gram, the average mass of a housefly is about 1 gram. A gram is a rather small unit. Most of the things that we will measure will have masses of 10 to 10,000 grams. For example, this book has a mass of about 2,300 grams.

Now that we know what the metric unit for mass is, we need to know a little more about the concept of mass. Some people might think of mass as weight. That's not exactly true. Mass and weight are 2 different measurements. **Mass** measures how much matter exists in an object. **Weight**, on the other hand, measures how hard gravity pulls on that object.

For example, if I were to use my bathroom scale and weigh myself, I would find that I weigh 150 pounds. However, if I were to take that scale to the moon and weigh myself, I would find that I weigh only 25 pounds. Does that mean I'm thinner on the moon than I am at home? Of course not. It means that on the moon, gravity is not as strong as it is in my house on Earth.

On the other hand, if I were to measure my mass at home, I would find it to be 68,000 grams. If I were to measure my mass on the moon, it would still be 68,000 grams. That's the difference between mass and weight. Since weight is a measure of how hard gravity pulls, an object weighs different amounts depending on the gravity that is present. Because mass measures how much matter is in an object, it does not depend on the gravity present.

Unfortunately, there are many other unit systems in use today besides the metric system. In fact, the metric system is probably not the system with which you are most familiar. You are probably most familiar with the English system. The English unit for mass is (believe it or not) called the **slug**. Although we will not use the slug often, it is important to understand what it means, especially when you study physics. The English system uses pounds as a measurement of how much material an object has. However, pounds are not a measure of mass; they are a measure of weight. The metric unit for weight is called the **Newton**.

There is more to measurements than mass, however. We might also want to measure how big an object is. For this, we must use the metric system's unit for distance, the **meter**. You are probably familiar with a yardstick. A meter is slightly longer than a yardstick. The English unit for distance is the foot. There are many other units of distance like inches, yards, and miles, but we'll talk about those a little later.

We also need to be able to measure how much space an object occupies. This measurement is commonly called **volume** and is measured in the metric system with the unit called the **liter**. The main unit for measuring volume in the English system is the **gallon**. To give you an idea of the size of a liter, it takes just under 4 liters to make a gallon.

Finally, we have to be able to measure the passage of time. When studying matter, we will see that it has the ability to change. The shape, size, and chemical properties of certain substances change over time. It is important to be able to measure time so that we can determine how quickly the changes take place. In both the English and metric systems, time is measured in **seconds**. That is a good thing, isn't it? We know of only one way to measure time!

Since it is very important for you to be able to recognize which units correspond to which measurements, table 1.1 summarizes what you have just read. The letters in parentheses are the commonly used abbreviations for the units listed. You should memorize this table.

TABLE 1.1
Physical Quantities and Their Base Units

Physical Quantity	Base Metric Unit	Base English Unit
Mass	gram (g)	slug (sl)
Distance	meter (m)	foot (ft)
Volume	liter (L)	gallon (gal)
Time	second (s)	second (s)

MANIPULATING UNITS

Let's suppose we asked you to measure the width of your home's kitchen using the English system. What unit would you use? Most likely, you would express your measurement in feet. However, suppose instead we asked you to measure the diameter of a penny. Would you still use the foot as your measurement unit? Probably not. Since you know the English system already, you would probably recognize that inches are also a unit for distance; since a penny is relatively small, you would use inches instead of feet. In the same way, if we asked you to measure the distance between 2 cities, you would probably express your measurement in terms of miles, not feet. This is why we used the term *base English unit* in table 1.1. Even though the English system's normal unit for distance is the foot, there are alternative units for length when measuring very short or very long distances. The same holds true for other English units. For example, volume can be measured in cups, pints, and ounces. We choose the unit based on the object to be measured.

The metric system also has alternative units for measuring small things compared to measuring big things. These alternative units are constructed by placing a prefix in front of the metric base unit. You will soon see the metric system is easier to use and understand than the English system because the prefixes always have the same relationship to the base unit, regardless of what measurement is used.

To use a larger or smaller scale in the metric system, simply add a prefix to the base unit. For example, in the metric system, the prefix *centi-* means one-hundredth, or 0.01. So, if we wanted to measure the length of a sewing needle in the metric system, we would probably express the measurement with the *centimeter* unit. Since a centimeter is one-hundredth of a meter, it can be used to measure relatively small things. On the other hand, the prefix *kilo-* means 1,000. If we want to measure the distance between 2 states, we would probably use the *kilometer*. Since each kilometer is 1,000 times longer than the meter, it can be used to measure long things.

The beauty of the metric system is that these prefixes mean the same thing *regardless of the physical quantity being measured!* If we were measuring something with a very large mass (such as a car or boat), then we would use the kilogram unit. One kilogram is the same as 1,000 grams. In the same way, if we were measuring something that had a large volume, we might use the kiloliter, which is 1,000 liters. Adding a *kilo-* prefix to a unit

multiplies the scale of the unit by 1,000—that is, 1,000 times larger than the base unit scale.

Compare this incredibly logical system of units to the chaotic English system. If we want to measure something short, we use the inch unit, which is equal to one-twelfth of a foot. On the other hand, if we want to measure something with small volume, we might use the quart unit, which is equal to one-fourth of a gallon. In the English system, every alternative unit has a different relationship to the base unit, and we must remember all of those crazy numbers. We have to remember that there are 12 inches in a foot, 3 feet in a yard, and 5,280 feet in a mile, while at the same time remembering that for volume there are 8 ounces in a cup, 2 cups in a pint, 2 pints in a quart, and 4 quarts in a gallon. That's a lot to memorize! Thankfully, the majority of science operates around the metric system.

In the metric system, all we have to remember is what the prefix means. Since the *centi-* prefix means one-hundredth, then we know that 1 centimeter is one-hundredth of a meter, 1 centiliter is one-hundredth of a liter, and 1 centigram is one-hundredth a gram. Since the *kilo-* prefix means 1,000, we know that there are 1,000 meters in a kilometer, 1,000 grams in a kilogram, and 1,000 liters in a kiloliter. Doesn't that make a lot more sense?

Another advantage to the metric system is that the prefixes are all based on a factor of 10. Table 1.2 summarizes the most commonly used prefixes and their numerical meanings. The prefixes in boldface type are the ones that we will use over and over again. Memorize those 3 prefixes and their meanings before you take the test for this module. The commonly used abbreviations for these prefixes are listed in parentheses.

TABLE 1.2
Common Prefixes Used in the Metric System

Prefix	Numerical Meaning
micro (μ)	0.00001
milli (m)	0.001
centi (c)	0.01
deci (d)	0.1
deca (D)	10
hecta (H)	100
kilo (k)	1,000
mega (M)	1,000,000

Remember that each of these prefixes, when added to a base unit, makes an alternative scale for measurement. If you wanted to measure the length of something small, you would have all sorts of options for which unit to use. If you wanted to measure the length of someone's foot, you could use the decimeter. Since the decimeter is one-tenth of a meter, it measures things that are only slightly smaller than a meter. On the other hand, if you wanted to measure the length of a sewing needle, you could use the centimeter because a sewing needle is significantly smaller than a meter. Or if you want to measure the thickness of a piece of paper, you might use the millimeter since it is one-thousandth of a meter, which is a *really* small unit.

You can see that the metric system is much more logical and versatile than the English system. That is, in part, why scientists and most countries in the world use it as their main

system of measurement. With the exception of the United States, almost every country in the world uses the metric system as its standard system of units. Since scientists in the United States frequently work with scientists from other countries around the world, they must use and understand the metric system. Throughout all of the modules of this chemistry course, the English system of measurement will be presented only for illustration purposes. Since scientists must thoroughly understand the metric system, it will be the main system of units that we will use.

CONVERTING BETWEEN UNITS

Now that you understand what prefixes are and how they are used in the metric system, you must become familiar with converting between units within the metric system. In other words, if you measure the length of an object in centimeters, you should also be able to convert your answer to any other distance unit. For example, if you measure the length of a sewing needle in centimeters, you should be able to convert that length to millimeters, decimeters, meters, etc. Accomplishing this task is relatively simple as long as you remember the skills you learned in multiplying fractions. Suppose we asked you to complete the following problem:

$$\frac{7}{64} \times \frac{64}{13} =$$

There are 2 ways to figure out the answer. The first way would be to multiply the numerators and the denominators together and then simplify the fraction. If you did it that way, it would look something like this:

$$\frac{7}{64} \times \frac{64}{13} = \frac{448}{832} = \frac{7}{13}$$

You could get the answer much more quickly, however, if you remember that when multiplying fractions, common factors in the numerator and the denominator cancel each other out. The 64 in the numerator cancels with the 64 in the denominator, and the only factors left are the 7 in the numerator and the 13 in the denominator. In this way, you reach the final answer in one less step:

$$\frac{7}{\cancel{64}} \times \frac{\cancel{64}}{13} = \frac{7}{13}$$

Another skill in multiplying fractions you will need is illustrated in this next equation:

$$\frac{7}{13} \times \frac{3}{3} = \frac{21}{39}$$

Notice what happens when the original fraction is multiplied by a value equal to 1. The answer is equal to the original fraction, but it looks different. Multiplying a fraction by a value of 1 does not change the value of the fraction.

We will use these fraction concepts in converting between units. Suppose we measure the length of a pencil to be 15.1 centimeters, but the person who wants to know the length of the pencil would like us to tell him the answer in meters. How would we convert between centimeters and meters? First, we would need to know the relationship

between centimeters and meters. According to table 1.2, *centi-* means 0.01. So 1 centimeter is the same thing as 0.01 meters. In mathematical form, we would say:

$$1 \text{ centimeter} = 0.01 \text{ meter}$$

Look what happens when we divide both sides by 1 centimeter:

$$\frac{1 \text{ centimeter}}{1 \text{ centimeter}} = \frac{0.01 \text{ meter}}{1 \text{ centimeter}}$$

The fraction on the right equals 1! Do you see how that works? If 1 centimeter equals 0.01 meters, then the numerator and the denominator on the right side are saying the same thing, which is why that fraction equals 1. So if we use this fraction and multiply it by a measurement, then we will not change the value of the measurement. We call this fraction a conversion factor.

Now that we know how centimeters and meters relate to one another, we can convert from one to another. First, we write down the measurement that we know:

$$15.1 \text{ centimeters}$$

We then realize that any number can be expressed as a fraction by putting it over the number 1. So we can rewrite our measurement as:

$$\frac{15.1 \text{ centimeters}}{1}$$

Now we can convert that measurement into meters by multiplying it with the conversion factor we determined above. We have to do it the right way so that the original measurement unit cancels out when multiplied by the conversion factor. Here's how we do it:

$$\frac{15.1 \text{ centimeters}}{1} \times \frac{0.01 \text{ meter}}{1 \text{ centimeter}} = .151 \text{ meter}$$

Given Unit Conversion Factor Wanted Unit

So 15.1 centimeters is the same as 0.151 meters. This conversion method, called the **factor-label method**, works for 2 reasons. First, since 0.01 meters is the same as 1 centimeter, multiplying our measurement by 0.01 meters over 1 centimeter is the same as multiplying by 1. Since nothing changes when we multiply by 1, we haven't altered the value of our measurement at all. Second, by putting the 1 centimeter in the denominator of the second fraction, we allow the centimeters unit to cancel (just like the 64 canceled in the previous discussion). Once the centimeters unit has canceled, the only thing left is meters, so we know that our measurement is now in meters.

This is how we will do all of our factor-label setups. We will first write the measurement we are given in fraction form by putting it over 1. We will then put the unit we do not want in the denominator of the conversion factor and put the unit we do want in the numerator. Finally, the numerical meaning of any prefixes needs to go on the opposite side of the conversion factor to get the fraction to equal 1. We will see many examples of this method, so don't be concerned if you are a little confused right now.

Here is a list of steps for the factor-label method.

1. Create a fraction out of the given measurement by placing it over 1.
2. Place the original measurement unit in the denominator of the conversion factor.
3. Place the wanted unit in the numerator of the conversion factor.
4. Place the numerical meaning of any prefixes on the opposite side of the conversion factor.
5. Multiply the given measurement fraction by the conversion factor.

It may seem odd that words can be treated exactly the same as numbers, but measuring units have just that property. Whenever a measurement is used in any mathematical equation, the units for that measurement must be included in the equation. Those units are then treated the same way numbers are treated. We will come back to this point in an upcoming section of this module.

We will be using the factor-label method for many other types of problems throughout this course, so it is very, very important for you to learn it. Also, since we will be using it so often, we should start abbreviating things so that they will be easier to write down. We will use the abbreviations for the base units listed in table 1.1, along with the prefix abbreviations listed in table 1.2. For example, kilograms will be abbreviated as kg, while milliliters will be abbreviated as mL.

Since the factor-label method is so important in our studies of chemistry, let's see how it works in example 1.1.

EXAMPLE 1.1

A student measures the mass of a rock to be 14,351 grams. What is the rock's mass in kilograms?

First, we use the definition of *kilo-* to determine the relationship between grams and kilograms:

$$1 \text{ kg} = 1,000 \text{ g}$$

Then we put our given measurement in fraction form:

$$\frac{14,351 \text{ g}}{1}$$

Then we multiply our measurement by a fraction that contains the relationship noted above, making sure to put the 1,000 g in the denominator so that the unit of grams will cancel out:

$$\frac{14,351 \cancel{\text{g}}}{1} \times \frac{1 \text{ kg}}{1,000 \cancel{\text{g}}} = 14.351 \text{ kg}$$

Given Unit Conversion Factor Wanted Unit

So, 14,351 grams is the same as **14.351 kilograms**.

Because we will use it over and over again, you must master this powerful technique. Also, you will see toward the end of this module that the factor-label method can become extremely complex; therefore, it is very important that you take the time now to answer “On Your Own” questions 1.1–1.3. Once you have solved the problems on your own, check your answers using the solutions provided at the end of the module.

ON YOUR OWN

- 1.1 A student measures the mass of a bag of sand to be 9,321 g. What is the bag's mass in kg?
- 1.2 If a glass contains 0.465 L of water, what is the volume of water in mL?
- 1.3 On a professional basketball court, the distance from the 3-point line to the basket is 724.0 cm. What is this distance in meters?

CONVERTING BETWEEN UNIT SYSTEMS

As you may have guessed, the factor-label method can also be used to convert *between* systems of units as well as within systems of units. If a measurement is done in the English system, the factor-label method can be used to convert that measurement to the metric system, or vice versa. To be able to do this, we must learn the relationships between metric and English units, as shown in table 1.3. Although these relationships

are important, we will not use them very often, so you don't need to memorize them.

TABLE 1.3
Relationships between English and Metric Units

Measurement	English/Metric Relationship
Distance	1.00 inch = 2.54 cm
Mass	1.00 slug = 14.59 kg
Volume	1.00 gallon = 3.78 L

We can use this information to form conversion factors for the factor-label method the same way we did for the metric system conversions. Study example 1.2 to see how this works.

EXAMPLE 1.2

The length of a tabletop is measured to be 37.8 inches. What is that length in cm?

To solve this problem, we first put the given measurement in its fraction form:

$$\frac{37.8 \text{ in}}{1}$$

We then multiply this fraction by the conversion relationship so that the inches unit cancels:

$$\frac{37.8 \text{ in}}{1} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 96.012 \text{ cm}$$

Given Unit Conversion Factor Wanted Unit

So a measurement of 37.8 inches is equivalent to **96.012 cm**.

Give yourself a little more practice with the factor-label method by answering “On Your Own” questions 1.4–1.5.

MORE COMPLEX UNIT CONVERSIONS AND PROBLEM SOLVING

Now that we have seen some simple applications of the factor-label method, let’s look at more complex problems. For example, suppose we measure the volume of a liquid to be 4,523 centiliters but would like to convert this measurement into kiloliters. This is a more complicated problem because the relationship between cL and kL is not direct. In all of the previous examples, we knew the relationship between the unit we had and the unit to which we wanted to convert. In this problem, we need to add an extra step to convert from one to the other.

We would say there is an indirect relationship between the 2 units. We know how many cL are in a L and how many L are in a kL, so we can use these 2 relationships to do a 2-step conversion. First, we can convert centiliters into liters:

$$\frac{4,523 \cancel{\text{cL}}}{1} \times \frac{0.01 \text{ L}}{1 \cancel{\text{cL}}} = 45.23 \text{ L}$$

Given Unit Conversion Factor Wanted Unit

Then we can convert liters into kiloliters:

$$\frac{4,523 \cancel{\text{L}}}{1} \times \frac{1 \text{ kL}}{1,000 \cancel{\text{L}}} = 0.04523 \text{ kL}$$

Given Unit Conversion Factor Wanted Unit

We are forced to do this 2-step process because of the indirect relationship between the 2 prefix units in the metric system. However, we can always convert between 2 prefix units if we first convert to the base unit. To speed up this kind of conversion, we can combine these 2 steps into 1 line:

$$\frac{4,523 \cancel{\text{cL}}}{1} \times \frac{0.01 \cancel{\text{L}}}{1 \cancel{\text{cL}}} \times \frac{1 \text{ kL}}{1,000 \cancel{\text{L}}} = 0.04523 \text{ kL}$$

Given Unit Conversion Factors Wanted Unit

You will be seeing mathematical equations like this one as we move through the subject of chemistry, so it is important for you to understand what’s going on. The first fraction in the equation above represents the measurement that we were given. Since there isn’t a relationship between the unit we were given and the unit to which we will convert, we first convert the given unit to the base unit. This is accomplished with the second fraction in the equation. When the first fraction is multiplied by the second fraction, the cL unit cancels and is replaced by the L unit. The third fraction then cancels the L unit and replaces it with the kL unit, which is the unit we want. This gives us our final answer.

ON YOUR OWN

- 1.4 How many kilograms are in 8.465 slugs?
- 1.5 If an object occupies a volume of 6.1236 liters, how many gallons does it occupy?

We solved all of the previous problems using the same steps.

1. The problem gave a measurement with a unit and asked for a different unit.
2. We planned and laid out conversion factors. Think of the sequence of unit conversions you will need to get from the initial unit to the final unit for the answer. For each change of unit in your plan, you will need a conversion factor.
3. Finally, we set up the factor-label method using the steps we learned earlier.

Try to use the above approach in example 1.3.

EXAMPLE 1.3

The mass of an object is measured to be 0.030 kg. What is the object's mass in mg?

Given Unit: kg

Wanted Unit: mg

Plan Conversion Factors: Because the relationship between milligrams and kilograms is indirect, we will plan a 2-step conversion. First, we know that:

$$1 \text{ kg} = 1,000 \text{ g}$$

We also know that:

$$1 \text{ mg} = 0.001 \text{ g}$$

So we need to multiply our original measurement in fraction form with both of these relationships. We must do it in such a way as to cancel out the kg and replace it with g, and then cancel out the g and replace it with mg.

Set Up Problem:

$$\frac{.030 \text{ ~~kg~~}}{1} \times \frac{1,000 \text{ ~~g~~}}{1 \text{ ~~kg~~}} \times \frac{1 \text{ ~~g~~ mg}}{0.001 \text{ ~~g~~}} = 30,000 \text{ mg}$$

Given Unit Conversion Factors Wanted Unit

The object's mass is 0.030 kg, which is the same as **30,000 mg**.

ON YOUR OWN

- 1.6 A balloon is blown up so that its volume is 1,500 mL. What is its volume in kL?
- 1.7 If the length of a race car track is 2.0 km, what is it in cm?
- 1.8 How many mg are in 0.01 Mg?

The factor-label method is one of the most important tools you can learn for the study of chemistry (and physics, for that matter). Therefore, you must become a veritable expert at it. Try your hand at “On Your Own” questions 1.6–1.8 so that you can get some more practice.

DERIVED UNITS

We mentioned previously that units can be used in mathematical expressions in the same way that numbers can be used. Just as there are rules for adding, subtracting, multiplying, and dividing numbers, there are also rules governing those operations when using units. You will have to become very adept at using units in mathematical expressions, so let's discuss those rules now.

Adding and Subtracting Units: When adding and subtracting units, the most important thing to remember is that the units we are adding or subtracting must be identical. For example, we cannot add grams and liters. The result would not make sense physically. Since gram is a unit of mass and liter is a unit of volume, there is no way we can add or subtract them. We also cannot add or subtract kilograms and grams. Even though both units measure mass, we cannot add or subtract them unless the units are *identical*. If we did want to add or subtract them, we would have to convert the kilograms into grams or convert the grams into kilograms. It doesn't matter which way we go, as long as the units we add or subtract are identical.

Once we have identical units, we can add and subtract them using the rules of algebra. Since $2x + 3x = 5x$, we know that $2 \text{ cm} + 3 \text{ cm} = 5 \text{ cm}$. In the same way, $3.1 \text{ g} - 2.7 \text{ g} = 0.4 \text{ g}$. When adding or subtracting units, we add or subtract the numbers they are associated with and then simply carry the unit along into the answer.

Multiplying and Dividing Units: When multiplying and dividing units, it doesn't matter whether or not the units are identical. Unlike addition and subtraction, we can multiply or divide any unit by any other unit. In algebra:

$$3x \times 4y = 12xy$$

When multiplying units:

$$3 \text{ kg} \times 4 \text{ mL} = 12 \text{ kg} \times \text{mL}$$

Similarly, in algebra:

$$6x \div 2y = 3 \frac{x}{y}$$

When dividing units:

$$6 \text{ g} \div 2 \text{ mL} = 3 \frac{\text{g}}{\text{mL}}$$

So when multiplying or dividing units, you multiply or divide the numbers and then do exactly the same thing to the units.

Let's use the rules we've just learned to explore a few other things about units. First, let's see what happens when we multiply measurements that have the same units. Suppose we wanted to measure the surface area of a rectangular table. From geometry, we know that the area of a rectangle is the length times the width. So, let's suppose we measure the length of a table to be 1.1 meters and the width to be 2.0 meters. Its area would be:

$$1.1 \times 2.0 = 2.2$$

What would the units be? In algebra, we would say that:

$$1.1x \times 2.0x = 2.2x^2$$

Therefore:

$$1.1 \text{ m} \times 2.0 \text{ m} = 2.2 \text{ m}^2$$

This tells us that m^2 (square meter) is a unit for area.

Let's take this one step further. Suppose we measure the length, width, and height of a small box to be 1.2 cm, 3.1 cm, and 1.4 cm, respectively. What would the volume of the box be? From geometry, we know that volume is length times width times height, so the volume would be:

$$1.2 \text{ cm} \times 3.1 \text{ cm} \times 1.4 \text{ cm} = 5.208 \text{ cm}^3$$

Therefore, cm^3 (usually called cubic centimeters or cc's) is a unit for volume. If you've ever listened to doctors or nurses talking about how much liquid to put in a syringe when administering a shot, they usually use cc's as the unit. When a doctor tells a nurse, "Give the patient 4 cc's of penicillin," he is telling the nurse to inject a 4- cm^3 volume of penicillin into the patient.

Wait a minute. Wasn't the metric unit for volume the liter? Yes, but another metric unit for volume is the cm^3 . In addition, m^3 (cubic meters) and km^3 (cubic kilometers) are possible units for volume. This is a very important point. Often, several different units exist for the same measurement. The units you use will depend, to a large extent, on what information you are given in the first place. We'll see more about this fact later.

Units like cm^3 are called **derived units** because they are derived from math calculations with basic units that make up the metric system. Many of the units you will use in chemistry are derived units. We will discuss one very important physical quantity with derived units in an upcoming section of this module, but first you need to understand exactly how to use derived units in mathematical equations.

Let's suppose we want to take the volume that we previously determined for the box and convert it from cubic centimeters to cubic meters. You might think the conversion would look something like this:

$$\frac{5.208 \text{ cm}^3}{1} \times \frac{0.01 \text{ m}}{1 \text{ cm}}$$

This conversion might look correct, but there is a major problem with it. Remember what the factor-label method is designed to accomplish. In the end, the old units are supposed to cancel out, leaving the new units in their place. The way this conversion is set up, the old units *do not cancel!* When we multiply these 2 fractions together, the cm in the denominator does not cancel out the cm^3 in the numerator. When multiplying fractions, the numerator and denominator must be *identical* for them both to cancel. The cm in the denominator above must be replaced with cm^3 .

How is this done? It's quite simple. Do you remember the math equation for volume? It is length times width times height. You have to provide a conversion factor for each unit of the length, width, and height.

$$\frac{5.208 \text{ cm}^3}{1} \times \frac{0.01 \text{ m}}{1 \text{ cm}} \times \frac{0.01 \text{ m}}{1 \text{ cm}} \times \frac{0.01 \text{ m}}{1 \text{ cm}}$$

We can shorten this setup by using an exponent:

$$\frac{5.208 \text{ cm}^3}{1} \times \left(\frac{0.01 \text{ m}}{1 \text{ cm}} \right)^3$$

Inside the parentheses, the m becomes m³, the cm becomes cm³, the 0.01 becomes 0.000001, and the 1 stays as 1:

$$\frac{5.208 \text{ cm}^3}{1} \times \frac{0.000001 \text{ m}^3}{1 \text{ cm}^3} = 0.000005208 \text{ m}^3$$

Given Unit Conversion Factor **Wanted Unit**

Now since both the numerator and denominator have a unit of cm³, that unit cancels and is replaced with the m³. So a volume of 5.208 cm³ is equivalent to a volume of 0.000005208 m³.

Since cubic meters, cubic centimeters, and the like are measurements of volume, you might have already guessed that there must be a relationship between these units and the other volume units we discussed earlier. In fact, 1 cm³ is the same thing as 1 mL. This is a very important relationship and is something you will have to know before you can finish this module. Commit it to memory now:

1 cubic centimeter is the same as 1 milliliter (1 cm³ = 1 mL).

Let's combine this fact with the mathematics we just learned and perform a very complicated unit conversion. If you can understand example 1.4 and successfully complete "On Your Own" questions 1.9–1.11, then you have mastered unit conversion. If things are still a bit shaky for you, don't worry. There are plenty of practice problems at the end of this module to help you practice your unit conversion skills.

EXAMPLE 1.4

The length, width, and height of a small box are measured to be 1.1 in, 3.2 in, and 4.6 in, respectively. What is the box's volume in liters?

Given Unit: in (length)

Wanted Unit: L (volume)

Plan Conversion Factors: First, determine volume. To solve this problem, we first use the geometric equation for the volume of a box:

$$V = l \times w \times h$$

Set Up Problem:

$$V = 1.1 \text{ in} \times 3.2 \text{ in} \times 4.6 \text{ in} = 16.192 \text{ in}^3$$

Plan Conversion Factors: Second, convert to liters. Now that we know the volume, we just have to convert from in³ to L. This is a little more difficult than it sounds, however. Since there is

no direct relationship between in³ and L, we must go through a series of conversions to get to the desired unit. First, we can convert our unit from the English system to the metric system by using the following relationship:

$$1 \text{ in} = 2.54 \text{ cm}$$

To do this, we will have to cube the fraction we multiply by so that we have cm³ and in³.

Set Up Problem:

$$\frac{16.192 \text{ in}^3}{1} \times \left(\frac{2.54 \text{ cm}}{1 \text{ in}} \right)^3 = \frac{16.192 \text{ in}^3}{1} \times \left(\frac{16.387 \text{ cm}^3}{1 \text{ in}^3} \right) = 265.338 \text{ cm}^3$$

Now that we have the metric volume unit, we can use the fact that a cm³ is the same as a mL:

$$265.338 \text{ cm}^3 = 265.338 \text{ mL}$$

Now we can convert from mL to L:

Set Up Problem:

$$\frac{265.338 \text{ mL}}{1} \times \frac{0.001 \text{ L}}{1 \text{ mL}} = 0.265338 \text{ L}$$

The volume, then, is **0.265338 L**.

Alternative approach: First, convert inches to centimeters, then get the volume in cubic centimeters, then convert to milliliters, and finally convert to liters.

Set Up Problem:

We convert initial inch units to cm:

$$\frac{1.1 \text{ in}}{1} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 2.794 \text{ cm} \qquad \frac{3.2 \text{ in}}{1} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 8.128 \text{ cm}$$

$$\frac{4.6 \text{ in}}{1} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 11.684 \text{ cm}$$

We calculate volume in cm³:

$$V = 2.794 \text{ cm} \times 8.128 \text{ cm} \times 11.684 \text{ cm} = 265.338 \text{ cm}^3$$

Finally, we convert cm³ to L (remember that cm³ is equal to mL):

$$\frac{265.338 \text{ mL}}{1} \times \frac{0.001 \text{ L}}{1 \text{ mL}} = \mathbf{0.265338 \text{ L}}$$

You will see that this is the same answer as above. We simply changed the order of how we did it.

MAKING MEASUREMENTS

Now that we've learned about measurement units, we need to spend a little time learning how to make measurements. After all, being able to manipulate units in mathematical equations isn't going to help us unless we can make measurements with those units to begin with. To learn how to make measurements properly, we have to know how to use measuring instruments. Let's start with a simple measuring instrument: the ruler.

Suppose we wanted to measure the length of this purple line with an English ruler. We would make the measurement something like this:

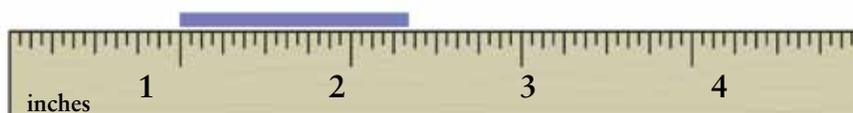


Illustration by Megan Whitaker

First, notice that we did not start the measurement at the beginning of the ruler. Instead, we lined up the ribbon with the first inch mark because it is slightly more accurate. It is very difficult to line up the edge of a ruler with the edge of the object you are measuring. This is especially true when the ruler is old and the edges are worn. So, the first rule for measuring with a ruler is to start at 1, not 0.

How would you read this measurement? First, you need to see what the scale on the ruler is. If you count the number of dashes between 1 inch and 2 inches, you will find that there are 15 of them. Every dash is worth one-sixteenth of an inch because 15 dashes break up the area between 1 inch and 2 inches into 16 equal regions.

Now that we know the scale is marked off in sixteenths of an inch, we can see that the ribbon is a little bigger than $1 \frac{5}{16}$ of an inch. Is that the best we can do? Of course not! Because the edge of the ribbon falls between $\frac{5}{16}$ ($\frac{10}{32}$) and $\frac{6}{16}$ ($\frac{12}{32}$) of an inch, we can estimate that it is approximately $\frac{11}{32}$. The proper length of the ribbon is $1 \frac{11}{32}$ inches. Generally, chemists do not like fractions in their final measurements, so we will convert $\frac{11}{32}$ into its decimal form to get a measurement of 1.34375 inches. Later on we will see that this measurement has far too many digits in it, but for now we will assume that it is okay.

Let's measure that same ribbon with a metric ruler:

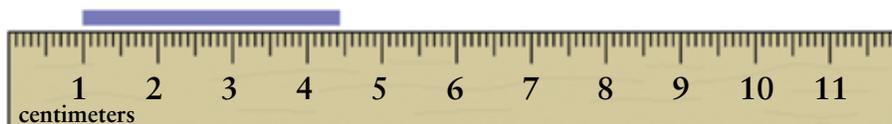


Illustration by Megan Whitaker

Now what measurement do we get? There are 9 small dashes between each cm mark; therefore, the scale of this ruler is one-tenth of a cm, or 0.1 cm. This is typical of metric

ON YOUR OWN

- 1.9 Should you be impressed if someone says she can hold her breath for 0.00555 hours? You must first convert this to seconds to answer. (HINT: 1 hour = 60 minutes and 1 minute = 60 seconds.)
- 1.10 How many cm^3 are in 0.0091 kL?
- 1.11 The area of a room is 32 m^2 . What is the area of the room in mm^2 ?

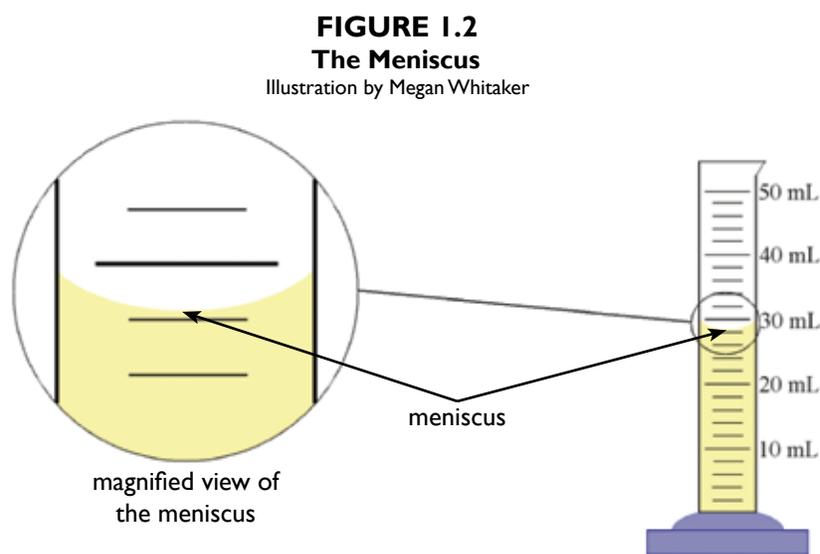
rulers. They are almost always marked off in tenths since all the prefixes in the metric system are multiples of 10. If you think about it, $1 \text{ mm} = 0.1 \text{ cm}$, so you could also say that each small dash is 1 mm. The ribbon is between 3.4 cm and 3.5 cm long. (The printing of your specific page could alter your measurement slightly.) Using our method of approximating between the dashes, we would say that the ribbon is 3.41 cm long.

How do we deal with the part of the line between the dashes? Here is a rule you will need to follow with any scientific instrument. You will always estimate one digit beyond what the instrument is marked. We can say 3.41 cm is the length of the line. Our instrument is marked off in 0.1 cm, so we will always report the measurement to the nearest 0.01 cm. What if you measure 3.43 cm and someone else measures 3.41 cm? There is nothing wrong with this because the last digit of any measurement is an estimate.

Whenever you are using a measuring device marked with a scale, be sure to use it the way we have used the rulers here. First, determine what the dashes on the scale mean. Then, try to estimate between the dashes if the object you are measuring does not exactly line up with a dash. That gives you as accurate a measurement as possible. You should always strive to read the scale to *the next decimal place* if possible. For example, on a metric ruler, the scale is marked off in 0.1 cm, so you should read the ruler to 0.01 cm, as discussed above.

One physical quantity that chemists measure frequently is volume because they spend a great deal of time mixing liquids. When chemists measure volume, one of the most useful tools is the **graduated cylinder**. This device looks a lot like a rain gauge. It is a hollow cylinder with markings on it. These markings, called graduations, measure the volume of liquid that is poured into the cylinder.

In the last experiment you will perform in this module, you will use a graduated cylinder (or a suitable substitute) to measure volume; so you need to know how to do this. When liquid is poured into a cylinder, the liquid tends to creep up the edges of the cylinder. This is because there are attractive forces between the liquid and the cylinder. Therefore, liquid poured into a graduated cylinder does not have a flat surface. Instead, it looks something like figure 1.2:



The curved surface of the yellow liquid is called the **meniscus** (muh nis' kus). To determine the volume of the liquid in any graduated cylinder, you must read the level of the liquid from the bottom of the meniscus. On the graduated cylinder in figure 1.2, there are 4 dashes between each marking of 10 mL, splitting the distance between the 10 mL marks into 5 divisions. This means that each dash must be worth 2 mL. If you look at the bottom of the meniscus in the figure, you will see that it is between 28 and 30 mL. Is the volume 29 mL then? No, not quite.

In the 2 examples we considered before, it was hard to guess how far between the dashes the object's edge was because the dashes were very close together. In this example, the dashes are farther apart, so we can be a bit more precise in our final answer. For the volume to be 29 mL, the bottom of the meniscus would have to be exactly halfway between 28 and 30. Clearly, the meniscus is much closer to 28 than to 30. So the volume is really between 28 and 29, and probably a little closer to 28. We would estimate that the proper reading is 28.3 mL. It could be as low as 28.1 or as high as 28.5, so 28.3 is a good compromise. Remember, you need to try to read the scale to the next decimal place. Since the scale is marked off in increments of 2 mL, you must try to read the answer to 0.1 mL. Experiment 1.3 will give you some practice at this kind of estimation.

ACCURACY, PRECISION, AND SIGNIFICANT FIGURES

Now that we've learned a bit about taking measurements, we need to discuss when measurements are good and when they are not. In chemistry, there is always some uncertainty in the value of a measurement. We can describe the measurement in 2 ways—its accuracy and its precision. Even though these words are used interchangeably in daily life, there's an important distinction between them.

Accuracy—An indication of how close a measurement is to the true value.

Precision—An indication of the scale on the measuring device that was used.

In other words, the more correct a measurement is, the more accurate it is. The smaller the scale on the measuring instrument, the more precise the measurement.

So let's look again at the ribbon measurement as an example. Suppose we used a ruler with a scale marked off in 0.1 cm and measured a length of 3.45 cm. What does that number mean? It means that the length of the ribbon, as far as we could tell, was somewhere between 3.445 and 3.455 cm long. We could not determine the length of the ribbon any better than that because our ruler was not *precise* enough to do any better.

On the other hand, if we found a ruler with a scale marked in increments of 0.01 cm, we could get a more precise measurement. For example, we could get a measurement of 3.448 cm because estimating between the dashes gives us one more decimal place. Since the ruler is marked off in hundredths, we can get a measurement out to the thousandths place. That extra digit in the thousandths place nails down the length better. It would be impossible to obtain so precise a measurement from the ruler we used in the example above, so the new ruler provides a more precise measurement. These examples show that the precision of our measurements depends completely on the measuring devices we use. The smaller the scale on the instrument, the more precise our measurement will be.

However, suppose we used the second ruler improperly. Maybe we read the scale

incorrectly or didn't line up the ribbon to the ruler very well and got a measurement of 3.118 cm. Even though this measurement is more *precise* than the one we made with the first ruler, it is significantly less *accurate* because it is way off of the correct value of 3.448 cm. The *accuracy* of our measurement depends on how carefully and correctly we use the measuring device. In other words, a measurement's *precision* depends upon the instrument, whereas a measurement's *accuracy* depends upon the person doing the measurement.

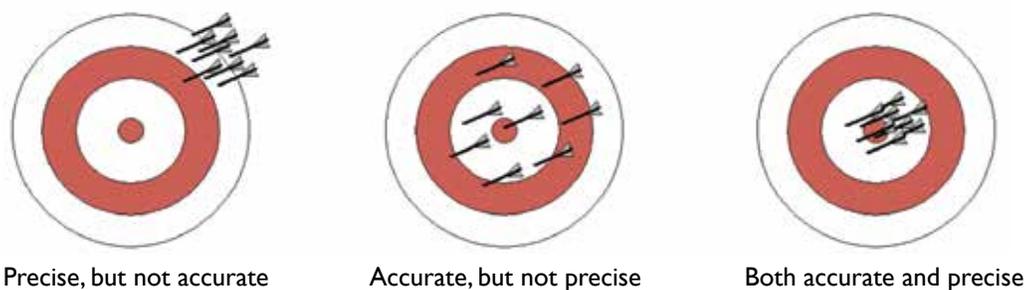
Since a measurement's precision depends on the instrument used, the only way to improve precision is to get a better, more precise instrument. However, there are other ways to improve accuracy. First, make sure you understand the proper methods of using each instrument. Second, practice making measurements, which will help your skill and your accuracy.

The most practical way to improve your accuracy in measurement is to make your measurement several times and average the results. This averages out all the little differences that can occur between measurements, even while using the same instrument. An even better way of assuring accuracy is to have several *different* people make the measurements and average all their answers together. The more individual measurements are made, the more accurate the average of them will be.

Let's illustrate this whole idea of accuracy and precision in another way. Suppose you were throwing darts at a dart board. Figure 1.3 shows 3 possible outcomes that we will discuss.

FIGURE 1.3
Accuracy and Precision

Illustration by Megan Whitaker



The first target on the left has all the darts clumped together, but they are way off the bull's-eye. If we made several measurements with a precise device but used it wrongly every time or the device had a flaw, it would give us numerous measurements that were very close to one another but far from the true value. This example shows a lot of precision, but not much accuracy.

The middle target has the darts surrounding the bull's-eye, but they are far from one another. If we used a measuring device that is not very precise but used it correctly, it would give us similar results. This may be because we had a hard time estimating between the marks on the scale of the instrument. This example shows accuracy, but not much precision.

The last target on the right has all darts clumped together, and they are on the bull's-eye. This is an example of measurements that are both accurate and precise. They are accurate because they average out to the correct value of the bull's-eye, and they are precise because they are very close to one another, indicating that a precise measuring device was used. This example is what every chemist desires when making measurements.

Since accuracy and precision are very important, we need to know how to evaluate

the accuracy and precision of our measurements. The way to determine the accuracy of a measurement is to compare it to the correct value. If you have no idea what the correct value is, determining your measurement's accuracy is difficult. It is not impossible, but we will not spend time in this course on this topic.

Determining the precision of a measurement is quite easy. To determine the precision of a measurement and an instrument, you merely need to look at its **significant figures**. But what is a significant figure? It is a digit that is read from an instrument. That purple line was 3.41 cm long. All of those digits came from the ruler, so that measurement has 3 significant figures. The volume of liquid in the graduated cylinder was 28.3 mL. Again, all of those digits came from the instrument, so they are significant figures. A significant figure is a measured digit!

If 2 instruments measure the same thing, the one which gives a significant figure in the smallest decimal place is the more precise instrument. For example, 2.545 cm is more precise than 2.54 cm because 2.545 cm has more significant figures. There are 4 rules you need to remember to determine the number of significant figures:

1. All nonzero digits are significant.
2. All zeros in front of the first 1–9 digit are not significant.
3. All zeros between 2 significant figures are significant.
4. All zeros at the end of a number *and* to the right of the decimal point are significant.

TABLE 1.4
Examples of Significant Figure Rules

Number	Example	Number of Significant Figures
Nonzero digits	9,341	4
Zeros in front of first 1–9 digit	0.000564	3
Zero between 2 significant figures	120.043	6
Zero at end of number <i>and</i> to right of decimal point	510.0 510	4 2

Counting significant figures is very important in science and in our ability to understand measurements, so you must have a firm grasp on this concept. Read through example 1.5 and follow the logic. After that, try “On Your Own” question 1.12.

EXAMPLE 1.5

Count the significant figures in each of the following numbers:

- a. 3.234 b. 6.016 c. 105.340 d. 0.00450010 e. 2,330

- (a) Since every digit is nonzero, every digit is a significant figure (first rule). This means there are **4 significant figures** in this number.
- (b) Three of the digits are nonzeros and therefore are significant figures (first rule). The 0 is also significant because it is between 2 significant figures (second rule). So there are **4 significant figures** in the number.

- (c) Four of the digits are nonzeros and therefore are significant figures (first rule). The 0 between the 1 and 5 is significant because it's between 2 significant figures (second rule). The last 0 is also significant because it is at the end of the number and to the right of the decimal point (third rule). The answer is there are **6 significant figures** in the number.
- (d) The 3 nonzero digits are all significant figures, as are the zeros between the 5 and 1. The 0 at the end is also a significant figure. However, the first 3 zeros are not significant because they are not between 2 significant figures and they are not at the end of the number to the right of the decimal. They are important, but they are not measured. There are **6 significant figures** in the number.
- (e) The 3 nonzero digits are significant figures. The zero is not significant because it is before, not at the end of, the number to the right of the decimal. There are **3 significant figures** in the number.

ON YOUR OWN

1.12 How many significant figures are in the following measurements?

- | | |
|--------------|---------------|
| a. 3.0220 cm | d. 61.054 kg |
| b. 0.0060 m | e. 300,000 mm |
| c. 1.00450 L | |

As we have seen, the precision of our instrument determines the number of significant figures we can report in a measurement. In our graduated cylinder example, we decided that we could reasonably approximate our measurement to somewhere between 28 and 29 mL. If we can do that, then we can report our

answer to the nearest tenth of a mL, giving us a 3-significant-figure answer of 28.3 mL. If, instead, our graduated cylinder had been marked off in tenths of a mL, we could probably approximate the measurement to somewhere between 28.2 and 28.3 mL, allowing us to have a 4-significant-figure answer like 28.26 mL.

In the example which used an English ruler, we said that our final answer, 1.34375 inches, had too many digits in it. Now hopefully you can see why. According to this number, our ruler was precise enough to measure distance of one-hundred-thousandth of an inch! That's far too much precision. Most English rulers are, at best, precise to 0.01 inches. Therefore, the proper length of the ribbon that you should report is 1.34 in.

Suppose we had another ribbon to measure:



Illustration by Megan Whitaker

How would we report its measurement? Would we say that this ribbon is 3 cm long? Actually, that measurement is not quite right. The ribbon does seem to end right on the 4 cm line, so there is no need to do any approximations here. Why, then, is 3 cm a wrong answer for the length of the ribbon?

The problem with reporting the length of the ribbon as 3 cm is that we are not being as precise as we can be. Since the ruler's scale is marked off in 0.1 cm, we can safely report our answers to the hundredths of a cm. If the ribbon's edge had fallen between 2 dashes,

we could have approximated as we did above. Therefore, the precision of the ruler is to the hundredths place. Therefore, if the object's edge falls right on one of the dashes, do not throw away precision. Report this length as 3.00 cm. This tells someone who reads the measurement that the ribbon's length was measured to a precision in the hundredths of a cm.

If you report the length as 3 cm, that means the ribbon could be as short as 2.5 cm or as long as 3.4 cm. Both of those measurements round to 3. But the ruler we used was much more precise. It determined the length of the ribbon to be 3.00 cm. So the ribbon is somewhere between 2.995 and 3.004 cm long. Keeping all of the significant figures that you can is a very important part of doing chemistry experiments. You will get some practice at this in experiment 1.3.

Reporting the precision of a measurement is just as important as reporting the number itself. Why? Let's suppose that the No-Weight Cookie Co. just produced a diet cookie that they claim has only 5 calories per cookie. To confirm this claim, researchers did several careful experiments and found that, in fact, there were 5.4 calories per cookie. Does this result mean that the No-Weight Cookie Co. lied about the number of calories in its cookies? No. When the company reported that there are 5 calories per cookie, the precision of their claim indicated that there could be anywhere from 4.5 to 5.4 calories per cookie. The researchers' finding was more precise than the company's claim; nevertheless, the company's claim was accurate.

SCIENTIFIC NOTATION

Since reporting the precision of a measurement is so important, we need a notation system that allows us to do this no matter what number is involved. As numbers get very large, it becomes more difficult to report their precision properly. For example, suppose we measured the distance between 2 cities as 100.0 km. According to our rules of precision, reporting 100.0 km as the distance means that our measuring device was marked off in units of 1 km, and we estimated between the marks to come up with 100.0 km. However, suppose our measurement wasn't that precise. Suppose the instrument we used could determine the distance only to within 10 km? How could we write down a distance of 100 km and indicate that the precision was only to within 10 km?

The answer to this question lies in the technique of **scientific notation**. In scientific notation, we write numbers so that no matter how large or how small they are, they always include a decimal point. Remember, a number can be represented in many different ways. The number 4, for example, could be written as 2×2 or 4×1 or simply 4. Each one of these is an appropriate representation of the number 4. In scientific notation, we always represent a number as a something times a power of 10. For example, 50 could be written in scientific notation as 5×10 . The number 150 could be written as 1.5×100 or 15×10 .

Do you see why this helps us in writing down the precision of our original measurement? Instead of writing the distance as 100 km, we could write it as 1×100 or 1.0×100 or 1.00×100 . How does this help? According to our rules of significant figures, the 0 in 1.0 is significant because it is at the end of the number and to the right of the decimal. By writing down our measurement as 1.0×100 , we indicate that the 0 was measured and that the measurement is precise to within 10 km. There is no way to do that with normal decimal notation because neither of the zeros in 100 is significant. Scientific notation, then, gives us a way to *make* zeros significant if they need to be. If our measurement of 100 km was precise to within 1 km, we could indicate that by reporting

EXAMPLE 1.6

Convert the following numbers into scientific notation:

- a. 20,300 b. 3,151,367 c. 234,000 d. 0.000002340 e. 0.000875

- (a) The decimal place must be moved to the left by 4 digits to get it next to the 2. Since we are dealing with a big number, we have to multiply by a 10 raised to a positive power. Therefore, the answer is 2.03×10^4 . Since the last 2 zeros to the right of the 3 are *not* significant as the number is written, we must drop them in our answer. All zeros are significant in 2.0300×10^4 .
- (b) The decimal place must be moved to the left 6 places and the number is big, so the answer is 3.151367×10^6 .
- (c) The decimal place must be moved to the left 5 places, and since it is a big number, the answer is 2.34×10^5 . The last 3 zeros were dropped because as written, they are not significant.
- (d) The decimal must be moved 6 places to the right. Since this is a small number, we are dealing with a negative exponent, so the answer is 2.340×10^{-6} . In this case, the final 0 cannot be dropped because, based on our rules of significant figures, a 0 at the end of a number and to the right of the decimal point is significant.
- (e) The decimal point must be moved 4 places to the right. Since it is a small number, the answer is 8.75×10^{-4} .

Convert the following numbers from scientific notation back into decimal form.

- (a) 3.45×10^{-5} (b) 2.3410×10^7 (c) 1.89×10^{-9} (d) 3.0×10

- (a) Since the 10 is raised to a negative power, the decimal point must be moved to make it small. The power of -5 tells us that we move it 5 spaces, so the answer is **0.0000345**.
- (b) Since the power of 10 is positive, we must move the decimal point to make the number bigger. The power of 7 tells us we must move it 7 places, so the answer is **23,410,000**. Note that we cannot indicate that the 0 after the 1 is significant in this notation. It is clearly significant in the original number, so it is impossible to properly represent the precision of this number in decimal form.
- (c) We must move the decimal point 9 places and make the number smaller, so the answer is **0.00000000189**.
- (d) Since the exponent is not listed, we assume it's 1. That means that we move the decimal point 1 place so that the number gets bigger, so the answer is **30**. Once again, there is no way to indicate that the 0 is significant, as it is in the original number.

ON YOUR OWN

1.13 Convert the following numbers from decimal form to scientific notation.

- a. 26,089,000 c. 0.00009870
b. 12,000,000,003 d. 0.980

1.14 Convert the following numbers from scientific notation to decimal form.

- a. 3.456×10^{14} c. 3.45×10^{-5}
b. 1.2341×10^3 d. 3.10×10^{-1}

USING SIGNIFICANT FIGURES IN MATHEMATICAL PROBLEMS

Now that we have the ability to write down any measurement with its proper precision, there is only one more topic on significant figures that we need to discuss. We need to know how to use our concepts of significant figures when we work mathematical problems. Suppose we had 2 measurements and wanted to add them together. Since each measurement has its own precision, the final answer would also

have a certain precision. How do we know the precision of our answer?

For example, suppose we measured the total length of a knife to be 25.46 cm. Later, someone else measured the length of the knife handle with a less precise ruler and got 7.8 cm. If we wanted to determine the length of the knife's blade, either we could measure it, or we could say that the blade's length was the total length of the knife minus the length of the handle, or $25.46 \text{ cm} - 7.8 \text{ cm}$. If we do the subtraction, we get 17.66 cm. This answer is too precise because the knife handle was measured with a less precise ruler. The answer is limited to the least precise instrument, so the proper answer is 17.7 cm.

To add, subtract, multiply, or divide measurements, we will use 2 rules about using significant figures in mathematical equations. You will be using these rules over and over again throughout this course, so you will be expected to know them:

1. *Adding and Subtracting with Significant Figures:* When adding and subtracting measurements, round your answer so that it has the same precision as the least precise measurement in the calculation.
2. *Multiplying and Dividing with Significant Figures:* When multiplying and dividing measurements, round the answer so that it has the same number of significant figures as the measurement with the fewest significant figures.

Example 1.7 shows how these rules work.

EXAMPLE 1.7

A student measures the mass of a jar that is filled with sand and finds it to be 546.2075 kg. A note on the jar says, "When empty, this jar has a mass of 87.61 kg." What is the mass of the sand in the jar?

Since 546.2075 kg is the mass of both the jar and the sand, and since 87.61 kg is the mass of the jar alone, the mass of the sand must be the difference between the 2:

$$\begin{array}{r} 546.2075 \text{ kg} \\ - 87.61 \text{ kg} \\ \hline 458.5975 \text{ kg} \end{array}$$

However, since the precision of the jar's mass only goes out to the hundredths place, that's the best we can do in our final answer. Therefore, the mass of the sand is **458.60 kg**. Note that this number has more significant figures than 87.61. That doesn't matter because in addition and subtraction, we do not count significant figures; we look only at precision.

A woman runs 3.012 miles in 0.430 hours. What is her average speed?

We can find her average speed by dividing the distance traveled by the time:

$$\text{Speed} = 3.012 \text{ miles} \div 0.430 \text{ hour} = 7.004651163$$

The 3.012 miles has 4 significant figures, while 0.430 hours has 3. Therefore, our final answer must have 3 significant figures, making it **7.00 miles/hour**

Now that we have learned these rules, you will be expected to use them *in all further mathematical operations!* Whether you are working an “On Your Own” problem, a practice problem, a test problem, or an experiment, you will use these rules. In the examples and answers for all previous problems, these rules have not been followed, but they will be from now on. By the time you finish the next couple of modules, keeping track of significant figures and precision should be second nature to you.

There is one more point that you must understand about significant figures before you get some practice using the rules. When making unit conversions, you might be tempted to round everything to 1 significant figure because of the conversion relationships. For example, when you convert 121 g into kg, you use the following equation:

$$\frac{121 \text{ g}}{1} \times \frac{1 \text{ kg}}{1,000 \text{ g}}$$

Note that the 1 kg, the 1 in the denominator of the first fraction, and 1,000 g all look like they have only 1 significant figure. You might be tempted to round your answer to 1 significant figure. However, that would not be correct. The reason is simple: These 3 numbers all come from definitions. They are infinitely precise. The 1 kg is really 1.000... kg, and the 1,000 g is really 1,000.000... g because *exactly* 1 kilogram is defined to be *exactly* 1,000 g. In the same way, the 1 on the bottom of the first fraction is really 1.000... because it is an integer. The *only* number in this equation that has a limited number of significant figures is the 121 g (it is a measurement), so the answer is 0.121 kg.

In general, then, the prefixes used in the metric system as well as the integers used in fractions are infinitely precise and have an infinite number of significant figures. As a result, we ignore them when determining the significant figures in a problem. This is a very important rule:

The definitions of the prefixes in the metric system and the integers used in fractions are not considered when determining the significant figures in the answer.

Get some practice making measurements, using them in mathematical equations, and keeping track of significant figures by performing experiment 1.3.

EXPERIMENT 1.3

PURPOSE: To compare conversions to measurements.

MATERIALS:

- Book (not oversized)
- Metric and English rulers
- Safety goggles

QUESTION: How do measurements compare to conversions?

HYPOTHESIS: Write a hypothesis about how close you expect your conversions to be to measurements.

PROCEDURE:

1. Lay the book on a table and measure its length in inches. Read the ruler as shown in the measurement section above, estimating any answer that falls between the markings on the scale. Next, convert the fraction to a decimal (as we did in the measurement section above) and round it to the hundredths place because that's the precision of an English ruler.
2. Measure the width of the book in the same way.
3. Now that you have the length and width measured, multiply them together to get the surface area of the book. Since you are multiplying inches by inches, your area unit should be in^2 . Remember to count the significant figures in each of the measurements and round your final answer so that it has the same number of significant figures as the measurement with the least number of significant figures.
4. Use the relationship given in table 1.3 to convert the length measurement into cm. Do the same thing to the width measurement, making sure to keep the proper number of significant figures. Note that the relationship between inches and centimeters is exact. The 2.54 cm should not be taken into account when considering the significant figures because 1 inch is exactly 2.54 cm.
5. Use the metric ruler to measure the length and width of the book in centimeters. Once again, do it as shown in the measurements section above. If the scale of the ruler is marked off in 0.1 cm, then your length and width measurements should be written to the hundredths of a centimeter. Compare these answers to the length and width you calculated by converting from inches. They should be nearly the same. If they are different by only a small percentage, there is no problem. However, if they differ by more than a small percentage, recheck your measurements and your conversions.
6. Multiply the length and width measurements you took with the metric ruler to calculate the surface area of the book in cm^2 . Use the relationship between inches and centimeters to convert your answer into in^2 . Remember, since you are using a derived unit, the conversion is more complicated. You might want to review example 1.4.

7. Now compare the converted value for the surface area to the one you calculated in step 3 using your English measurements. Once again, they should be equal or close to equal. If not, you have either measured wrongly or made a mistake in your conversion.
8. Clean up and return everything to the proper place.

CONCLUSION: Write something about how well you made your measurements.

MEASURING TEMPERATURE

In chemistry, we will be measuring and making calculations with temperature, so we need to know what units are used in the measurement. The temperature unit that you are probably most familiar with is **Fahrenheit** (abbreviated as °F). We hear this unit used often by weather reporters talking about tomorrow's weather and how warm or cold it will be. Although this is a very common temperature unit, it is not used by chemists. Instead, chemists use one of 2 other temperature units: **Celsius** (sel' see us; abbreviated °C) or **Kelvin** (kel' vuhn; abbreviated K). First we need to see how these units are defined, and then we will see how they relate to Fahrenheit and why chemists use them.

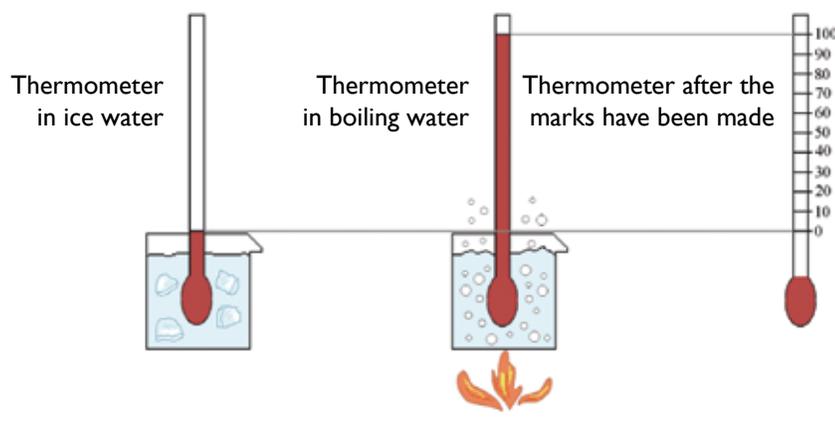
When we measure temperature, we are measuring how much the liquid within the thermometer is expanding. Therefore, we must find a way to relate that measurement to a unit which means temperature. We do this in the following way:

1. Immerse a thermometer in a mixture of ice and water.
2. Make a mark where we see the liquid in the thermometer and assign a value to that mark. In the Celsius temperature scale, we call it exactly 0°C. In the Fahrenheit scale, we give that mark a value of exactly 32°F.
3. Immerse the thermometer in a pot of boiling water.
4. Make a mark where we see the liquid in the thermometer and assign it a value of exactly 100°Celsius or exactly 212°Fahrenheit.
5. Divide the distance between the 2 marks into equal divisions. Then we have a temperature scale.

This method for defining a temperature scale is illustrated in figure 1.4.

FIGURE 1.4
Making a Celsius Thermometer

Illustration by Megan Whitaker



This process of using certain physical measurements to define the scale of a measuring device is called calibration. This particular calibration makes use of a surprising fact in chemistry:

If ice and water are thoroughly mixed, the temperature of the mixture will stay the same (0.0°C or 32.0°F), regardless of the amount of ice or water present.

This might sound rather surprising, but it is true. Even though you might think that a little water with a lot of ice is colder than a lot of water with a little ice, they are actually the same temperature! Equally surprising is this fact:

Boiling water is always at the same temperature (100.0°C or 212.0°F at standard atmospheric pressure) whether it is boiling rapidly or hardly boiling at all.

Now that we know how the Celsius temperature unit is defined, we can learn how it relates to the Fahrenheit unit. It makes sense that Fahrenheit and Celsius relate to one another since they both measure the same thing: temperature. They are related by a very simple equation:

$$^{\circ}\text{C} = \frac{5}{9} (^{\circ}\text{F} - 32) \quad \text{Equation 1.1}$$

In this equation, °C represents the temperature in degrees Celsius, and °F stands for the temperature in degrees Fahrenheit. So if you must use a Fahrenheit thermometer in your experiments, you can use this equation to convert your measurements into Celsius. Please note one very important thing about this equation: The 5, 9, and 32 are all exact. Therefore, you need not consider their significant figures. They have infinite precision and an infinite number of significant figures. The only significant figures you must consider are those of the original measurement. Using this equation, we will report our answer using the multiplication rule even though there is a subtraction in the equation. Example 1.8 will help show you how this is done.

EXAMPLE 1.8

A student uses a Fahrenheit thermometer to do a chemistry experiment but then must convert his answer to Celsius. If the temperature reading was 50.0°F, what is the temperature in Celsius?

To solve this one, we simply use equation 1.1:

$$^{\circ}\text{C} = \frac{5}{9} (50.0 - 32)$$

$$^{\circ}\text{C} = 10.0$$

There are 3 significant figures in the original measurement. Since the other numbers in this equation are exact, the answer must have 3 significant figures. Therefore, the answer is **10.0°C**.

We usually say that room temperature is about 25°C. What is this temperature in Fahrenheit?

To solve this one, we must first rearrange equation 1.1 using algebra. Once we do this, we get the following equation:

$$^{\circ}\text{F} = \frac{9}{5} (^{\circ}\text{C}) + 32$$

$$^{\circ}\text{F} = \frac{9}{5} (25) + 32$$

$$^{\circ}\text{F} = 77$$

The presence of only 2 significant figures in the original number means only 2 significant figures in the end; therefore, the temperature is **77°F**.

Try “On Your Own” questions 1.15–1.16 to see whether or not you fully understand this type of conversion.

What about the other unit mentioned earlier? The Kelvin temperature unit is a special unit that we will use quite a bit in later modules. It is special because we can never reach a temperature of 0 Kelvin or lower, for reasons we will see in a later module. This fact makes the Kelvin temperature scale different from most others. After all, anything colder than ice water has a negative temperature in Celsius units. This means that temperatures less than 0 are quite common in the Celsius scale. Although not quite as common, it is possible to reach temperatures less than 0 in the Fahrenheit scale as well. It is impossible for anything in nature to reach 0 Kelvin or below. Since we can never get to or go below 0 Kelvin, the Kelvin temperature scale is often called an **absolute temperature scale**.

Once we have a temperature in units of degrees Celsius, converting it to Kelvin is simple. All we do is add 273.15 to the measurement. In mathematical terms, we would use this equation:

$$\text{K} = ^{\circ}\text{C} + 273.15 \quad \text{Equation 1.2}$$

K is the temperature in units of Kelvin, and °C is the temperature in units of Celsius. In this equation, the 273.15 is *not* exact. Its precision plays a role. Note that since this equation involves adding, we use the rules of addition and subtraction when determining the significant figures involved. Those rules are *different* from the ones for multiplication and division, so be aware of that. Example 1.9 will show you how to use this equation.

ON YOUR OWN

- 1.15 Normal body temperature is 98.6°F. What is this temperature in Celsius?
- 1.16 Rubbing alcohol boils at 180.5°C. What is the boiling temperature of water in Fahrenheit?

EXAMPLE 1.9**What is the boiling temperature of water in Kelvin?**

This conversion is a snap. We just realize that water boils at 100.0°C . If we put that temperature into equation 1.2, we get this:

$$\text{K} = 100.0 + 273.15 = 373.15$$

The original temperature goes out to the tenths place, while 273.15 goes out to the hundredths place. The rules for significant figures in adding tell us that the answer must have the same precision as the least precise number in the equation. Therefore, our final answer is **373.2 K**.

The lowest temperature that has ever been recorded in the United States of America is -80.0°F . What is this temperature in Kelvin?

Since the only way we can get to Kelvin is by adding 273.15 to the temperature in Celsius, we must first convert $^{\circ}\text{F}$ into $^{\circ}\text{C}$:

$$^{\circ}\text{C} = \frac{5}{9} (-80.0 - 32)$$

$$^{\circ}\text{C} = -62.2$$

Now that we have the answer in $^{\circ}\text{C}$, we can easily convert to Kelvin:

$$\text{K} = -62.2 + 273.15 = 210.95$$

Our final answer is **211.0 K**. Our rules for adding tell us that the precision must be kept to the tenths place because that is the same precision as the least precise number in the equation. That's why our final answer goes out to the tenths place. So you see, even very, very cold temperatures in the Celsius and Fahrenheit temperature scales are still rather large numbers in the Kelvin temperature scale!

ON YOUR OWN

1.17 What is the Fahrenheit equivalent of 0.00 Kelvin? (Use 3 significant figures for this measurement.)

Now cement your knowledge of temperature conversions with “On Your Own” question 1.17.

THE NATURE OF A SCIENTIFIC LAW

One way to approach chemistry or any other science is to look around you and

try to think of logical explanations for what you see. You would certainly observe, for instance, that different substances have different forms and appearances. For example, some substances are gases, some are liquids, and some are solids. Some are hard and shiny, but others are soft and dull. You would also observe that different substances behave differently. Iron rusts, but gold does not; copper conducts electricity, but sulfur doesn't. How can you explain these and a vast number of other observations?

God made our world far too complex for us to understand by looking and thinking alone. We need to ask specific questions and conduct experiments to find answers. Scientists develop laws through experimentation and observation. After experimenting on or observing some facet of nature, they formulate a hypothesis to explain their observations. A **hypothesis** is no more than an educated guess that attempts to explain some aspect of the world around us. For example, when early scientists observed rotting meat, they always saw maggots crawling around on it. This led them to form the hypothesis that maggots are created from rotting meat.

Once a hypothesis has been formulated, scientists test it with more rigorous experiments. For example, after forming the hypothesis that maggots are created from rotting meat, early scientists did experiments to make sure that the maggots were not coming from something else. They would put rotting meat on a shelf high in the air to make sure that no maggot could crawl up to it. Even when the rotting meat was put high in the air, maggots still appeared on it. To early scientists, such experiments confirmed their hypothesis. Although there was no way for maggots to crawl up to the rotting meat, they did indeed appear on it. Many similar experiments convinced early scientists that their original hypothesis was correct.

Once a hypothesis is confirmed by more rigorous experimentation, it is considered a **theory**. After numerous experiments, the theory may be considered a **scientific law**. A scientific law is really nothing more than an educated guess that has been confirmed over and over again by experimentation. The problem with putting too much faith in a scientific law is that the experiments that established it might be flawed, making the scientific law itself flawed.

science and creation

We study science to learn more about creation and, ultimately, the Creator. God is the one who holds it all in His hands and uses it for His glory. Early scientists who were experimenting with rotting meat and maggots called their theory the theory of spontaneous generation. As the centuries passed, many more experiments were done to test the theory. Those experiments seemed to support the idea that life, such as maggots, could be created from nonlife, such as rotting meat. All of the experiments done to confirm this theory, however, were flawed. For example, Francesco Redi, an Italian physician, showed that if the rotting meat was completely isolated from the outside world, no maggots would appear; however, microscopic organisms did. French scientist Louis Pasteur eventually performed careful experiments that overturned the theory of spontaneous generation. His work showed that even microscopic organisms could not arise from nonlife but came to the meat by dust particles that blew in the wind.

The point of this story is to illustrate that when you read about scientific results, you must keep in mind that scientific theories are not laws of nature and can never be absolutely proven. There is always a chance that a new experiment might give results that can't be explained by a present theory. All a theory can do is provide the best explanation available at the time. If new experiments uncover results that present theories can't explain, the theories will have to be modified or perhaps even replaced. Science experiments are a good way to try to understand the nature of God's creation, but they are not absolute truth. Remember that.

salt and light

Bishop Robert Grosseteste (1175–1253) was an English statesman, philosopher, theologian, and scientist. He was one of the first scientists to establish a framework for what would later become the scientific method. Besides being a great scientist, Grosseteste was a strong Christian. He taught that the purpose of inquiry was not to come up with great inventions, but instead to learn the reasons behind the facts. In other words, he wanted to explain why things happened the way they did. That's the essence of science. He said that "just as the light of the sun irradiates the organ of vision and things visible, enabling the former to see and the latter to be seen, so too the irradiation of a spiritual light brings the mind into relation with that which is intelligible" (Stevenson 1899, 52).



Stained glass window by William Morris (to the designs of Burne-Jones) [Photo: public domain]

EXPERIMENTATION AND THE SCIENTIFIC METHOD

As we go about studying chemistry, you will conduct experiments to help you understand the concepts being presented in this course. You should use the scientific method to document the results of your experiments. The scientific method includes techniques for investigating, acquiring new knowledge, or correcting and integrating previous knowledge.

1. **Purpose:** What is your goal? What do you want to answer? Example: I would like to know what phase water is at room temperature.
2. **Hypothesis/Prediction:** What do you think will happen in this experiment? Example: I believe that water is liquid at room temperature.
3. **Experiment:** Test the hypothesis with systematic observations, measurements, and laboratory techniques.
4. **Analysis:** Determine what the results of the experiment show and decide on the next actions to take.
5. **Conclusion:** Describe what the results of the experiment show and the next actions to take.

As we do experiments, we will use this experimental method to document our findings.

Now that you have read this module and answered the "On Your Own" questions, it is time for you to shore up your new skills and knowledge with the practice problems and review questions at the end of this module. As you go through them, check your answers with the solutions provided and be sure you understand any mistakes you made. If you need more practice problems, you can find them in appendix B. Once you are confident in your abilities, take the test. If you do not score at least 70% on the test, then you should probably review this module before you proceed to the next one.

SUMMARY OF KEY EQUATIONS AND TABLES IN MODULE I

Equation 1.1: $^{\circ}\text{C} = \frac{5}{9} (^{\circ}\text{F} - 32)$ Converting $^{\circ}\text{F}$ to $^{\circ}\text{C}$

Equation 1.2: $\text{K} = ^{\circ}\text{C} + 273.15$ Converting $^{\circ}\text{C}$ to K

Table 1.1: Physical Quantities and Their Base Units

Table 1.2: Common Prefixes Used in the Metric System

Table 1.3: Relationships between English and Metric Units

Table 1.4: Examples of Significant Figure Rules

ANSWERS TO THE “ON YOUR OWN” QUESTIONS

$$1.1 \quad \frac{9,321 \cancel{\text{g}}}{1} \times \frac{1 \text{ kg}}{1,000 \cancel{\text{g}}} = 9.321 \text{ kg}$$

$$1.2 \quad \frac{0.465 \cancel{\text{L}}}{1} \times \frac{1 \text{ mL}}{1,000 \cancel{\text{L}}} = 465 \text{ mL}$$

$$1.3 \quad \frac{724.0 \cancel{\text{cm}}}{1} \times \frac{0.01 \text{ m}}{1 \cancel{\text{cm}}} = 7.240 \text{ m}$$

$$1.4 \quad \frac{8.465 \cancel{\text{st}}}{1} \times \frac{14.59 \text{ kg}}{1 \cancel{\text{st}}} = 123.5 \text{ kg}$$

$$1.5 \quad \frac{6.1236 \cancel{\text{L}}}{1} \times \frac{1 \text{ gal}}{3.78 \cancel{\text{L}}} = 1.62 \text{ gal}$$

$$1.6 \quad \frac{1,500 \cancel{\text{mL}}}{1} \times \frac{0.001 \cancel{\text{L}}}{1 \cancel{\text{mL}}} \times \frac{1 \text{ kL}}{1,000 \cancel{\text{L}}} = 0.0015 \text{ kL}$$

$$1.7 \quad \frac{2 \cancel{\text{km}}}{1} \times \frac{1,000 \cancel{\text{m}}}{1 \cancel{\text{km}}} \times \frac{1 \text{ cm}}{0.01 \cancel{\text{m}}} = 200,000 \text{ cm}$$

$$1.8 \quad \frac{.01 \cancel{\text{Mg}}}{1} \times \frac{1,000,000 \cancel{\text{g}}}{1 \cancel{\text{Mg}}} \times \frac{1 \text{ mg}}{0.001 \cancel{\text{g}}} = 10,000,000 \text{ mg}$$

$$1.9 \quad \frac{0.00555 \cancel{\text{hr}}}{1} \times \frac{60 \cancel{\text{min}}}{1 \cancel{\text{hr}}} \times \frac{60 \text{ sec}}{1 \cancel{\text{min}}} = 19.98 \text{ sec}$$

That is not a long time to hold one's breath. So we would not be impressed.

$$1.10 \quad \frac{0.0091 \cancel{\text{kL}}}{1} \times \frac{1,000 \cancel{\text{L}}}{1 \cancel{\text{kL}}} \times \frac{1 \text{ mL}}{0.001 \cancel{\text{L}}} = 9,100 \text{ mL} = 9,100 \text{ cm}^3 \text{ (mL and cm}^3 \text{ equivalent)}$$

1.11 The relationship between m and mm is easy:

$$1 \text{ mm} = 0.001 \text{ m}$$

To set up the conversion, we start with:

$$\frac{32 \text{ m}^2}{1} \times \frac{1 \text{ mm}}{0.001 \text{ m}}$$

This expression *does not cancel m²*. There is an m² on the top of the first fraction and only an m on the bottom of the second fraction. To cancel m² (which we must do to get the answer), we have to square the conversion fraction:

$$\frac{32 \text{ m}^2}{1} \times \left(\frac{1 \text{ mm}}{0.001 \text{ m}} \right)^2$$

Then we get:

$$\frac{32 \text{ m}^2}{1} \times \frac{1 \text{ mm}^2}{0.000001 \text{ m}^2} = 32,000,000 \text{ mm}^2$$

- 1.12 (a) All 3 nonzero digits are significant figures, as are both zeros. One 0 is between 2 significant figures, and the other is at the end of the number to the right of the decimal. There are **5 significant figures**.
- (b) The first 3 zeros are not significant because they are not between 2 significant figures. The 6 is a significant figure, as is the last 0 because it is at the end of the number to the right of the decimal. So there are **2 significant figures**.
- (c) All digits are significant figures here. The first 2 zeros are between significant figures, and the last 0 is at the end of the number to the right of the decimal. Therefore, there are **6 significant figures**.
- (d) All digits are significant figures. The 0 is between 2 significant figures. There are **5 significant figures**.
- (e) All the zeros are not significant in this number. **There is 1 significant figure**.

1.13 (a) 2.6089×10^7 (b) $1.2000000003 \times 10^{10}$
 (c) 9.870×10^{-5} (d) 9.80×10^{-1}

1.14 (a) 345,600,000,000,000 (b) 1,234.1
 (c) 0.0000345 (d) 0.310

1.15 To solve this one, we simply need to use equation 1.1:

$$^{\circ}\text{C} = \frac{5}{9} (^{\circ}\text{F} - 32)$$

$$^{\circ}\text{C} = \frac{5}{9} (98.6 - 32)$$

$$^{\circ}\text{C} = 37.0$$

Our measurement starts with 3 significant figures, and the other numbers in the equation are exact, so we must end up with 3 significant figures. The answer is **37.0°C**.

- 1.16 This one requires that we use algebra to rearrange equation 1.1 so that we can solve for °F:

$$^{\circ}\text{F} = \frac{9}{5}(^{\circ}\text{C}) + 32$$

$$^{\circ}\text{F} = \frac{9}{5}(180.5) + 32$$

$$^{\circ}\text{F} = 356.9$$

Since 180.5 has 4 significant figures and everything else in the equation is exact, our answer is **356.9°F**.

- 1.17 The only way we can convert to Fahrenheit is if we have a temperature in Celsius. Before we can get the answer, we must first convert 0.00 K to degrees Celsius by rearranging equation 1.2:

$$^{\circ}\text{C} = \text{K} - 273.15$$

$$^{\circ}\text{C} = 0.00 - 273.15 = -273.15$$

Since we are subtracting, we look at precision. The original measurement goes out to the hundredths place, as does 273.15. Therefore, our answer should go out to the hundredths place. Now we can convert to Fahrenheit:

$$^{\circ}\text{F} = \frac{9}{5}(^{\circ}\text{C}) + 32$$

$$^{\circ}\text{F} = \frac{9}{5}(-273.15) + 32$$

$$^{\circ}\text{F} = -459.67$$

Since 273.15 is the only number in the equation that is not exact, the answer must have the same number of significant figures. The answer is **-459.67°F**.

STUDY GUIDE FOR MODULE 1

REVIEW QUESTIONS

- Which of the following contains no matter?
 - A baseball
 - A balloon full of air
 - Heat
 - A light ray
- List the base metric units used to measure length, mass, time, and volume.
- In the metric system, what does the prefix *milli-* mean?
- All conversion factors, when in the form of a fraction, must equal _____.
- Which has more liquid: a glass holding 0.05 kL or a glass holding 12,000 mL?
- How long is the bar in the picture below?

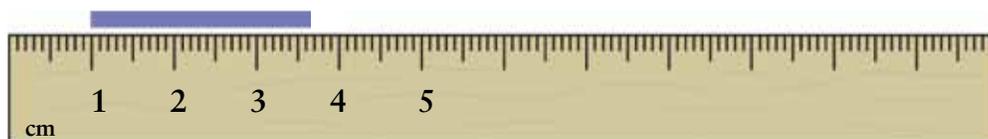


Illustration by Megan Whitaker

- Two students measure the mass of an object that is known to be 50.0 grams. The first student measures the mass to be 49.8123 grams. The second measures the mass to be 50.1 grams. Which student was more precise? Which student was more accurate?
- Explain what a significant figure is.
- How many significant figures are in the following numbers?
 - 120350
 - 10.020
 - 0.000000012
 - 7.20×10^2
- A student measures the mass of object A to be 50.3 grams and measures the mass of object B to be 200.24 grams. She then reports the combined mass to be 251 grams. Is this student correct? Why or why not?

11. What would be the units on the following calculations? You do not have to do the math since this question only wants to know the units.
 - a. $8 \text{ cm} + 2 \text{ cm} =$
 - b. $4 \text{ g} \div 2 \text{ mL} =$
12. Which answer for question 11 will be a derived unit?
13. What are the 2 basic rules for using scientific notation?
14. Which is colder: 50.0 grams of water at 0.00°C or 50.0 g of water at 32.00°F ?

PRACTICE PROBLEMS

Be sure to use the proper number of significant figures in all of your answers!

1. Convert 2.4 mL into L.
2. Convert 69.00 km into m.
3. Convert 0.091 kg into cg.
4. If an object has a volume of 69.2 mL, how many kL of space does it occupy?
5. A box is measured to be 23 cm by 45 cm by 38 cm. What is its volume in cubic meters?
6. A nurse injects 71.0 cc of medicine into a patient. How many liters is that?
7. Convert the following decimal numbers into scientific notation:
 - a. 12.45000
 - b. 3,040,000
 - c. 6,100.500
 - d. 0.001234
8. Convert the following numbers back into decimal:
 - a. 6×10^9
 - b. 3.0450×10^{-3}
 - c. 1.56×10^{21}
 - d. 4.50000×10^{-7}
9. Convert 85.6°C into Fahrenheit.
10. The temperature of the moon during its day is 396 K. What is that in Celsius? In Fahrenheit?
11. The average low temperature of International Falls, MN, in January is -7.0°F . What is that in $^\circ\text{C}$?



ATOMS AND MOLECULES

In the first module, we discussed matter. In this module, we will explore some of the most fundamental aspects of matter. We need to know what matter is composed of, why matter exists in such a wide variety of substances, and how matter is able to change from one substance to another.

Pretend you were given a radio and asked to tell what it is made of. What would you do? Most likely, you would begin to take it apart. You would find a way to remove the back cover and look inside. Of course, you would find a lot of things inside the radio, so you would probably have to take it apart piece by piece until you had broken it down into all of its individual components. You could then learn what each component was and make a list of everything that makes up the radio. By taking the radio apart, you would be able to learn about its composition.

first reactions

Adam was permitted to name things created by God. We continue to exercise that honor in science. Chemistry uses a periodic table of elements to display the properties of these basic elements. As you begin to understand the elements' physical and chemical properties, you will see that the interactions of these basic building blocks are the foundation for the rest of chemistry.

FIGURE 2.1
An Old Radio
© iStockphoto.com



Taking a radio apart tells you more than just what it's made of, however. If you wanted to learn how that radio works, you could learn how each individual component works. That knowledge would help you immensely in understanding how the entire radio work and help you gain some understanding about how other electronic devices, such as televisions, work. To truly understand the nature of something, we have to break it down into all of its parts. Once we have done that, we can learn about each individual component, and that will allow us to understand whatever thing we are investigating.

This has long been the chemist's method for trying to understand the nature of matter. Chemists believe that if you can break matter down into its fundamental components and learn everything there is to learn about those components, you will then know everything there is to know about matter. In this module, we will begin to explore how chemists over the ages have pursued this method and what they have learned as a result.

EARLY ATTEMPTS TO UNDERSTAND MATTER

During the earliest years of scientific investigation, ancient Greek philosophers thought that the matter inside a substance was continuous. By this they meant that you could divide any substance in half over and over again. You would eventually have to use tweezers and a very powerful magnifying glass, but no matter how small the substance became, you could still divide it in half again. In other words, they thought that substances were composed of long, unbroken blobs of matter. This concept was known as the **continuous theory of matter**.

About 4 centuries before Jesus Christ walked this earth, the Greek philosopher Democritus was walking toward the shore of a beach and was suddenly struck with an amazing thought. He noticed that from a distance, the sand along the shore looked like it was one long, continuous blob of yellow. As he walked closer to the beach, however, he could see that the sand was not at all continuous. Instead, it was made up of hundreds of thousands of individual particles called grains.

FIGURE 2.2

Rocky Ocean Beach

Wikimedia Commons, public domain



Democritus then looked out at the ocean and wondered to himself whether or not water was made exactly the same way. From the shore, the water seemed to be a huge

mass of unbroken matter. But suppose he could magnify the water. Would he eventually see grains of water? Could it possibly be that the continuous theory of matter was based on an illusion? Perhaps matter was not really continuous. Maybe it just appeared that way because he could not magnify it enough to see the individual particles that make it up. Democritus wrote a great deal about this idea, which is referred to as the **discontinuous theory of matter**. Instead of being continuous, he argued, matter is composed of tiny individual particles. Our eyes simply cannot see these particles, so the apparent continuity of matter is an illusion. Democritus's musings represent man's first attempt to think in terms of atoms.

Since no experiments could be imagined to test the validity of either the continuous or the discontinuous theory of matter, scientists could not come to any firm conclusions about which theory was more scientifically sound. Each theory had its fervent supporters and its ardent critics. Unfortunately, neither side had any empirical evidence (facts gained by experiment and observation) to support its theory, so scientific progress on this point stalled for many centuries.

THE LAW OF MASS CONSERVATION

Progress in understanding the nature of matter picked up in the late 1700s. During that time, Antoine Lavoisier (lah vwah zyay'), considered by some to be the founder of chemistry, developed a theory that stunned the scientific world. This theory is now considered a scientific law and is called **the law of mass conservation**.

The law of mass conservation—Matter cannot be created or destroyed; it can only change forms.

When you burn wood in the fireplace, the matter in the wood is not destroyed. Instead, it changes from wood into other substances. Today, we know that those substances are carbon dioxide, water, and ash.

This may not sound very surprising. However, to the chemistry community in Lavoisier's day, this was a startling revelation. You see, early chemists were guided by what they observed. Since the human eye cannot observe everything in nature, this led them to some very wrong conclusions.

For example, early chemists would watch wood burning in the fire and would see the wood slowly disappear. They were unable to see the carbon dioxide and water vapor forming from the wood because those substances are colorless gases. As a result, they simply thought that the matter in the wood was disappearing—that matter was being destroyed. Because their experiments were not accurate, early chemists developed scientific laws that were quite wrong.

Today we have developed methods that allow us to detect the presence of colorless gases, so we can measure the mass of the carbon dioxide and water vapor produced when wood burns. We can also determine that when wood burns, the oxygen in the air around the wood is used up. This is why a fire cannot burn unless there is a plentiful supply of oxygen. If we measure the mass of the wood and the oxygen before a fire starts and we trap all of the carbon dioxide and water vapor produced by the fire after it starts, we will find out that the mass of the wood and oxygen used by the fire is exactly equal to the

mass of the ash, water vapor, and carbon dioxide produced by the fire. Experiment 2.1 allows you to see that even when matter goes through substantial change, the total mass is not affected.

EXPERIMENT 2.1

PURPOSE: To investigate the conservation of mass.

MATERIALS:

- 100-mL beaker (A juice glass can be used instead.)
- 250-mL beaker (A juice glass can be used instead.)
- Watch glass (A small saucer can be used instead. It must cover the mouth of the 100-mL beaker or juice glass listed above.)
- Teaspoon
- Lye (This is commonly sold with the drain cleaners in hardware stores and supermarkets. Make sure that your bottle is labeled 100% lye or something similar. If you cannot find lye, you can order it online. It may be listed under its chemical name, sodium hydroxide. *Always use this chemical in a well-ventilated area.*)
- White vinegar
- Several leaves of red (often called purple) cabbage
- Water
- Small pot for boiling water
- Measuring cup
- Stove
- Mass scale
- Stirring rod (A spoon will work.)
- Safety goggles
- Rubber cleaning gloves

QUESTION: Is mass conserved during a chemical reaction?

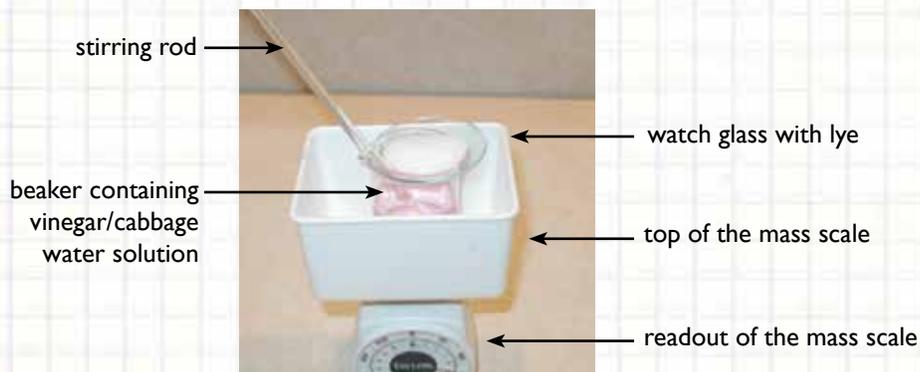
HYPOTHESIS: Propose an answer to the question.

PROCEDURE:

1. Add 2 cups of water to the pot.
2. Place the cabbage leaves into the water.
3. Place the pot on the stove and turn on the burner. Let the water boil for 5 minutes.
4. While you are waiting for the water to boil, use the scale to measure the mass of your beaker.
5. Add approximately 60 mL of vinegar to the beaker. The beaker should have volume levels marked, so you can just fill it to the 60 mL mark. You needn't add exactly 60 mL. Anywhere from 50 to 60 will be fine. If you are not using a beaker, add about $\frac{1}{4}$ cup. When chemists make measurements like these, where they don't need to get

an exact amount, they call them qualitative measurements. On the other hand, when chemists make measurements as exactly as possible, they call them quantitative measurements. Remember these terms because we will use them again.

6. Put on the gloves.
7. Measure 1 teaspoon of lye and put it on the watch glass. You can take the gloves off, but **whenever you are touching anything that has lye on it (as you will do later), put the gloves back on because lye is caustic and can cause chemical burns.**
8. Once the water in the pot has been boiling for 5 minutes, take the pot off the stove and allow it to cool for a few minutes.
9. Carefully pour about 50 mL of the solution in the pot into the 250-mL beaker. If some cabbage leaves get into the beaker, that is fine.
10. Allow the solution in the 250-mL beaker to cool.
11. Once the solution is cool enough that you can comfortably pick up the beaker in your hand, pour about 10 mL of the solution into the 100-mL beaker. The vinegar in the beaker should turn pink. This is due to chemicals (anthocyanins) that come from the cabbage leaves. As you will learn in module 9, anthocyanins turn pink in the presence of acids such as vinegar.
12. Place the stirring rod in the beaker so that it stands in the beaker.
13. Place the watch glass on top of the beaker so that it covers the mouth of the beaker.
14. Place the beaker (with the stirring rod in it and the watch glass covering it) on the scale, and measure the mass of the total assembly. Your setup should look like the picture below:



15. After you have measured the total mass of the experimental setup, dump the lye into the vinegar.
16. Stir the solution with the stirring rod. Do not lift the stirring rod out of the solution because you would risk losing some of the solution.
17. You should quickly observe a color change because the lye (a base) neutralizes the vinegar (an acid). Since anthocyanins are pink in the presence of acids and the pink color is replaced by a greenish-yellow color, you know that the vinegar is being changed. You will learn a lot more about this process in module 9. For now, just realize that you are observing a chemical reaction, where the vinegar and lye are forming 2 completely different chemicals.
18. Notice also that most of the lye seems to have disappeared. Of course, it didn't really disappear. It simply reacted with the vinegar and changed into a chemical that is now dissolved in the water.

19. Rest the stirring rod against the side of the beaker and place the watch glass back on the top of the beaker so that (aside from the lye and the color change), the set-up looks like it did in step 14.
20. Carefully touch the sides of the beaker. The beaker should be noticeably warmer than when you put it on the scale. This is further evidence that a chemical reaction occurred because energy in the form of heat was released.
21. Read the mass of the entire assembly again.
22. Clean up and return everything to the proper place.

CONCLUSION: The mass in step 21 should be the same as what you read in step 14. This demonstrates the law of mass conservation. Although vinegar and lye were changed into 2 different substances, although heat energy was released, and although the lye seemed to disappear, the mass of the entire assembly did not change.

Once chemists accepted the law of mass conservation, they started using it to analyze many of the changes they observed in matter. Example 2.1 illustrates one way that the law of mass conservation can be used in analyzing how matter changes.

EXAMPLE 2.1

A chemist notices that when given enough energy, a certain white powder changes into 2 different substances: tin and oxygen. A chemist watches 151 grams of powder undergo this change. The chemist easily collects the tin that was formed and measures its mass to be 119 grams. Unfortunately, he could not collect the oxygen that was formed and therefore could not determine its mass. Use the law of mass conservation to determine how much oxygen was made.

The law of mass conservation tells us that we cannot create or destroy matter. If the chemist started with 151 grams of matter, no matter what change occurred, he must end up with 151 grams of matter. Since we know that only 2 substances—tin and oxygen—were formed in the course of the experiment, all 151 grams must be accounted for between the tin and the oxygen. Mathematically, we would say:

$$\text{Mass of powder} = \text{Mass of tin} + \text{Mass of oxygen}$$

Rearranging the equation gives us:

$$\text{Mass of oxygen} = \text{Mass of powder} - \text{Mass of tin}$$

Plug in our numbers:

$$\text{Mass of oxygen} = 151 \text{ grams} - 119 \text{ grams} = 32 \text{ grams}$$

Notice the significant figures here. Although both 151 and 119 have 3 significant figures, the answer can have only 2 because we are subtracting and thus do not count significant figures. Instead, we

look at the decimal place. Since both numbers have their last significant figure in the ones place, the answer must also have its last significant figure in the ones place. Therefore, **32 grams of oxygen** were formed when 151 grams of the white powder underwent the observed change.

ON YOUR OWN

- 2.1 A compound is decomposed into its constituent elements of carbon, nitrogen, and hydrogen. If, in the decomposition, 12.0 grams of carbon, 14.0 grams of nitrogen, and 3.0 grams of hydrogen were formed, what was the mass of the compound before it was decomposed?

Make sure that you understand how to use the law of mass conservation by answering “On Your Own” question 2.1

ELEMENTS: THE BASIC BUILDING BLOCKS OF MATTER

Once chemists began using the law of mass conservation, they noticed some-

thing quite interesting: Some substances can undergo change in such a way as to make many other substances, each less massive than the original. For example, they noticed that a black powder called iron sulfide can undergo a change that forms iron and sulfur. Starting with 88 grams of iron sulfide would make 56 grams of iron and 32 grams of sulfur. Similarly, 85 grams of a white powder known as sodium nitrate could undergo a change that would produce 23 grams of sodium, 48 grams of oxygen, and 14 grams of nitrogen. In both of these cases, the change that occurred seemed to divide the matter that was in the powder into smaller bundles that each became a new substance. Chemists called this type of change **decomposition**.

On the other hand, chemists noticed that there were some substances that would not undergo decomposition in any way. For example, once they had made sulfur from iron sulfide, there was nothing they could do to force the sulfur to decompose into several less massive substances. No matter what they tried to do, the sulfur would not change into different substances that were less massive than the original sample of sulfur. It seemed that the matter in the sulfur could not be made into smaller substances. This told chemists that there was something special about sulfur.

Remember the example of the radio used earlier? If you began to take apart a radio, you would eventually get to the point where you had a lot of parts (screws, wires, and dials, for example) that you could not disassemble any further. When you got to that point, you would say that you had broken the radio down into all of its components.

When chemists saw that iron sulfide could be decomposed into iron and sulfur but that the iron and sulfur could not be decomposed any further, they assumed that this meant iron and sulfur were the components of iron sulfide. These components could not be decomposed any further, so they represented the simplest form of the matter within the iron sulfide. Chemists called these components **elements**.

Element—Any substance that cannot be broken down (decomposed) into simpler components.

Chemists soon discovered that there were many, many substances that could not be decomposed into smaller substances; there were many, many elements. Since elements were considered the components of all forms of matter, chemists wanted to catalog them

carefully. They compiled the most important information about each element into a table that was called the **periodic table of elements**. As time went on, this table became the chemist's most important tool. It remains so even to this day. The modern form of the periodic table of elements is shown on the next page.

The periodic table of elements can also be found on the inside front cover of this book. Feel free to photocopy it and keep it handy so you can easily look at it as you read the rest of this and other modules. In addition, you will use it heavily while doing all problems and tests. You might want to put it in a clear plastic sheet protector and use it as a bookmark so you will always have it handy while you read. The periodic table of elements contains a huge amount of information, so we will be referring to it over and over again. It will help you quite a bit to look at the table every time we refer to it so that you will see how to use it.

The periodic table of elements is often called simply the periodic table. Notice that it has 118 boxes. Each box represents an element, and there are currently 118 elements. Believe it or not, only 92 of these elements (those with boxes numbered 1–92) appear naturally in creation. The 26 other elements (numbered 93–118) have been manufactured by scientists in a lab. We say there are currently 118 elements because sometimes a scientist will figure out how to manufacture a new one. For example, element 118 was manufactured in 2002 by experiments conducted at Dubna, Russia, at the Flerov Laboratory of Nuclear Reactions by workers from the Joint Institute for Nuclear Research in Russia and the Lawrence Livermore National Laboratory in the United States. The researchers collided a Californium (Cf) atom with a calcium (Ca) atom.

Think for a minute about what this means: Since elements are the building blocks of matter, everything you see around you is composed of elements which are listed on the table. If you were to break down anything in nature into its individual components of matter, then each of those components would be one of the elements listed on the table.

Each box in the table contains 1–3 letters that make up the chemical abbreviation for each element. For example, box 6 has a C in it. This is the chemical abbreviation for the element carbon. In the same way, box 20 contains a Ca. This is the chemical abbreviation for the element calcium. If an abbreviation contains more than one letter, the first letter is always capitalized and the others are not capitalized.

The abbreviations for some elements are temporary until they get their assigned names and symbols from the International Union of Pure and Applied Chemistry (IUPAC). You'll notice the letter U often in higher atomic numbers. The IUPAC temporarily names new elements by assigning each digit a Latin root. For example, the element roentgenium, atomic number 111, was temporarily called Uuu—*un* for one—so Uuu is 1–1–1. Then they add the suffix *-ium*. So element 111 was ununium. Can you guess the temporary name for element 115? It would be Uup—ununpentium, or 1–1–5!

In 1937, the first artificially created element, technetium, was made. In 1940, neptunium was made, followed quickly by element 94, plutonium. Between 1940 and 1961, elements 95–103 were created. Elements 104–106 were created by 1974, and it wasn't until 1981 that element 107 was confirmed. It took 15 years to create elements 108–112, and then another 9 years to synthesize 113–118. So now the heaviest element, 119, waits to be confirmed, and if it is, the periodic table will get a new row!

While this is generally the rule followed for the periodic table, you will soon see that

THE PERIODIC TABLE OF ELEMENTS

1A	2A	3A	4A	5A	6A	7A	8A
1 H 1.01	4 Be 9.01	5 B 10.8	6 C 12.0	7 N 14.0	8 O 16.0	9 F 19.0	2 He 4.00
3 Li 6.94	12 Mg 24.3	13 Al 27.0	14 Si 28.1	15 P 31.0	16 S 32.1	17 Cl 35.5	10 Ne 20.2
11 Na 23.0	19 K 39.1	20 Ca 40.1	31 Ga 69.7	32 Ge 72.6	33 As 74.9	34 Se 79.0	18 Ar 39.9
19 K 39.1	37 Rb 85.5	38 Sr 87.6	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	36 Kr 83.8
21 Sc 45.0	39 Y 88.9	40 Zr 91.2	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	54 Xe 131.3
22 Ti 47.9	41 V 50.9	42 Cr 52.0	80 Hg 200.6	85 At (210)	86 Rn (222)	87 Ra (226.0)	35 Br 79.9
23 V 50.9	43 Tc (98)	44 Ru 101.1	78 Pt 195.1	88 Fr (223)	113 Uut (284)	114 Fl (289)	37 Rb 85.5
24 Cr 52.0	44 Rh 102.9	45 Pd 106.4	77 Ir 192.2	89 Ac (227)	112 Cn (285)	115 Uup (288)	39 Y 88.9
25 Mn 54.9	45 Ru 101.1	46 Rd 106.4	76 Os 190.2	90 Th 232.0	116 Lv (293)	117 Uus (294)	40 Zr 91.2
26 Fe 55.8	46 Rh 102.9	47 Ag 107.9	75 Re 186.2	91 Pa 231.0	118 Uuo (294)	118 Uuo (294)	41 Nb 92.9
27 Co 58.9	47 Rh 102.9	48 Pd 106.4	74 W 183.8	92 U 238.0	119 Uuq (295)	119 Uuq (295)	42 Mo 96.0
28 Ni 58.7	48 Rh 102.9	49 Pd 106.4	73 Ta 181.0	93 Np (237)	120 Uuq (296)	120 Uuq (296)	43 Tc (98)
29 Cu 63.5	49 Rh 102.9	50 Pd 106.4	72 Hf 178.5	94 Pu (244)	121 Uuq (297)	121 Uuq (297)	44 Ru 101.1
30 Zn 65.4	50 Pd 106.4	51 Ag 107.9	71 Re 186.2	95 Am (243)	122 Uuq (298)	122 Uuq (298)	45 Ru 101.1
31 Ga 69.7	51 Ag 107.9	52 Ag 107.9	70 Os 190.2	96 Cm (247)	123 Uuq (299)	123 Uuq (299)	46 Rd 106.4
32 Ge 72.6	52 Ag 107.9	53 Ag 107.9	69 Ir 192.2	97 Bk (247)	124 Uuq (300)	124 Uuq (300)	47 Ag 107.9
33 As 74.9	53 Ag 107.9	54 Ag 107.9	68 Pt 195.1	98 Cf (251)	125 Uuq (301)	125 Uuq (301)	48 Pd 106.4
34 Se 79.0	54 Ag 107.9	55 Ag 107.9	67 Au 197.0	99 Es (252)	126 Uuq (302)	126 Uuq (302)	49 In 114.8
35 Br 79.9	55 Ag 107.9	56 Ag 107.9	66 Hg 200.6	100 Fm (257)	127 Uuq (303)	127 Uuq (303)	50 Sn 118.7
36 Kr 83.8	56 Ag 107.9	57 Ag 107.9	65 Tb 158.9	101 Md (258)	128 Uuq (304)	128 Uuq (304)	51 Sb 121.8
37 Rb 85.5	57 Ag 107.9	58 Ag 107.9	64 Gd 157.2	102 No (259)	129 Uuq (305)	129 Uuq (305)	52 Te 127.6
38 Sr 87.6	58 Ag 107.9	59 Ag 107.9	63 Eu 152.0	103 Lr (262)	130 Uuq (306)	130 Uuq (306)	53 I 126.9
39 Y 88.9	59 Ag 107.9	60 Ag 107.9	62 Sm 150.4	104 Lu 175.0	131 Uuq (307)	131 Uuq (307)	54 Xe 131.3
40 Zr 91.2	60 Ag 107.9	61 Ag 107.9	61 Pm (145)	105 La 138.9	132 Uuq (308)	132 Uuq (308)	55 Cs 132.9
41 Nb 92.9	61 Ag 107.9	62 Ag 107.9	60 Nd 144.2	106 Ce 140.1	133 Uuq (309)	133 Uuq (309)	56 Ba 137.3
42 Mo 96.0	62 Ag 107.9	63 Ag 107.9	59 Pr 140.9	107 Bh (270)	134 Uuq (310)	134 Uuq (310)	57 La 138.9
43 Tc (98)	63 Ag 107.9	64 Ag 107.9	58 Ce 140.1	108 Hs (277)	135 Uuq (311)	135 Uuq (311)	58 Ce 140.1
44 Ru 101.1	64 Ag 107.9	65 Ag 107.9	57 Re 186.2	109 Mt (276)	136 Uuq (312)	136 Uuq (312)	59 Pr 140.9
45 Rh 102.9	65 Ag 107.9	66 Ag 107.9	56 Os 190.2	110 Ds (281)	137 Uuq (313)	137 Uuq (313)	60 Nd 144.2
46 Rd 106.4	66 Ag 107.9	67 Ag 107.9	55 Ir 192.2	111 Rg (280)	138 Uuq (314)	138 Uuq (314)	61 Pm (145)
47 Ag 107.9	67 Ag 107.9	68 Ag 107.9	54 Pt 195.1	112 Cn (285)	139 Uuq (315)	139 Uuq (315)	62 Sm 150.4
48 Pd 106.4	68 Ag 107.9	69 Ag 107.9	53 Au 197.0	113 Uut (284)	140 Uuq (316)	140 Uuq (316)	63 Eu 152.0
49 In 114.8	69 Ag 107.9	70 Ag 107.9	52 Hg 200.6	114 Fl (289)	141 Uuq (317)	141 Uuq (317)	64 Gd 157.2
50 Sn 118.7	70 Ag 107.9	71 Ag 107.9	51 Sb 121.8	115 Uup (288)	142 Uuq (318)	142 Uuq (318)	65 Tb 158.9
51 Sb 121.8	71 Ag 107.9	72 Ag 107.9	50 Sn 118.7	116 Lv (293)	143 Uuq (319)	143 Uuq (319)	66 Dy 162.5
52 Te 127.6	72 Ag 107.9	73 Ag 107.9	49 In 114.8	117 Uus (294)	144 Uuq (320)	144 Uuq (320)	67 Ho 164.9
53 I 126.9	73 Ag 107.9	74 Ag 107.9	48 Pd 106.4	118 Uuo (294)	145 Uuq (321)	145 Uuq (321)	68 Er 167.3
54 Xe 131.3	74 Ag 107.9	75 Ag 107.9	47 Ag 107.9	119 Uuq (295)	146 Uuq (322)	146 Uuq (322)	69 Tm 168.9
55 Cs 132.9	75 Ag 107.9	76 Ag 107.9	46 Rd 106.4	120 Uuq (296)	147 Uuq (323)	147 Uuq (323)	70 Yb 173.0
56 Ba 137.3	76 Ag 107.9	77 Ag 107.9	45 Pd 106.4	121 Uuq (297)	148 Uuq (324)	148 Uuq (324)	71 Lu 175.0
57 La 138.9	77 Ag 107.9	78 Ag 107.9	44 Rh 102.9	122 Uuq (298)	149 Uuq (325)	149 Uuq (325)	72 Hf 178.5
58 Ce 140.1	78 Ag 107.9	79 Ag 107.9	43 Tc (98)	123 Uuq (299)	150 Uuq (326)	150 Uuq (326)	73 Ta 181.0
59 Pr 140.9	79 Ag 107.9	80 Ag 107.9	42 Cr 52.0	124 Uuq (300)	151 Uuq (327)	151 Uuq (327)	74 W 183.8
60 Nd 144.2	80 Ag 107.9	81 Ag 107.9	41 V 50.9	125 Uuq (301)	152 Uuq (328)	152 Uuq (328)	75 Re 186.2
61 Pm (145)	81 Ag 107.9	82 Ag 107.9	40 Zr 91.2	126 Uuq (302)	153 Uuq (329)	153 Uuq (329)	76 Os 190.2
62 Sm 150.4	82 Ag 107.9	83 Ag 107.9	39 Y 88.9	127 Uuq (303)	154 Uuq (330)	154 Uuq (330)	77 Ir 192.2
63 Eu 152.0	83 Ag 107.9	84 Ag 107.9	38 Sr 87.6	128 Uuq (304)	155 Uuq (331)	155 Uuq (331)	78 Pt 195.1
64 Gd 157.2	84 Ag 107.9	85 Ag 107.9	37 Rb 85.5	129 Uuq (305)	156 Uuq (332)	156 Uuq (332)	79 Au 197.0
65 Tb 158.9	85 Ag 107.9	86 Ag 107.9	36 Kr 83.8	130 Uuq (306)	157 Uuq (333)	157 Uuq (333)	80 Hg 200.6
66 Dy 162.5	86 Ag 107.9	87 Ag 107.9	35 Br 79.9	131 Uuq (307)	158 Uuq (334)	158 Uuq (334)	81 Tl 204.4
67 Ho 164.9	87 Ag 107.9	88 Ag 107.9	34 Se 79.0	132 Uuq (308)	159 Uuq (335)	159 Uuq (335)	82 Pb 207.2
68 Er 167.3	88 Ag 107.9	89 Ag 107.9	33 As 74.9	133 Uuq (309)	160 Uuq (336)	160 Uuq (336)	83 Bi 209.0
69 Tm 168.9	89 Ag 107.9	90 Ag 107.9	32 Ge 72.6	134 Uuq (310)	161 Uuq (337)	161 Uuq (337)	84 Po (209)
70 Yb 173.0	90 Ag 107.9	91 Ag 107.9	31 Ga 69.7	135 Uuq (311)	162 Uuq (338)	162 Uuq (338)	85 At (210)
71 Lu 175.0	91 Ag 107.9	92 Ag 107.9	30 Zn 65.4	136 Uuq (312)	163 Uuq (339)	163 Uuq (339)	86 Rn (222)
	92 Ag 107.9	93 Ag 107.9	29 Cu 63.5	137 Uuq (313)	164 Uuq (340)	164 Uuq (340)	
	93 Ag 107.9	94 Ag 107.9	28 Ni 58.7	138 Uuq (314)	165 Uuq (341)	165 Uuq (341)	
	94 Ag 107.9	95 Ag 107.9	27 Co 58.9	139 Uuq (315)	166 Uuq (342)	166 Uuq (342)	
	95 Ag 107.9	96 Ag 107.9	26 Fe 55.8	140 Uuq (316)	167 Uuq (343)	167 Uuq (343)	
	96 Ag 107.9	97 Ag 107.9	25 Mn 54.9	141 Uuq (317)	168 Uuq (344)	168 Uuq (344)	
	97 Ag 107.9	98 Ag 107.9	24 Cr 52.0	142 Uuq (318)	169 Uuq (345)	169 Uuq (345)	
	98 Ag 107.9	99 Ag 107.9	23 V 50.9	143 Uuq (319)	170 Uuq (346)	170 Uuq (346)	
	99 Ag 107.9	100 Ag 107.9	22 Ti 47.9	144 Uuq (320)	171 Uuq (347)	171 Uuq (347)	
	100 Ag 107.9	101 Ag 107.9	21 Sc 45.0	145 Uuq (321)	172 Uuq (348)	172 Uuq (348)	
	101 Ag 107.9	102 Ag 107.9	20 Ca 40.1	146 Uuq (322)	173 Uuq (349		

there are exceptions. What would you expect the abbreviation for sodium to be? Probably S or So, right? That would be consistent with the nomenclature described above, but it's not right. The chemical abbreviation for sodium is Na. Where did that come from?

Sometimes chemists do not use the English name of the element to determine its abbreviation; they use its Latin name. The Latin name for sodium is *natrium*, so its abbreviation is Na. In addition, chemists sometimes abbreviate elements with just the first letter of the Latin name. For example, the abbreviation for potassium is K because the Latin name for potassium is *kalium*. This system has come about because chemists all over the world discovered elements at different times in history. Since chemists in some countries used Latin as the scientific language, they used Latin to name the elements. In other countries, chemists used English. Other elements are simply named after the person who discovered them or where the discovery occurred.

As time goes on, you will begin to learn the names that go with the abbreviations on the table. For now, you will be expected to memorize a few of them so that you can understand some of the things you read in this module. By the end of this module, you will be expected to know the first 30 elements on the table, H through Zn, plus Br and I. You need not know the numbers that appear with them on the periodic table. You need only know that H is the abbreviation for hydrogen, He is the abbreviation for helium, etc. Table 2.3 at the end of this module lists all element names and their corresponding abbreviations so that you can begin to learn them.

The next thing you might notice is that each box on the table contains 2 numbers. These numbers tell us some very important things about each element. We will not get to the meaning of those numbers for a couple of modules, so for now you can simply ignore them. Also, you might think that the table is structured in a funny way. The boxes are not simply piled on top of each other or arranged in nice, even rows. We will see in a later module that the table is laid out this way to group the elements that act alike in the same column.

Finally, notice the heavy, jagged line that starts at the element boron (B) and moves diagonally down the table to astatine (At). This very important line separates the elements into 2 classes: **metals** and **nonmetals**. Elements that are called metals are usually malleable (can be easily bent and shaped), have luster (are shiny), and are able to conduct electricity. On the other hand, nonmetals are typically brittle (break easily when you try to bend or shape them), lack luster, and do not conduct electricity.

Elements that appear on the table to the left of the jagged line (Zn, Ca, and Pt, for example) are classified as metals. Elements that appear on the table to the right of the jagged line (B, N, and Br, for example) are nonmetals. As usual, this rule has a few exceptions. The most important exception (and one you will be expected to remember) is hydrogen.

Even though hydrogen is left of the jagged line, it is always considered a nonmetal.

You will have to wait for a while before you have enough information to understand why hydrogen is an exception to this rule. For now, you just need to memorize the fact that it is.

The other exception to this rule is a class of elements known as **metalloids**. These elements have some metal properties and some nonmetal properties. For example, silicon

(Si) is brittle, but it has luster. In addition, it conducts electricity under certain conditions but not under other conditions. For now, you do not have to consider metalloids. As far as you are concerned, there are only 2 classes of elements—metals and nonmetals—and these classes are separated by the heavy jagged line on the table.

ON YOUR OWN

- 2.2 Which of the following elements is (are) metal(s)? S, Ca, At, Tl, Hs, Xe, H
- 2.3 If you want to make a wire for an electrical system and you have only the following elements to work with, which should you choose? Ag, P, Br
- 2.4 Which of the following would you expect to be brittle? Co, Cs, C

The classification of elements into metals and nonmetals is important because we will use it again and again to identify the nature of different substances. It is important, therefore, that you can recognize whether an element is a metal or a nonmetal by looking at the table. In addition, it is important for you to know the 3 properties that describe metals and nonmetals. Answer “On Your Own” questions 2.2–2.4 to make sure you understand this classification scheme.

COMPOUNDS

Since chemists classify as elements all substances that cannot be decomposed into simpler substances, we also need to classify the substances that can be decomposed into simpler substances. We call these **compounds**.

Compounds—Substances that can be decomposed into elements by chemical means.

Based on these definitions, we can say that matter comes in 2 forms: elements and compounds. Elements are the basic building blocks of all substances. Compounds are formed when elements group together to form a new substance. Alternatively, we could say that elements are formed when compounds are broken down into their components. Either way of looking at these 2 forms of matter is correct.

In 1794, French chemist Joseph Proust (proost) demonstrated a fact about compounds that later became known as **the law of definite proportions**.

The law of definite proportions—The proportion of elements in any compound is always the same.

Proust discovered that when several elements come together to form a compound, they always come together in the same proportion. For example, water is made when the element hydrogen (H) and the element oxygen (O) react together. Whenever this happens, it always requires 8 grams of oxygen for every 1 gram of hydrogen. If the reaction started with 10 grams of hydrogen, it would need 80 grams of oxygen. When the reaction was done, 90 grams of water would be formed, in agreement with the law of mass conservation. No matter how much hydrogen the reaction started with, the mass of oxygen had to be 8 times greater, or some hydrogen would be left over.

You can think about the law of definite proportions in the same way that you think about recipes. If a cake recipe says that the cake serves 8 people, you could make a cake to serve 40 people by using 5 times the ingredients called for in the recipe. The law of

definite proportions tells us that there is a single recipe that governs how much of each element is needed to make a particular compound. If you don't add elements together in exactly the proportion the recipe calls for, then there will be some leftover ingredients at the end. Figure 2.3 tries to illustrate this idea.

FIGURE 2.3
The Law of Definite Proportions

A chemist decides that she wants to make table salt. She has already determined that table salt is a compound composed of 2 elements: sodium (Na) and chlorine (Cl). What she doesn't know is what proportion of each element she must use. To determine this, she does a few experiments. First, she reacts 10.0 grams of sodium with 10.0 grams of chlorine. This is a very dangerous reaction because sodium (a solid metal) is very reactive and chlorine (a gas) is poisonous. When the reaction is over, she finds that she has 16.5 grams of table salt and 3.5 grams of sodium left over.



10.0 g sodium and 10.0 g chlorine makes 16.5 g table salt and 3.5 g leftover sodium.

Sodium: Dennis s.k. Collection/Dnn87 (CC BY-SA) . Illustration of chlorine by David Weiss. Salt: © Crestock.com (nenovbrothers).

First, we see that this experiment confirms the law of mass conservation. Since the chemist started out with 20.0 grams (10.0 g + 10.0 g) of matter, she also had to end up with 20.0 grams (16.5 g + 3.5 g) of matter. Since the sodium and chlorine were not added in the correct proportion, the substance that she ended up with was not pure table salt. Instead, only 16.5 grams of table salt could be made. The rest of the 20.0 grams, then, is accounted for by sodium that could not react because too much was added in the first place.

Second, we see that the *chlorine* must have run out. There must have been plenty of sodium left over. Since there was no leftover chlorine, that must mean that the chlorine was completely used up in the reaction. This little fact allows us to determine the proper proportions (the recipe) of sodium and chlorine a chemist must have to make table salt. Since 3.5 grams of sodium were left over, then that mass of sodium was not used in the reaction. To find out how much sodium was used, we simply subtract the leftover amount from the amount we started with:

$$\text{Amount of sodium used} = \text{Starting amount} - \text{Amount left over}$$

$$\text{Amount of sodium used} = 10.0 \text{ g} - 3.5 \text{ g} = 6.5 \text{ g}$$

We see that only 6.5 grams of sodium were used when reacted with 10.0 grams of chlorine. The proper proportion in which to add sodium and chlorine is 6.5 grams of sodium for every 10.0 grams of chlorine.

To test this recipe, our chemist then reacts 10.0 grams of chlorine with 6.5 grams of sodium. According to the law of definite proportions, these 2 elements should now react to form table salt with no leftovers since she is adding the elements together in the proper proportions:



6.5 g sodium and 10.0 g chlorine makes 16.5 g table salt and no leftover sodium.

So she did, indeed, determine the proper recipe for making table salt: 10.0 grams chlorine plus 6.5 grams sodium makes 16.5 grams table salt. To convince herself that the law of definite proportions really works, she tripled the masses of sodium and chlorine and tried again:



19.5 g sodium and 30.0 g chlorine makes 49.5 g table salt.

Once again, the experiment ended with no leftovers because the elements were added together in the proper proportion.

As a final experiment, the chemist adds 6.5 grams of sodium to 15.0 grams of chlorine. In this case, the law of definite proportions shows that she used too much chlorine because 6.5 grams of sodium needs only 10.0 grams of chlorine. There should, therefore, be 5.0 grams of chlorine left over in the end.



6.5 g sodium and 15.0 g chlorine makes 16.5 g table salt and 5.9 g leftover chlorine.

The law of definite proportions, coupled with the law of mass conservation, can be used to solve problems like example 2.2. After studying the example, answer “On Your Own” questions 2.5–2.6.

EXAMPLE 2.2

A chemist reacts 15.0 grams of calcium (Ca) with 15.0 grams of oxygen (O). This reaction makes 21.0 grams of a compound known as lime. Along with the lime, there is also some leftover oxygen. If the chemist wants to make 55 grams of lime and have no leftover oxygen or calcium, how much of each element should he use?

To solve this problem, we first need to determine the recipe for making lime. To do that, we need to figure out how much oxygen was used in the reaction. Since there was no leftover calcium, we know that all 15.0 grams of calcium were used. How much leftover oxygen was there? We can find out by using the law of mass conservation:

Total mass before reaction = Total mass after reaction

$15.0 \text{ g} + 15.0 \text{ g} = 21.0 \text{ g} + \text{Mass of leftover oxygen}$

$\text{Mass of leftover oxygen} = 30.0 \text{ g} - 21.0 \text{ g} = 9.0 \text{ g}$

Since 9.0 grams of oxygen were left over:

$\text{Mass of oxygen used} = \text{Starting mass} - \text{Mass left over}$

$$\text{Mass of oxygen used} = 15.0 \text{ g} - 9.0 \text{ g} = 6.0 \text{ g}$$

So the proper recipe to use in making lime is this: 15.0 grams of calcium plus 6.0 grams of oxygen makes 21.0 grams of lime. The problem asks how much of each element is needed to make 55 grams of lime. We obviously need to increase the recipe, but by how much? We have to find what number, when multiplied by 21.0, gives us 55. That will tell us by what factor we need to increase our recipe:

$$21.0 \text{ g} \times x = 55 \text{ g}$$

$$x = \frac{55 \text{ g}}{21.0 \text{ g}} = 2.6$$

Notice that in the expression for x , the unit of grams cancels. As a result, x has no units. Numbers like this are called **dimensionless quantities**.

Now we know that to make 55 grams of lime, we must increase our recipe by a factor of 2.6. If the original recipe was 15.0 grams calcium + 6.0 grams oxygen, then the new recipe is:

$$\text{Mass of calcium} = 15.0 \text{ g} \times 2.6 = 39 \text{ g}$$

$$\text{Mass of oxygen} = 6.0 \text{ g} \times 2.6 = 16 \text{ g}$$

Therefore, to make 55 grams of lime, the chemist needs to react **39 grams of calcium with 16 grams of oxygen**.

ON YOUR OWN

2.5 When 24.0 grams of the element carbon (C) react with 8.08 grams of the element hydrogen (H), natural gas (the same stuff that is burned in a gas stove or a gas furnace) is produced. There is no leftover carbon or hydrogen when these quantities are used. How much carbon and hydrogen would you need to make 128.4 grams of natural gas?

2.6 A chemist decomposes 60.0 grams of a purple powder into 32.8 grams of the element chlorine (Cl) and an unmeasured amount of the element cobalt (Co). The chemist uses the results of this experiment to help him determine the recipe for making this purple powder from its elements. He then decides to make 2.00 kg of it. How much cobalt and how much chlorine gas will he need?

THE LAW OF MULTIPLE PROPORTIONS

One more law that deals with the nature of compounds is the **law of multiple proportions**.

The law of multiple proportions—If 2 elements combine to form different compounds, the ratio of masses of the *second* element that react with a fixed mass of the *first* element will be a simple, whole-number ratio.

This law shows that sometimes there are different ways of combining elements. If this is the case, there is a simple relationship between the recipes for the 2 compounds. Although the law sounds a bit hard to understand when you read it, an example will make it very clear.

When 12.0 grams of carbon (C) and 32.0 grams of oxygen (O) combine, they make 44.0 grams of carbon dioxide, a gas that humans exhale. On the other hand, under the right conditions, 12.0 grams of carbon can combine with just 16.0 grams of oxygen to make 28.0 grams of carbon monoxide, a completely different compound. To illustrate just how different these compounds are, carbon monoxide is poisonous to humans. When 12.0 grams of carbon combine with 32.0 grams of oxygen, they form a compound that humans exhale each time they breathe. However, when carbon and oxygen combine in a different proportion (12.0 grams and 16.0 grams), they make a compound that is lethal to humans.

The law of multiple proportions indicates that there is a simple relationship between the amount of oxygen in the recipe for carbon dioxide and the amount of oxygen in the recipe for carbon monoxide. It first says that we must use a *fixed* amount of the first element (carbon). We have already done that; in each recipe, the amount of carbon is the same (12.0 grams). Then we determine the ratio of the mass of oxygen used to make the first compound and the mass of oxygen used to make the second compound. It took 32.0 grams of oxygen to make carbon dioxide but only 16.0 grams of oxygen to make carbon monoxide. The ratio of the mass of oxygen in carbon dioxide to the mass of oxygen in carbon monoxide, then, is 32.0:16.0, or 2:1. That's a simple, whole-number ratio.

Let's look at one more example just to make sure you understand this law. When 36.0 grams of carbon react with 12.0 grams of hydrogen, they produce 48.0 grams of methane, which is also called natural gas. On the other hand, when 36.0 grams of carbon react with 8.0 grams of hydrogen, they produce 44.0 grams of propane, which is the gas typically burned by outdoor gas grills. Once again, 2 elements (carbon and hydrogen) are reacting in different proportions to make different gases. We have a fixed amount of carbon (36.0 grams). If we then take the ratio of the masses of hydrogen that react with this fixed mass of carbon, we get 12.0:8.0, or 3:2. Once again, it's a simple, whole-number ratio. That's what the law of multiple proportions says should happen.

What's the big deal about the law of multiple proportions? It was the main evidence for an atomic theory proposed by John Dalton in the early 1800s. Dalton used his theory to predict the law of multiple proportions, and other scientists then did experiments to demonstrate that the law was correct. This provided excellent evidence for the validity of Dalton's theory.

DALTON'S ATOMIC THEORY

The law of definite proportions, combined with the law of mass conservation, led John Dalton, an English chemist, to propose a theory he thought would help chemists better understand the nature of matter. With this theory, Dalton thought that he could explain *why* these laws were always obeyed. He called it his atomic theory, and it forms the basis for our modern-day interpretation of the nature of matter. Dalton's theory included 4 vital assumptions:

1. All elements are composed of small, indivisible particles called atoms.
2. All atoms of the same element have exactly the same properties.
3. Atoms of different elements have different properties.
4. Compounds are formed when atoms are joined together. Since atoms are indivisible, they can join together only in simple, whole-number ratios.

Remember, these 4 ideas were assumptions. He could not offer definitive proof that they were true. However, he said that if you assume that they are true, you can explain why the laws of mass conservation and definite proportions are always obeyed. Also, he predicted the law of multiple proportions (as you see in statement 4), which was later demonstrated to be true.

How do these assumptions explain the laws of mass conservation and definite proportions? Consider the law of mass conservation. According to Dalton's assumptions, all compounds and elements that undergo change are simply rearranging their atoms. Since those atoms are indivisible, they obviously cannot be destroyed, so the total number of atoms in the entire system must stay the same. If the total number of atoms stays the same, the total amount of matter stays the same, and therefore, the mass stays the same.

Now let's consider the law of definite proportions. Dalton's theory assumes that atoms can join together to make compounds only in simple, whole-number ratios. In other words, a compound might have 1 carbon atom and 2 oxygen atoms, while another compound might have 1 carbon atom and 1 oxygen atom. A compound cannot have 1 carbon atom and 1.5 oxygen atoms because atoms are indivisible. It can have either 1 atom or 2 atoms, never 1.5 atoms. If atoms combine in this way, their masses would always combine in the same proportion, which is exactly what the law of definite proportions says.

Of course, the real evidence for Dalton's theory came with the fact that he predicted the law of multiple proportions. Since Dalton believed that atoms combine in simple, whole-number ratios, he predicted that those ratios should be seen in the recipes for different compounds. When scientists studied this, they found out that he was right. This helped Dalton's atomic theory gain acceptance in the scientific community. After all, it is nice for a theory to be able to explain facts that are already known. However, if a theory *predicts* facts that have not already been determined, it adds a lot of weight to the theory.

salt and light

John Dalton (1766-1844) was a Christian, chemist, physicist, and meteorologist. He is best known for the development of atomic theory, including the idea that atomic weights of elements were proportional to each other. These advances in science are so important to how we view the atom that Dalton is often referred to as the father of modern chemistry.



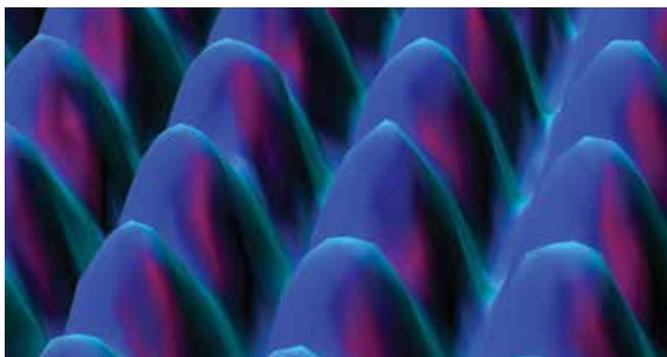
Portrait by Charles Turner (Public Domain)

As time has gone on, chemists have seen that a great many of the facts that we know about nature can be explained quite nicely if we assume that Dalton's atoms do indeed exist. Atoms are far too small to be seen, so we can never really *prove* that they exist. However, if we assume they exist, we can explain a wealth of facts that we have accumulated over the years. Chemists today simply accept the idea that atoms exist, even though their existence can never be proven.

If we could see these atoms, what would they look like? Figure 2.4 gives you an idea of what a group of atoms might look like if we were able to actually see them.

FIGURE 2.4
A Scanning Tunneling Electron Microscope Image of the Surface of a Nickel Foil

Image courtesy of the IBM research division



Something important about this figure is that you are *not* looking at a picture of atoms. For us to see an object (or take a picture of it), visible light must bounce off the object and hit our eyes (or the camera's film). That cannot happen with atoms because atoms are too small for visible light to bounce off them. You are not seeing a picture here. You are seeing a *computer image* based on some data.

In this figure, a scanning tunneling electron microscope was used to examine the surface of a nickel foil. In this kind of microscope, an electrified probe is placed very close to the surface of the metal, and the flow of electricity from the probe to the metal is monitored by a computer. Based on a theory called quantum mechanical tunneling, the computer then draws what the theory says the surface of the metal must look like for the electricity to flow the way that it did. If the theory is correct (and if the computer used the theory correctly), this is a silhouette of the nickel atoms on the surface of the foil. If the theory is incorrect, then who knows what this image represents? This image is certainly not proof that atoms exist. However, it does give more evidence for their existence.

The image in figure 2.4 is pretty much how Dalton pictured atoms. Today we are fairly sure that atoms have a more detailed structure than what is shown in the figure. The scanning tunneling electron microscope does not have enough resolution to show that detailed structure. However, as you will learn in module 3, other experiments have been done to probe the detailed structure of the atom, and today we have a theory that helps us picture the detailed structure of an atom. Before we can get to that, we need to learn a little more about matter and how it changes.

One final note about Dalton's atomic theory: Although chemists do accept the fact that atoms exist, we have also come to realize that 2 of Dalton's assumptions are not

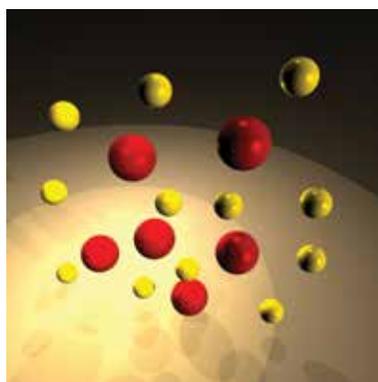
quite right. For example, the first assumption contains a small mistake. Under the right circumstances, atoms can, indeed, be split apart. Atoms are not truly indivisible. For the purposes of this course, however, we can assume that Dalton was right on this point because we won't be splitting atoms. His second assumption is not quite right either. Certain atoms within an element can be heavier than other atoms in that same element. These atoms are called isotopes and will be discussed in detail in module 3.

MOLECULES: THE BASIC BUILDING BLOCKS OF COMPOUNDS

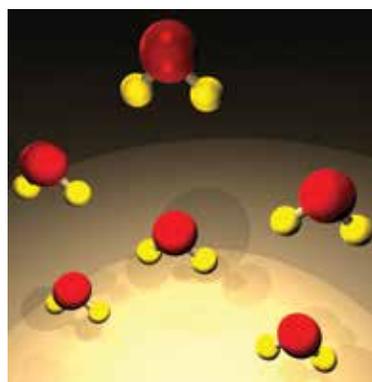
We have already said that elements are simply vast collections of identical atoms. If we could see a sample of the element helium (He), the gas put in balloons to make them float, at the atomic level, we would see a bunch of identical particles jumbled together much like grains of sand on a beach. Democritus (way back in the fourth century B.C.) was right. Matter is discontinuous: It contains little “grains” that we call atoms.

What about compounds, however? If we looked at the water in the ocean on the atomic level, would we see a bunch of atoms jumbled together? No, not exactly. Water is a compound that is formed when 2 hydrogen (H) atoms join with 1 oxygen (O) atom. If we looked at water on an atomic level, we would not see just a bunch of hydrogen and oxygen atoms jumbled together; we would instead see little groups of atoms that each contained 2 hydrogen atoms and 1 oxygen atom. Those *groups* would be jumbled up together like grains of sand. The groups that are formed when atoms join together to form compounds are called **molecules**. In other words, while all elements are composed of identical atoms, all compounds are composed of identical molecules. Figure 2.5 shows the difference between a mixture of atoms and a compound made of molecules.

FIGURE 2.5
Atoms and Molecules
Illustrations by David Weiss



A bunch of oxygen atoms (red) and hydrogen atoms (yellow) grouped together. This is not a compound. It is just a mixture of atoms.



The same oxygen and hydrogen atoms grouped together to make water molecules. This is what water would look like on the atomic scale.

The basic building blocks of matter, then, are atoms.

If a substance consists of identical atoms, then it is called an element.

On the other hand, when atoms join together, they form molecules.

If a substance consists of identical molecules, then it is called a compound.

ABBREVIATING AND CLASSIFYING COMPOUNDS

We've already seen that each element has its own name and abbreviation. What about compounds? Each compound needs to have a name, and it would be nice to be able to abbreviate those names. How do we do this? We name elements based on the atoms that make them up. For example, the element fluorine (F) is a substance made up of individual fluorine atoms. An element's name is the same as the atoms that make it up.

Compounds, then, should be named after the molecules that make them up. So how do we name a molecule? To learn how to name molecules, we first have to learn how to abbreviate them. Since molecules are made up of atoms, we can use the abbreviations on the periodic table to help us abbreviate molecules as well. For example, a water molecule is formed when 2 hydrogen (H) atoms join with 1 oxygen (O) atom. The abbreviation for a water molecule, then, is H_2O .

How do we get this abbreviation? The letters in the abbreviation come from the atoms which make up the molecule. Since water includes hydrogen atoms and oxygen atoms, the abbreviations for these 2 atoms, H and O, must be in the abbreviation for the molecule. In addition to the letters in the molecule's abbreviation, there are also numbers. These numbers, written as subscripts, tell us how many of each atom is in the molecule. The subscript 2 after the H tells us that water contains 2 hydrogen atoms. Since there is 1 oxygen atom in a water molecule, why isn't there a 1 subscript after the oxygen? That's because chemists love to leave off ones. In chemistry, if a number should exist but is not written down, then it is assumed to be 1. So to abbreviate a molecule, we simply use the abbreviation for each atom that makes up the molecule and follow it with a subscript that indicates the number of those atoms that are in the molecule. If the number is 1, we don't write it down.

Before we do some examples and problems, 2 terms must be introduced. Chemists do not use the term *abbreviation* when dealing with atoms and molecules. Instead, they use the terms *chemical symbol* and *chemical formula*. A chemical symbol is the abbreviation for an atom. So H is the chemical symbol for hydrogen. A chemical formula simply tells how many of each atom make up a particular molecule. Therefore, H_2O is the chemical formula for a water molecule. Sometimes chemists drop the word *chemical* and simply use the terms *symbol* and *formula*. You will have to get used to these terms, so we will begin using them now. Example 2.3 shows how these formulas work, and "On Your Own" questions 2.7–2.8 will give you some practice with them.

EXAMPLE 2.3

What is the chemical formula for a molecule that contains 1 atom of sodium, 1 atom of nitrogen, and 3 atoms of oxygen?

The chemical symbols for sodium, nitrogen, and oxygen are Na, N, and O, respectively. Since there is only 1 Na atom and 1 N atom, we needn't put number subscripts after them. However, there are 3 O atoms, so a subscript of 3 must follow the O:



How many atoms are in the molecule whose formula is Na_2HPO_4 ?

The formula indicates that there are 2 Na atoms, 1 H atom, 1 P atom, and 4 O atoms. How do we know that H is one atom's symbol, P is another atom's symbol, and O is another's? Remember, each atom's symbol only contains 1 capital letter. Therefore, every time you see a capital letter, you know you are dealing with a new atom. The total number of atoms, then, is simply the sum of all the individual atoms:

$$\text{Total number of atoms} = 2 + 1 + 1 + 4 = 8$$

So there are a total of **8 atoms** in this molecule.

ON YOUR OWN

- 2.7** Give the formula for each of the following molecules:
- A molecule that has 1 atom of potassium and 1 atom of bromine
 - A molecule that is made up of 5 carbon atoms, 10 hydrogen atoms, and 5 oxygen atoms
 - A molecule with 2 lithium atoms, 1 sulfur atom, and 4 oxygen atoms
- 2.8** How many of each atom exist in the following molecules?
- | | |
|-----------------------------|--------------------------------------|
| a. RbCl | b. KMnO_4 |
| c. Na_3PO_4 | d. $\text{HC}_2\text{H}_3\text{O}_2$ |

CLASSIFYING MATTER AS IONIC OR COVALENT

Once you have determined a compound's chemical formula, you are well on your way to figuring out its name. There is one other thing you must do to name a compound properly: You must classify it. Just as elements are classified into 2 groups—metals and nonmetals, compounds are classified based on their properties as well. There are many, many different classification schemes for compounds, and you will learn several of them as this course continues. For now, we will look at the broadest, simplest classification scheme.

Compounds can be classified as either ionic or covalent. You will learn later about subclasses within the covalent class, but this is a good start for now. A compound is considered ionic if, when dissolved in water, it conducts electricity. Pure water cannot conduct electricity, but when an ionic compound is dissolved in water, it can.

“Now wait a minute,” you might be saying. “Aren't we always told not to use water around electricity? If water cannot conduct electricity, then why are we concerned about it?” As we said earlier, *pure* water cannot conduct electricity. However, the water that comes out of your sink tap is not pure water. It has several ionic compounds dissolved in it. The most notable ionic compound in water is sodium fluoride, which helps protect your teeth against decay. Almost every city in the United States adds sodium fluoride (or something equivalent) to its drinking water for precisely that reason. Even rain is not pure water. It also has ionic compounds dissolved in it. About the only way to get pure water is

to buy or make distilled water. Distilled water is about 99% pure water. We will talk later in this module about what pure water is.

If a compound does not allow water to conduct electricity after it is dissolved, it is a covalent compound. How can we tell whether a compound is ionic or covalent? One way would be to dissolve it in water and see if it conducts electricity. That's what you will do in experiment 2.2.



If you purchased the MicroChem kit mentioned in the Student Notes, you can perform experiment 3 in that kit in addition to experiment 2.2.

EXPERIMENT 2.2

PURPOSE: To determine the electrical conductivity of compounds dissolved in water.

SUPPLIES

- Distilled water (Available at grocery stores. Half a gallon is plenty.)
- Baking soda
- Sugar
- 9-volt battery (***Do not use an electrical outlet in place of the battery. The electrical energy in a wall socket will hurt you and can kill you!***)
- 2 pieces of wire (preferably insulated), each of which is at least 15 cm long
- Scissors or wire cutters to strip insulation from wire (if it is insulated)
- Tape (preferably black electrical tape)
- Optional: Stores like Radio Shack sell a cap that will fit on 9V batteries. The cap has wires that you can use to replace the wires, scissors or wire cutters, and tape listed above.
- ½ teaspoon
- 100-mL beaker or small glass
- Safety goggles

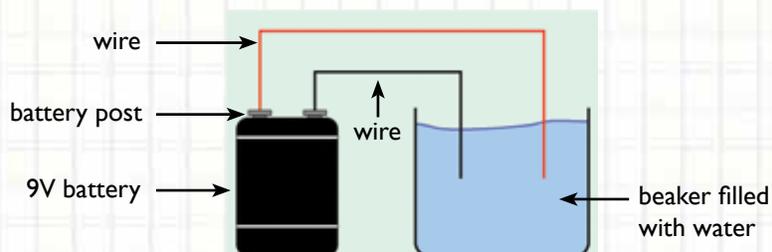
HYPOTHESIS: Ionic compounds will conduct electricity in water; covalent compounds will not conduct electricity in water.

PROCEDURE:

1. Rinse out your beaker or glass with tap water to get rid of any detergent residues left in it from washing. Then rinse it twice with distilled water to get out any impurities left behind from the tap water.
2. Add 80 mL of distilled water to the beaker. Treat this volume as a qualitative measurement.
3. If your wires are insulated, strip the insulation off of each wire end so that you have about 2 cm of bare wire at all ends. When that is done, attach one wire's end to one post of the battery and the other wire's end to the other post. The best way to do

this is to wad up the bare wire into a ball and shove the ball inside the battery post. Then, while pressing firmly, tape down the insulated part of the wire at the top of the post. The main concern is that the bare wire must be touching the battery post firmly so that it makes good electrical contact.

- Once each wire is attached to its own battery post, immerse the other ends of the wires into the water contained in your glass or beaker. Make sure the ends do not touch each other. In the end, your experiment should look something like the drawing below:



- Watch the bare ends of the wire closely for a little while. Nothing exciting happens, right?
- Now remove the wires from the water, add half a teaspoon of baking soda to the water, and mix it around a bit to get it to dissolve. The water should be cloudy at first, but the undissolved baking soda will eventually settle, and the solution will be relatively clear again.
- Immerse the ends of the wire as you did before. Now what's happening? If the other ends of your wires are firmly touching their individual battery posts, you should see bubbles forming on the wire ends that are immersed in the solution.
- Why is the bubbling happening? You'll learn in physics that when electricity is conducted, particles called electrons (ee lek' trons) move from one end of the battery to the other. This motion contains a lot of kinetic energy. The way we have set the experiment up, the electrons must travel through one wire, through the water, and through the other wire to reach the other end of the battery. When this motion takes place, some of that energy can be given to individual water molecules. Under the conditions of this experiment, the kinetic energy generated by the electricity is enough to cause the water to decompose into hydrogen and oxygen, its constituent elements. Since hydrogen and oxygen are gases, they form bubbles at the ends of the wires. For this energy to exist to begin with, electricity must be conducted. Since pure water cannot conduct electricity, nothing happened when you immersed the wires in the pure water. However, when you dissolved baking soda in the water, it became able to conduct electricity. This means that baking soda must be an ionic compound.
- Rinse the glass or beaker with tap water and distilled water as you did before.
- Repeat the experiment with table sugar. You will find that no bubbles form in this case. This must mean that no electricity can be conducted, so table sugar must be a covalent compound.
- Clean up and return everything to the proper place.

CONCLUSION: What did this experiment show you about ionic and covalent compounds? Write something about what you have learned.

In this experiment, we used the decomposition of water as an indicator for electrical conductivity. If bubbles appeared, that told us water was being decomposed, which meant that electricity was being conducted. This is another example of measuring something indirectly, as we discussed in module 1.

To classify compounds as ionic or covalent, you could go through an experiment like this for each compound you want to classify. Fortunately, there is an easier way. To classify compounds, all you have to do is look at the chemical formula and follow 2 rules:

1. If a compound contains at least 1 metal atom and at least 1 nonmetal atom, the compound is ionic.
2. If a compound is made up solely of nonmetal atoms, the compound is covalent.

Following these rules, you can now see why our experiment turned out the way it did. The chemical formula for table sugar is $C_{12}H_{22}O_{11}$. The atoms C, H, and O are all nonmetals, so table sugar is a covalent compound. It will not conduct electricity when dissolved in water. On the other hand, the chemical formula for baking soda is $NaHCO_3$. The atom Na is a metal, while H, C, and O are nonmetals; therefore, baking soda is an ionic compound and will conduct electricity when dissolved in water.

We've already told you that there are exceptions to nearly every rule in chemistry. Although all ionic compounds conduct electricity when dissolved in water, there are some compounds that are not ionic but will nevertheless conduct electricity when mixed with water. HCl is an excellent example. It is clearly a covalent compound because both hydrogen and chlorine are nonmetals. When mixed with water, however, HCl *will* conduct electricity because it chemically reacts with the water to produce ions. You will learn about this later. At this point in your chemistry education, you can ignore such exceptions. As far as you are concerned, if a compound conducts electricity when dissolved in water, it is an ionic compound.

In the same way, although the vast majority of compounds made up solely of nonmetal atoms are covalent, there are some exceptions. For example, ammonium hydroxide (NH_4OH) is made up solely of nonmetal atoms. However, it is an ionic compound. You will learn why this is the case later. Once again, at this point in your chemistry education, you can ignore such exceptions and just follow the rules as discussed in this module. Make sure you understand these rules by classifying the molecules in "On Your Own" question 2.9.

ON YOUR OWN

2.9 Classify the following compounds as ionic or covalent:

- | | |
|------------------|-------------------|
| a. SiF_4 | b. $CaBr_2$ |
| c. $LiC_2H_3O_2$ | d. $C_6H_{12}O_6$ |

NAMING COMPOUNDS

Now that you can classify and interpret a molecule's chemical formula, you are finally ready to learn how to name compounds. The first thing to remember is that we use a different system to name ionic compounds than we use to name covalent compounds. Since the ionic-compound naming system is a little easier, we will begin there. To name an ionic compound:

1. Start with the name of the first atom in the molecule.
2. Replace the name of the next atom in the molecule with its *-ide* name.
3. Putting those 2 names together gives the compound's name.

This is why sodium chloride is the name for NaCl. The symbol Na means sodium, and Cl stands for chlorine. According to the rules, we replace *chlorine* with *chloride*. Putting these names together, we get sodium chloride. In the same way, we would call K_2O potassium oxide. Pretty easy, huh? Later we will see that this gets more complicated when an ionic compound includes more than 2 different kinds of atoms. For now, we will stick with ionic compounds that have just 2 different kinds of atoms in them, so naming ionic compounds will be pretty simple at first.

You need to memorize table 2.1 for the test at the end of this module.

TABLE 2.1
The *-ide* Atom Names of Nonmetals

Atom	<i>-ide</i> Name	Atom	<i>-ide</i> Name
H	hydride	P	phosphide
C	carbide	S	sulfide
N	nitride	Cl	chloride
O	oxide	Br	bromide
F	fluoride	I	iodide

Naming covalent compounds is a little trickier because covalent compounds are more complicated than ionic compounds. When metal atoms and nonmetal atoms form ionic compounds, they can form only one type of molecule. For example, when sodium (Na) and chlorine (Cl) get together, the only molecule they can form is NaCl. They cannot form $NaCl_2$ or Na_2Cl or anything like that. Therefore, when we say the name *sodium chloride*, we know it means only one possible compound: NaCl.

Compare this to the situation with covalent molecules. When carbon (a nonmetal) joins with oxygen (a nonmetal), multiple compounds can be formed: CO or CO_2 . In this case, if we used the ionic compound naming system, we would call both of these molecules carbon oxide. Clearly, we can't have the same name for 2 different molecules. To fix this, we add prefixes in front of the name of each atom in the compound. This way, the number of atoms in each molecule is explicitly stated in its name. The prefixes we use in this naming system are summarized in table 2.2, which you need to memorize.

TABLE 2.2
Prefixes for Naming Covalent Compounds

Prefix	Meaning	Prefix	Meaning
mono	1	hexa	6
di	2	hepta	7
tri	3	octa	8
tetra	4	nona	9
penta	5	deca	10

These prefixes are inserted before each name in the compound to indicate how many of each type of atom are in the molecule. Once again, this general rule has one exception. If the prefix used on the first atom is *mono-*, we drop it. For example, CO_2 has the name *carbon dioxide*. Since there is only 1 carbon atom in the molecule, we should use the *mono-* prefix. However, since carbon is the first atom, we drop it. There are 2 oxygen atoms, so we must use the *di-* prefix in front of oxygen. Just like naming ionic compounds, we change the last atom's name to an *-ide* ending.

We can drop the *mono-* prefix only if it is on the first atom in the molecule. Therefore, CO is named carbon monoxide. Since there is only 1 oxygen atom, we need to use the *mono-* prefix, and since oxygen isn't the first atom in the molecule, we cannot drop it. Finally, we can never drop any prefix unless it is *mono-*, so the molecule C_2H_6 is named dicarbon hexahydride. See if you understand both of these naming systems by studying example 2.4 and then answering "On Your Own" questions 2.10–2.11.

EXAMPLE 2.4

Name the compound whose chemical formula is Al_2O_3 .

To name a compound, we must first determine whether it is ionic or covalent. Since aluminum (Al) lies on the left of the jagged line, it is a metal. This means we are working with an ionic compound. We therefore simply change *oxygen* to *oxide* and put the names together: **aluminum oxide**.

What is the name of the molecule PH_3 ?

Since P and H are both nonmetals, this is a covalent compound. We have to use the prefixes in the name. The prefix for phosphorus (P) is *mono-*, but since it is the first atom, we can drop it. The prefix for hydrogen is *tri-*, and we must add the *-ide* ending to hydrogen: **phosphorus trihydride**.

ON YOUR OWN

- 2.10 Name the following compounds:
 a. CF_4 b. Al_2S_3 c. P_2O_3 d. CaO
- 2.11 Give the chemical formulas for the following covalent compounds:
 a. sulfur hexafluoride
 b. pentacarbon decahydride

This lesson in naming compounds might lead you to believe that the process of giving names to compounds is easy. Don't be fooled into thinking this. We have merely scratched the surface of naming compounds. Notice that the only compounds we have named contain only 2 different types of atoms. That's because compounds with more than 2 types of atoms are more difficult to name. In addition, some

ionic molecules break the rules, and we have to develop a new naming system for them. Finally, there is a class of covalent compounds called organic compounds that has a whole new naming scheme. We will revisit some of these tougher topics later on, so don't dismiss as easy the very difficult process of naming compounds.

One other complication that crops up when we start naming compounds is the problem of common names. Think for a minute about why we don't call water by its

real chemical name: dihydrogen monoxide. This is because many compounds (like water) were named long before chemists knew about atoms and molecules and decided upon a standard naming system. These compounds were named without following any of the rules we just studied. Since old habits die hard, many compounds are still called by their common names rather than their proper chemical names. Ammonia (NH_3) is another example. The proper name for ammonia is nitrogen trihydride, but chemists called it ammonia long before our naming system was developed. Common names like these are here to stay, so we just have to get used to recognizing these names and associating them with the proper compound.

CLASSIFYING MATTER

Now that we have named compounds, we can finish this module with classification of matter. If we can properly classify the various types of matter we see in God's creation, it will be easier for us to determine how that matter is changing. First, every bit of matter that we see can be classified into 2 large groups: **mixtures** and **pure substances**.

Pure substance—A substance that contains only one element or compound.

Mixture—A substance that contains different compounds and/or elements.

These definitions may sound a bit confusing at first, but don't worry; you'll soon get the hang of them. Examples of pure substances include sulfur, iron, sugar, and table salt. In each of these substances, only 1 element or compound exists. The first 2 examples are elements since they are made up of atoms. In a sample of sulfur, only 1 type of atom exists: sulfur (S) atoms. In a sample of iron, only 1 type of atom exists: iron (Fe) atoms. The last 2 examples are compounds since they are made up of molecules. Sugar is made up of only 1 type of molecule: $\text{C}_{12}\text{H}_{22}\text{O}_{11}$. Likewise, table salt is composed entirely of NaCl molecules. You should recognize that sugar is a covalent compound (made up solely of nonmetals), whereas table salt is ionic (made up of a metal and a nonmetal).

Earlier, you used distilled water in one of your experiments. We noted that the distilled water was 99% pure. What that means is that 99% of all molecules in distilled water are H_2O molecules. If 100% of all the molecules in distilled water were H_2O molecules, then distilled water could be labeled a pure substance. This is what we mean when we say that pure substances contain only 1 element or compound.

Mixtures, on the other hand, contain more than 1 compound and/or element. If we were to dissolve table salt in pure water, the result would be a mixture of 2 compounds: H_2O and NaCl. This is what we mean when we say that mixtures have different compounds and/or elements. The H_2O molecules and the NaCl molecules in a saltwater mixture do not combine to make new molecules; they just occupy the same general area. That's what a mixture is. The important thing to realize about mixtures is that individual elements or compounds retain their individual properties while in a mixture. For example, pure water is tasteless. Table salt, on the other hand, has a distinct, sharp taste. A mixture of saltwater, then, tastes sharp like salt, but not as sharp as pure salt. Each compound retains its taste, and the tasteless water dilutes the sharp taste of the salt in the mixture.

This leads to another way we can distinguish between pure substances and

mixtures: The individual components in a mixture can be physically separated based on their individual properties. For example, consider figure 2.6.

FIGURE 2.6
Mixtures Can Be Separated Based on the Properties of Their Components

Photos by Flanders Video Productions



Sulfur (the yellow powder) and iron (the black powder) are mixed together.



A magnet (the white bar) is brought near the mixture.



The magnet pulls the iron out of the mixture and leaves the sulfur behind.



Sulfur and iron are mixed together and then heated so that they form the compound iron sulfide (the black powder).



A magnet (the white bar) is brought near the iron sulfide.



The magnet doesn't pick up anything because the iron no longer has its individual properties.

In the top portion of the figure, we are working with a *mixture* of sulfur and iron. Since this is a mixture, the sulfur retains its properties (its yellow color and nonmagnetism, for example), and the iron retains its properties (its black color and magnetism, for example). These properties can be used to separate the 2 elements. If a magnet is brought near the mixture, the iron will cling to the magnet, but the sulfur will not. We could separate the mixture back into its components (sulfur and iron) by using the magnet to pull out all of the iron. In the end, we would have 2 pure substances, iron and sulfur. A mixture, then, can be separated using the individual properties of its components.

Compare this to the lower portion of the figure, where the same 2 elements (sulfur and iron) are heated together so that they form the *compound* iron sulfide. Although iron sulfide contains both sulfur and iron, it is *not* a mixture; it is a pure substance. We can see that in several ways. For example, notice that there is no yellow in iron sulfide: it is black. The sulfur does not retain its characteristic yellow color. Notice also that when the magnet is brought near the iron sulfide, nothing clings to it. That's because the iron does not retain its ability to respond to a magnet. Both the sulfur and the iron have changed or given up their individual properties to become a compound. This is a pure substance.

Its elements cannot be separated based on their individual properties because they do not retain their individual properties when they form a compound.

The only way we could get the iron and sulfur out of iron sulfide is to *decompose* the iron sulfide back into a mixture of iron and sulfur and then separate the 2 substances. Therefore, mixtures can be *physically separated* into their individual components, but pure substances must be decomposed before their individual components can be physically separated. Experiment 2.3 gives you an opportunity to find another way of separating the components in a mixture.

EXPERIMENT 2.3

PURPOSE: To separate a mixture of sand and salt.

SUPPLIES:

- 2 beakers - 100 mL and 250 mL (or 2 glasses—1 large and 1 small)
- Sand (Kitty litter is an acceptable substitute, but don't use the kind that clumps.)
- Table salt
- Funnel
- Water
- Filter paper (You can cut a circle out of the bottom of a coffee-maker filter.)
- Stirring rod (or small spoon)
- Teaspoon
- Heat source such as a stove or alcohol burner
- Safety goggles

QUESTION: Can mixtures be separated?

HYPOTHESIS: Write down a hypothesis related to separating mixtures into elements and compounds

PROCEDURE:

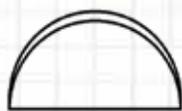
1. Pour a teaspoon of sand and a teaspoon of salt into the 100-mL beaker.
2. Use your stirring rod to mix them well. At this point, you have a mixture of sand and salt in your beaker. The molecules of sand (SiO_2) do not combine with the molecules of salt (NaCl) in any way. These 2 different types of molecules simply occupy the same general space now. In this experiment, you will learn how to separate this mixture back into sand and salt.
3. One way you could accomplish this separation would be to use fine tweezers and a powerful magnifying glass. If you looked at the mixture through the magnifying glass, you could see little grains of salt and little grains of sand. You could then use the tweezers to slowly pick the salt out of the sand. With enough patience and time, you could separate the mixture in that way. Unfortunately, this method is quite tedious and would take an enormous amount of time. To speed up the procedure, you will use a different method. Remember that in a mixture, the

individual molecules still retain their own unique properties. One property of salt is that it dissolves in water. However, sand does not. We will use this property to separate them.

- Add about 25 mL (1/8 cup) of water (a qualitative measurement) to the mixture and use your stirring rod to mix it in well. As the water mixes with the sand and salt, the salt will dissolve in water and the sand will not. Now you have an ugly mixture of sand, salt, and water in your beaker.
- To filter your mixture so that the saltwater will separate from the sand, first fold a circle of filter paper as shown below. In the last step of opening the triangle to make a cone, it's important to note that the triangle has 4 layers (from folding it in half 2 times). To get a cone that works as a filter, 3 layers need to be on one side and 1 layer on the other.



Start with a large piece of circular filter paper.



Fold it in half.



Fold it in half again so it makes a triangle.



Illustration by Megan Whitaker

Open the triangle to make a cone.

- Place the resulting cone of filter paper in the funnel. To get the filter paper to retain its cone shape and stick to the funnel, wet it with some water.
- Hold the funnel above the empty 250-mL beaker so that anything falling through the funnel will land in the beaker.
- Pour the mixture of water, salt, and sand into the filter paper. Make sure that the water level never rises above the top of the filter-paper cone. In addition, for best results, try to pour off just the liquid first and leave the majority of the sand in your beaker until the very end. As the liquid filters through the paper, you will see it fall into the 250-mL beaker. As long as you do not allow the liquid level to rise above the top of the filter-paper cone, the liquid in the beaker will be clear. Continue this process until everything in the beaker (including the sand) has been poured into the funnel.
- Once all of the liquid has filtered through the filter paper (this may take a while!), rinse the 100-mL beaker with 10 mL (1 tablespoon) of water and dump it all into the filter paper. Allow all of the liquid to drain through the filter paper.
- What you have now is pure sand on the filter paper and a mixture of salt and water in your beaker. To get rid of the water, simply start heating it so that it will boil. **Warning: If you are using glasses instead of beakers for this experiment, do not heat the glass. Instead, pour the salt water into a pot and boil it there. Regular glasses tend to break when they are heated to 100°C!**
- Allow all of the water to boil away. What's left behind? Pure table salt is left as a residue in the beaker.
- Clean up and return everything to the proper place.

CONCLUSION: You have separated your mixture of sand and salt back into its components. The sand is on the filter paper, and the salt is in the 250-mL beaker.



If you purchased the MicroChem kit mentioned in the Student Notes, you can learn another method used to separate mixtures by performing experiment 1.

This is the idea behind the concept of mixtures. Since the individual molecules within a mixture retain their unique properties, we can use those properties to separate them. Since salt dissolves in water and sand does not, we used that particular property as a means of separation in the experiment.

Mixtures and pure substances, then, are the 2 broad classifications of matter. However, they can both be broken down into 2 subclassifications as well. Pure substances can be either elements or compounds. Therefore, the classifications you learned before are really subclasses of the pure substance classification. Mixtures can be further classified as either **heterogeneous** (het' uh roh jee' nee us) or **homogeneous** (ho' moh jee' nee us).

Homogeneous mixture—A mixture with a composition that is always the same no matter what part of the sample is observed.

Heterogeneous mixture—A mixture with a composition that is different depending on what part of the sample is observed.

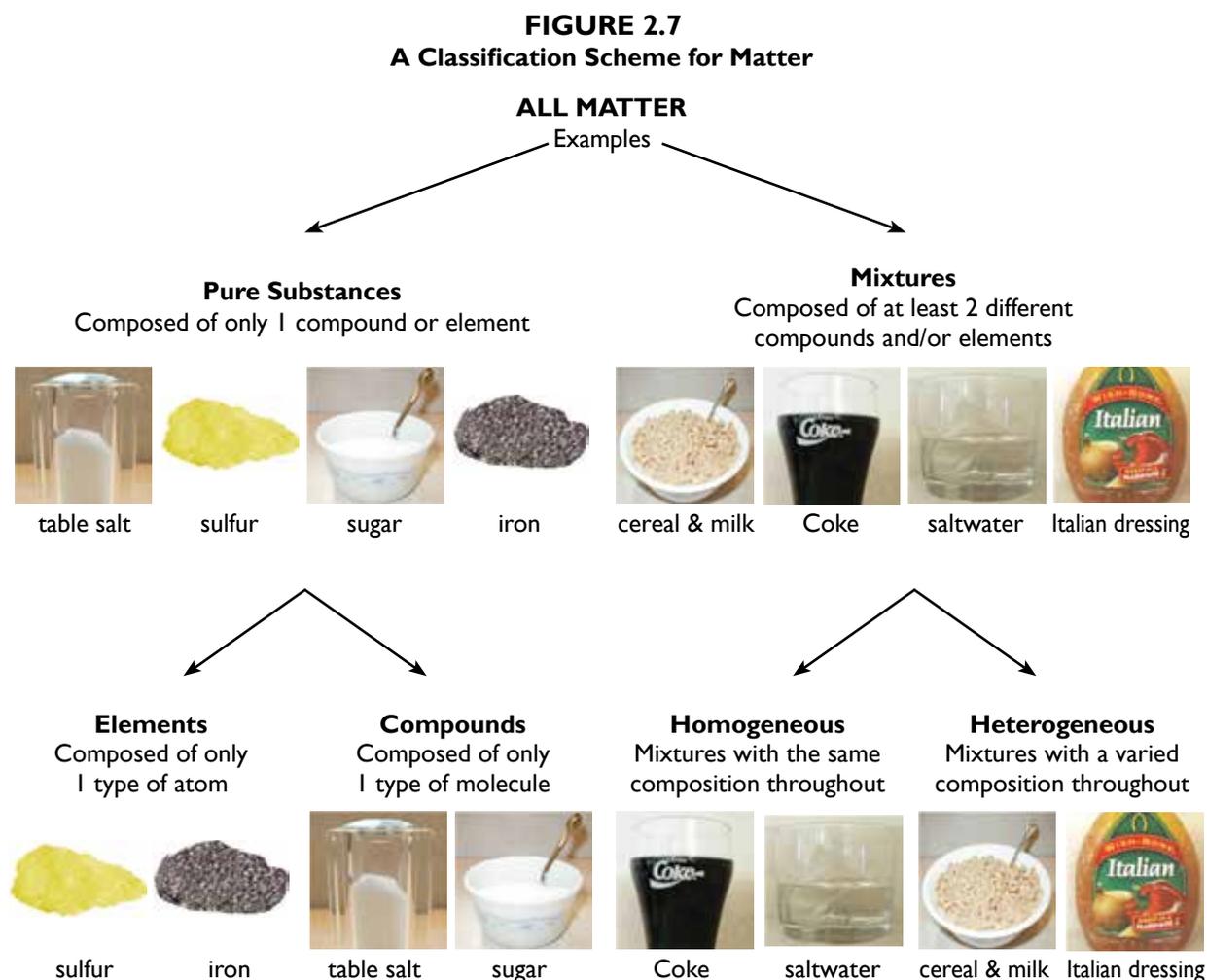
These definitions make sense when you think about the prefixes *homo-* and *hetero-*. *Homo-* means the same, and *hetero-* means different. Therefore, homogeneous solutions are the same throughout the sample, whereas heterogeneous solutions are different in different parts of the sample.

The best way to explain the difference between heterogeneous and homogeneous mixtures is by example. In experiment 2.3, you made up a mixture of sand and salt. When you first dumped the sand and the salt together, it was a heterogeneous mixture. If you looked at one portion of the mixture, you might see mostly sand, whereas another portion of the mixture might be mostly salt. As you stirred the mixture, it started to become homogeneous because the sand and salt were mixing evenly throughout the sample. If you stirred the sand and salt enough, you would end up with a mixture in which the salt was evenly spread throughout the sand. That would be a homogeneous mixture.

Examples of heterogeneous mixtures include Italian salad dressing, a bowl of cereal and milk, and a cake with icing on it. In each case, these mixtures have a different composition depending on where you observe the sample. When you take Italian dressing out of the refrigerator, the seeds, herbs, and spices are mostly at the bottom of the bottle. The seeds, herbs, and spices are more concentrated in one part of the sample than the other. A bowl of cereal and milk typically has mostly cereal on the top and milk at the bottom because the cereal tends to float on top of the milk. Finally, an iced cake is a mixture of cake and icing, but the icing is completely on the top, while the cake is underneath.

Examples of homogeneous mixtures include saltwater, Coke, and steel. Each of these mixtures has the same composition throughout. A glass of saltwater has the same amount of salt and water at the top of the glass as at the bottom. Coke is a mixture of several ingredients, but there is no difference in taste or composition from the top of the Coke

can to the bottom. Finally, steel is a mixture of iron (Fe) and carbon (C). Iron gives steel its strength, and carbon helps the steel stay rigid. The amounts of carbon and iron are the same no matter what sample of steel is observed, so it is also a homogeneous mixture. Figure 2.7 sums up these classifications of matter.

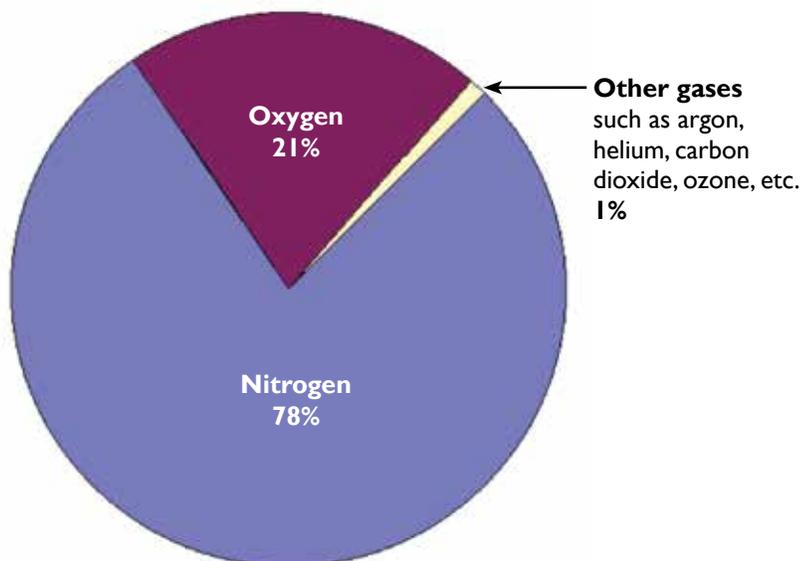


A couple of interesting facts related to the concept of mixtures are worth discussing. First, if you buy a carton of milk, the label says that the milk is homogenized. Now that you know this classification scheme, you can understand what that word means. When milk is taken from a cow, it is a heterogeneous mixture. The cream in the milk tends to rise to the top, and the water tends to fall to the bottom. The milk is thick and creamy if it comes off the top of the bottle, whereas it is thin and watery if it comes off the bottom of the bottle. Before you buy the milk, it is treated so that the cream stays evenly distributed throughout the milk. Therefore, the formerly heterogeneous mixture is turned into a homogeneous mixture. This is what the term *homogenized* means.

Second, did you know that air is a mixture? Most people think that since animals breathe in oxygen, the air around them must be pure oxygen. Nothing could be farther

from the truth. In fact, air is a mixture of several gases, including nitrogen, argon, water vapor, carbon dioxide, and oxygen. Oxygen actually makes up less than $\frac{1}{4}$ of the air that we breathe. Assuming there is no humidity, the air we breathe is 78% nitrogen, 21% oxygen, 0.9% argon, 0.03% carbon dioxide, and 0.07% other gases. Figure 2.8 shows this relationship in pie graph form.

FIGURE 2.8
The Composition of Dry Air



This mixture of gases is ideal for human life. Even though we need oxygen to live, we can have too much oxygen. Breathing elevated levels of oxygen for prolonged periods of time can have serious consequences, such as lung damage, chest pains, and even blindness! In addition, the frequency of natural forest fires increases significantly when the amount of oxygen in the air increases. For every 1% increase in the amount of oxygen in the air, the chance of a natural forest fire occurring goes up by 70%! Therefore, we need oxygen in our air, but not too much of it.

The air in our atmosphere does not have too much oxygen in it because the oxygen has been diluted with nitrogen. Nitrogen is the perfect gas to accomplish this dilution because it does not react with our bodies in any way but is important in the biological world through the nitrogen cycle. We breathe in nitrogen with every breath, but then we breathe it right out again because it does not interact with our bodies. Almost all other gases—carbon monoxide, sulfur dioxide, nitrogen dioxide, etc.—are poisonous to human life, so it is very good that

science and creation

Doesn't it seem like an amazing coincidence that the air on this planet just happens to have oxygen, but not too much oxygen? It has the perfect amount of oxygen for human life. The other gas that makes up the vast majority of the rest of our planet's air just happens to be one of the few gases that does not affect our bodies in any negative way. Clearly, such a perfect mixture of gases would never have come about by chance. God made the air around us perfect for sustaining human life. As we go on in this course, you will see that there are thousands and thousands of "coincidences" like this one that clearly show that this world has a Creator. This is why we say that real science points unerringly to God.

ON YOUR OWN

2.12 Classify the following as elements, compounds, homogeneous mixtures, or heterogeneous mixtures:

- | | |
|--------------------|--------------------|
| a. mercury | f. cesium |
| b. carbon monoxide | g. ocean water |
| c. chocolate | h. barium chloride |
| d. apple | i. chocolate chip |
| e. tea (no ice) | cookie |

those gases do not make up any significant fraction of the air we breathe.

See if you understand our matter classification scheme by answering “On Your Own” question 2.12.

In the end, then, we can see that Democritus’s musings on the nature of matter were correct. Chemists have shown throughout the years that matter does not come in continuous sheets; it comes in little packages known as atoms. When identical

atoms are jumbled together, the substance that results is called an element. When atoms join together, they make molecules. When these molecules are collected together, they form substances called compounds. Even though he didn’t know any of these details, Democritus’s ideas were essentially right. Not bad for someone who lived more than 2,000 years ago, is it?

SUMMARY OF TABLES IN MODULE 2

Table 2.1: The *-ide* Atom Names of Nonmetals

Table 2.2: Prefixes for Naming Covalent Compounds

Table 2.3: List of Element Names and Symbols

TABLE 2.3
Element Names and Symbols

The list is compiled in the order in which they appear on the periodic table.
The elements in bold are the ones you need to memorize for the test.

Name	Symbol	Name	Symbol	Name	Symbol
Hydrogen	H	Yttrium	Y	Iridium	Ir
Helium	He	Zirconium	Zr	Platinum	Pt
Lithium	Li	Niobium	Nb	Gold	Au
Beryllium	Be	Molybdenum	Mo	Mercury	Hg
Boron	B	Technetium	Tc	Thallium	Tl
Carbon	C	Ruthenium	Ru	Lead	Pb
Nitrogen	N	Rhodium	Rh	Bismuth	Bi
Oxygen	O	Palladium	Pd	Polonium	Po
Fluorine	F	Silver	Ag	Astatine	At
Neon	Ne	Cadmium	Cd	Radon	Rn
Sodium	Na	Indium	In	Francium	Fr
Magnesium	Mg	Tin	Sn	Radium	Ra
Aluminum	Al	Antimony	Sb	Actinium	Ac
Silicon	Si	Tellurium	Te	Thorium	Th
Phosphorus	P	Iodine	I	Protactinium	Pa
Sulfur	S	Xenon	Xe	Uranium	U
Chlorine	Cl	Cesium	Cs	Neptunium	Np
Argon	Ar	Barium	Ba	Plutonium	Pu
Potassium	K	Lanthanum	La	Americium	Am
Calcium	Ca	Cerium	Ce	Curium	Cm
Scandium	Sc	Praseodymium	Pr	Berkelium	Bk
Titanium	Ti	Neodymium	Nd	Californium	Cf
Vanadium	V	Promethium	Pm	Einsteinium	Es
Chromium	Cr	Samarium	Sm	Fermium	Fm
Manganese	Mn	Europium	Eu	Mendelevium	Md
Iron	Fe	Gadolinium	Gd	Nobelium	No
Cobalt	Co	Terbium	Tb	Lawrencium	Lr
Nickel	Ni	Dysprosium	Dy	Rutherfordium	Rf
Copper	Cu	Holmium	Ho	Dubnium	Db
Zinc	Zn	Erbium	Er	Seaborgium	Sg
Gallium	Ga	Thulium	Tm	Bohrium	Bh
Germanium	Ge	Ytterbium	Yb	Hassium	Hs
Arsenic	As	Lutetium	Lu	Meitnerium	Mt
Selenium	Se	Hafnium	Hf	Darmstadtium	Ds
Bromine	Br	Tantalum	Ta	Roentgenium	Rg
Krypton	Kr	Tungsten	W	Copernicium	Cn
Rubidium	Rb	Rhenium	Re	Flerovium	Fl
Strontium	Sr	Osmium	Os	Livermorium	Lv

ANSWERS TO THE “ON YOUR OWN” QUESTIONS

- 2.1 Since the compound contains only carbon, nitrogen, and hydrogen, the sum of those individual masses must add up to the mass of the compound:

$$\text{Mass of compound} = \text{Mass of nitrogen} + \text{Mass of carbon} + \text{Mass of hydrogen}$$

$$\text{Mass of compound} = 14.0 \text{ g} + 12.0 \text{ g} + 3.0 \text{ g} = 29.0 \text{ g}$$

By the law of mass conservation, the mass of the compound before the decomposition must have been **29.0 g**.

- 2.2 Since S, At, and Xe are to the right of the jagged line, they are nonmetals. In addition, H is always a nonmetal because it is an exception to the rule. **Ca, Tl, and Hs are metals.**
- 2.3 Electrical wires have to be able to conduct electricity and should bend relatively easily. You should choose a metal to work with. Metals lie to the left of the jagged line, and the only one that does is **Ag**.
- 2.4 Nonmetals are brittle, so they all crumble into pieces instead of bending. The only nonmetal on the list is **C**.
- 2.5 According to the law of mass conservation, 24.0 grams of carbon + 8.08 grams of hydrogen must make 32.1 grams of natural gas since there are no leftover amounts of carbon or hydrogen. To make 128.4 grams of natural gas, all we have to do is determine the factor by which we need to increase our recipe.

$$(32.1 \text{ g})x = 128.4 \text{ g}$$

$$x = \frac{128.4 \text{ g}}{32.1 \text{ g}} = 4.00$$

If we need to increase our recipe by a factor of 4.00, then all we need to do is multiply all of our ingredient amounts by 4.00:

$$\text{Mass of carbon} = 24.0 \text{ g} \times 4.00 = 96.0 \text{ g}$$

$$\text{Mass of hydrogen} = 8.08 \text{ g} \times 4.00 = 32.3 \text{ g}$$

So to make 128.4 grams of natural gas, we must add **96.0 grams of carbon to 32.3 grams of hydrogen**. *Note: You might notice that these 2 numbers do not add up to 128.4 grams as they should by the law of mass conservation. Instead, they add to 128.3 g. This is a result of the rounding that we must do to keep track of significant figures; it is considered normal.*

- 2.6 Before we can answer this question, we need to determine how much cobalt was in the original sample. That will then tell us the recipe. For this, we use the law of

mass conservation. Since we started with 60.0 grams, we must end with 60.0 grams as well. If 32.8 grams of mass is accounted for by the cobalt, then the rest must have gone to the chlorine:

$$\text{Mass of chlorine} = 60.0 \text{ g} - 32.8 \text{ g} = 27.2 \text{ g}$$

So the recipe for making 60.0 grams of the purple powder is to react 27.2 grams of cobalt with 32.8 grams of chlorine. But the chemist wants to make 2.00 kg of the stuff. First we have to convert that to grams so all of our units are the same. The amount 2.00 kg is the same as 2.00×10^3 g. Now we need to determine what to multiply our recipe by so it will make 2.00×10^3 g:

$$(60.0 \text{ g})x = 2.00 \times 10^3 \text{ g}$$

$$x = \frac{2.00 \times 10^3 \text{ g}}{60.0 \text{ g}} = 33.3$$

Now we just multiply our list of ingredients by this number, and we'll have the answer:

$$\text{Mass of cobalt} = 27.2 \text{ g} \times 33.3 = 906 \text{ g}$$

$$\text{Mass of chlorine} = 32.8 \text{ g} \times 33.3 = 1090 \text{ g}$$

To make 2.00 kg of the purple powder, react **906 grams of cobalt with 1090 grams of chlorine**.

- 2.7 (a) The chemical symbols for potassium and bromine are K and Br. Since there is only 1 atom of each, then there is no need to put any numbers in the formula. Therefore, **KBr** is the chemical formula.
- (b) The symbols for carbon, hydrogen, and oxygen are C, H, and O. Since there are 5 carbon atoms and 5 oxygen atoms, there must be a 5 subscript after the C and O. Since there are 10 hydrogen atoms, a 10 must follow the H. So **C₅H₁₀O₅** is the proper chemical formula.
- (c) The proper formula is **Li₂SO₄**.
- 2.8 (a) This molecule is made up of **1 rubidium atom** and **1 chlorine atom**.
- (b) Since there is no number after the K (potassium), we assume there is **1 potassium** atom. The same assumption holds true for the Mn, so there is **1 manganese** atom. The 4 after the oxygen means there are **4 oxygen atoms**.
- (c) In this molecule, there are **3 sodium atoms**, **1 phosphorus atom**, and **4 oxygen atoms**.
- (d) This molecule is made up of **2 carbon atoms**, **4 hydrogen atoms**, and **2 oxygen atoms**.
- 2.9 (a) Since Si and F are on the right of the jagged line, this molecule is comprised solely of nonmetals. This makes it **covalent**.

- (b) Because Ca (calcium) is on the left of the jagged line, it is a metal. Boron (Br), on the right of the jagged line, is a nonmetal. This means the compound is **ionic**.
- (c) This is also **ionic** because Li is a metal.
- (d) All atoms in this molecule are nonmetals, so the compound is **covalent**.
- 2.10 (a) This molecule is covalent, so we need to use prefixes. The prefix for carbon (C) is *mono-*, but we drop it because it is first. The prefix for fluorine (F) is *tetra-*, so we have **carbon tetrafluoride**.
- (b) Al is a metal, so this is an ionic compound. We use no prefixes: **aluminum sulfide**.
- (c) **diphosphorus trioxide**
- (d) **calcium oxide**
- 2.11 (a) SF_6 (b) C_5H_{10}
- 2.12 (a) Mercury is on the periodic table, so it is an **element**.
- (b) Carbon monoxide is the name for the molecule CO, and molecules form a **compound**.
- (c) Chocolate contains sugar, cocoa, and other ingredients that are mixed together so the composition is the same throughout. Chocolate would be a **homogeneous mixture**.
- (d) If you cut open an apple, you can see that there is a skin on the outside and pulp and seeds on the inside. Since the composition is different inside and outside, an apple is a **heterogeneous mixture**.
- (e) Tea has items dissolved in water. The composition is the same throughout, so it is a **homogeneous mixture**.
- (f) Cesium is on the periodic table, so it is an **element**.
- (g) Ocean water has more than just salt and water in it. There are all sorts of items in the ocean water. Also, the composition of ocean water is not the same throughout. This makes ocean water a **heterogeneous mixture**.
- (h) Magnesium chloride is the name for the molecule MgCl_2 , so it is a **compound**.
- (i) A chocolate chip cookie contains several different ingredients, and the composition varies throughout, so it is a **heterogeneous mixture**.

STUDY GUIDE FOR MODULE 2

REVIEW QUESTIONS

1. What is the difference between the continuous theory of matter and the discontinuous theory of matter?
2. What 2 laws were instrumental in the development of Dalton's atomic theory? What law did Dalton predict using his theory?
3. Describe how the law of mass conservation can be used to explain the result of experiment 2.1.
4. List the 4 assumptions of Dalton's atomic theory. Which of the assumptions are wrong? Why?
5.
 - a. What is the difference between an atom and an element?
 - b. What is the difference between a molecule and a compound?
 - c. What is the difference between an atom and a molecule?
 - d. What is the difference between an element and a compound?
6. What are the physical characteristics that distinguish metals from nonmetals?
7. How can you determine whether an atom is a metal or a nonmetal from the periodic table? Are there any exceptions?
8. How can you experimentally determine whether a compound is ionic or covalent?
9. How can you determine from the periodic table whether a compound is ionic or covalent?
10. Which law governs why carbon and oxygen can form 2 different molecules: CO and CO₂?
11. Why do chemists use 2 different naming systems for compounds?
12. If a substance can be physically separated into its components, is it a pure substance or a mixture?
13. If a substance can be given a chemical name, such as calcium bromide, is it a pure substance or a mixture?
14. What element makes up the majority of the air we inhale?
15. Describe the difference between quantitative and qualitative measurements.

16. A change that seems to break a molecule apart is called a _____.
17. A student writes this formula for sodium chloride: NACL. What is wrong with this formula?
18. What are the elements that lie along the jagged line called?
19. Describe a dimensionless quantity.
20. Why doesn't distilled water conduct electricity?

PRACTICE PROBLEMS

1. A chemist uses an experiment similar to experiment 2.2 to decompose 150.0 grams of water (H_2O). If he finds that he makes 16.7 grams of hydrogen from the decomposition, how much oxygen must he have produced?
2. To burn a piece of wood, oxygen gas has to be added. Ash, carbon dioxide, and water are products of the burning process. If 3.8 kg of oxygen gas is used to burn a piece of wood, then what was the mass of the wood if a chemist collects 3.12 kg water, 925 g of ash, and 15.14 kg of carbon dioxide?
3. Identify the following as metal or nonmetal: Br, Sc, P, Ra.
4. In an experiment to make hydrogen peroxide, a chemist mixes 40.0 grams of hydrogen with 40.0 grams of oxygen. She finds that 42.6 grams of hydrogen peroxide are produced. In addition, a lot of hydrogen is left over at the end of the experiment. How many grams of hydrogen should have been added to the 40.0 grams of oxygen so that there would be no leftovers at the end of the experiment?
5. In making calcium nitride, 100.0 grams of calcium plus 100.0 grams of nitrogen makes 152.5 grams of product along with some leftover nitrogen. How much nitrogen and calcium should be added together to make 1.0 kg of calcium nitride without any leftovers?
6. What is the chemical formula of a molecule that is made up of 12 carbon atoms, 24 hydrogen atoms, 10 oxygen atoms, and 2 sulfur atoms?
7. How many total atoms are in 1 molecule of $\text{C}_3\text{H}_6\text{C}_{12}\text{O}$?
8. Which of the following molecules are covalent?
a. C_2H_6 b. PH_3 c. BeS d. Li_2O
9. Name all of the compounds in problem 8.

10. What are the chemical formulas for the following compounds?
 - a. tetranitrogen hexahydride
 - b. dihydrogen monoxide

11. Classify the following as either a mixture or a pure substance:
 - a. Soil
 - b. Silver
 - c. Nitric acid, HNO_3
 - d. Lemonade

12. Reclassify everything in problem 11 as an element, compound, homogeneous mixture, or heterogeneous mixture.



ATOMIC STRUCTURE

In the last module, we began working with atoms. We have used them only in chemical formulas so far, and you might have wondered what they look like. Unfortunately, it's not possible for us to actually see atoms. But for a moment, let's suppose we could. If we could see an atom, what would it look like? That's what we're going to learn about in this module.

You might wonder how in the world we can determine what atoms look like if we cannot see them. Remember, chemists often use things that we can see to learn about things we cannot see. For example, if we were to ask you if it is windy outside, then how would you answer us? You would probably go to the window and look outside. But you cannot see the wind, so how do you know it is windy? You look for evidence of the wind, don't you? Maybe you see a flag flapping, tree branches bending, or paper blowing along the ground. All of these evidences point to the fact that it is windy outside. We can also use indirect observations to help us learn about atoms. Observing how matter responds to certain situations can help us learn more about the structure of the atom.

think about this

"For his invisible attributes, namely, his eternal power and divine nature, have been clearly perceived, ever since the creation of the world, in the things that have been made. So they are without excuse" (Romans 1:20). We cannot see God, but the Apostle Paul states that God is clearly seen in His creation. Everywhere we look, we can see God's attributes. When confronted with all of the evidence that points to God, mankind is without excuse when denying Him.

first reactions

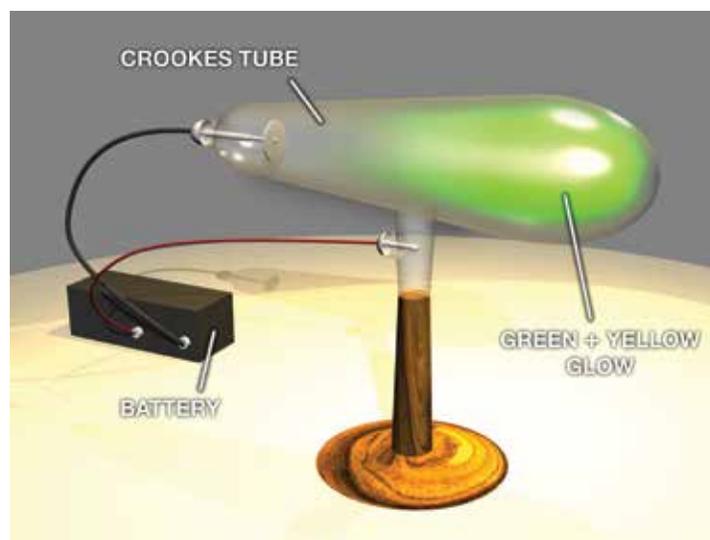
As you learn more about the world around you, you explore more deeply how things are created. It's important to remember, however, that we can never reach the limit of how God created out of nothing, nor the love that chose to create this world. During this journey of discovery, the limits of our knowledge of the smallest of things—the inside of the atom—will always lead us to new questions. Ecclesiastes 8:17 tells us, "No one can comprehend what goes on under the sun. Despite all their efforts to search it out, no one can discover its meaning. Even if the wise claim they know, they cannot really comprehend it" (NIV).

HISTORICAL OVERVIEW

Ever since Democritus first theorized about atoms some 2,400 years ago, scientists have wondered what they might look like. Democritus thought that each type of atom had its own unique shape: some were triangles, some were pentagons, and some were squares. These shapes, he thought, fitted together like jigsaw puzzle pieces to form molecules. As time went on, chemists began to do experiments that started to unlock the mysteries of the atom.

One of the first experiments that shed some light on the question of the atom's structure was performed by an English chemist named William Crookes (krooks). In the second half of the 1800s, Crookes studied how certain gases behaved when exposed to electricity. He filled a glass tube (which we often call a **Crookes tube**) with a tiny amount of gas and connected the tube to a battery, allowing electricity to pass through it. Crookes noticed that when the battery was hooked up, the end of the tube would glow a faint greenish–yellow color, as we see in figure 3.1.

FIGURE 3.1
A Crookes Tube
Illustration by David Weiss.

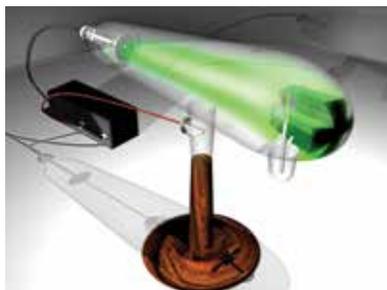


When the battery was not hooked up, there would be no glow. He concluded from this experiment that the battery was producing particles in the gas. These particles, he theorized, traveled through the gas and hit the end of the tube, causing the glow that he saw. Since these particles always seemed to travel from the post on the battery called the cathode (kath' ohd), Crookes called them cathode rays. As a result, a Crookes tube is sometimes referred to as a **cathode ray tube**.

To try to confirm his theory that these cathode rays were particles, he put an obstacle between the cathode and the end of the tube. He reasoned that if these cathode rays were, in fact, particles, they would be blocked by the obstacle and would not hit the end of the tube. Since they could not hit the end of the tube, they would not cause a glow. This is, indeed, just what he saw, as illustrated in figure 3.2.

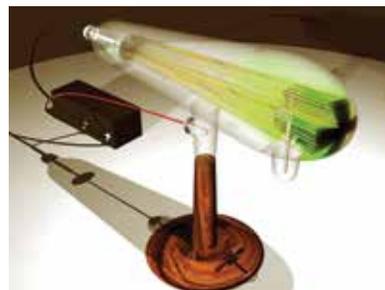
FIGURE 3.2
A Crookes Tube with an Obstacle

Illustrations by David Weiss



The Experiment

An obstacle was placed between the cathode and the wall of the tube.
 A shadow of the obstacle appeared at the end of the tube.



The Interpretation

The results seemed to indicate that particles were coming from the cathode. If they struck the obstacle, they could not make it to the wall, and that explains the shadow.

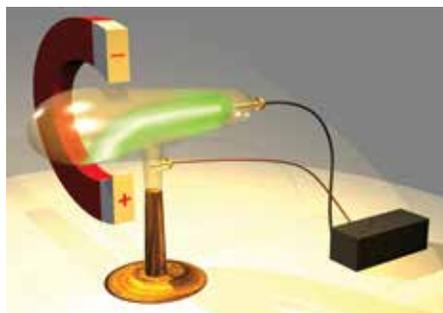
About 20 years after Crookes did his experiments, another English chemist, J. J. Thomson, revisited them to try to determine more about the nature of these cathode rays. Thomson's work ended up demonstrating that one of John Dalton's main assumptions about atoms was wrong. In the last module, we discussed Dalton's theory that assumed that atoms were indivisible. He reasoned that since atoms were the smallest unit of matter, they could not be broken down into smaller pieces. However, Thomson showed that regardless of what type of gas he put in the Crookes tube, cathode rays were always formed. He further showed that these particles were precisely the same, whether he filled the Crookes tube with helium, argon, nitrogen, or any other gas. This led Thomson to believe that these cathode rays were a part of every atom and that his experiments were breaking them away from the atoms.

Thomson's idea that cathode rays were small parts of the atom was further boosted by his discovery that these rays were electrically charged. He determined this by seeing how they reacted in the presence of magnetic and electric fields. To see why this fact was critical in tearing down Dalton's assumption that atoms are indivisible, you first have to understand a few things about electrical charge.

FIGURE 3.3
Crooke's Tube with Magnets

This diagram shows the path cathode rays take in the presence of a magnet. Notice how they bend toward the positive plate of the magnet.

Illustration by David Weiss



ELECTRICAL CHARGE

Everyone knows that there is something called electricity. We know that it can be used to operate lights, televisions, motors, etc. Most people do not understand what electricity is, however. In fact, most scientists don't have a really good understanding of electricity. That's because understanding electricity begins with understanding the concept of electrical charge. Unfortunately, scientists do not have a clear understanding of what electrical charge is or why it exists. We know how electrical charges behave and how they generate electricity, but we really don't know much about what they are or where they come from. Don't worry if you don't understand everything in the discussion that follows; we will point out the concepts that are important for you to understand at the end.

There are 2 types of electrical charge: **positive** and **negative**. We don't know what makes something negative or positive in charge, but we know that these charges exist. We also know that the vast majority of matter on earth does not have any overall electrical charge. We call these things electrically **neutral**. To learn a little more about what it means to be electrically positive or negative, perform experiment 3.1.

EXPERIMENT 3.1

PURPOSE: To investigate electrical charge.

MATERIALS:

- Comb
- Aluminum foil
- Cellophane (Scotch) tape
- Safety goggles

QUESTION: How do electrical charges behave?

HYPOTHESIS: Write down your hypothesis about what the electrical charges will do in the presence of each other.

PROCEDURE:

1. Take 2 pieces of cellophane tape, each approximately 5 inches long, and attach them to the top of a table. Attach them so that about 1/2 inch of tape hangs off the table.
2. Once the tape is firmly attached to the table, grab the ends of the tape that hang over the table edge and quickly pull both pieces so that they detach from the table.
3. Hold the tape so that each piece hangs down.
4. Turn the pieces of tape so that their sticky sides face each other, and slowly bring them together. As you move the sticky sides toward each other, you will see them begin to repel each other. The pieces of tape will float away from each other as you try to move them together.
5. Cut some aluminum foil into several very small pieces and lay those pieces in a small pile on the table.

6. Pass the comb over the foil pieces. Bring the comb as close as you can to the foil without actually touching it. Nothing should happen.
7. Comb your hair with the comb. For this experiment to work, your hair must be reasonably clean. If it is too greasy, the next step of the experiment will not work.
8. After combing your hair, pass the comb over the aluminum pieces as you did before. This time, the aluminum pieces should move on their own toward the comb. If this didn't happen, your hair was probably a bit too greasy. Wash your hair, let it dry completely, then try again.
9. Clean up and return everything to the proper place.

CONCLUSION: How do we explain what happened in the experiment? Write down some thoughts and observations.

How do we explain what happened in the experiment? Let's start with what happened in the first part. To explain the results, you must understand that when you pulled the tape off the table, the sticky substance pulled some electrical charges off the table. Since each piece was pulling off the same type of charge, each piece of tape ended up having the same type of electrical charge. As those like charges were brought together, they repelled each other. Therefore, we come to our first rule of electrical charge:

Like charges repel one another.

What about the second part of the experiment? In that part, you started with a comb and some aluminum foil, neither of which was electrically charged. In fact, both the comb and the aluminum had electrical charges in them, but there were as many positive electrical charges as negative ones. Just as $-1 + 1 = 0$, the negative and positive charges cancelled each other out to give no charge at all.

However, when you combed your hair, the comb picked up some stray negative electrical charges from your hair. This caused the comb to have more negative charges than positive charges, so it became negatively charged. When the comb was then placed near the aluminum foil, the negative charge of the comb repelled all of the negative charges in the aluminum foil. The negative charges in the aluminum foil got as far away from the comb as possible. But when that happened, the parts of the foil near the comb had more positive charges than negative ones, and those positive charges were attracted to the comb. Therefore, the foil pieces began to move toward the comb. Therefore, we come to the second rule of electrical charge:

Opposite charges attract each other.

The other important thing to remember about electrical charge was mentioned in the previous discussion:

Every substance on earth has electrical charges. However, if the number of positive charges equals the negative charges, the substance is neutral with no overall charge ($-1 + 1 = 0$). If there is more of one charge than the other, the substance takes on that electrical charge ($-2 + 1 = -1$).

Therefore, the comb started with no overall electrical charge because it had an equal number of positive and negative charges in it. However, when it picked up a few extra negative charges from your hair, it became negatively charged because there were more negative charges than positive charges in it.

If you don't understand everything you've just read, don't worry. The concepts of electrical charge are difficult to comprehend. Even the most brilliant chemist in the world still doesn't fully understand them. What you have to be able to understand are the 3 blue statements above. Those are the foundational concepts of electrical charge that we will use throughout this module.

ELECTRICAL CHARGE AND ATOMIC STRUCTURE

We now have enough information under our belts to understand why Crookes's and Thomson's experiments destroyed Dalton's postulate that atoms are indivisible. When the Crookes tube was filled with gas, the gas was always electrically neutral. The gas had as many negative charges as positive charges. However, when the battery was hooked up to the Crookes tube, negatively charged particles were produced. Thomson reasoned that those negatively charged particles must have been taken from the atoms of gas, much like the comb you used in the previous experiment pulled negative charges from your hair.

If negative charges could be pulled from an electrically neutral atom, then the atom must be composed of smaller components. Some of those components must be negatively charged, and the other components must be positively charged. Since the gas atoms started out neutral, there must be an equal number of positively and negatively charged components in atoms. Crookes and Thomson showed that these charged components could be separated, demonstrating that atoms could be broken down into smaller components. This destroyed Dalton's assumption that atoms are indivisible.

Thomson later did a famous experiment that measured the ratio of the charge and mass of these negatively charged particles. He showed that no matter what atoms they were pulled from, these particles always had the same ratio of charge to mass. He considered these particles to be something common to all atoms, and he called them **electrons** (ee lek' trahnz). As we have already said, since electrons are pulled from electrically neutral atoms, there must also be positive components in the atom. Thomson's student, Ernest Rutherford, discovered those components and called them **protons** (pro' tahns). In 1932, another English scientist, James Chadwick, discovered that there was an additional component to the atom. This component was electrically neutral (it had no charge), so Chadwick called it the **neutron** (new' trahn).

By 1932, scientists had determined that atoms were made up of 3 components: protons, electrons, and neutrons. What makes one atom different from another atom is the number of protons, neutrons, and electrons each contains. For example, look at the periodic table again. The first box on the periodic table represents the atom hydrogen. A hydrogen atom is composed of 1 proton, 1 electron, and 0 neutrons. The second box

on the table represents the atom helium, which contains 2 protons, 2 electrons, and 2 neutrons. Hydrogen and helium have totally different properties. For example, hydrogen is explosive when placed in contact with a flame or even a spark. On the other hand, helium will never burn, no matter how hot a flame it comes into contact with. This difference is the result of the different number of protons, neutrons, and electrons that make up the atoms.

DETERMINING THE NUMBER OF PROTONS AND ELECTRONS IN AN ATOM

You might wonder how we know that a hydrogen atom has 1 proton and a helium atom has 2. This is what the top number in each box of the periodic table tells us. We call that number the **atomic number**, and it tells how many protons are in an atom.

An atom's atomic number tells how many protons it contains.

In addition, since all atoms are electrically neutral, we know they must contain an equal number of protons (positive charges) and electrons (negative charges).

All atoms have an equal number of electrons and protons.

ON YOUR OWN

- 3.1 How many protons and electrons are in the following atoms?
- molybdenum (Mo)
 - beryllium (Be)
 - hafnium (Hf)

Therefore, if you look at the periodic table, you will see that sulfur (S) has an atomic number of 16. This tells you that a sulfur atom has 16 protons and 16 electrons in it. Make sure you understand this idea by answering “On Your Own” question 3.1.

DETERMINING THE NUMBER OF NEUTRONS IN AN ATOM

As you can see, when you are armed with the periodic table, it is very simple to determine the number of protons and electrons in an atom. Determining the number of neutrons in an atom is a little more difficult, however. To understand why this is the case, we must talk more about the periodic table. We said earlier that each box in the periodic table represents a single type of atom, but that's not really true. In fact, each box on the periodic table represents several different atoms, but all of those atoms are part of the same element. For example, the sixth box on the table represents carbon. There are 3 different types of carbon atoms. Each has the same number of protons and electrons (6), but they have different numbers of neutrons. When atoms have the same number of protons and electrons but different numbers of neutrons, we call them **isotopes** (eye' suh tohpz).

Isotopes—Atoms with the same number of protons but different numbers of neutrons.

The most common type of carbon atom is composed of 6 protons, 6 electrons, and 6 neutrons. There are 2 other types of atoms which belong to the element we call carbon. The second type of carbon atom still has 6 protons and 6 electrons, but it has 7 neutrons.

The third type of carbon atoms has 6 protons, 6 electrons, and 8 neutrons. All 3 of these types of atoms are made up of the same number of protons and electrons but different numbers of neutrons. Therefore, they are all isotopes.

What is the difference between these carbon isotopes? Not much. The one with 8 neutrons is a little heavier than the one with 7 neutrons. Similarly, the one with 7 neutrons is a bit heavier than the one with 6 neutrons. Other than that, their chemical natures are precisely the same. Any reaction that one of the isotopes can participate in, they all can participate in. In fact, if you were to have a lump of carbon (a black solid) in your hand, all 3 isotopes would be present. The isotope with 6 neutrons would make up about 98.9% of all atoms in the sample, the isotope with 7 neutrons would make up about 1.1%, and the isotope with 8 neutrons would make up less than $1.0 \times 10^{-12}\%$. These isotopes would appear identical in almost every way. They would all participate exactly the same in any chemical reaction. You could not find a way to separate them from one another chemically. This is a very important point:

Isotopes behave identically in their chemistry;
the main difference between them is their mass.

Every element is composed of several isotopes. This means that every box on the periodic table actually represents several different isotopes. Since that is the case, it is impossible to look at the periodic table and determine how many neutrons are in a given atom. To determine the number of neutrons in an atom, we must be given a little more information. This information is provided by the **mass number**.

Mass number—The total number of neutrons and protons in an atom.

This mass number is not always provided when discussing elements because often it is not necessary to consider the number of neutrons in an atom. However, if the mass number is given, then it is easy to determine how many neutrons are in that particular atom.

Remember, the most common isotope of carbon has 6 protons and 6 neutrons. Its mass number is $6 + 6 = 12$. We can include this mass number in 2 ways. When writing out the full name of the element (carbon), we can add a hyphen followed by the mass number and call this isotope carbon-12. On the other hand, if we just want to write the symbol for carbon (C), then the mass number must appear as a superscript *preceding* the symbol. The symbol for this isotope would be ^{12}C . What about the other isotopes of carbon? The one with 6 protons and 7 neutrons would have a mass number of $6 + 7 = 13$. Therefore, we could call it carbon-13 or ^{13}C . Similarly, the isotope with 6 protons and 8 neutrons would be called carbon-14 or ^{14}C .

FIGURE 3.4
Atomic Symbol (Magnesium)

Illustration by David Weiss



Before we go any further, let's make a point to clear up some confusion you might have. Some students interpret the term *isotope* incorrectly, thinking that there is one normal version of an atom and that all the others are isotopes. For example, they think that carbon-12 is the normal version of carbon and that carbon-13 and carbon-14 are isotopes. That's *not* the proper use of the term. *All 3 types of carbon are isotopes.* Look back at the definition. Carbon-12, carbon-13, and carbon-14 all have 6 protons, but they have different numbers of neutrons. They are *all* isotopes of carbon. Think about it this way: The term *isotope* is *relational*. It refers to how atoms relate to one another. Consider 3 men who have the same father. We would say that these men are brothers. The term *brother* is relational. It tells how the men relate to one another. There is not one normal man, while the other 2 men are the brothers! They are *all* brothers. In the same way, if several atoms all have the same number of protons (the same atomic number) but different numbers of neutrons, they are *all* isotopes.

Now that you have learned about isotopes, you have what you need to determine the makeup of any atom on the periodic table, provided you have enough information. Study example 3.1 and then tackle "On Your Own" questions 3.2–3.3 to be certain you understand how this is done.

EXAMPLE 3.1

What are the name and symbol of an atom that is made up of 10 protons, 10 electrons, and 11 neutrons?

Since the atom has 10 protons, its atomic number is 10. The periodic table shows that neon (Ne) is the element with that atomic number. To give the full name, we must also include its mass number. Since it has 10 protons and 11 neutrons, its mass number is $10 + 11 = 21$. Therefore, the name of the atom is **neon-21**, and its symbol is **^{21}Ne** .

How many protons, electrons, and neutrons make up a ^{129}I atom?

To determine the number of protons and electrons, all we have to do is find I (iodine) on the periodic table. We see that its atomic number is 53. This means it has **53 protons and 53 electrons**. In addition, we are given its mass number, which represents the total number of protons and neutrons. If the atom has 53 protons, and its total number of protons + neutrons is 129, then the number of neutrons it has is $129 - 53 = 76$ **neutrons**.

ISOTOPES AND NUCLEAR BOMBS

Before we leave this discussion of protons, electrons, neutrons, and isotopes, we should point out one more thing. If you pay attention to the news at all, you have probably heard about how the United States and other nuclear powers in the world are trying to keep nuclear technology away from certain countries and are concerned about those countries that are developing nuclear technology. Have you ever wondered exactly

what nuclear technology is? Most people would say that it refers to the ability to make a nuclear bomb. In fact, that's not true at all! Any person who has studied nuclear physics at the college level has all of the knowledge needed to build a nuclear bomb! The trick is not in making the nuclear bomb; almost anyone (with a little training) can do that. The real trick to making a nuclear bomb is in making the bomb's fuel.

Uranium is the usual fuel used to make nuclear bombs. However, the element uranium has 3 naturally occurring isotopes: ^{238}U , ^{235}U , and ^{234}U . A sample of natural uranium is about 99.3% ^{238}U , about 0.7% ^{235}U , and less than 0.01% ^{234}U . Only the ^{235}U isotope can be used to make a bomb. For a nuclear bomb to work, the uranium used in the bomb must have a significantly larger percentage of ^{235}U than natural uranium has. In other words, to be able to produce a nuclear bomb, the amount of ^{235}U in the uranium used must be artificially increased. The way to do this, of course, is to get rid of some of the ^{238}U in the uranium. This process, called **isotopic** (eye' suh top' ik) **enrichment**, is very tricky. After all, the only real difference between ^{235}U and ^{238}U is the mass. Chemically, they behave in exactly the same way. This makes it nearly impossible to separate them for isotopic enrichment.

Of course, nearly impossible and impossible are 2 different things. During World War II, scientists in a group called the Manhattan Project were able to figure out a way to partially separate ^{238}U from ^{235}U . The process is incredibly inefficient, but it is able to take natural uranium and enrich the ^{235}U content from 0.7% to a little more than 7%, which is enough to fuel a nuclear bomb. Ultimately, making a nuclear bomb is simple; however, getting the fuel to power it is nearly impossible, and the technique for making this fuel is difficult and not readily shared.

ATOMIC STRUCTURE IN MORE DETAIL

Now that you know how to determine the number of protons, neutrons, and electrons in an atom, you might ask how these particles arrange themselves in an atom. Are they jumbled together like a bunch of marbles in a sack? Do they arrange themselves in groups? Do they form any neat arrangements or designs? These are the same questions that chemists were asking themselves in the early 1900s.

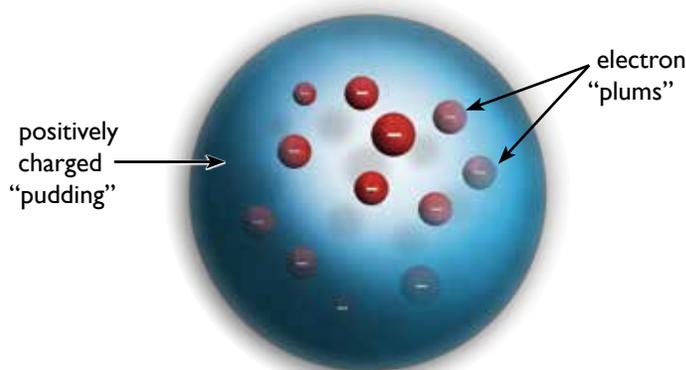
After electrons were first discovered (and before neutrons were discovered), chemists thought that an atom looked a lot like a dish of plum pudding. This concept, called the plum pudding model, was developed by British physicist, J.J. Thomson. It is illustrated in figure 3.5.

ON YOUR OWN

- 3.2 Give the number of protons, electrons, and neutrons that make up the following atoms:
 a. ^{22}Ne b. Fluorine-19 c. ^{235}U
- 3.3 An atom has 33 protons, 33 electrons, and 41 neutrons. What are its full name and symbol?

FIGURE 3.5
The Plum Pudding Model of the Atom

Illustration by David Weiss



In essence, chemists thought that the atom was made up of a pudding-like substance which had positive electrical charge. Electrons, they thought, were suspended in this "pudding" like plums, making the entire atom electrically neutral. Remember, it's impossible to see atoms; so chemists just had to imagine what an atom might look like. That's why we call this concept a **model** of the atom. A model is simply a constructed image of something we cannot see with our eyes.

Since we can't see atoms, how do we know whether or not the plum pudding model is correct? The best way to determine that is to compare it to experimental data. If the model is inconsistent with the experimental data, then the model is obviously incorrect. However, if the model is consistent with the data, then we don't necessarily know that the model is right, but at least we know that it might not be wrong. The trick is in designing an experiment that will produce data that really test all aspects of the model.

Ernest Rutherford (Thomson's student) designed the perfect experiment to test the validity of the plum pudding model. In earlier experiments, Rutherford had identified a type of radiation called alpha particles. He determined that these alpha particles have positive electrical charges and are emitted by certain isotopes that were called radioactive. You will learn a great deal more about radioactive isotopes if you take advanced chemistry. For now, just remember that **alpha particles** are positive particles emitted by certain radioactive isotopes.

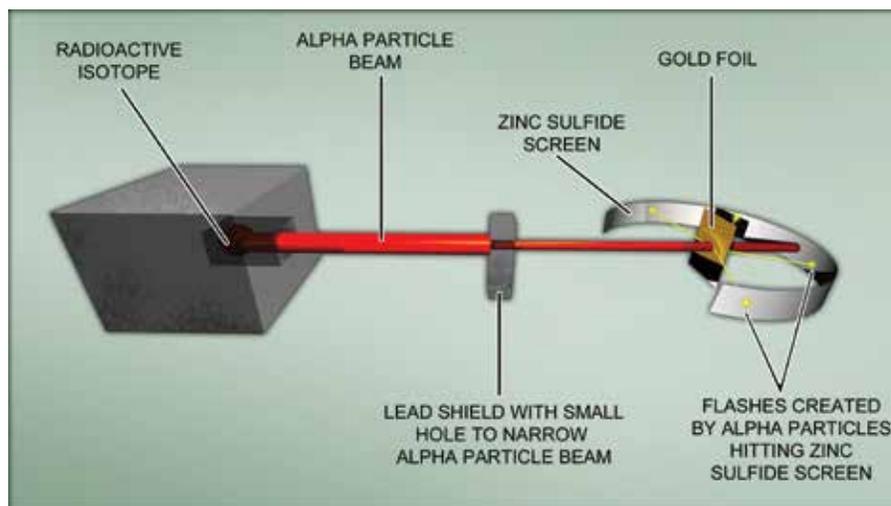
Rutherford decided that what he needed to do was to shoot some of the alpha particles at a thin metal foil. He reasoned that if the plum pudding model of the atom were correct, the alpha particles should pass through the foil without changing their direction much. After all, the positive alpha particles would be attracted by the negative "plums" (the electrons) but repelled by the positive "pudding." Since the plums were randomly distributed around the pudding, the alpha particles would be attracted and repelled in many different directions, and the result would be that they would travel straight through the atom.

If you have trouble understanding why the alpha particles should travel straight through a plum pudding atom, think about it this way. Suppose you were trying to move a large boulder by tying many ropes to it and having several people pull on the ropes

while others pushed on the boulder. If everyone pulled and pushed in different, random directions, the boulder wouldn't go anywhere. To move the boulder, everyone would need to pull in one direction and push in roughly the opposite direction. The same principle applies here. When traveling through a plum pudding atom, the alpha particles would be pulled and pushed in different, random directions. This would result in the alpha particles not being moved from their original path.

We hope that you can see the ingenuity behind Rutherford's experiment. The data provided would be an excellent test of the model in question. If the alpha particles did not behave as Rutherford expected, then clearly the plum pudding model was wrong. However, if they did behave as expected, then that would be one more bit of evidence in favor of the plum pudding model. In the end, the experiment looked something like figure 3.6.

FIGURE 3.6
Rutherford's Experiment
Illustration by David Weiss



In this experiment, Rutherford put a radioactive isotope in a thick lead box with a single hole in it. The radioactive isotope inside the box was emitting alpha particles (a type of nuclear radiation). Most of the alpha particles were stopped inside the box, but any that were emitted in the direction of the hole escaped. This made a beam of alpha particles that could be pointed toward a target. The beam was narrowed by another piece of lead with a hole in it. In the end, Rutherford built an alpha-particle shotgun out of the lead box that directed a thin beam of alpha particles to hit a thin piece of gold foil.

Rutherford surrounded the target with a zinc sulfide screen. When an alpha particle hits zinc sulfide, a bright green glow is emitted. Rutherford could see where the alpha particles went after they collided with the gold foil by simply looking at where the zinc sulfide screen was glowing. Wherever it was glowing very brightly, many alpha particles were hitting it. Wherever it was glowing faintly, only a few alpha particles were hitting it.

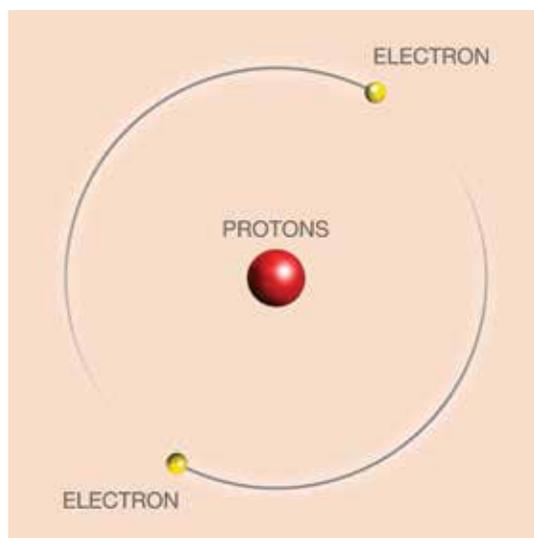
Rutherford ran his experiment and, to his astonishment, the zinc sulfide screen glowed everywhere! The glow was the brightest directly in front of the target, but alpha particles were hitting the zinc sulfide screen at all other positions as well. Rutherford noticed that

the farther away from the front of the target, the dimmer the glow was. Nevertheless, alpha particles were colliding with the target and bouncing in every conceivable direction.

This, of course, spelled the end of the plum pudding model of the atom. After all, there was no way to imagine how the alpha particles could possibly bounce off of the target and change direction if the plum pudding model of the atom were correct. Rutherford tried his experiment with several different types of targets to make sure that his conclusion applied to all atoms. In the end, the results were the same, regardless of what target was used.

Of course, it wasn't enough for Rutherford to simply determine that the plum pudding model of the atom was wrong. Now that he had all this experimental data, he needed to try to develop a new model of the atom that was consistent with all of the information he had up to that point. Rutherford developed such a model and called it the **planetary model** of the atom, but it is usually referred to today as the **Rutherford model** of the atom. This model is illustrated in figure 3.7.

FIGURE 3.7
The Rutherford Model of the Atom
Illustration by David Weiss

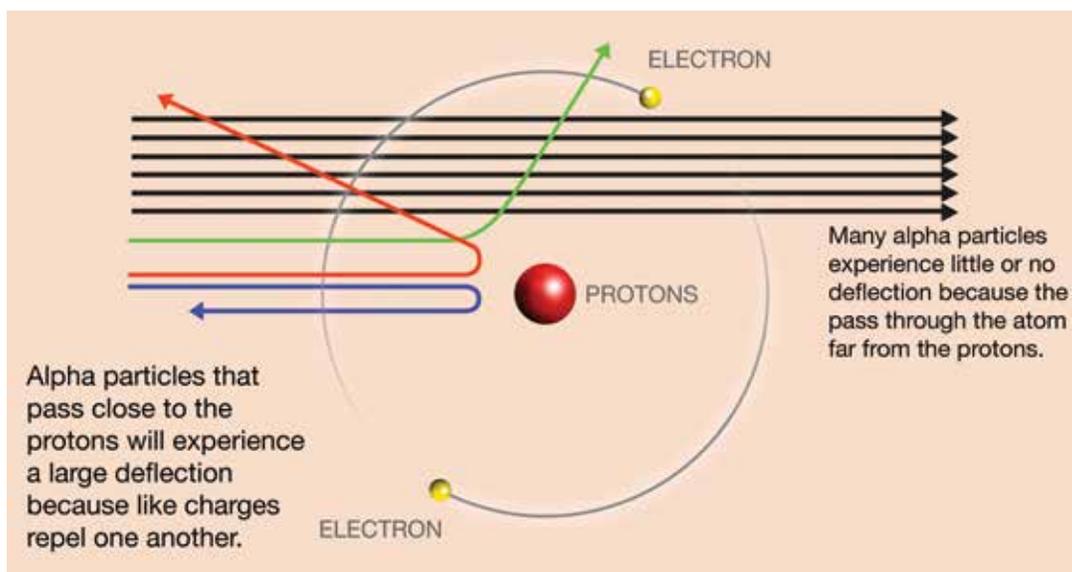


In Rutherford's model, the protons were clustered together at the center of the atom, called the **nucleus** (new' klee us), while the electrons orbited around the nucleus in the same way that the planets in our solar system orbit around the sun. This similarity between his view of the atom and our solar system is why Rutherford called it the planetary model of the atom. Notice that Rutherford's model does not contain neutrons because neutrons had not been discovered yet.

Realize that this idea of what the atom looks like was not something Rutherford just dreamed up. It was the only feasible model that explained the data he got from his experiment. Remember, Rutherford's experiment showed that when positive particles were shot at atoms, most of those particles passed through the atoms undisturbed. However, some were deflected away from their original path. Sometimes the deflection was enormous, and the alpha particle would bounce backward after hitting an atom.

Rutherford's model is consistent with all of this data. To help you understand why, look at figure 3.8.

FIGURE 3.8
Why Rutherford's Model Is Consistent with His Data
 Illustration by David Weiss



In the figure, the arrows represent alpha particles that are passing through the atom. Unlike the plum pudding model, this model has the positive electrical charge concentrated at the center of the atom. The negative charges (the electrons) are orbiting around the protons. Since the electrons are not concentrated in one spot, they do not affect the path of the alpha particles greatly. The protons cause most of the deflection of the alpha particles. The vast majority of alpha particles that pass through the atom, then, will not be deflected because they will pass through the atom far from the protons. These alpha particles, represented by the straight arrows, will continue to travel straight, hitting the zinc sulfide screen in front of the target.

However, if an alpha particle passes through the atom close to the protons, it will be repelled by the protons' positive charge and will be deflected from its original path. The closer it passes to the protons, its path will bend. If it is headed directly toward the proton, it will be repelled so much that it will turn around and head in the opposite direction or appear to bounce back. These deflected alpha particles are represented by the curved arrows.

The vast majority of alpha particles are not deflected because they do not pass very close to the protons and they strike the zinc sulfide screen directly in front of the target, as Rutherford saw in his experiment. Rutherford's observation that the number of alpha particles that were deflected is very small also makes sense in terms of this model because the number of particles that pass close to the protons is quite small. In the end, then, Rutherford's model was consistent with all of the data produced in his experiment.

The one thing Rutherford's model was not consistent with, however, was a current theory in physics. Physicists during Rutherford's time had noticed that whenever electrically charged particles moved in circles, they emitted light. When they emitted this light, they would lose energy and slow down. According to this theory, the Rutherford model of the atom would never be stable. This theory would predict that the electrons moving in a circle would continue to emit light until they lost all of their energy. It would also predict that as the electrons lost energy, they would spiral toward the protons in the atom until they collided with them.

Even though Rutherford's model was not consistent with a current theory in physics, it was the only model of the atom that could be devised that was consistent with his experimental data. It became the accepted model of the atom until it was later modified by Niels Bohr (boor). To understand Bohr's model of the atom and what experiments led up to it, you must first learn a little about light.

THE NATURE OF LIGHT

In module 1, you learned that since light neither has mass nor takes up space, it is not matter. However, you haven't yet learned what light is and what its properties are. That's the purpose of this discussion.

The first thing you must realize about the nature of light is that scientists don't completely understand it. There are currently 2 models for light: the particle model and the wave model. The particle model states that light is made up of small particles called **photons**. Any beam of light that you see is made up of billions of individual particles (photons) that are all traveling in the same direction. The wave model, on the other hand, says that a beam of light is a wave that travels much like a wave on the ocean.

To make things even more confusing, sometimes light behaves like a particle, and sometimes it behaves like a wave. Scientists really aren't sure they understand this phenomenon, but they accept it. Currently, theory states that light acts like both a particle and a wave. The situation that light is in determines whether it will behave as a particle or a wave. Scientists call this theory the **particle/wave duality theory**, and it is used frequently in describing light.

Particle/wave duality theory—The theory that light sometimes behaves as a particle and sometimes behaves as a wave.

If you don't understand the particle/wave duality theory, don't worry. Scientists don't understand how light can be both a particle and a wave, but for now we are forced to

think about this

We associate light with illumination. At first, you might think of the physical properties of light that you experience every day. However, light is used to describe many things, including God, our testimony, and righteousness, to name a few. For example, Psalm 18:28 says, "For it is you who light my lamp; the LORD my God lightens my darkness." In this verse, light is used as a metaphor to describe God as He directs our path. Without God illuminating our path, we will stumble and fall into darkness. Also, Matthew 5:16 says, "In the same way, let your light shine before others, so that they may see your good works and give glory to your Father who is in heaven." Here the apostle Matthew uses light to describe how Christians are to contrast with a dark and sinful world.

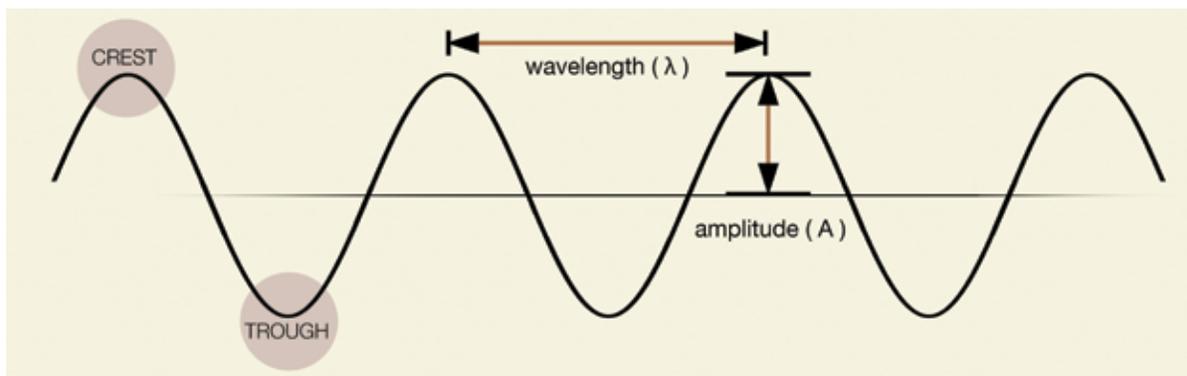
believe it because that's what the experimental data indicate.

The wave model of light is by far the most useful way to picture light, so we will start there. Consider the wave in figure 3.9.

FIGURE 3.9

A Wave

Illustration by David Weiss



This is one way to picture light. It moves through space just like a wave moves through water. The high points of the wave are called **crests**, and the low points are called **troughs**. If we were to describe the above wave, we would need to specify 2 things. First, we have to say how far apart the waves are. We use the term **wavelength** to accomplish this.

Wavelength—The distance between the crests (or troughs) of a wave.

Since wavelength is a measurement of distance, its units will be meters, centimeters, etc. The wavelength of a wave is usually given the symbol λ (the lower-case Greek letter *lambda*), as illustrated in the figure.

To describe the wave completely, we not only have to tell how far apart the waves are; we also need to tell you how big the waves are. This is referred to as the **amplitude** of the wave.

Amplitude—A measure of the height of the crests or the depths of the troughs on a wave.

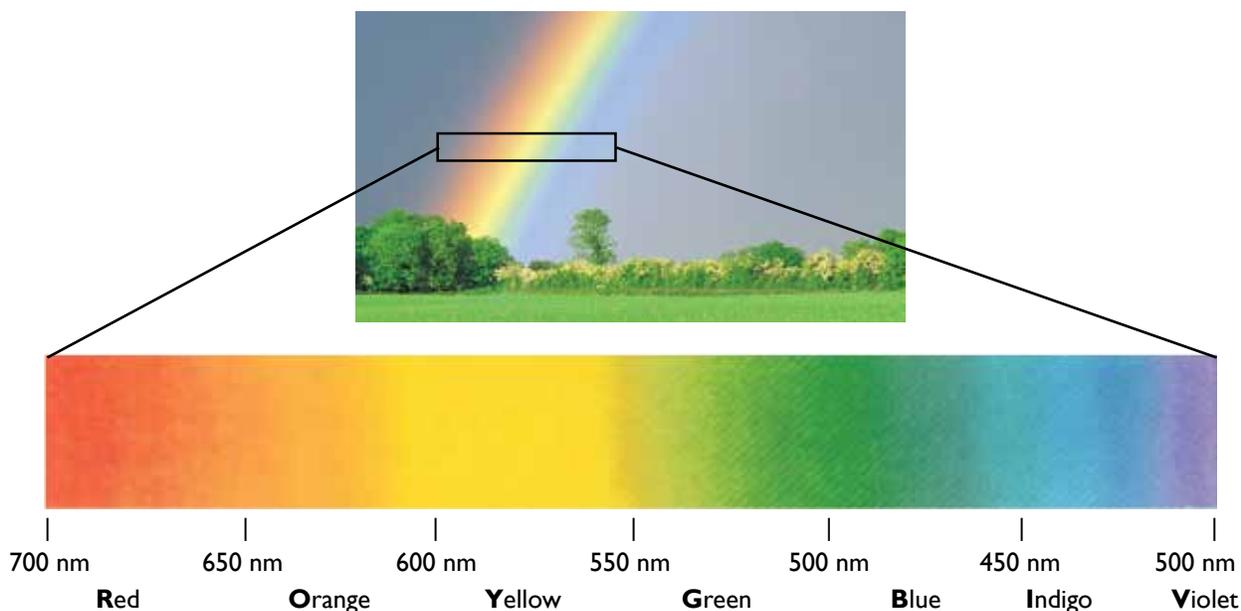
Amplitude is usually given the symbol A , as illustrated above. Amplitude is not a distance measurement, however. The units for amplitude vary from wave to wave, and we will not get into those distinctions here. If you take physics, you will learn about the units of amplitude. For now, just remember that amplitude means how big the wave is. The larger the amplitude, the bigger the wave.

How does all of this apply to light? The wavelength of light corresponds to its *color*. Perhaps you have seen what happens when sunlight hits a prism. Once it passes through the prism, the apparently white light turns into a rainbow of colors. This is because through laws you will learn in physics, a prism breaks light into its different wavelengths. Each color you see represents a range of wavelengths.

Think about the last time you saw a rainbow. What colors did you see? Red was at the top of the rainbow; orange was next; yellow followed orange; and then there were green, blue, indigo, and violet. These are the basic colors of light, and they always appear in a rainbow in that order. Why?

FIGURE 3.10
The Visible Spectrum

Rainbow photo ©GoodShoot/SuperStock, Inc./Visible spectrum illustration by Megan Whitaker



For example, red light is made up of waves whose wavelengths are between 650 and 700 nanometers long. The prefix *nano-* means 10^{-9} , so red light’s wavelengths are very small. However, of all the colors, red has the longest wavelength. Figure 3.10 summarizes each basic color and the approximate range of wavelengths that make up that color. This range of light wavelengths is typically called the **visible spectrum**. You will learn why it is called this a bit later.

As you will learn in physics, the wavelength of light determines how the light passes through a prism. Since wavelength determines color, the colors of light in a rainbow always appear in the same place because they always pass through the water droplets in the air in exactly the same way.

Although you needn’t memorize the wavelengths that each color corresponds to, you *do* need to memorize the *relative* size of the wavelengths in question—that is, you need to know that red light has the largest wavelengths, orange light has smaller wavelengths,

think about this

In Genesis 9:11-17, God promised Noah that the rainbow would symbolize His promise to never again destroy the earth with a flood. A rainbow is formed by little water droplets suspended in the air after a rainfall. These water droplets act like little prisms, separating light into different wavelengths. What an amazing God we serve who can take billions of tiny little prisms and create an eternal promise for us!

and so on. This is easy if you think about the colors as a single name. The first letter of each color is listed in the figure in boldface type to emphasize that when you put them together in the proper order, you come up with a man's name: ROY G. BIV. If you think of ROY G. BIV every time you think of the colors in the visible spectrum, you will always know that red light has the longest wavelength and violet light has the smallest, and you will also know the order of all colors between them.

Now that we know what the wavelength of light means, what does the amplitude of light mean? Simply put, the amplitude of a light wave tells how bright the light wave is. A dim light bulb produces light waves with a small amplitude, while a bright light bulb produces light with a large amplitude. If we had 2 light bulbs of the same color but one was dimmer than the other, we could say that the light bulbs were both producing light of the same wavelength but of different amplitudes. Similarly, if we had 2 light bulbs of different color but the same brightness, we could say that the light bulbs produced light waves of equal amplitude but of different wavelengths.

There's a bit more to describing a light wave than just its amplitude and wavelength; the speed of the light wave is also important. After all, we know that light travels from place to place. If we turn on a light bulb, the light rushes from the light bulb and fills up the room. Therefore, light does travel. But how fast does it travel? Have you ever watched a thunderstorm and seen a lightning flash but heard the thunder crash a few moments later? This happens because the light generated by the lightning flash travels more quickly than the sound it makes. The light reaches your eyes before the sound reaches your ears. For every 5 seconds between the time you see the lightning flash and the time you hear the thunderclap, the lightning is striking 1 mile away from you.

Light, then, travels fast—much faster than sound. In fact, light travels in air with a speed of 3.0×10^8 meters per second (m/s). In units you are more familiar with, that translates to 9.8×10^8 feet per second (ft/s). That's really moving! To put it another way, light travels 190,000 miles every second. One thing that's very interesting about this number is that it does not change with the wavelength or amplitude of the light wave. Regardless of the color or the brightness of light, its speed in air is always 3.0×10^8 m/s. Since this value doesn't change, we call it a **physical constant**.

Physical constant—A measurable quantity in nature that does not change.

You will be introduced to several physical constants in this course, and the speed of light is the first one. Each physical constant has its own symbol. The symbol for the speed of light is c . When we say that $c = 3.0 \times 10^8$ m/s, we are giving you the speed of light. You do not need to memorize this number. We will always provide you with any physical constants you need to solve the assigned problems.

Now that we know about the speed of light, we can finally speak of another very important means of describing a wave: **frequency**.

Frequency—The number of wave crests (or troughs) that pass a given point each second.

To best understand frequency, suppose you are standing in the ocean. If you look out at the waves, the distance between each crest is the wavelength. On the other hand, if you

stand still and count the number of wave crests that hit you over a certain interval of time, that is frequency.

Think for a moment about the unit associated with frequency. Suppose you stand still in the water at the beach and count that 2 ocean wave crests hit you every second. You would report your measurement as 2 wave crests per second. However, since everyone at the beach can see that you are counting waves, you might just say “2 per second.” What does *per second* mean? The unit meters per second is mathematically expressed as meters/second. *Per second* just means 1/second. This means that the standard unit for frequency is simply 1/second. This unit is usually called the **Hertz** (abbreviated **Hz**), in honor of Heinrich Rudolf Hertz, a German physicist who greatly expanded our knowledge of light and its relationship to electricity.

The frequency, wavelength, and speed of light are all related to each other by a simple equation:

$$f = \frac{c}{\lambda} \quad \text{Equation 3.1}$$

In this equation, f stands for frequency, c is the speed of light, and λ stands for wavelength. Notice how the units work out. Speed is in m/s, while wavelength is in meters. When we divide the speed by the wavelength, meters cancel, and we are left with 1/s, which is also called Hertz. This is a very important equation, and you should memorize it. See how we use it in example 3.2.

salt and light

Imagine holding a world record that no one has been able to surpass for nearly 100 years.

Edward William Morley (1838-1923)

has a record in science that no one has been able to surpass. He was able to determine the atomic weight of oxygen to more

precision using wet chemistry techniques than anyone else has ever accomplished. Today, atomic masses are determined by physical means, using sophisticated instruments with computers. However, Morley was able to figure the atomic mass of oxygen using techniques involving solutions, reactions, and typical glassware found in a chemistry laboratory. His work set the standard for laboratory precision that can still be followed today.

We have included Edward Morley at this point due to his work in another area of science. He was known for helping Albert Michelson measure the speed of light. This research was a significant building block for Albert Einstein to develop his theory of relativity. The partnership between Morley and Michelson was full of troubles, but through it all, Morley stayed true to his belief in God even though Michelson was an atheist. Christians working in the field of chemistry may need to be able to work alongside atheists at some point in their careers. It is important for you to develop your Christian walk at an early age so as to not be influenced by others.

Morley, who was homeschooled until the age of 19, wrote these words while a professor at Western Reserve University: “I believe that Jesus Christ shall come with the clouds of heaven to judge the world in righteousness and that those who have believed in Him shall inherit eternal life through the grace of God” (Graves 1996, 155). That is a very bold statement of his faith, so Edward Morley is a great example of a man of strong faith who had great skills as a scientist.



Photo: Public Domain

EXAMPLE 3.2

A violet light wave has a wavelength of 405 nm. What is its frequency? ($c = 3.0 \times 10^8 \frac{\text{m}}{\text{s}}$, and the prefix *nano-* means 10^{-9}).

To solve this problem, we first must recognize that we use equation 3.1:

$$f = \frac{c}{\lambda}$$

To find the frequency, we just need to divide the speed, which is given, by the wavelength, which is also given. However, before we can use the equation, we must make sure our units work out. The wavelength is in nm, while the speed is in m/s. For the equation to work, our units must agree. Either m needs to change to nm or vice versa. We choose to change nm into m. According to what's given:

$$1 \text{ nm} = 10^{-9} \text{ m}$$

$$\frac{405 \text{ nm}}{1} \times \frac{10^{-9} \text{ m}}{1 \text{ nm}} = 4.05 \times 10^{-7} \text{ m}$$

Try to put that equation into your calculator to see if you can get the right answer. If your calculator is a scientific calculator, you can put the numbers into the calculator in scientific notation. However, most calculators want you to input scientific notation in a very specific way. If you don't do it correctly, your answers will be off by a factor of 10 or more.

Now that our units agree, we can put the numbers into the equation:

$$f = \frac{3.0 \times 10^8 \frac{\text{m}}{\text{s}}}{4.05 \times 10^{-7} \text{ m}} = 7.4 \times 10^{14} \frac{1}{\text{s}}$$

Remember, since the speed of light as given has only 2 significant figures, we can only report our answer to 2 significant figures as well. The frequency is **$7.4 \times 10^{14} \text{ Hz}$** . What does the answer mean? It means that when you look at violet light, 7.4×10^{14} crests of the light wave hit you every second! Once again, make sure you can put this equation into your calculator and get the right result. If you cannot, consult your calculator manual or contact the Apologia Help Line.

A light wave has a frequency of $4.70 \times 10^{14} \text{ Hz}$. What are its wavelength and color according to the chart in figure 3.10? ($c = 3.0 \times 10^8 \text{ m/s}$; *nano-* means 10^{-9})

Once again, we use equation 3.1:

$$f = \frac{c}{\lambda}$$

But now we have to rearrange it so that we are solving for wavelength:

$$\lambda = \frac{c}{f}$$

Now we can plug in the numbers:

$$\lambda = \frac{3.0 \times 10^8 \frac{\text{m}}{\text{s}}}{4.70 \times 10^{14} \frac{1}{\text{s}}} = 6.4 \times 10^{-7} \text{ m}$$

The chart in figure 3.10 lists wavelength in nm. If we are to refer to the chart, we must convert to nm:

$$\frac{6.4 \times 10^{-7} \text{ m}}{1} \times \frac{1 \text{ nm}}{10^{-9} \text{ m}} = 640 \text{ nm}$$

The wavelength is **640 nm**. According to figure 3.10, then, this light is **orange light**.

ON YOUR OWN

($c = 3.0 \times 10^8 \frac{\text{m}}{\text{s}}$; *nano-* means 10^{-9})

- 3.4 What color is light that has a frequency of 6.4×10^{14} Hz?
- 3.5 If red light has wavelengths from 7.00×10^2 nm to 6.50×10^2 nm, what is the frequency range for red light?

Make sure you understand how to use equation 3.1 by answering “On Your Own” questions 3.4–3.5.

THE ELECTROMAGNETIC SPECTRUM

Now that you’ve had some experience using equation 3.1, go back and just look at it for a moment. Notice how the equation shows that frequency and wavelength are related. As wavelength gets larger, equation 3.1 tells us that we

will be dividing by a large number. What does that tell us about frequency? It tells us that for large wavelengths, frequency is small. Conversely, if wavelength is small, frequency is large. When 2 quantities behave like this, we say that they are **inversely related** to one another. When studying light, this inverse relationship between frequency and wavelength is very important to remember.

When wavelength is large, frequency is small.

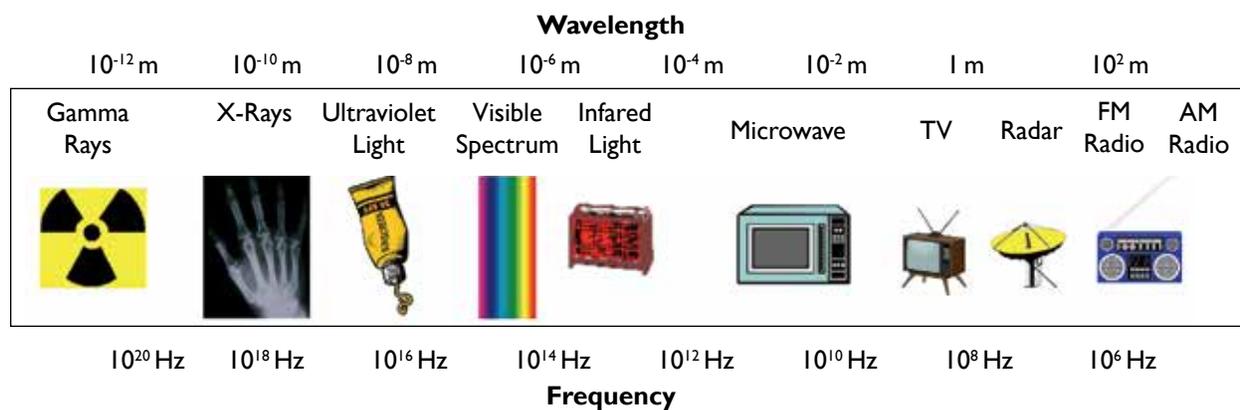
When wavelength is small, frequency is large.

This inverse relationship between frequency and wavelength should make sense if you think about it. Imagine yourself once again standing in the ocean. As you look at the waves coming toward you, if their crests are far apart (they have a large wavelength), they will not hit you very often (they will have a small frequency). However, if the wave crests are close together (small wavelength), they will hit you very often (large frequency).

You may not realize that the light we see with our eyes (the visible spectrum) is only a small part of the light that comes to us from the sun. The sun bathes our planet with light of many, many different wavelengths and frequencies. Our eyes perceive only a small

fraction of the total amount of this light. Figure 3.11 is a more complete representation of all light that comes from the sun, which we call the **electromagnetic spectrum**.

FIGURE 3.11
The Electromagnetic Spectrum
Illustrations from MasterClips collection



If you examine figure 3.11, you will first realize just how small a part of the electromagnetic spectrum we actually see with our eyes. The rainbow under the words *visible spectrum* represents that part of the electromagnetic spectrum. Everything else shown in figure 3.11, from wavelengths of 10^{-14} m all the way to wavelengths of 10^4 m, is still light. Our eyes are sensitive only to the light contained in the visible spectrum, which means that is the only light we can see.

However, as you can tell by the figure, we find uses for much of the light we cannot see.

For example, the signals that are sent out from a radio station to our radio antennae are, in fact, waves of light. We cannot see these light waves, but they are there nevertheless. Since they are there, we can use them to carry radio signals (FM and AM radio signals, for example). The only difference between these light waves and the ones we see with our eyes is the wavelength. Wavelengths of light used to carry radio signals are longer than those of visible light. In addition, television antennae also use light with wavelengths longer than those of visible light. What may be even more surprising to you is that microwave ovens use light to cook food. Microwaves also have longer wavelengths than those of visible light.

Notice that while we often use several of the light waves with wavelengths *longer* than visible light, we do not frequently use any of the light waves with wavelengths *shorter* than visible light. There is a very good reason for this. Light, of course, has energy. Its energy is kinetic since the light is moving. The energy of a light wave is directly proportional to its frequency. This is an important enough fact that you need to memorize it.

think about this

“And God said, ‘This is the sign of the covenant that I make between me and you and every living creature that is with you, for all future generations: I have set my bow in the cloud, and it shall be a sign of the covenant between me and the earth. When I bring clouds over the earth and the bow is seen in the clouds, I will remember my covenant that is between me and you and every living creature of all flesh. And the waters shall never again become a flood to destroy all flesh.’” Genesis 9:12-15

As a light wave's frequency increases, its energy increases.

As its frequency decreases, its energy decreases.

Since frequency and wavelength are inversely related to each other, we could also say this:

As a light wave's wavelength increases, its energy decreases.

As its wavelength decreases, its energy increases.

Based on these 2 facts, then, light waves with wavelengths shorter than those of visible light have higher frequencies and therefore higher energy. This is a very important consideration because when light strikes something (your eye, your skin, etc.), it deposits its energy into what it strikes. When visible light strikes your eye, it deposits its energy there. As you'll see in the next experiment, your eye uses that energy to transmit signals to the brain, and that's what causes you to see. Light with wavelengths shorter than visible light has enough energy to kill living tissue. When ultraviolet light (the first thing in figure 3.11 with wavelengths shorter than those of visible light) strikes your skin, it can kill some of your cells. If your skin is exposed to too much ultraviolet light, a large number of cells will die, and you will get a burn on your skin, which we call a sunburn. You can avoid a **sunburn** by covering your skin with sunscreen, a lotion containing chemicals that absorb the ultraviolet light before it hits your skin.

Because of the nature of ultraviolet light, God has built into this wonderful planet a very efficient system that filters out the majority of these rays. You may have heard some talk about the ozone layer. We will go into this in detail in the next module, but for now we'll say the ozone layer is part of the filtering system that God has designed to shield us from this destructive type of light. Although it is incredibly efficient, the ozone layer does allow some ultraviolet light through, and that is why you can get sunburned if you stay out in the sun too long without protection.

Gamma rays and X-rays are more energetic than ultraviolet light, so they are even more dangerous to living tissue. Nevertheless, we still use both of these forms of light occasionally in medical procedures. X-rays, of course, are used to examine bones and other aspects of our internal anatomy for diagnostic purposes. Although these X-rays will kill some living tissue, that risk is worth the benefit of being able to diagnose internal problems without surgery. Therefore, as long as you do not get X-rays frequently, your risk associated with the nature of the light is low, and the benefit you get from the diagnosis of internal problems is high. Gamma rays are also used in some medical applications. For example, certain types of cancer can be treated by exposing the cancerous site to gamma rays. Even though this leads to the death of healthy tissue, it can also lead to the death of cancerous tissue, so the risk of gamma ray exposure can be worth the medical benefit.

Before we leave this section, let's talk about one point of terminology. Notice that figure 3.11 is titled "The Electromagnetic Spectrum." That's because we know that light is an electromagnetic phenomenon. It is produced by the interaction of electrically

charged particles. Light is often called **electromagnetic** (ee lek' troh mag net' ik) **radiation**. Although that sounds like a scary term, it just means light, but it includes all wavelengths of light, not just the visible wavelengths.

THE RELATIONSHIP BETWEEN FREQUENCY AND ENERGY

You need to know one more equation that deals with light. This equation relates the frequency of a light wave to its energy. As already mentioned, as a light wave's frequency increases, its energy does as well. The mathematical equation that governs this relationship is:

$$E = h \times f \quad \text{Equation 3.2}$$

In this equation, E is the energy of the light wave and f is the frequency. The symbol h refers to another physical constant, called Planck's (Plahnks) constant. Named after its discoverer, German physicist Max Planck, this constant allows us to relate energy and frequency. Its value is $6.63 \times 10^{-34} \frac{\text{J}}{\text{Hz}}$. You needn't memorize the number; it will be given in any problem for which it is needed. Instead of learning the value, look at the units. Notice that if we put those units into equation (3.2), then we will get the following:

$$E = \frac{\text{J}}{\text{Hz}} \times \text{Hz}$$

When this happens, the Hertz unit cancels, leaving Joules, an energy unit.

Remember, Hertz means 1/second. With that in mind, let's look at the unit for Planck's constant once again. If we divide the unit Joule by 1/second, we get this:

$$\frac{\text{J}}{\text{Hz}} = \frac{\text{J}}{\frac{1}{\text{s}}} = \text{J} \div \frac{1}{\text{s}} = \text{J} \times \text{s}$$

“Joule-second” is another way of expressing the unit for Planck's constant.

Now that you understand the units for Planck's constant, see how we use it in example 3.3.

EXAMPLE 3.3

A light wave has a frequency of 2.3×10^{16} Hz. What is its energy?

$$h = 6.63 \times 10^{-34} \frac{\text{J}}{\text{Hz}}; c = 3.0 \times 10^8 \frac{\text{m}}{\text{s}}$$

This is easy. We simply plug our numbers into equation 3.2:

$$E = h \times f$$

$$E = 6.63 \times 10^{-34} \frac{\text{J}}{\text{Hz}} \times 2.3 \times 10^{16} \text{ Hz}$$

$$E = 1.5 \times 10^{-17} \text{ J}$$

We can keep only 2 significant figures because our frequency had only 2. This might seem like a small amount of energy, but it corresponds to the energy of ultraviolet light, so it is enough to kill living tissue!

A light wave has a wavelength of 1.21×10^{-12} m. What is its energy?

This problem is a little harder because the only equation we can use to calculate the energy of light is equation 3.2, and it uses *frequency*, not *wavelength*. To solve this, then, we must first turn the wavelength we've been given into frequency. We can do this with equation 3.1:

$$f = \frac{c}{\lambda}$$

$$f = \frac{3.0 \times 10^8 \frac{\text{m}}{\text{s}}}{1.21 \times 10^{-12} \text{ m}} = 2.5 \times 10^{20} \frac{1}{\text{s}} = 2.5 \times 10^{20} \text{ Hz}$$

Now that we have the frequency, we can use equation 3.2:

$$E = h \times f$$

$$E = 6.63 \times 10^{-34} \frac{\text{J}}{\text{Hz}} \times 2.5 \times 10^{20} \text{ Hz}$$

$$E = 1.7 \times 10^{-13} \text{ J}$$

The energy is 1.7×10^{-13} J.

ON YOUR OWN

($h = 6.63 \times 10^{-34} \frac{\text{J}}{\text{Hz}}$; $c = 3.0 \times 10^8 \frac{\text{m}}{\text{s}}$;
nano- means 10^{-9})

- 3.6 If a light wave has energy of 3.4×10^{-14} J, what is its frequency?
- 3.7 If a visible light wave has energy of 3.3×10^{-9} J, what color is it?
 (You can use figure 3.10 if you need to.)

Make sure you understand how to use equation 3.2 by answering “On Your Own” questions 3.6–3.7.

HOW THE EYE DETECTS COLOR

Before we go on to discussing the Bohr model of the atom, perform experiment 3.2. Hopefully, it will give you some insight into how humans perceive color.

EXPERIMENT 3.2

PURPOSE: To determine how our eyes detect color.

MATERIALS:

- 2 plain white sheets of paper (There cannot be lines on them.)
- A bright red marker (A crayon will also work, but a marker is better.)

QUESTION: How does the eye detect color?

HYPOTHESIS: Write down your hypothesis about what will happen to our eyes if we stare at one color for 60 seconds.

PROCEDURE:

1. Take one of the sheets of paper and make a thick cross on it with the red marker. The cross should be about 6 inches long, and the 2 legs that make it up should be about $\frac{3}{4}$ inch thick.
2. Color the entire cross so that you have a large, solid bright red cross in the middle of a white sheet of paper.
3. Put the clean sheet of white paper underneath the sheet with the cross on it. Make sure the cross faces you so that you can see it.
4. Stare at the cross for a full 60 seconds. You can blink if you need to, but do not take your eyes off the cross.
5. After a full 60 seconds of staring at the cross, quickly pull the top sheet of paper out of the way so that you can see only the clean sheet of paper on the bottom. What happens? For about 90% of students, a blue-green cross will appear for a few seconds on the blank sheet of paper. After a few moments, it should vanish. This optical illusion will not work for some people, especially if they have a tendency toward color blindness.
6. Clean up and return everything to the proper place.

CONCLUSION: Write something about what happened in the experiment.

What happened in the experiment? To see light, your eyes are equipped with cells called **rods** and **cones**. These cells transmit electrical signals to the brain whenever they are hit by certain energies of light. The brain receives the electrical transmissions and uses them to form an image in your mind. The rods are sensitive mostly to low levels of light and are not very sensitive to color. However, the cones respond to certain specific energies of light. Some cones are sensitive only to low-energy visible light (red), while others are sensitive to medium-energy visible light (green), and still others are sensitive to the highest-energy visible light (blue). When colored light hits these cells, they will send signals to the brain only if the light that they are sensitive to is hitting them. If blue light hits your eyes, the cones that respond to blue light will send signals to your brain, but the cones that respond to green and red light will not. As a result, your brain interprets the

signals coming from your eyes and puts a blue image in your head. Similarly, if red light hits your eyes, the cones that respond to red light will send signals to your brain, but the cones that respond to green and blue will not. As a result, your brain constructs a red image.

If a mixture of colors hits your eyes, the cones send signals to your brain in proportion to the amount of light to which they are sensitive. For example, suppose you are looking at a purple ball. When that purple hits your eyes, the cones that respond to red light send some signals to your brain, and the cones that respond to blue light send some signals to your brain. Your brain then constructs a purple image of the ball in response. If the purple color of the ball is on the blue side of purple, then the cones that respond to blue light send more signals to the brain than the cones that respond to red light. On the other hand, if the purple is more red than blue, the cones that respond to red light send more signals to the brain than those that respond to blue light. All of the colors that you see are the result of the brain receiving signals from 3 types of cone cells and adding those 3 primary colors (red, green, and blue) with the weight indicated by the amount of signal coming from each type of cell.

What happened in the experiment? While you were looking at the red cross, your cones that respond to low-energy (red) light were sending signals to the brain, but your other cones weren't doing anything. Cones get tired pretty quickly, and when they have sent the same signal to the brain for a period of several seconds, they eventually just shut off. The brain, sensing that no more signals are coming from the cones, assumes that they have shut off simply because they are tired, and it holds the same image in your mind until new signals come along. As you were staring at the cross, your low-energy cones eventually shut off. Since no more signals were coming from the low-energy cones and since no signals had ever come from the medium and high-energy cones, the brain was receiving no more signals. It therefore assumed you were still looking at the cross and continued to hold the image in your mind.

When you yanked the top sheet away, white light began to hit your eyes where only red light had hit them before. Since white light contains all energies, your medium- and high-energy cones began to receive light and transmit signals to your brain. Your low-energy cones were still shut off, so they didn't send any signals, even though they should have. The brain started receiving new signals, but only from the high- and medium-energy cones. It constructed an image of green (medium-energy) and blue (high-energy) light. Eventually, your low-energy cones realized that they had to start transmitting again, and once they did, the brain realized that the eyes were seeing all energies of light and so formed a white image in your mind.

The way we perceive color, then, is based on the energy of the light that hits our eyes. Isn't it marvelous how well designed the eye is to handle such a complex operation? That should tell you something about how marvelous its Designer is!

THE BOHR MODEL OF THE ATOM

During our discussion of Rutherford's model of the atom, we mentioned that it was inconsistent with electromagnetic theory. Nevertheless, since that model seemed to be the only way to explain Rutherford's experimental data, it became the accepted model of the day. As chemists sought to understand more about atomic structure, they did more experiments.

One of the experiments that shed some light on atomic structure involved adding energy to atoms by heating them or passing electricity through them. When this was done, the atoms emitted light, some of which was in the visible spectrum. The interesting thing was that each type of atom seemed to emit its own unique color or colors of visible light. For example, when mercury atoms were heated, a pale blue glow would occur. On the other hand, when electricity was passed through neon atoms, a bright pink light was emitted.

As chemists investigated these experiments further, the mystery deepened. They learned that the colors which were being emitted during these experiments were the result of a handful of individual wavelengths of visible light being emitted by the atom. For example, when electricity is passed through hydrogen gas, the atoms emit a purple light. This light is composed of exactly 4 wavelengths: 410 nm (violet), 434 nm (indigo), 486 nm (blue), and 656 nm (red–orange). These wavelengths mix together to form a purple color.

What makes these results interesting is the fact that atoms seem to emit *individual* wavelengths of light. This is very unusual. Most colors that we see are the result of a *range* of wavelengths, not *individual* ones. For example, if we were to make a red light, we would probably paint a normal light bulb red. The paint would keep all of the wavelengths except the red ones from getting through. The light coming from this painted bulb would not be made up of just one individual wavelength. Instead, all wavelengths from 700 to 650 nm would be there. The fact that atoms emit only a handful of individual wavelengths indicates that something unique is at work.

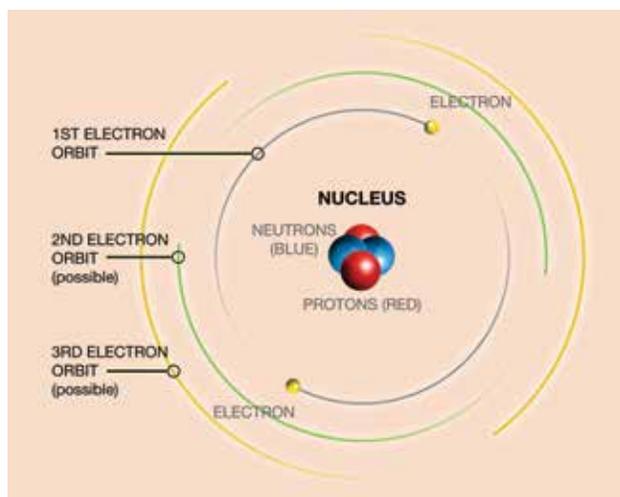
Another interesting fact is that each element produces its own *unique* set of wavelengths when heated or electrified. While hydrogen emits visible light with wavelengths of 410 nm, 434 nm, 486 nm, and 656 nm, tantalum (Ta) emits visible light of only a single wavelength: 535 nm. This little fact is quite useful. After all, if you need to determine the atoms in a substance, one way to do it is to heat the substance up and determine the wavelengths of light that are emitted. Since each element in the substance will emit its own unique wavelengths of light, you should be able to match the elements with their respective wavelengths and eventually determine the elemental composition of the substance.

This is the main way that astronomers determine the elemental makeup of distant stars. Stars are bright because the substances that make them up are very hot and they therefore emit light. They emit so many different wavelengths of light that, to our eyes, the light appears white. However, a scientific instrument known as a **spectrometer** (spek trahm' uht er) can analyze the light and determine all of the individual wavelengths that make it up. Then, matching the elements to their expected individual wavelengths can determine the elemental composition of the star. This method of analysis, known as **spectroscopy** (spek trahs' kuh pee), can be applied to chemicals as easily as it is applied to stars. It is one of the most useful tools a chemist has to analyze unknown substances.

The ability of atoms to emit individual wavelengths of light also helped Niels Bohr develop his own model of the atom. Bohr took the planetary model developed by Rutherford and added his own little twist. Rutherford assumed that all of the electrons orbited around the protons in a single orbit. Bohr suggested, instead, that there were several possible orbits that the electrons could be in. Furthermore, he suggested, the electrons could jump from one orbit to another. In addition, by the time Bohr had come up with his model of the atom, the neutron had been discovered by Chadwick. Bohr put

the neutrons and protons together in the center of the atom and called it the **nucleus**. This model is shown in figure 3.12.

FIGURE 3.12
The Bohr Model of the Atom
 Illustration by David Weiss



One very important advance that Bohr made in developing his model was his assumption that the electrons could only jump into and out of specific orbits. So, if an electron decided to jump out of orbit 1, it could not go just anywhere. It had to go into another allowed orbit, such as orbit 2. And even though it could be in orbit 1 and jump to orbit 2, *it could not be anywhere in between*. This type of assumption is called a **quantum** (kwahnt' uhm) assumption, and at the time it was considered rather odd. After all, if the electron could go to either orbit, why couldn't it go somewhere in between? Bohr had no answer to this question (indeed, even modern chemists have only a rough answer), but it was a necessary assumption in his model.

One of the first questions you might ask is this: How do the electrons jump from one orbit to another? The orbits that are farther away from the nucleus require more energy. If an electron wants to jump from an orbit close to the nucleus (orbit 1) to an orbit far away from the nucleus (orbit 3), it can do so by absorbing some energy. This is what happens when you heat a substance or pass electricity through it. When an electron moves from an orbit close to the nucleus to an orbit far from the nucleus, we say that the electron has been **excited**.

If, on the other hand, an electron would like to move from an orbit far away from the nucleus to an orbit close to the nucleus, the electron must somehow get rid of its extra energy. In other words, while in the orbit far from the nucleus, the electron has more energy than it would in an orbit close to the nucleus. To get to that closer orbit, the electron must find a way to get rid of that excess energy. It does this by emitting light. When an electron moves from an orbit far from the nucleus to an orbit close to the nucleus, we say it has **de-excited**.

When you heat a substance and its atoms emit light, you are witnessing a 2-step process. Electrons in the atoms first gain energy (because of the heat) and move into

orbits that are far away from the nucleus. Then, after they sit in the far orbits for a short time, they decide they would rather be back in their original orbit, so they emit light to get rid of the excess energy and go back to their original orbit.

This is what makes the flames of a fire. When wood is burned, the chemicals in the wood are first changed into carbon dioxide gas and water vapor via a combustion reaction, as we have discussed before. However, heat is also produced in the reaction. This heat excites the electrons in the gases surrounding the fire, and they move to orbits farther away from their nucleus. However, after being there for a while, they decide to de-excite back to their original orbit. To do this, they must emit light. This is where the pleasant yellow–orange glow of the flames comes from.

If you light a natural gas oven, you will notice that the flames are blue, not orange. This is because the natural gas company intentionally contaminates its natural gas with a substance whose atoms emit a blue glow when excited. The difference in the flame color is due to the fact that different atoms are being excited. You can examine another similar phenomenon if you burn some wrapping paper. The dyes in most wrapping papers contain elements that are different from other papers. So while most papers burn with the same yellow–orange glow that wood burns with, wrapping paper typically burns with green, blue, or red flames because of the elements that make up the dyes in it.

Do you remember that Bohr assumed that an atom's electrons could only be in certain orbits? Since each orbit has a specific energy associated with it, this means that an electron can have only certain energies. If it is in orbit 1, it has a certain energy (E_1). If it is in orbit 2, it has another, higher energy (E_2). But since the electron can never be between orbit 1 and orbit 2, it can also never have any of the energies between E_1 and E_2 . This assumption makes no sense at all because common sense tells us that any object should be able to attain any energy it wants, provided there is some way of transferring the necessary energy. Bohr said there were only certain energies that the electrons in the atom could have. This, another quantum assumption in Bohr's model, is the one we now see as critical to understanding atomic structure.

Bohr did not just dream up his model for the atom. In fact, for the hydrogen atom, he had detailed mathematical descriptions of the electron moving in its various orbits. He could use his mathematics to determine the energy of each electron orbit. He could determine how much energy the electron needed to jump from one orbit to another. Once he knew the energy necessary for the electron to jump from one orbit to the next, he could then determine the wavelength of light that the electrons must emit to de-excite. He found that the wavelengths he calculated were *exactly* equal to the wavelengths that chemists observed for de-exciting hydrogen atoms! This indicated that Bohr was on the right track, and his model for the atom quickly gained acceptance.

In fact, even though today's accepted model of the atom is slightly different from Bohr's model, the detailed mathematics that he developed are still taught in college chemistry today. These mathematics, which are a bit too detailed to tackle in this course, provide good insight into the physical mechanisms that hold the atom together. If you end up taking advanced chemistry, you will learn the mathematics of the Bohr model of the atom.

Additionally, the stylized pictures of atoms you often see are based on the Bohr model of the atom. Although science has moved past Bohr's model, the scientific community recognizes his achievements to this day.

THE QUANTUM MECHANICAL MODEL OF THE ATOM

Today's model of the atom, called the **quantum mechanical model**, is a bit more complicated than Bohr's model, but it still retains many of Bohr's concepts. For example, chemists still believe that electrons orbit around the nucleus and that there are many, many different orbits that electrons can continually jump between. In addition, we still say that the electrons need energy to go into orbits that are far away from the nucleus and that they need to release energy in the form of light to get back to orbits close to the nucleus. Finally, Bohr's quantum assumption about energy is still followed. We assume that there are only certain energies that an electron in an atom can have. An electron might have an energy of E_1 in one orbit and an energy of E_2 in the next higher orbit, but the electron can have no energy between E_1 and E_2 .

Where the quantum mechanical model mainly differs from the Bohr model is the type of orbits that the electrons can occupy. In the Bohr model, electrons orbit the nucleus in circular orbits. In the quantum mechanical model, this is not the case. The way we theorize that electrons orbit the nucleus today is a little more complex. We assume that electrons do not orbit in fixed circles but instead orbit in clouds we call **orbitals** (or' buh tuls). When orbiting in a fixed circle, the electron is always the same distance from the nucleus. In today's model of the atom, electrons can, at different times, be at different distances from the nucleus but still be in the same orbital. However, if you were to watch the electron for quite a while, you would see that it stayed within a certain boundary.

If you are having trouble understanding this, think about it this way. Suppose you had a dog in your backyard and you drove a stake in the center of the yard and tied a rope to it. When the dog was tied to the rope, he could be anywhere in the backyard within a circle determined by the length of the rope. Although the dog could be anywhere inside the circle, he could not go outside the circle. This is like an electron in an orbital. The orbital has a general shape, and the electron can be anywhere within that shape. On the other hand, suppose the rope got wet during the winter and froze while it was stretched out. Then, because the rope was rigid and would always stay stretched out, the dog could not move anywhere except on the edge of the circle. This is like an electron orbiting in the Bohr model. It cannot be anywhere except on the edge of the circular orbit. That's the difference between electron orbits (the Bohr model) and electron orbitals (the quantum mechanical model).

The other difference between the Bohr model and the quantum mechanical model is the shape of the orbitals. In the Bohr model, all electrons had to orbit in circles. In the quantum mechanical model, there are several differently shaped orbitals that electrons can use to orbit around the nucleus. The shapes of these orbitals can become very complex, so we will limit our discussion to the first 3 types. The majority of atoms in nature use only these orbitals, so that will be enough for an introduction into this rather esoteric concept.

The simplest type of orbital an electron can occupy is called the **s orbital**. As illustrated in figure 3.13, this orbital is spherical, with the nucleus at the center. An electron that occupies an s orbital can be anywhere inside the sphere but cannot venture outside of the sphere.

FIGURE 3.13**s Orbitals**

Illustration by Megan Whitaker



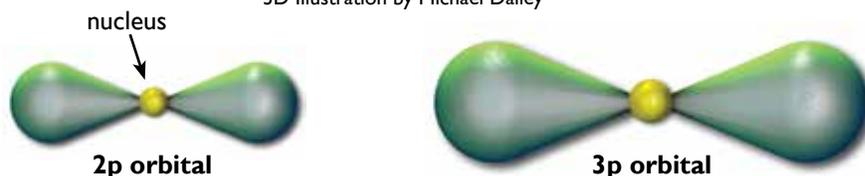
Notice in figure 3.13 that there are 3 s orbitals, each of differing size. This is because in the quantum mechanical atom, just like in the Bohr atom, electrons can orbit far away from the nucleus or close to the nucleus depending on their energy. In the Bohr atom, the orbital number (1, 2, 3, etc.) determined how far they were from the nucleus. In the same way, the quantum mechanical model of the atom allows electrons to orbit far from the nucleus or close to it. If an electron orbits far away from the nucleus in a spherical pattern, it will occupy a large s orbital. If it orbits close to the nucleus in a spherical pattern, it will occupy a small s orbital. We distinguish small s orbitals from large ones with the number that appears next to the orbital letter. A 1s orbital is always smaller than a 2s orbital, which is always smaller than a 3s orbital.

Just like in the Bohr model, the farther away from the nucleus the electron is, the more energy it must have. The electrons in a 1s orbital have less energy than the electrons in a 2s orbital. Since the size of the orbital determines the energy of the electrons, the number that appears next to the orbital letter is often called the energy level of the atom.

The next orbital shape we will consider is the **p orbital**. This orbital is dumbbell-shaped, as illustrated in figure 3.14. The nucleus is in the center of the dumbbell. Just as there are different sizes of s orbitals that correspond to different energies of electrons, there are also different sizes of p orbitals that correspond to different electron energies. One interesting thing about p orbitals is that there are none of them on the first energy level of the atom. This means there is no such thing as a 1p orbital. The first p orbital is the 2p orbital.

FIGURE 3.14**p Orbitals**

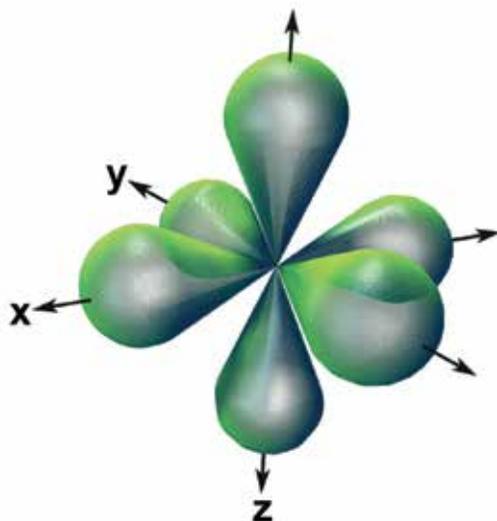
3D Illustration by Michael Dailey



Another interesting thing about p orbitals is that for every energy level (or size of the orbital), there are 3 different p orbitals. They are all shaped the same, but they are oriented differently in space. One p orbital is oriented horizontally (as drawn in figure 3.14), while the second one is oriented vertically. The third is oriented 90 degrees from both of the others.

In other words, each of the 3 p orbitals is oriented along a different axis in 3-dimensional space. Figure 3.15 illustrates how the 3 p orbitals all exist together in an atom.

FIGURE 3.15
Three Different p Orbitals in the Same Energy Level of an Atom
3D Illustration by Michael Dailey



So, as the figure shows, there can be 3 different p orbitals for each energy level. There are three 2p orbitals, three 3p orbitals that are all bigger than the 2p orbitals, three 4p orbitals that are all bigger than the 3p orbitals, and so on. All of this is a little obscure and may be hard to picture in your mind. If that is the case, don't be too concerned. At the end of this discussion, we will summarize what you need to know about the quantum mechanical model.

Finally, the last type of orbital we will use in this course is the d orbital. The shapes of the d orbitals are very complex, so we will not attempt to illustrate them. What you need to know about d orbitals is that there are 5 different d orbitals for each energy level. Also, the lowest-energy (smallest) d orbitals are located on the third energy level.

One important thing to note about the different shaped orbitals is that within the same atomic energy level, the different shaped orbitals require different amounts of energy. It is rather easy for an electron to whirl around the nucleus in a sphere, so s orbitals are the lowest-energy of all orbitals. On the other hand, a dumbbell is a little more complex, so p orbitals require more energetic electrons than s orbitals. Finally, since d orbitals have even more complex shapes, they require more energy than either s or p orbitals.

One more important thing to know about orbitals is that each individual orbital can hold only 2 electrons. After 2 electrons, the orbital is full, and any new electrons must find another orbital to fill. Since there are 3 different p orbitals on an energy level, 6 electrons could fit into all 3 p orbitals on that energy level (2 per orbital). In the same way, since there are 5 different d orbitals on an energy level, up to 10 electrons can go into the d orbitals of a given energy level.

Think a bit about all of these orbitals and how electrons use them. If an electron is orbiting as close as it can to the nucleus, we say that it is in the first energy level. On the first energy level, only 1 type of orbital exists: the s orbital. If an electron moves as

close as it possibly can to the nucleus, then it must do so in a spherically shaped way (1s orbital). However, if the electron goes up to the next energy level (2), it is a little farther away from the nucleus and has the option of moving around the nucleus in a spherical way (2s orbital) or in 1 of any 3 dumbbell-shaped orbitals (2p orbitals). If the electron goes away from the nucleus even farther into the third atomic energy level, then it has the choice of being spherically shaped (3s orbital), 1 of 3 dumbbell-shaped orbitals (3p orbitals), or 1 of 5 more complex orbitals (3d orbitals).

This is the modern-day view of the atom. The atom has a nucleus of tightly packed protons and neutrons. Its electrons whirl around that nucleus in several different energy levels (1, 2, 3 . . .). Depending on the energy level, the electrons have several different shapes of orbitals (s, p, d, and other orbitals) in which they can travel around the nucleus. Each specific energy level (1, 2, 3 . . .) has its own energy requirements for the electrons. In addition, each orbital shape has additional energy requirements for the electrons. Based on the energy an electron has, it will go into the energy level and orbital whose energy requirements it best meets. Each orbital has a maximum capacity of 2 electrons. This is the summary of things that you must understand about the quantum mechanical model of the atom.

BUILDING ATOMS IN THE QUANTUM MECHANICAL MODEL (ELECTRON CONFIGURATIONS)

If all of the previous discussion confused you a bit, don't worry too much. In this section, we will try to apply what we have learned about the quantum mechanical model of the atom to describe what an atom might look like. After you have seen this application, you might gain a better understanding of what you read in the last section.

Before we start describing what atoms look like, there is one more fact of nature that you must learn:

All forms of matter try to stay in their lowest possible energy state.

This is a fact of nature that we will use over and over again. Think about it this way: Matter is basically lazy. It doesn't want to work any harder than it needs to. If at all possible, matter will always try to stay in its lowest-energy configuration.

For example, remember when we were discussing the light that was emitted from excited atoms? We said that when an atom is heated, its electrons absorb some of that kinetic energy and use it to travel to an orbital farther away from the nucleus. The electrons then emit light to go back down to their original orbital. Why do they do that? Once they jump to the higher-energy orbital, why don't they just stay there? They don't stay there because once they are in a higher-energy orbital, they are not in their lowest possible energy state. Therefore, the electrons release energy in the form of light to get back to their lowest energy state.

Since the lowest possible energy state is so important to matter, we give it a special name. We call it the **ground state** of the substance.

Ground state—The lowest possible energy state for a given substance.

Be sure you understand that the ground state is different for every substance on earth. For example, the ground state of a hydrogen atom is different from the ground state of a helium atom, as we will soon see.

Now that we know this very important fact about nature, we are finally ready to start building up atoms. First, let's describe the simplest atom in nature: the hydrogen-1 (^1H) atom. By looking at the periodic table, we can tell how many protons, neutrons, and electrons are in this atom. On the table, hydrogen's atomic number is 1. This means there are 1 proton and 1 electron in a hydrogen atom. Also, since the mass number of this hydrogen atom is 1, then there are $1 - 1 = 0$ neutrons in the atom. We already know that a hydrogen atom has 1 proton and no neutrons in its nucleus. What about the electron? Where is it?

Since all matter in nature tries to get to its lowest energy state, the electron will go into the lowest energy orbital available. Energy level 1 is the lowest energy level, and energy level 1 has only an s orbital in it, so the electron goes into the 1s orbital. This means that the electron is orbiting relatively close to the nucleus and the orbit has the shape of a sphere. Therefore, the overall picture of a hydrogen atom in its ground state is 1 electron orbiting close to 1 proton, and the electron's orbital is spherical.

Now let's move to another simple atom, the helium-4 (^4He) atom. After looking at the table, we see that it has 2 protons and 2 electrons. In addition, since its mass number is 4, it has $4 - 2 = 2$ neutrons. The 2 electrons will try to go into the lowest energy orbits available. Since energy level 1 is the lowest energy level and since it has only an s orbital in it, then both electrons will go into the 1s orbital. Our picture for the helium-4 atom in its ground state is 2 electrons orbiting close to a nucleus tightly packed with 2 protons and 2 neutrons. The electrons whirl about this nucleus in a spherical orbit.

Continuing, we can consider a lithium-6 (^6Li) atom. After looking at the table and dealing with the mass number, we see that it has 3 protons, 3 electrons, and 3 neutrons. The 3 electrons will try to go into the lowest energy orbits available. Since energy level 1 is the lowest energy level and it has only an s orbital, 2 electrons will go into the 1s orbital. However, any single orbital can hold only 2 electrons. The third electron in the atom cannot fit into the 1s orbital. It must, therefore, find the next lowest energy orbital to go into. Since there is no more room on the first energy level, it goes into energy level 2. At this point, it can either go into a 2s orbital or a 2p orbital. Since s orbitals are lower in energy than p orbitals, it will go into the 2s orbital. Our picture for the lithium-6 atom is 2 electrons orbiting close to and 1 electron orbiting farther away from the nucleus tightly packed with 3 protons and 3 neutrons. All 3 electrons whirl about this nucleus in spherical orbitals.

Do you see how the thinking goes here? Let's try a much more complicated atom, and we will then introduce an abbreviated notation that will make things a little easier to describe. Consider a ^{22}Ne atom. According to our rules, it has 10 protons, 10 electrons, and 12 neutrons. Where do the electrons go?

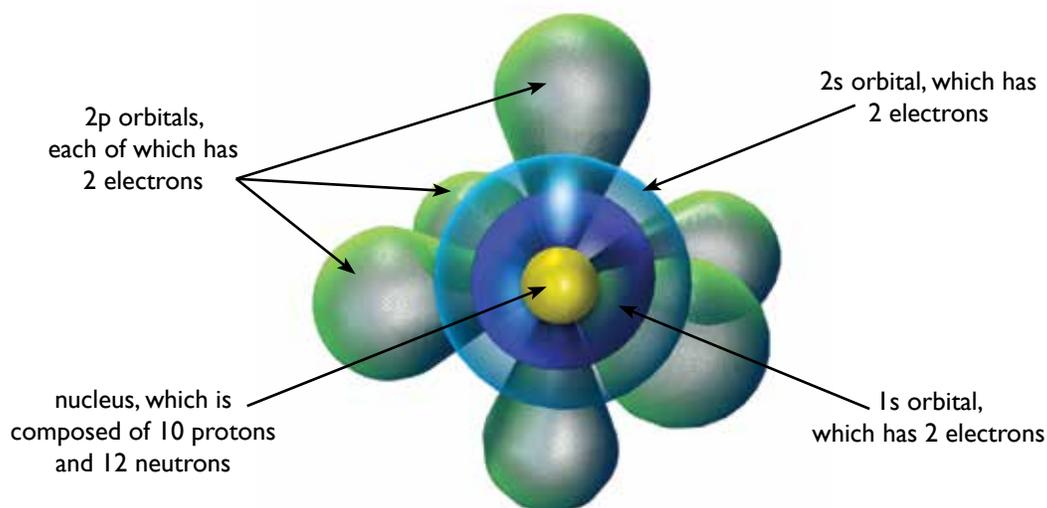
The electrons will look for the lowest-energy orbitals available. The first 2 will occupy the 1s orbital since energy level 1 is the lowest energy level available. Since only 1 s orbital is available on energy level 1, the rest of the electrons will have to go into higher energy levels. On energy level 2, both s and p orbitals are available. Since the s orbitals are lower in energy, the next 2 electrons will fill that orbital. So far, we have placed 4

electrons in their orbitals, and we have 6 to go. On energy level 2, there are 3 p orbitals that can also be filled. Each p orbital can hold 2 electrons, so the 6 remaining electrons can go into the 2p orbitals.

Our picture of the ^{22}Ne atom is a tightly-packed nucleus of 10 protons and 12 neutrons with 10 electrons whirling around it. Two of the electrons are orbiting the nucleus closely in a spherical pattern. Two more electrons are whirling around the nucleus a bit farther away, also in a spherical pattern. However, 6 more electrons are whirling about in dumbbell-shaped clouds. This is the way a modern-day chemist would describe a ^{22}Ne atom. Figure 3.16 illustrates what this would look like.

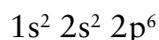
FIGURE 3.16
The Quantum Mechanical View of a ^{22}Ne Atom

3D Illustration by Michael Dailey



Hopefully you now have a reasonably good grasp of the quantum mechanical model of the atom. If you understand how we came up with the description of the ^{22}Ne atom, then you are all set. If you don't quite understand it, then you need to reread the material or seek out another source because you can't go further in this course without understanding this important concept.

Since the quantum mechanical model of the atom is complicated, we will simplify it a bit with some notation. We will learn in the next module that all of the chemical behavior of an atom is governed solely by the number of electrons it has. As a result, the nucleus of an atom is relatively unimportant to a chemist. When giving the quantum mechanical description of an atom, chemists do not bother to mention the nucleus. Also, rather than saying in words where all of the electrons go, we develop an abbreviation to help us. The abbreviation for the neon atom we just discussed is this:



This says all a chemist needs to know about the structure of a neon atom. The numbers to the left of the letters represent the energy level, the letters represent the shape of the orbital, and the superscript numbers tell how many electrons are in the orbital. The abbreviation above tells that there are 2 electrons in the 1s orbital, 2

Looking at figure 3.17, you should first see that for the purposes of this discussion, we pretend that helium sits right next to hydrogen on the table. The next thing you should see in the figure is that the columns have been separated. We call the first 2 columns the s orbital block because the last electrons in the atoms of that block end up in an s orbital. Since each s orbital can have only 2 electrons in it, the s orbital block can be only 2 columns wide. The next group represents those atoms whose last electrons end up in d orbitals. Since there are 5 d orbitals and each can have 2 electrons, there are 10 elements in each row that have their last electrons in d orbitals. Finally, the last block is the p orbital block. These elements have their last electrons in a p orbital. Since the 3 p orbitals can each hold 2 electrons, the p orbital block is 6 columns wide. The 2 lowest rows on the periodic table (the lanthanides and actinides) are not going to appear on your test. However, if you are interested, their last electrons go into f orbitals, which are even more complex than d orbitals.

Since the columns represent the electron orbitals in the atoms, the rows of the table represent the energy level that those electrons are in. For example, notice that Ne is in row 2. This means that the last electron in Ne goes in the second energy level. The fact that Ne is in a column that is a part of the p orbital block tells us that the last electron is in a p orbital. Neon's last electron goes in the 2p orbital. We figured this out earlier with energy arguments, but figuring it out this way is easier.

To use the arrangement of the periodic table to determine electron configurations, you just have to look at the periodic table and find the atom whose electron configuration you wish to determine. Then, starting with hydrogen, walk through the table, assigning electrons in each orbital one box at a time. When you get to the box that represents the atom you are interested in, you're done. The only trick you have to remember is that in the d orbital block, the row that the elements are on is one number higher than the energy level that the electrons are in. So, when filling up d orbitals, subtract 1 from the row number to get the energy level. If this all sounds a bit confusing, example 3.4 should clear things up a bit.

EXAMPLE 3.4

Write the electron configuration for a carbon atom.

First, we look at the periodic table and find out that carbon (C) is element 6. To get to this atom, we start walking through the table from the beginning. The first 2 elements are H and He. Their boxes are in row 1 and the s orbital group. The first 2 electrons go into the 1s orbital. Going on through the table, we get to Li and Be. They are in row 2 and the s orbital group, so we put 2 electrons in the 2s orbital. Continuing on, we hit B and then C, where we want to stop. This means that we went through 2 boxes in the row 2, p orbital block. So there are 2 electrons in the 2p set of orbitals. Since we've hit carbon, we are done. So the electron configuration is:



Remember, the superscripts in the electron configuration represent the number of electrons in each orbital. The sum of the superscripts must equal the number of electrons in the atom. Carbon has 6 electrons, and our superscripts add up to 6. Our answer, therefore, checks out.

Write the electron configuration for a Cl atom.

To get to Cl (17), we have to go through several boxes. We must walk through all of row 1, which has 2 boxes in the s orbital block. That gives us $1s^2$. We must also go through all of row 2, which has 2 boxes in the s orbital block and 6 boxes in the p orbital block. This gives us $2s^2 2p^6$. We also must go through the s orbital block in row 3. This gives us $3s^2$. Finally, to get to Cl, we go through 5 boxes in the row 3, p orbital block. This gives us $3p^5$. The electron configuration, then, is:



The superscripts add up to the number of electrons in a Cl atom (17), so our answer checks out.

Give the electron configuration of a Ag atom.

To get to Ag, we have to walk through several boxes on the table. We must walk through all of row 1, which has 2 boxes in the s orbital block. That gives us $1s^2$. We must also go through all of row 2, which has 2 boxes in the s orbital block and 6 boxes in the p orbital block. This gives us $2s^2 2p^6$. Row 3 has 2 boxes in the s orbital group and 6 in the p orbital box, so that gives us $3s^2 3p^6$. Row 4 has 2 boxes in the s orbital block, 10 in the d orbital block, and 6 in the p orbital block. Remembering that for d orbitals, we must subtract 1 from the row number, this gives us $4s^2 3d^{10} 4p^6$. In row 5, we must go through the 2 boxes in the s orbital block, so that gives us $5s^2$. In row 5, we also have to go through 9 boxes to get to Ag. Once again, we subtract 1 from the row number with d orbitals, so this gives us $4d^9$. In the end, then, our electron configuration is:



Our superscripts add up to 47, so our answer checks out.

This is all you do to get the electron configuration of an atom. With the periodic table as a guide, it is not a hard task at all. However, let's give you a couple of words of advice:

1. Some chemistry books teach a different method to get electron configurations. Their method requires memorization and is not the way real chemists do electron configurations. That's why we don't teach it their way. Don't get confused if you look at other books, however. Their method might be different, but the answers should be the same.
2. Remember that these electron configurations are not just a worthless exercise. They have very, very deep meaning to chemists. They not only tell the structure of a particular atom, but, as we will see in the next module, they also tell a great deal about the chemistry of the atom.

See if you can do electron configurations in “On Your Own” questions 3.8–3.10. Feel free to use the periodic table, but do not use figure 3.17. You should be able to do electron configurations from a normal table, not one specially enhanced for electron configurations.

ON YOUR OWN

- 3.8 Write the electron configuration for silicon (Si).
- 3.9 Write the electron configuration for selenium (Se).
- 3.10 What is the electron configuration of silver (Ag)?

ABBREVIATED ELECTRON CONFIGURATIONS

Before we leave electron configurations, let’s give you a handy abbreviation that will keep you from writing a lot when you give your electron configurations. The last column in the table, column 8A, is very important. We will see why it is so important in the next module. For now, just realize that most chemists gauge an atom by how far it is away from this column in the table. One way you can abbreviate electron configurations is to look for the nearest atom in 8A that has a *lower* atomic number than the atom you are interested in. Then, you simply add any extra electrons on top of that atom’s configuration. See how we do this in example 3.5.

EXAMPLE 3.5

Write the abbreviated electron configuration for Sr.

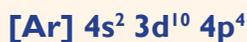
To abbreviate an electron configuration, we just find the nearest 8A element that has a lower atomic number than the atom we are interested in. In this case, Kr fits the bill. We can then say that the only difference between Sr and Kr is that there are 2 more boxes in the row 5, s orbital block. We can write the electron configuration as:



Putting Kr in square brackets tells a chemist that you are abbreviating the electron configuration in terms of the 8A element Kr.

Write the abbreviated electron configuration for Se.

The nearest 8A element that has a lower atomic number than Se is Ar. Don’t be fooled here. Kr is the closest 8A element, but it does not have a *lower* atomic number. The only differences between Se and Ar are that there are 2 boxes in the row 4, s orbital group, 10 boxes in the row 4, d orbital group, and 4 boxes in the row 4, p orbital group. The abbreviated electron configuration for Se is:



ON YOUR OWN

- 3.11 What is the abbreviated electron configuration for Br?
- 3.12 What is the abbreviated electron configuration for In?

Answer “On Your Own” questions 3.11–3.12 to cement the idea of abbreviated electron configurations in your mind.

As we said before, you will soon learn that the meaning of electron configurations is very deep, so your ability to predict them is a necessity.

THE AMAZING DESIGN OF ATOMS

Before we leave this module, there are a couple of interesting things to tell you about atoms. First, the atom is made up mostly of empty space. Since matter is made up entirely of atoms, this also means that matter is mostly empty space. This may sound crazy. After all, when you hit your head against the wall, the wall certainly doesn't feel like empty space. Nevertheless, the statement is true.

Nuclear chemistry experiments have demonstrated that the nucleus of a hydrogen atom has a radius of about 1×10^{-15} meters. On the other hand, the mathematics of the quantum mechanical model of the atom tell us that the radius of the 1s orbital in that same atom (the closest orbital to the nucleus) is about 5×10^{-11} meters. On the atomic scale, the electrons are *very far away* from the nucleus. If you don't understand the implications of these numbers, think about it this way: If the nucleus of the hydrogen atom were 1 inch in diameter, the electron would be orbiting it about 3/4 of a mile away! Between the nucleus and the electron there would be absolutely nothing. No air, no matter. Nothing. Just empty space. Therefore, the vast majority of an atom is simply empty space.

The most interesting aspect of an atom is not that it is made up of mostly empty space. The most interesting aspect of the atom is how well designed it is. You see, for all atoms except ^1H to exist, they must have neutrons. The neutron is just slightly more massive than the proton. It's a good thing, too. Nuclear chemistry experiments indicate that the neutron must be more massive than the proton, or it would not exist. Calculations indicate that if the mass difference between the neutron and the proton were changed by as little as 0.2%, all neutrons would spontaneously decay into protons, and the only atom in the universe would be ^1H . If that were the case, no life could exist in the universe! (See the Apologia book extras website mentioned in the Student Notes for more information.)

The mass of the electron is also quite crucial for the existence of life. Remember, the electrons whirl around the nucleus in the orbitals that we just discussed. The *size* of those orbitals depends on the strength of the electrical attraction between the protons and electrons, as well as the mass of the electron. If the electron's mass were larger, the size of the orbitals would be smaller. If the electron's mass were smaller, the size of the orbitals would be larger. Why does that matter?

If the electron's mass were larger, the electron orbitals would be smaller. As a result, atoms would hold on to their electrons more tightly, and it would be difficult to form ionic compounds. Many of the chemical processes that occur in living organisms depend on the existence of ions from ionic compounds. Without these ionic compounds, the chemical processes that depend on them would cease, and living organisms would die.

On the other hand, most of the compounds that make up living organisms are

covalent compounds. As mentioned earlier, atoms in covalent compounds share their electrons. If the electron mass were smaller, the electron orbitals would be larger and, as a result, atoms would hold onto their electrons rather loosely. This would make it very easy to form ionic compounds but very hard to form covalent compounds. If that were the case, living organisms could not even form, much less exist!

In the end, then, the electron mass is just right to produce electron orbitals that are neither too big nor too small. The balance between ionic and covalent molecules is perfect for the formation and existence of life, due to the mass of the electron. Estimates indicate that a variation of as little as 2% in the mass of the electron would make it impossible for life as we know it to exist! (Corey 2002)

What is really amazing about the perfect charge balance between the proton and the electron is that these 2 particles are incredibly different. The proton is about 2,000 times heavier than the electron. In addition, the proton seems to be made up of 3 smaller particles called quarks, while as far as we can tell, the electron is not made up of any smaller particles. Nevertheless, although these 2 particles are quite different, they have *perfectly* balanced charges!

The most important thing to realize is that these atoms are the *simplest* building blocks of matter. Even these very simple building blocks have delicately balanced parts that all work together to give the atom its properties. If one of these parts were even slightly altered, the delicate balance would be destroyed, and the chemistry necessary for life would be forever lost. How is it, then, that these 2 incredibly different particles are able to carry *exactly* the right amount of charge to make a stable atom? How is it that the neutron has *precisely* the right amount of extra mass to make the atom stable? How is it that the electron has *precisely* the right mass to make the existence of living organisms possible? Without these 3 incredible “coincidences,” life could never exist in the universe. Nevertheless, we are here.

As noted before, there are many such incredible “coincidences” in nature. There are so many that it might seem impossible for anyone who really understands physics or chemistry to believe that all this we are studying just came about by chance! The universe is too perfectly designed for that. From the smallest atom to the largest galaxy, incredible design features abound in our universe. These design features are the fingerprints of God, making all who objectively study the universe aware of His existence.

SUMMARY OF KEY EQUATIONS IN MODULE 3

Equation 3.1: $f = \frac{c}{\lambda}$ Frequency and wavelength relationship

Equation 3.2: $E = h \times f$ Frequency and energy relationship

ANSWERS TO THE “ON YOUR OWN” QUESTIONS

- 3.1 (a) On the periodic table, Mo has an atomic number of 42.
So a **Mo atom has 42 protons and 42 electrons.**
- (b) On the periodic table, Be has an atomic number of 4.
So a **Be atom has 4 protons and 4 electrons.**
- (c) On the periodic table, Hf has an atomic number of 72.
So an **Hf atom has 72 protons and 72 electrons.**
- 3.2 (a) On the periodic table, Ne has an atomic number of 10. This means it has **10 protons and 10 electrons.** Its mass number, according to the problem, is 22. If it has 22 total protons + neutrons, and it has 10 protons, then it has $22 - 10 =$ **12 neutrons.**
- (b) On the periodic table, fluorine (F) has an atomic number of 9. This means it has **9 protons and 9 electrons.** Its mass number, according to the problem, is 19. If it has 19 total protons + neutrons, and it has 9 protons, then it has $19 - 9 =$ **10 neutrons.**
- (c) On the periodic table, U has an atomic number of 92. This means it has **92 protons and 92 electrons.** Its mass number, according to the problem, is 235. If it has 235 total protons + neutrons, and it has 92 protons, then it has $235 - 92 =$ **143 neutrons.**
- 3.3 First, since it has 33 protons, we know its atomic number is 33. The element on the table which has an atomic number of 33 is arsenic (As). The problem states that this atom also has 41 neutrons; its mass number is $33 + 41 = 74$. Therefore, **the atom is arsenic-74, which has a symbol of ^{74}As .**

- 3.4 To solve this, we use equation 3.1:

$$f = \frac{c}{\lambda}$$

But we have to rearrange it so that we are solving for wavelength:

$$\lambda = \frac{c}{f}$$

Now we can plug in the numbers:

$$\lambda = \frac{3.0 \times 10^8 \frac{\text{m}}{\text{s}}}{6.4 \times 10^{14} \frac{1}{\text{s}}} = 4.7 \times 10^{-7} \text{ m}$$

Figure 3.10 lists wavelength in nm. If we refer to the chart, we must convert to nm:

$$\frac{4.7 \times 10^{-7} \cancel{\text{m}}}{1} \times \frac{1 \text{ nm}}{10^{-9} \cancel{\text{m}}} = 4.7 \times 10^2 \text{ nm}$$

According to figure 3.10, this light is **blue light**.

- 3.5 To solve this problem, we need to simply convert the wavelength range into a frequency range. To do this, we just convert the 2 wavelengths in the range into frequencies:

$$\frac{7.00 \times 10^2 \cancel{\text{nm}}}{1} \times \frac{10^{-9} \text{ m}}{1 \cancel{\text{nm}}} = 7.00 \times 10^{-7} \text{ m}$$

Now that our units agree, we can put the numbers into the equation:

$$f = \frac{3.0 \times 10^8 \frac{\cancel{\text{m}}}{\text{s}}}{7.00 \times 10^{-7} \cancel{\text{m}}} = 4.3 \times 10^{14} \frac{1}{\text{s}}$$

Now we convert the other part of the range:

$$\frac{6.5 \times 10^2 \cancel{\text{nm}}}{1} \times \frac{10^{-9} \text{ m}}{1 \cancel{\text{nm}}} = 6.50 \times 10^{-7} \text{ m}$$

Now that our units agree, we can put the numbers into the equation:

$$f = \frac{3.0 \times 10^8 \frac{\cancel{\text{m}}}{\text{s}}}{6.50 \times 10^{-7} \cancel{\text{m}}} = 4.6 \times 10^{14} \frac{1}{\text{s}}$$

So the frequency range is **between 4.3×10^{14} Hz and 4.6×10^{14} Hz**.

- 3.6 This problem is a direct application of equation 3.2. Since we know both E and h, we simply have to use algebra and rearrange the equation to solve for f:

$$f = \frac{E}{h}$$

$$f = \frac{3.4 \times 10^{-14} \cancel{\text{J}}}{6.63 \times 10^{-34} \frac{\cancel{\text{J}}}{\text{Hz}}} = 5.1 \times 10^{19} \text{ Hz}$$

Figure 3.11 shows that this is the frequency of X-rays. If you have ever had a medical X-ray, light of approximately this energy was projected onto you to produce the picture that the doctor or dentist examined.

- 3.7 In this problem, we are given the energy and asked to determine the color. The only thing that can tell us color is wavelength. Equation 3.2 only relates energy and frequency. So we must first determine the frequency, and then we can use equation 3.1 to figure out the wavelength.

To determine the frequency:

$$f = \frac{E}{h}$$

$$f = \frac{3.3 \times 10^{-19} \frac{\text{J}}{\text{J}}}{6.63 \times 10^{-34} \frac{\text{J}}{\text{Hz}}} = 5.0 \times 10^{14} \text{ Hz}$$

Now that we have the frequency, we can use algebra to rearrange equation 3.1 so that we can determine the wavelength:

$$\lambda = \frac{c}{f}$$

$$\lambda = \frac{3.0 \times 10^8 \frac{\text{m}}{\text{s}}}{5.0 \times 10^{14} \text{ Hz}} = \frac{3.0 \times 10^8 \frac{\text{m}}{\text{s}}}{5.0 \times 10^{14} \frac{\text{m}}{\text{s}}} = 6.0 \times 10^{-7} \text{ m}$$

To use figure 3.10, we must convert this answer to nm because that's the unit used in the figure. Remember, 1 nm = 10^{-9} m:

$$\frac{6.0 \times 10^{-7} \text{ m}}{1} \times \frac{1 \text{ nm}}{10^{-9} \text{ m}} = 6.0 \times 10^2 \text{ nm}$$

According to figure 3.10, light with this wavelength is **yellow light**.

- 3.8 To get to Si (14), we have to go through several boxes. We must walk through all of row 1, which has 2 boxes in the s orbital block. That gives us $1s^2$. We must also go through all of row 2, which has 2 boxes in the s orbital block and 6 boxes in the p orbital block. This gives us $2s^2 2p^6$. We also must go through the s orbital block in row 3. This gives us $3s^2$. Finally, to get to P, we go through 2 boxes in the row 3, p orbital block. This gives us $3p^2$. The electron configuration, then, is:



The superscripts add up to the number of electrons in a Si atom (14), so our answer checks out.

- 3.9 To get to Se, we have to walk through several boxes on the table. We must walk through all of row 1, which has 2 boxes in the s orbital block. That gives us $1s^2$. We must also go through all of row 2, which has 2 boxes in the s orbital block and 6 boxes in the p orbital block. This gives us $2s^2 2p^6$. Row 3 has 2 boxes in the

s orbital group and 6 in the p orbital box, so that gives us $3s^23p^6$. Row 4 has 2 boxes in the s orbital block and 10 in the d orbital block that we must go through. Remembering that for d orbitals we must subtract 1 from the row number, this gives us $4s^23d^{10}$. Finally, to get to Se, we must walk through 4 boxes in the row 4, p orbital block. That gives us $4p^4$. Therefore, the final electron configuration is:



Our superscripts add up to 34, so our answer checks out.

3.10 To get to Ag, we have to walk through several boxes on the table. We must walk through all of row 1, which has 2 boxes in the s orbital block. That gives us $1s^2$. We must also go through all of row 2, which has 2 boxes in the s orbital block and 6 boxes in the p orbital block. This gives us $2s^22p^6$. Row 3 has 2 boxes in the s orbital group and 6 in the p orbital box, so that gives us $3s^23p^6$. Row 4 has 2 boxes in the s orbital block, 10 in the d orbital block, and 6 in the p orbital block. Remembering that for d orbitals we must subtract 1 from the row number, this gives us $4s^23d^{10}4p^6$. In row 5, we must go through the 2 boxes in the s orbital block, so that gives us $5s^2$. In row 5, we also have to go through 9 d orbital boxes to get to Ag. Once again, we subtract 1 from the row number with d orbitals, so this gives us $4d^9$. Our electron configuration is:



Our superscripts add up to 47, so our answer checks out.

3.11 To abbreviate an electron configuration, we just find the nearest 8A element that has a lower atomic number than the atom we are interested in. In this case, Ar fits the bill. We can then say the only differences between Br and Ar are that there are 2 more boxes in the row 4, s orbital block, 10 boxes in the row 4, d orbital block, and 5 boxes in the row 4, p orbital block. We can write the electron configuration as:



3.12 The nearest 8A element that has a lower atomic number than In is Kr. The only differences between In and Kr are that there are 2 boxes in the row 5, s orbital group, 10 boxes in the row 5, d orbital group, and 1 box in the row 5, p orbital group. The abbreviated electron configuration for In is:



STUDY GUIDE FOR MODULE 3

REVIEW QUESTIONS

1. What was Rutherford's main contribution to the study of atomic structure?
2. What do we call the experimental apparatus that William Crookes used in his experiments, and what did he discover with it?
3. If 2 electrically charged particles repel one another, what can we conclude about their charges?
4. If a substance has 9 positive charges and 11 negative charges, is it positively charged, negatively charged, or neutral?
5. Give the location within the atom for a proton, electron, and a neutron.
6. Why is it so hard to separate one isotope from another?
7. What were the differences between the plum pudding model of the atom and the planetary model of the atom?
8. What is meant by saying that wavelength and frequency are inversely proportional to each other?
9. If you have an orange light bulb and a violet one, which emits waves with the largest wavelength? Which emits light of higher frequency? Which emits the higher energy light? (Do *not* look at the figures in the book. Just remember the name you were told to remember!)
10. Two light bulbs emit light of the same color. However, one is much brighter than the other. What can you say about the wavelengths and amplitudes of the waves being emitted by each bulb?
11. If an atom absorbs energy, what happens to its electrons? If the atom is emitting light, what are its electrons doing?
12. What is the concept of ground state and why is it so important in chemistry?
13. The 3 fundamental particles that make up the atom are the proton, neutron, and electron. Order them in terms of decreasing mass.
14. Define and give one example of a physical constant.
15. Describe what an orbital is.

PRACTICE PROBLEMS

$$(c = 3.0 \times 10^8 \frac{\text{m}}{\text{s}}; \text{ nano- means } 10^{-9}; h = 6.63 \times 10^{-34} \frac{\text{J}}{\text{Hz}})$$

- Give the number of protons, electrons, and neutrons in the following atoms:
 - ^{90}Zr
 - ^{22}Ne
 - ^{58}Ni
 - ^{222}Rn
- Which of the following atoms are isotopes?

$^{22}\text{Na}, ^{22}\text{Ne}, ^{23}\text{Na}, ^{22}\text{Mg}, ^{24}\text{Na}$
- What is the symbol of the atom made up of 6 protons, 8 neutrons, and 6 electrons?
- If light has a frequency of 2.4×10^{14} Hz, what is its wavelength?
- What is the energy of a light wave with a frequency of 9.8×10^{20} Hz?
- If a light wave has a wavelength of 712 nm, what is its energy?
- One electron is in a 2s orbital, while another is in a 3p orbital. Which has the higher energy? What shape is each electron's orbital?
- Give the full electron configurations of the following atoms:
 - V
 - N
 - Cs
- Give the abbreviated electron configurations for the following atoms:
 - Ti
 - S
 - Rb
- What is wrong with the following electron configurations?
 - $1s^2 2s^2 2p^6 3s^2 3p^7 4s^2 4d^9$
 - $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$