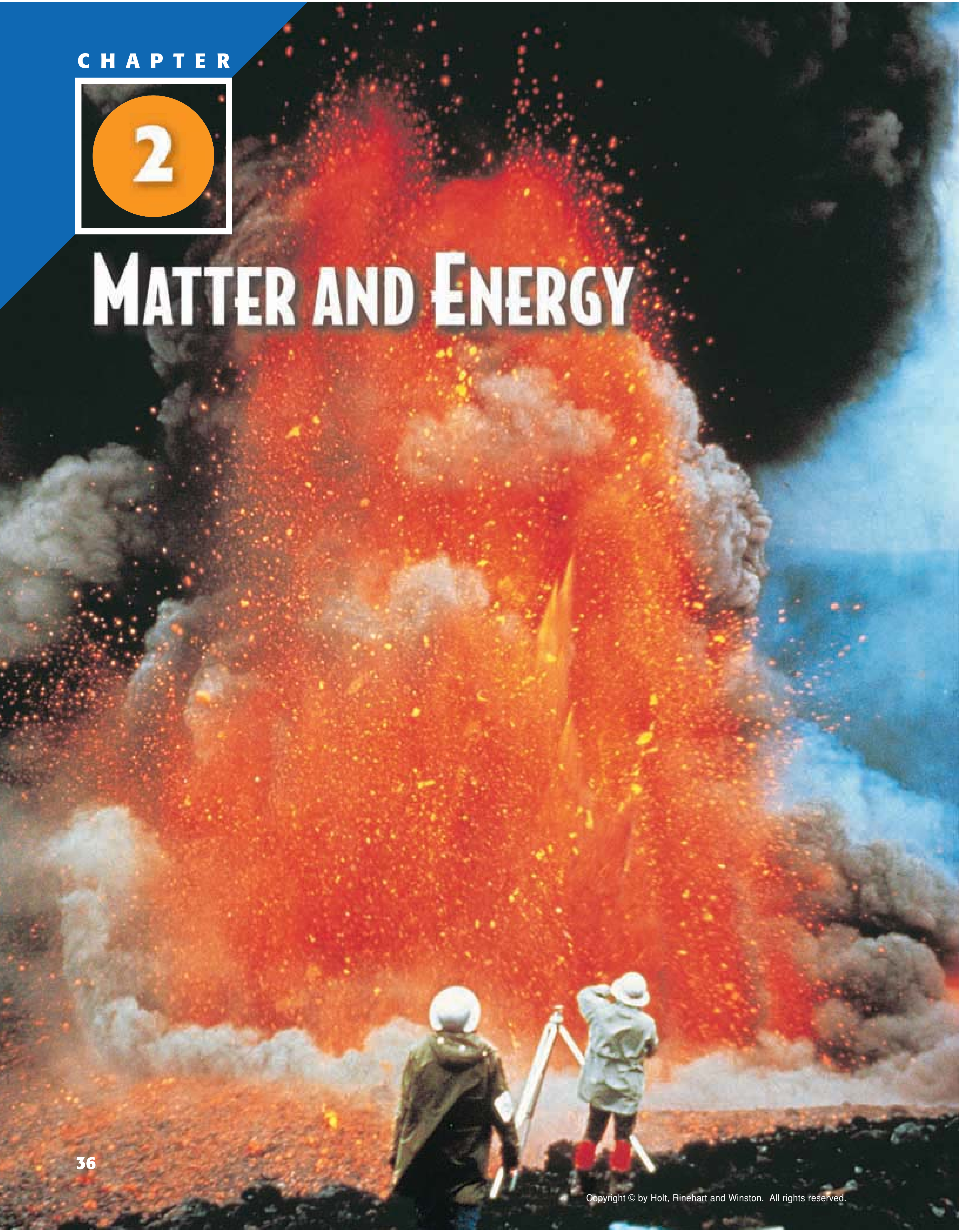
CHAPTER

2

MATTER AND ENERGY



The photo of the active volcano and the scientists who are investigating it is a dramatic display of matter and energy. Most people who view the photo would consider the volcano and the scientists to be completely different. The scientists seem to be unchanging, while the volcano is explosive and changing rapidly. However, the scientists and the volcano are similar in that they are made of matter and are affected by energy. This chapter will show you the relationship between matter and energy and some of the rules that govern them.

START-UP ACTIVITY

SAFETY PRECAUTIONS

Chemical Changes and Energy

PROCEDURE

1. Place a small **thermometer** completely inside a **jar,** and close the **lid.** Wait 5 min, and record the temperature.
2. While you are waiting to record the temperature, soak one-half of a **steel wool pad** in **vinegar** for 2 min.
3. Squeeze the excess vinegar from the steel wool. Remove the thermometer from the jar, and wrap the steel wool around the bulb of the thermometer. Secure the steel wool to the thermometer with a **rubber band.**
4. Place the thermometer and the steel wool inside the jar, and close the lid. Wait 5 min, and record the temperature.

ANALYSIS

1. How did the temperature change?
2. What do you think caused the temperature to change?
3. Do you think vinegar is a reactant or product? Why?

Pre-Reading Questions

1. **When ice melts, what happens to its chemical composition?**
2. **Name a source of energy for your body.**
3. **Name some temperature scales.**
4. **What is a chemical property? What is a physical property?**

CONTENTS 2

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SECTION 1

Energy

KEY TERMS

- energy
- physical change
- chemical change
- evaporation
- endothermic
- exothermic
- law of conservation of energy
- heat
- kinetic energy
- temperature
- specific heat

OBJECTIVES

1. **Explain** that physical and chemical changes in matter involve transfers of energy.
2. **Apply** the law of conservation of energy to analyze changes in matter.
3. **Distinguish** between heat and temperature.
4. **Convert** between the Celsius and Kelvin temperature scales.

Energy and Change

If you ask 10 people what comes to mind when they hear the word *energy,* you will probably get 10 different responses. Some people think of energy in terms of exercising or playing sports. Others may picture energy in terms of a fuel or a certain food.

If you ask 10 scientists what comes to mind when they hear the word *energy,* you may also get 10 different responses. A geologist may think of energy in terms of a volcanic eruption. A biologist may visualize cells using oxygen and sugar in reactions to obtain the energy they need. A chemist may think of a reaction in a lab, such as the one shown in **Figure 1.**

The word *energy* represents a broad concept. One definition of **energy** is the capacity to do some kind of work, such as moving an object, forming a new compound, or generating light. No matter how energy is defined, it is always involved when there is a change in matter.

energy
the capacity to do work

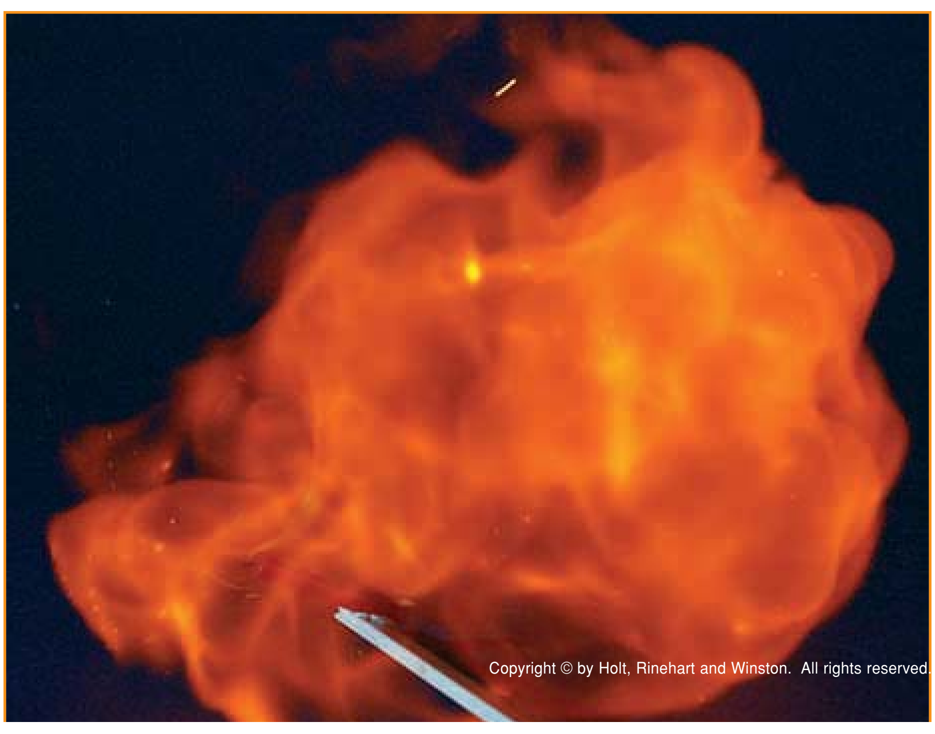


Figure 1
Energy is released in the explosive reaction that occurs between hydrogen and oxygen to form water.

Changes in Matter Can Be Physical or Chemical

Ice melting and water boiling are examples of **physical changes.** A physical change affects only the physical properties of matter. For example, when ice melts and turns into liquid water, you still have the same substance represented by the formula H_2O. When water boils and turns into a vapor, the vapor is still H_2O. Notice that in these examples the chemical nature of the substance does not change; only the physical state of the substance changes to a solid, liquid, or gas.

physical change

a change of matter from one form to another without a change in chemical properties

In contrast, the reaction of hydrogen and oxygen to produce water is an example of a **chemical change.** A chemical change occurs whenever a new substance is made. In other words, a chemical reaction has taken place. You know water is different from hydrogen and oxygen because water has different properties. For example, the boiling points of hydrogen and oxygen at atmospheric pressure are –252.8°C and –182.962°C, respectively. The boiling point of water at atmospheric pressure is 100°C. Hydrogen and oxygen are also much more reactive than water.

chemical change

a change that occurs when one or more substances change into entirely new substances with different properties

Every Change in Matter Involves a Change in Energy

All physical and chemical changes involve a change in energy. Sometimes energy must be supplied for the change in matter to occur. For example, consider a block of ice, such as the one shown in **Figure 2.** As long as the ice remains cold enough, the particles in the solid ice stay in place.

However, if the ice gets warm, the particles will begin to move and vibrate more and more. For the ice to melt, energy must be supplied so that the particles can move past one another. If more energy is supplied and the boiling point of water is reached, the particles of the liquid will leave the liquid's surface through **evaporation** and form a gas. These physical changes require an input of energy. Many chemical changes also require an input of energy.

evaporation

the change of a substance from a liquid to a gas

Sometimes energy is released when a change in matter occurs. For example, energy is released when a vapor turns into a liquid or when a liquid turns into a solid. Some chemical changes also release energy. The explosion that occurs when hydrogen and oxygen react to form water is a release of energy.

Figure 2
Energy is involved when a physical change, such as the melting of ice, happens.

Endothermic and Exothermic Processes

endothermic

describes a process in which heat is absorbed from the environment

Any change in matter in which energy is absorbed is known as an **endothermic** process. The melting of ice and the boiling of water are two examples of physical changes that are endothermic processes.

Some chemical changes are also endothermic processes. **Figure 3** shows a chemical reaction that occurs when barium hydroxide and ammonium nitrate are mixed. Notice in **Figure 3** that these two solids form a liquid, slushlike product. Also, notice the ice crystals that form on the surface of the beaker. As barium hydroxide and ammonium nitrate react, energy is absorbed from the beaker's surroundings. As a result, the beaker feels colder because the reaction absorbs energy as heat from your hand. Water vapor in the air freezes on the surface of the beaker, providing evidence that the reaction is endothermic.

exothermic

describes a process in which a system releases heat into the environment

Any change in matter in which energy is released is an **exothermic** process. The freezing of water and the condensation of water vapor are two examples of physical changes that are exothermic processes.

Recall that when hydrogen and oxygen gases react to form water, an explosive reaction occurs. The vessel in which the reaction takes place becomes warmer after the reaction, giving evidence that energy has been released.

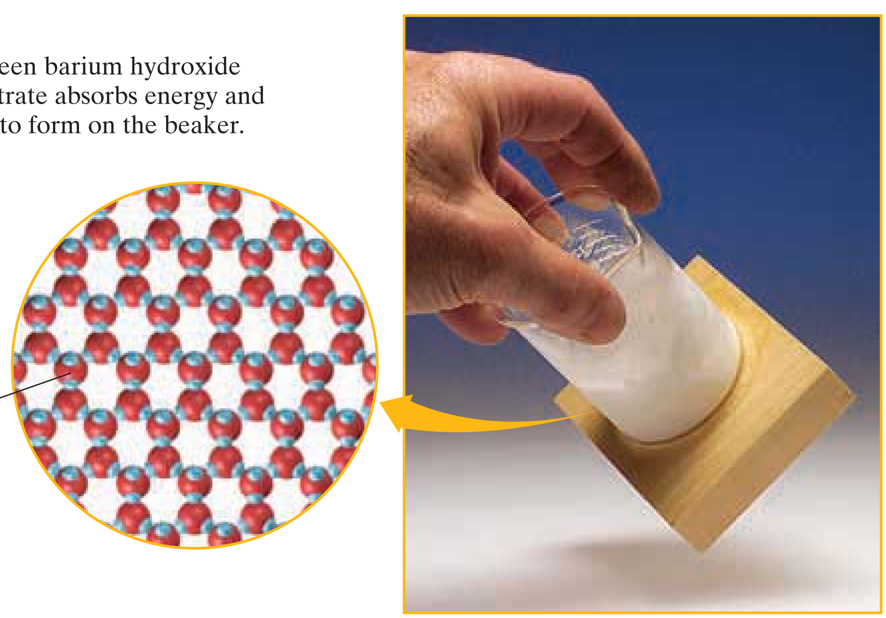
law of conservation of energy

the law that states that energy cannot be created or destroyed but can be changed from one form to another

Endothermic processes, in which energy is absorbed, may make it seem as if energy is being destroyed. Similarly, exothermic processes, in which energy is released, may make it seem as if energy is being created. However, the **law of conservation of energy** states that during any physical or chemical change, the total quantity of energy remains constant. In other words, energy cannot be destroyed or created.

Accounting for all the different types of energy present before and after a physical or chemical change is a difficult process. But measurements of energy during both physical and chemical changes have shown that when energy seems to be destroyed or created, energy is actually being transferred. The difference between exothermic and endothermic processes is whether energy is absorbed or released by the substances involved.

Figure 3
The reaction between barium hydroxide and ammonium nitrate absorbs energy and causes ice crystals to form on the beaker.



Conservation of Energy in a Chemical Reaction

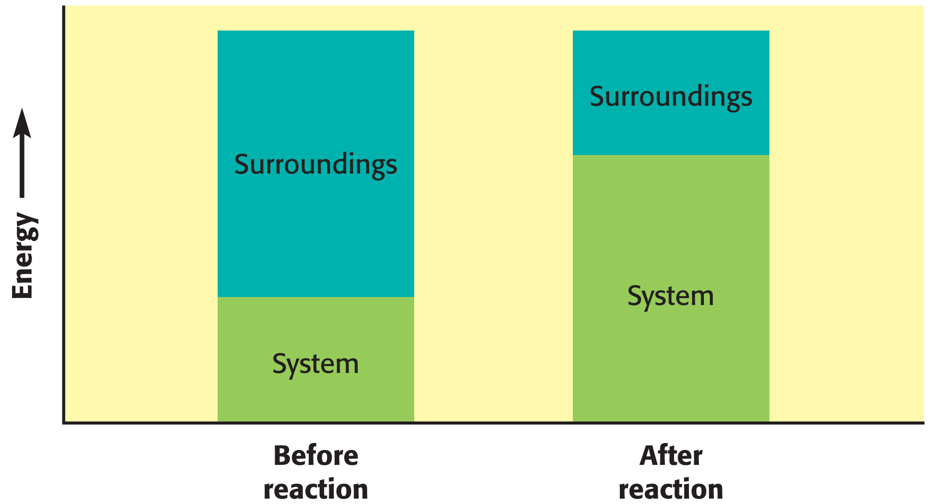


Figure 4 Notice that the energy of the reactants and products increases, while the energy of the surroundings decreases. However, the total energy does not change.

Energy Is Often Transferred

Figure 4 shows the energy changes that take place when barium hydroxide and ammonium nitrate react. To keep track of energy changes, chemists use the terms *system* and *surroundings*. A system consists of all the components that are being studied at any given time. In **Figure 4,** the system consists of the mixture inside the beaker. The surroundings include everything outside the system. In **Figure 4,** the surroundings consist of everything else including the air both inside and outside the beaker and the beaker itself. Keep in mind that the air is made of various gases.

Energy is often transferred back and forth between a system and its surroundings. An exothermic process involves a transfer of energy from a system to its surroundings. An endothermic process involves a transfer of energy from the surroundings to the system. However, in every case, the total energy of the systems and their surroundings remains the same, as shown in **Figure 4.**

Energy Can Be Transferred in Different Forms

Energy exists in different forms, including chemical, mechanical, light, heat, electrical, and sound. The transfer of energy between a system and its surroundings can involve any one of these forms of energy. Consider the process of *photosynthesis*. Light energy is transferred from the sun to green plants. Chlorophyll inside the plant's cells (the system) absorbs energy—the light energy from the sun (the surroundings). This light energy is converted to chemical energy when the plant synthesizes chemical nutrients that serve as the basis for sustaining all life on Earth.

Next, consider what happens when you activate a light stick. Chemicals inside the stick react to release energy in the form of light. This light energy is transferred from the system inside the light stick to the surroundings, generating the light that you see. A variety of animals depend on chemical reactions that generate light, including fish, worms, and fireflies.

Heat

heat

the energy transferred between objects that are at different temperatures; energy is always transferred from higher-temperature objects to lower-temperature objects until thermal equilibrium is reached

Heat is the energy transferred between objects that are at different temperatures. This energy is always transferred from a warmer object to a cooler object. For example, consider what happens when ice cubes are placed in water. Energy is transferred from the liquid water to the solid ice. The transfer of energy as heat during this physical change will continue until all the ice cubes have melted. But on a warm day, we know that the ice cubes will not release energy that causes the water to boil, because energy cannot be transferred from the cooler objects to the warmer one. Energy is also transferred as heat during chemical changes. In fact, the most common transfers of energy in chemistry are those that involve heat.

Energy Can Be Released As Heat

The worst industrial disaster in U.S. history occurred in April 1947. A cargo ship named the *Grandcamp* had been loaded with fertilizer in Texas City, a Texas port city of 50 000 people. The fertilizer consisted of tons of a compound called *ammonium nitrate.* Soon after the last bags of fertilizer had been loaded, a small fire occurred, and smoke was noticed coming from the ship's cargo hold. About an hour later, the ship exploded.

The explosion was heard 240 km away. An anchor from the ship flew through the air and created a 3 m wide hole in the ground where it landed. Every building in the city was either destroyed or damaged. The catastrophe on the *Grandcamp* was caused by an exothermic chemical reaction that released a tremendous amount of energy as heat.

Figure 5
Billowing black smoke filled the sky over Texas City in the aftermath of the *Grandcamp* explosion, shown in this aerial photograph.

All of this energy that was released came from the energy that was stored within the ammonium nitrate. Energy can be stored within a chemical substance as chemical energy. When the ammonium nitrate ignited, an exothermic chemical reaction took place and released energy as heat. In addition, the ammonium nitrate explosion generated **kinetic energy,** as shown by the anchor that flew through the air.

kinetic energy

the energy of an object that is due to the object's motion

Energy Can Be Absorbed As Heat

In an endothermic reaction, energy is absorbed by the chemicals that are reacting. If you have ever baked a cake or a loaf of bread, you have seen an example of such a reaction. Recipes for both products require either baking soda or baking powder. Both baking powder and baking soda contain a chemical that causes dough to rise when heated in an oven.

The chemical found in both baking powder and baking soda is sodium bicarbonate. Energy from the oven is absorbed by the sodium bicarbonate. The sodium bicarbonate breaks down into three different chemical substances, sodium carbonate, water vapor, and carbon dioxide gas, in the following endothermic reaction:

$$2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$$

The carbon dioxide gas causes the batter to rise while baking, as you can see in **Figure 6.**

Figure 6
Baking a cake or bread is an example of an endothermic reaction, in which energy is absorbed as heat.

Heat Is Different from Temperature

You have learned that energy can be transferred as heat because of a temperature difference. So, the transfer of energy as heat can be measured by calculating changes in temperature. Temperature indicates how hot or cold something is. **Temperature** is actually a measurement of the average kinetic energy of the random motion of particles in a substance.

temperature
a measure of how hot (or cold) something is; specifically, a measure of the average kinetic energy of the particles in an object

For example, imagine that you are heating water on a stove to make tea. The water molecules have kinetic energy as they move freely in the liquid. Energy transferred as heat from the stove causes these water molecules to move faster. The more rapidly the water molecules move, the greater their average kinetic energy. As the average kinetic energy of the water molecules increases, the temperature of the water increases. Think of heat as the energy that is transferred from the stove to the water because of a difference in the temperatures of the stove and the water. The temperature change of the water is a measure of the energy transferred as heat.

Temperature Is Expressed Using Different Scales

Thermometers are usually marked with the Fahrenheit or Celsius temperature scales. However, the Fahrenheit scale is not used in chemistry. Recall that the SI unit for temperature is the *Kelvin,* K. The zero point on the Celsius scale is designated as the freezing point of water. The zero point on the Kelvin scale is designated as *absolute zero,* the temperature at which the minimum average kinetic energies of all particles occur.

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In chemistry, you will have to use both the Celsius and Kelvin scales. At times, you will have to convert temperature values between these two scales. Conversion between these two scales simply requires an adjustment to account for their different zero points.

$$t(°\text{C}) = T(\text{K}) - 273.15\ \text{K} \qquad T(\text{K}) = t(°\text{C}) + 273.15°\text{C}$$

The symbols t and T represent temperatures in degrees Celsius and in kelvins, respectively. Also, notice that a temperature change is the same in kelvins and in Celsius degrees.

Transfer of Heat May Not Affect the Temperature

The transfer of energy as heat does not always result in a change of temperature. For example, consider what happens when energy is transferred to a solid such as ice. Imagine that you have a mixture of ice cubes and water in a sealed, insulated container. A thermometer is inserted into the container to measure temperature changes as energy is added to the ice-water mixture.

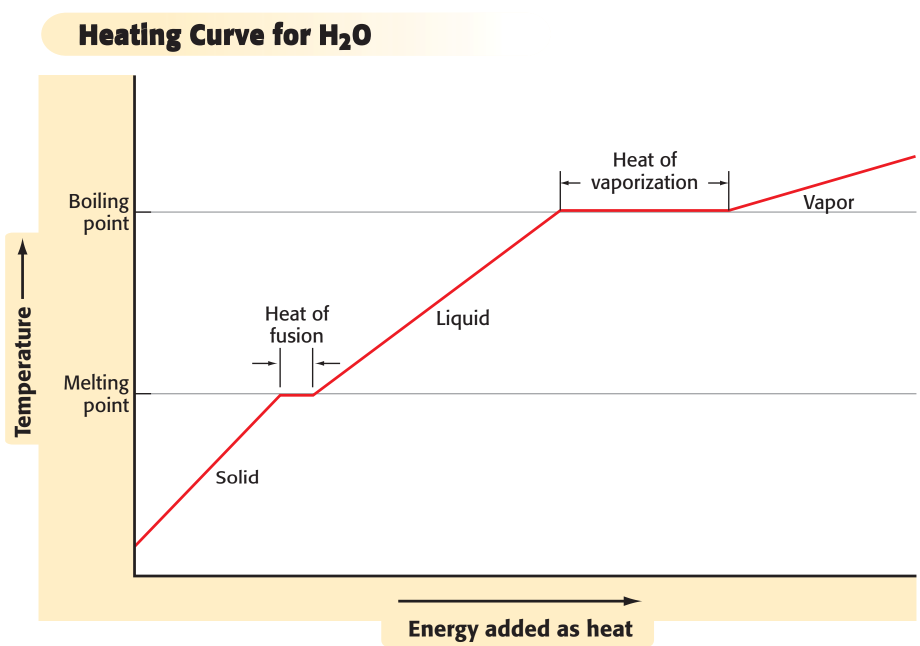
As energy is transferred as heat to the ice-water mixture, the ice cubes will start to melt. However, the temperature of the mixture remains at 0°C. Even though energy is continuously being transferred as heat, the temperature of the ice-water mixture does not increase.

Once all the ice has melted, the temperature of the water will start to increase. When the temperature reaches 100°C, the water will begin to boil. As the water turns into a gas, the temperature remains at 100°C, even though energy is still being transferred to the system as heat. Once all the water has vaporized, the temperature will again start to rise.

Notice that the temperature remains constant during the physical changes that occur as ice melts and water vaporizes. What happens to the energy being transferred as heat if the energy does not cause an increase in temperature? The energy that is transferred as heat is actually being used to move molecules past one another or away from one another. This energy causes the molecules in the solid ice to move more freely so that they form a liquid. This energy also causes the water molecules to move farther apart so that they form a gas.

Figure 7 shows the temperature changes that occur as energy is transferred as heat to change a solid into a liquid and then into a gas. Notice that the temperature increases only when the substance is in the solid, liquid, or gaseous states. The temperature does not increase when the solid is changing to a liquid or when the liquid is changing to a gas.

Figure 7
This graph illustrates how temperature is affected as energy is transferred to ice as heat. Notice that much more energy must be transferred as heat to vaporize water than to melt ice.



Transfer of Heat Affects Substances Differently

Have you ever wondered why a heavy iron pot gets hot fast but the water in the pot takes a long time to warm up? If you transfer the same quantity of heat to similar masses of different substances, they do not show the same increase in temperature. This relationship between energy transferred as heat to a substance and the substance's temperature change is called the **specific heat.**

specific heat
the quantity of heat required to raise a unit mass of homogeneous material 1 K or 1°C in a specified way given constant pressure and volume

The specific heat of a substance is the quantity of energy as heat that must be transferred to raise the temperature of 1 g of a substance 1 K. The SI unit for energy is the *joule* (J). Specific heat is expressed in joules per gram kelvin (J/g·K).

Metals tend to have low specific heats, which indicates that relatively little energy must be transferred as heat to raise their temperatures. In contrast, water has an extremely high specific heat. In fact, it is the highest of most common substances.

During a hot summer day, water can absorb a large quantity of energy from the hot air and the sun and can cool the air without a large increase in the water's temperature. During the night, the water continues to absorb energy from the air. This energy that is removed from the air causes the temperature of the air to drop quickly, while the water's temperature changes very little. This behavior is explained by the fact that air has a low specific heat and water has a high specific heat.

1 Section Review

UNDERSTANDING KEY IDEAS

1. What is energy?
2. State the law of conservation of energy.
3. How does heat differ from temperature?
4. What is a system?
5. Explain how an endothermic process differs from an exothermic process.
6. What two temperature scales are used in chemistry?

PRACTICE PROBLEMS

7. Convert the following Celsius temperatures to Kelvin temperatures.
   a. 100°C **c.** 0°C
   b. 785°C **d.** –37°C
8. Convert the following Kelvin temperatures to Celsius temperatures.
   a. 273 K **c.** 0 K
   b. 1200 K **d.** 100 K

CRITICAL THINKING

9. Is breaking an egg an example of a physical or chemical change? Explain your answer.
10. Is cooking an egg an example of a physical or chemical change? Explain your answer.
11. What happens in terms of the transfer of energy as heat when you hold a snowball in your hands?
12. Why is it impossible to have a temperature value below 0 K?
13. If energy is transferred to a substance as heat, will the temperature of the substance always increase? Explain why or why not.

SECTION

2 Studying Matter and Energy

KEY TERMS

- scientific method
- hypothesis
- theory
- law
- law of conservation of mass

OBJECTIVES

1. **Describe** how chemists use the scientific method.
2. **Explain** the purpose of controlling the conditions of an experiment.
3. **Explain** the difference between a hypothesis, a theory, and a law.

The Scientific Method

Science is unlike other fields of study in that it includes specific procedures for conducting research. These procedures make up the **scientific method,** which is shown in **Figure 8.** The scientific method is not a series of exact steps, but rather a strategy for drawing sound conclusions.

scientific method

a series of steps followed to solve problems, including collecting data, formulating a hypothesis, testing the hypothesis, and stating conclusions

A scientist chooses the procedures to use depending on the nature of the investigation. For example, a chemist who has an idea for developing a better method to recycle plastics may research scientific articles about plastics, collect information, propose a method to separate the materials, and then test the method. In contrast, another chemist investigating the pollution caused by a trash incinerator would select different procedures. These procedures might include collecting and analyzing samples, interviewing people, predicting the role the incinerator plays in producing the pollution, and conducting field studies to test that prediction.

No matter which approach they use, both chemists are employing the scientific method. Ultimately, the success of the scientific method depends on publishing the results so that others can repeat the procedures and verify the results.

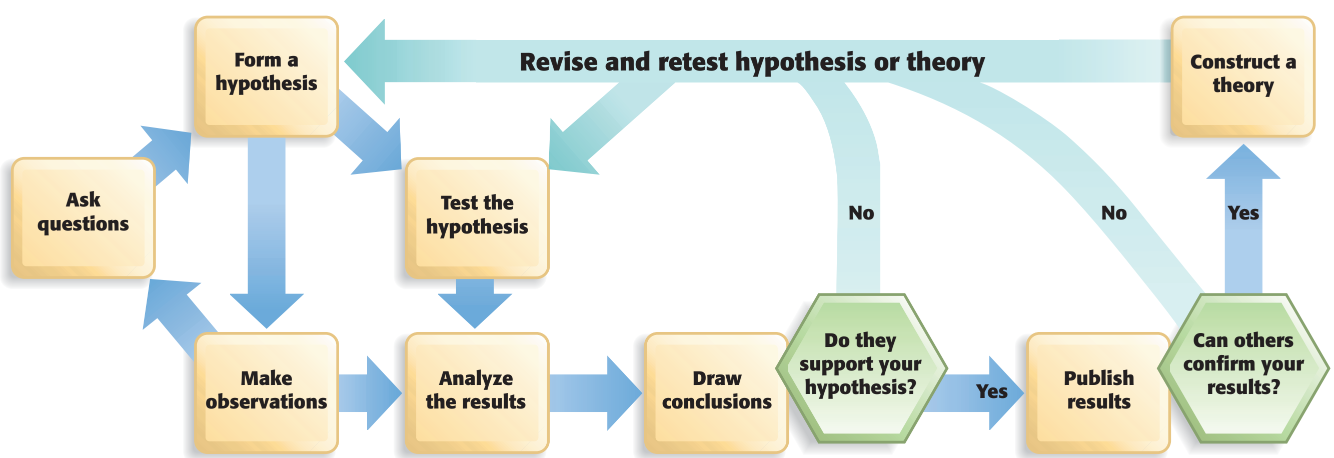


Figure 8
Each stage of the scientific method represents a number of different activities. Scientists choose the activities to use depending on the nature of their investigation.

Quick LAB

Using the Scientific Method

SAFETY PRECAUTIONS

PROCEDURE

1. Have someone prepare **five sealed paper bags,** each containing **an item commonly found in a home.**
2. Without opening the bags, try to determine the identity of each item.
3. Test each of your conclusions whenever possible. For example, if you concluded that one of the items is a refrigerator magnet, test it to see if it attracts small metal objects, such as paper clips.

ANALYSIS

1. How many processes that are part of the scientific method shown in **Figure 8** did you use?
2. How many items did you correctly identify?

Experiments Are Part of the Scientific Method

The first scientists depended on rational thought and logic. They rarely felt it was necessary to test their ideas or conclusions, and they did not feel the need to experiment. Gradually, experiments became the crucial test for the acceptance of scientific knowledge. Today, experiments are an important part of the scientific method.

An experiment is the process by which scientific ideas are tested. For example, consider what happens when manganese dioxide is added to a solution of hydrogen peroxide. Tiny bubbles of gas soon rise to the surface of the solution, indicating that a chemical reaction has taken place. Now, consider what happens when a small piece of beef liver is added to a solution of hydrogen peroxide. Tiny gas bubbles are produced. So, you might conclude that the liver contains manganese dioxide. To support your conclusion, you would have to test for the presence of manganese dioxide in the piece of liver.

Experiments May Not Turn Out As Expected

Your tests would reveal that liver does not contain any manganese dioxide. In this case, the results of the experiment did not turn out as you might have expected. Scientists are often confronted by situations in which their results do not turn out as expected. Scientists do not view these results as a failure. Rather, they analyze these results and continue with the scientific method. Unexpected results often give scientists as much information as expected results do. So, unexpected results are as important as expected results.

In this case, the liver might contain a different chemical that acts like manganese dioxide when added to hydrogen peroxide. Additional experiments would reveal that the liver does in fact contain such a chemical. Experimental results can also lead to more experiments. Perhaps the chemical that acts like manganese dioxide can be found in other parts of the body.

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Scientific Discoveries Can Come from Unexpected Observations

Not all discoveries and findings are the results of a carefully worked-out plan based on the scientific method. In fact, some important discoveries and developments have been made simply by accident. An example in chemistry is the discovery of a compound commonly known as Teflon®. You are probably familiar with Teflon as the nonstick coating used on pots and pans, but it has many more applications.

Teflon is used as thermal insulation in clothing, as a component in wall coverings, and as a protective coating on metals, glass, and plastics. Teflon's properties of very low chemical reactivity and very low friction make it valuable in the construction of artificial joints for human limbs. As you can see in **Figure 9,** Teflon is also used as a roofing material.

Teflon was not discovered as a result of a planned series of experiments designed to produce this chemical compound. Rather, it was discovered when a scientist made a simple but puzzling observation.

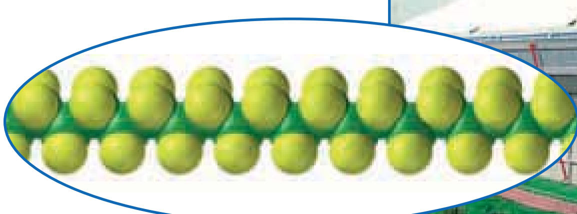
Teflon Was Discovered by Chance

In 1938, Dr. Roy Plunkett, a chemist employed by DuPont, was trying to produce a new coolant gas to use as a refrigerant. He was hoping to develop a less expensive coolant than the one that was being widely used at that time. His plan was to allow a gas called *tetrafluoroethene* (TFE) to react with hydrochloric acid. To begin his experiment, Plunkett placed a cylinder of liquefied TFE on a balance to record its mass.

He then opened the cylinder to let the TFE gas flow into a container filled with hydrochloric acid. But no TFE came out of the cylinder. Because the cylinder had the same mass as it did when it was filled with TFE, Plunkett knew that none of the TFE had leaked out. He removed the valve and shook the cylinder upside down. Only a few white flakes fell out.

Curious about what had happened, Plunkett decided to analyze the white flakes. He discovered that he had accidentally created the proper conditions for TFE molecules to join together to form a long chain. These long-chained molecules were very slippery. After 10 years of additional research, large-scale manufacturing of these long-chained molecules, known as Teflon or polytetrafluoroethene (PTFE), became practical.

Figure 9
Teflon was used to make the roof of the Hubert H. Humphrey Metrodome in Minneapolis, Minnesota.



Synthetic Dyes Were Also Discovered by Chance

If you have on an article of clothing that is colored, you are wearing something whose history can be traced to another unexpected chemistry discovery. This discovery was made in 1856 by an 18-year-old student named William Perkin, who was in his junior year at London's Royal College of Chemistry.

At that time, England was the world's leading producer of textiles, including those used for making clothing. The dyes used to color the textiles were natural products, extracted from both plants and animals. Only a few colors were available. In addition, the process to get dyes from raw materials was costly. As a result, only the wealthy could afford to wear brightly colored clothes for everyday use.

Mauve, a deep purple, was the color most people wanted for their clothing. In ancient times, only royalty could afford to own clothes dyed a mauve color. In Perkin's time, only the wealthy people could afford mauve.

STUDY TIP

Learning Terminology

Important terms and their definitions are listed in the margins of this book. Knowing the definitions of these terms is crucial to understanding chemistry. Ask your teacher about any definition that does not make sense.

To determine your understanding of the terms in this chapter, explain their definitions to another classmate.

Making an Unexpected Discovery

At first, Perkin had no interest in brightly colored clothes. Rather, his interest was in finding a way to make quinine, a drug used to treat malaria. At the time, quinine could only be made from the bark of a particular kind of tree. Great Britain needed huge quantities of the drug to treat its soldiers who got malaria in the tropical countries that were part of the British Empire. There was not enough of the drug to keep up with demand.

The only way to get enough quinine was to develop a synthetic version of the drug. During a vacation from college, Perkin was at home experimenting with ways of making synthetic quinine. One of his experiments resulted in a product that was a thick, sticky, black substance. He immediately realized that this attempt to synthesize quinine did not work. Curious about the substance, Perkin washed his reaction vessel with water. But the sticky product would not wash away. Perkin next decided to try cleaning the vessel with an alcohol. What he saw next was an unexpected discovery.

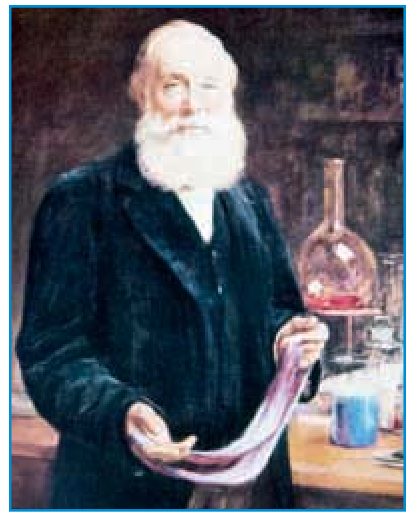


Figure 10
Through his accidental discovery of aniline purple, William Perkin found an inexpensive way to make the color mauve. His discovery brought on the beginning of the synthetic dye industry.

Analyzing an Unexpected Discovery

When Perkin poured an alcohol on the black product, it turned a mauve color. He found a way to extract the purple substance from the black product and determined that his newly discovered substance was perfect for dyeing clothes. He named his accidental discovery "aniline purple," but the fashionable people of Paris soon renamed it mauve.

Perkin became obsessed with his discovery. He left the Royal College of Chemistry and decided to open a factory that could make large amounts of the dye. Within two years, his factory had produced enough dye to ship to the largest maker of silk clothing in London. The color mauve quickly became the most popular color in the fashion industry throughout Europe. Perkin expanded his company and soon started producing other dyes, including magenta and a deep red. As a result of his unexpected discovery, Perkin became a very wealthy man and retired at the age of 36 to devote his time to chemical research. His unexpected discovery also marked the start of the synthetic dye industry.

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Scientific Explanations

Questions that scientists seek to answer and problems that they hope to solve often come after they observe something. These observations can be made of the natural world or in a laboratory. A scientist must always make careful observations, not knowing if some totally unexpected result might lead to an interesting finding or important discovery. Consider what would have happened if Plunkett had ignored the white flakes or if Perkin had overlooked the mauve substance.

Once observations have been made, they must be analyzed. Scientists start by looking at all the relevant information or data they have gathered. They look for patterns that might suggest an explanation for the observations. This proposed explanation is called a **hypothesis.** A hypothesis is a reasonable and testable explanation for observations.

hypothesis
a theory or explanation that is based on observations and that can be tested

Chemists Use Experiments to Test a Hypothesis

Once a scientist has developed a hypothesis, the next step is to test the validity of the hypothesis. This testing is often done by carrying out experiments, as shown in **Figure 11.** Even though the results of their experiments were totally unexpected, Plunkett and Perkin developed hypotheses to account for their observations. Both scientists hypothesized that their accidental discoveries might have some practical application. Their next step was to design experiments to test their hypotheses.

To understand what is involved in designing an experiment, consider this example. Imagine that you have observed that your family car has recently been getting better mileage. Perhaps you suggest to your family that their decision to use a new brand of gasoline is the factor responsible for the improved mileage. In effect, you have proposed a hypothesis to explain an observation.

Figure 11
Students conduct experiments to test the validity of their hypotheses.

Figure 12
Any number of variables may be responsible for the improved mileage that a driver notices. A controlled experiment can identify the variable responsible.

Scientists Must Identify the Possible Variables

To test the validity of your hypothesis, your next step is to plan your experiments. You must begin by identifying as many factors as possible that could account for your observations. A factor that could affect the results of an experiment is called a *variable*. A scientist changes variables one at a time to see which variable affects the outcome of an experiment.

Several variables might account for the improved mileage you noticed with your family car. The use of a new brand of gasoline is one variable. Driving more on highways, making fewer short trips, having the car's engine serviced, and avoiding quick accelerations are other variables that might have resulted in the improved mileage. To know if your hypothesis is right, the experiment must be designed so that each variable is tested separately. Ideally, the experiments will eliminate all but one variable so that the exact cause of the observed results can be identified.

Each Variable Must Be Tested Individually

Scientists reduce the number of possible variables by keeping all the variables constant except one. When a variable is kept constant from one experiment to the next, the variable is called a *control* and the procedure is called a *controlled experiment*. Consider how a controlled experiment would be designed to identify the variable responsible for the improved mileage.

You would fill the car with the new brand of gasoline and keep an accurate record of how many miles you get per gallon. When the gas tank is almost empty, you would do the same after filling the car with the brand of gasoline your family had been using before. In both trials, you should drive the car under the same conditions. For example, the car should be driven the same number of miles on highways and local streets and at the same speeds in both trials. You then have designed the experiment so that only one variable—the brand of gasoline—is being tested.

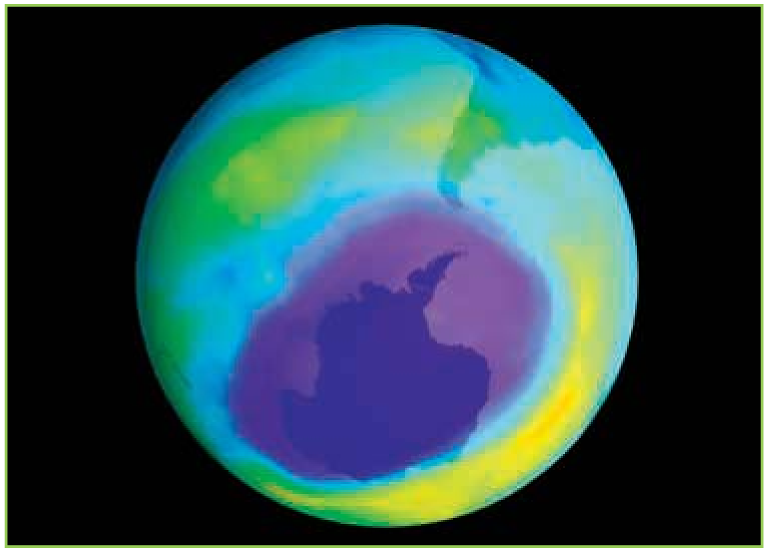


Figure 13
In 1974, scientists proposed a theory to explain the observation of a hole in the ozone layer over Antarctica, which is shown in purple. This hole is about the size of North America.

Data from Experiments Can Lead to a Theory

As early as 1969, scientists observed that the ozone layer was breaking down. Ozone, O_3, is a gas that forms a thin layer high above Earth's surface. This layer shields all living things from most of the sun's damaging ultraviolet light. In 1970, Paul Crutzen, working at the Max Planck Institute for Chemistry, showed the connection between nitrogen oxides and the reduction of ozone in air. In 1974, F. Sherwood Rowland and Mario Molina, two chemists working at the University of California, Irvine, proposed the hypothesis that the release of chlorofluorocarbons (CFCs) into the atmosphere harms the ozone layer. CFCs were being used in refrigerators, air conditioners, aerosol spray containers, and many other consumer products.

Repeated testing has supported the hypothesis proposed by Rowland and Molina. Any hypothesis that withstands repeated testing may become part of a **theory.** In science, a theory is a well-tested explanation of observations. (This is different from common use of the term, which means "a guess.") Because theories are explanations, not facts, they can be disproved but can never be completely proven. In 1995, Crutzen, Rowland, and Molina were awarded the Nobel Prize in chemistry in recognition of their theory of the formation and decomposition of the ozone layer.

theory

an explanation for some phenomenon that is based on observation, experimentation, and reasoning

Theories and Laws Have Different Purposes

Some facts in science hold true consistently. Such facts are known as laws. A **law** is a statement or mathematical expression that reliably describes a behavior of the natural world. While a theory is an attempt to explain the cause of certain events in the natural world, a scientific law describes the events.

For example, the **law of conservation of mass** states that the products of a chemical reaction have the same mass as the reactants have. This law does not explain why matter in chemical reactions behaves this way; the law simply describes this behavior. In some cases, scientific laws may be reinterpreted as new information is obtained. Keep in mind that a hypothesis *predicts* an event, a theory *explains* it, and a law *describes* it.

law

a summary of many experimental results and observations; a law tells how things work

law of conservation of mass

the law that states that mass cannot be created or destroyed in ordinary chemical and physical changes

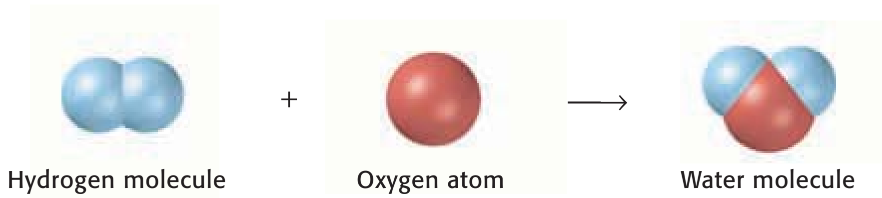


Figure 14
Models can be used to show what happens during a reaction between a hydrogen molecule and an oxygen atom.

Models Can Illustrate the Microscopic World of Chemistry

Models play a major role in science. A *model* represents an object, a system, a process, or an idea. A model is also simpler than the actual thing that is modeled. In chemistry, models can be most useful in understanding what is happening at the microscopic level. In this book, you will see numerous illustrations showing models of chemical substances. These models, such as the ones shown in **Figure 14,** are intended to help you understand what happens during physical and chemical changes.

Keep in mind that models are simplified representations. For example, the models of chemical substances that you will examine in this book include various shapes, sizes, and colors. The actual particles of these chemical substances do not have the shapes, sizes, or brilliant colors that are shown in these models. However, these models do show the geometric arrangement of the units, their relative sizes, and how they interact.

One tool that is extremely useful in the construction of models is the computer. Computer-generated models enable scientists to design chemical substances and explore how they interact in virtual reality. A chemical model that looks promising for some practical application, such as treating a disease, might be the basis for the synthesis of the actual chemical.

2 Section Review

UNDERSTANDING KEY IDEAS

1. How does a hypothesis differ from a theory?
2. What is the scientific method?
3. Do experiments always turn out as expected? Why or why not?
4. What is a scientific law, and how does it differ from a theory?
5. Why does a scientist include a control in the design of an experiment?
6. Why is there no single set of steps in the scientific method?
7. Describe what is needed for a hypothesis to develop into a theory.

CRITICAL THINKING

8. Explain the statement "No theory is written in stone."
9. Can a hypothesis that has been rejected be of any value to scientists? Why or why not?
10. How does the phrase "cause and effect" relate to the formation of a good hypothesis?
11. How would a control group be set up to test the effectiveness of a new drug in treating a disease?
12. Suppose you had to test how well two types of soap work. Describe your experiment by using the terms *control* and *variable.*
13. Why is a model made to be simpler than the thing that it represents?

SECTION

3 Measurements and Calculations in Chemistry

KEY TERMS

- **accuracy**
- **precision**
- **significant figure**

OBJECTIVES

1. **Distinguish** between accuracy and precision in measurements.
2. **Determine** the number of significant figures in a measurement, and apply rules for significant figures in calculations.
3. **Calculate** changes in energy using the equation for specific heat, and round the results to the correct number of significant figures.
4. **Write** very large and very small numbers in scientific notation.

Accuracy and Precision

When you determine some property of matter, such as density, you are making calculations that are often not the exact values. No value that is obtained from an experiment is exact because all measurements are subject to limits and errors. Human errors, method errors, and the limits of the instrument are a few examples. To reduce the impact of error on their work, scientists always repeat their measurements and calculations a number of times. If their results are not consistent, they will try to identify and eliminate the source of error. What scientists want in their results are *accuracy* and *precision*.

Measurements Must Involve the Right Equipment

Selecting the right piece of equipment to make your measurements is the first step to cutting down on errors in experimental results. For example, the beaker, the buret, and the graduated cylinder shown in **Figure 15** can be used to measure the volume of liquids. If an experimental procedure calls for measuring 8.6 mL of a liquid, which piece of glassware would you use? Obtaining a volume of liquid that is as close to 8.6 mL as possible is best done with the buret. In fact, the buret in **Figure 15** is calibrated to the nearest 0.1 mL.

Even though the buret can measure small intervals, it should not be used for all volume measurements. For example, an experimental procedure may call for using 98 mL of a liquid. In this case, a 100 mL graduated cylinder would be a better choice. An even larger graduated cylinder should be used if the procedure calls for 725 mL of a liquid.

The right equipment must also be selected when making measurements of other values. For example, if the experimental procedure calls for 0.5 g of a substance, using a balance that only measures to the nearest 1 g would introduce significant error.

Figure 15
All these pieces of equipment measure volume of liquids, but each is calibrated for different capacities.

Figure 16
a Darts within the bull's-eye mean high accuracy and high precision.

b Darts clustered within a small area but far from the bull's-eye mean low accuracy and high precision.

c Darts scattered around the target and far from the bull's-eye mean low accuracy and low precision.

Accuracy Is How Close a Measurement Is to the True Value

When scientists make and report measurements, one factor they consider is accuracy. The **accuracy** of a measurement is how close the measurement is to the true or actual value. To understand what accuracy is, imagine that you throw four darts separately at a dartboard.

accuracy
a description of how close a measurement is to the true value of the quantity measured

The bull's-eye of the dartboard represents the true value. The closer a dart comes to the bull's-eye, the more accurately it was thrown. **Figure 16a** shows one possible way the darts might land on the dartboard. Notice that all four darts have landed within the bull's-eye. This outcome represents high accuracy.

Accuracy should be considered whenever an experiment is done. Suppose the procedure for a chemical reaction calls for adding 36 mL of a solution. The experiment is done twice. The first time 35.8 mL is added, and the second time 37.2 mL is added. The first measurement was more accurate because 35.8 mL is closer to the true value of 36 mL.

Precision Is How Closely Several Measurements Agree

Another factor that scientists consider when making measurements is precision. **Precision** is the exactness of a measurement. It refers to how closely several measurements of the same quantity made in the same way agree with one another. Again, to understand how precision differs from accuracy, consider how darts might land on a dartboard.

precision
the exactness of a measurement

Figure 16b shows another way the four darts might land on the dartboard. Notice that all four darts have hit the target far from the bull's-eye. Because these darts are far from what is considered the true value, this outcome represents low accuracy. However, notice in **Figure 16b** that all four darts have landed very close to one another. The closer the darts land to one another, the more precisely they were thrown. Therefore, **Figure 16b** represents low accuracy but high precision. In **Figure 16c,** the four darts have landed far from the bull's-eye and each in a different spot. This outcome represents low accuracy and low precision.

Significant Figures

When you make measurements or perform calculations, the way you report a value tells about how you got it. For example, if you report the mass of a sample as 10 g, the mass of the sample may be between 8 g and 12 g or may be between 9.999 g and 10.001 g. However, if you report the mass of a sample as 10.0 g, you are indicating that you used a measuring tool that is precise to the nearest 0.1 g. The mass of the sample can only be between 9.95 g and 10.05 g.

significant figure

a prescribed decimal place that determines the amount of rounding off to be done based on the precision of the measurement

Scientists always report values using significant figures. The **significant figures** of a measurement or a calculation consist of all the digits known with certainty as well as one estimated, or uncertain, digit. Notice that the term *significant* does not mean "certain." The last digit or significant figure reported after a measurement is uncertain or estimated.

Significant Figures Are Essential to Reporting Results

Reporting all measurements in an experiment to the correct number of significant figures is necessary to be sure the results are true. Consider an experiment involving the transfer of energy as heat. Imagine that you use a thermometer calibrated in one-degree increments. Suppose you report a temperature as 37.5°C. The three digits in your reported value are all significant figures. The first two are known with certainty, but the last digit is estimated. You know the temperature is between 37°C and 38°C, and you estimate the temperature to be 37.5°C.

Now assume that you use the thermometer calibrated in one-tenth degree increments. If you report a reading of 36.54°C, the four digits in your reported value are all significant figures. The first three digits are known with certainty, while the last digit is estimated. Using this thermometer, you know the temperature is certainly between 36.5°C and 36.6°C, and estimate it to be 36.54°C.

Figure 17 shows two different thermometers. Notice that the thermometer on the left is calibrated in one-tenth degree increments, while the one on the right is calibrated in one-degree increments.

Figure 17
If the thermometer on the left is used, a reported value contains three certain figures, whereas the thermometer on the right can measure temperature to two certain figures.

Rules for Determining Significant Figures

1. Nonzero digits are always significant.
- For example, 46.3 m has three significant figures.
- For example, 6.295 g has four significant figures.

2. Zeros between nonzero digits are significant.
- For example, 40.7 L has three significant figures.
- For example, 87 009 km has five significant figures.

3. Zeros in front of nonzero digits are not significant.
- For example, 0.0095 87 m has four significant figures.
- For example, 0.0009 kg has one significant figure.

4. Zeros both at the end of a number and to the right of a decimal point are significant.
- For example, 85.00 g has four significant figures.
- For example, 9.070 000 000 cm has 10 significant figures.

5. Zeros both at the end of a number but to the left of a decimal point may not be significant. If a zero has not been measured or estimated, it is not significant. A decimal point placed after zeros indicates that the zeros are significant.
- For example, 2000 m may contain from one to four significant figures, depending on how many zeros are placeholders. For values given in this book, assume that 2000 m has one significant figure.

Calculators Do Not Identify Significant Figures

When you use a calculator to find a result, you must pay special attention to significant figures to make sure that your result is meaningful. The calculator in **Figure 18** was used to determine the density of isopropyl alcohol, commonly known as rubbing alcohol. The mass of a sample that has a volume of 32.4 mL was measured to be 25.42 g. Remember that the mass and volume of a sample can be used to calulate its density, as shown below.

$$D = \frac{m}{V}$$

The student in **Figure 18** is using a calculator to determine the density of the alcohol by dividing the mass (25.42 g) by the volume (32.4 mL). Notice that the calculator displays the density of the isopropyl alcohol as 0.7845679012 g/mL; the calculator was programmed so that all numbers are significant.

However, the volume was measured to only three significant figures, while the mass was measured to four significant figures. Based on the rules for determining significant figures in calculations described in **Skills Toolkit 1,** the density of the alcohol should be rounded to 0.785 g/mL, or three significant figures.

Figure 18
A calculator does not round the result to the correct number of significant figures.

2 SKILLS Toolkit

Rules for Using Significant Figures in Calculations

1. In multiplication and division problems, the answer cannot have more significant figures than there are in the measurement with the smallest number of significant figures. If a sequence of calculations is involved, do not round until the end.

$$\begin{array}{r} 12.257\ \text{m} \\ \times\ 1.162\ \text{m} \longleftarrow \text{four significant figures} \\ \hline 14.2426234\ \text{m}^2 \xrightarrow{\text{round off}} 14.24\ \text{m}^2 \end{array}$$

$$\frac{3.05\ \text{g}}{8.472\ \text{mL}} = 0.36000944\ \text{g/mL} \xrightarrow{\text{round off}} 0.360\ \text{g/mL} \quad \longleftarrow \text{three significant figures}$$

2. In addition and subtraction of numbers, the result can be no more certain than the least certain number in the calculation. So, an answer cannot have more digits to the right of the decimal point than there are in the measurement with the smallest number of digits to the right of the decimal. When adding and subtracting you should not be concerned with the total number of significant figures in the values. You should be concerned only with the number of significant figures present to the right of the decimal point.

$$\begin{array}{r} 3.95\ \text{g} \\ 2.879\ \text{g} \\ +\ 213.6\ \text{g} \\ \hline 220.429\ \text{g} \xrightarrow{\text{round off}} 220.4\ \text{g} \end{array}$$

Notice that the answer 220.4 g has four significant figures, whereas one of the values, 3.95 g, has only three significant figures.

3. If a calculation has both addition (or subtraction) and multiplication (or division), round after each operation.

Exact Values Have Unlimited Significant Figures

Some values that you will use in your calculations have no uncertainty. In other words, these values have an unlimited number of significant figures. One example of an exact value is known as a *count value*. As its name implies, a count value is determined by counting, not by measuring. For example, a water molecule contains exactly two hydrogen atoms and exactly one oxygen atom. Therefore, two water molecules contain exactly four hydrogen atoms and two oxygen atoms. There is no uncertainty in these values.

Another value that can have an unlimited number of significant figures is a *conversion factor*. There is no uncertainty in the values that make up this conversion factor, such as 1 m = 1000 mm, because a millimeter is defined as exactly one-thousandth of a meter.

You should ignore both count values and conversion factors when determining the number of significant figures in your calculated results.

SAMPLE PROBLEM A

Determining the Number of Significant Figures

A student heats 23.62 g of a solid and observes that its temperature increases from 21.6°C to 36.79°C. Calculate the temperature increase per gram of solid.

1 Gather information.

- The mass of the solid is 23.62 g.
- The initial temperature is 21.6°C.
- The final temperature is 36.79°C.

2 Plan your work.

- Calculate the increase in temperature by subtracting the initial temperature (21.6°C) from the final temperature (36.79°C).

$$\text{temperature increase} = \text{final temperature} - \text{initial temperature}$$

- Calculate the temperature increase per gram of solid by dividing the temperature increase by the mass of the solid (23.62 g).

$$\frac{\text{temperature increase}}{\text{gram}} = \frac{\text{temperature increase}}{\text{sample mass}}$$

3 Calculate.

$$36.79°\text{C} - 21.6°\text{C} = 15.19°\text{C} = 15.2°\text{C}$$

$$\frac{15.2°\text{C}}{23.62\ \text{g}} = 0.644\frac{°\text{C}}{\text{g}} \text{ rounded to three significant figures}$$

4 Verify your results.

- Multiplying the calculated answer by the total number of grams in the solid equals the calculated temperature increase.

$$0.644\frac{°\text{C}}{\text{g}} \times 23.62\ \text{g} = 15.2°\text{C}$$

rounded to three significant figures

PRACTICE HINT

Remember that the rules for determining the number of significant figures in multiplication and division problems are different from the rules for determining the number of significant figures in addition and subtraction problems.

PRACTICE

PROBLEM SOLVING SKILL

1 Perform the following calculations, and express the answers with the correct number of significant figures.

a. 0.1273 mL – 0.000008 mL

b. (12.4 cm × 7.943 cm) + 0.0064 cm^2

c. (246.83 g/26) – 1.349 g

2 A student measures the mass of a beaker filled with corn oil to be 215.6 g. The mass of the beaker is 110.4 g. Calculate the density of the corn oil if its volume is 114 cm^3.

3 A chemical reaction produces 653 550 kJ of energy as heat in 142.3 min. Calculate the rate of energy transfer in kilojoules per minute.

Specific Heat Depends on Various Factors

Recall that the specific heat is the quantity of energy that must be transferred as heat to raise the temperature of 1 g of a substance by 1 K. The quantity of energy transferred as heat during a temperature change depends on the nature of the material that is changing temperature, the mass of the material, and the size of the temperature change.

For example, consider how the nature of the material changing temperature affects the transfer of energy as heat. One gram of iron that is at 100.0°C is cooled to 50.0°C and transfers 22.5 J of energy to its surroundings. In contrast, 1 g of silver transfers only 11.8 J of energy as heat under the same conditions. Iron has a larger specific heat than silver. Therefore, more energy as heat can be transferred to the iron than to the silver.

Calculating the Specific Heat of a Substance

internet connect

www.scilinks.org
Topic: Specific Heat
SciLinks code: HW4119

SCILINKS Maintained by the National Science Teachers Association

Specific heats can be used to compare how different materials absorb energy as heat under the same conditions. For example, the specific heat of iron, which is listed in **Table 1,** is 0.449 J/g•K, while that of silver is 0.235 J/g•K. This difference indicates that a sample of iron absorbs and releases twice as much energy as heat as a comparable mass of silver during the same temperature change does.

Specific heat is usually measured under constant pressure conditions, as indicated by the subscript p in the symbol for specific heat, c_p. The specific heat of a substance at a given pressure is calculated by the following formula:

$$c_p = \frac{q}{m \times \Delta T}$$

In the above equation, c_p is the specific heat at a given pressure, q is the energy transferred as heat, m is the mass of the substance, and ΔT represents the difference between the initial and final temperatures.

Table 1 Some Specific Heats at Room Temperature

Element	Specific heat (J/g•K)	Element	Specific heat (J/g•K)
Aluminum	0.897	Lead	0.129
Cadmium	0.232	Neon	1.030
Calcium	0.647	Nickel	0.444
Carbon (graphite)	0.709	Platinum	0.133
Chromium	0.449	Silicon	0.705
Copper	0.385	Silver	0.235
Gold	0.129	Water	4.18
Iron	0.449	Zinc	0.388

SAMPLE PROBLEM B

Calculating Specific Heat

A 4.0 g sample of glass was heated from 274 K to 314 K and was found to absorb 32 J of energy as heat. Calculate the specific heat of this glass.

1 Gather information.

- sample mass (m) = 4.0 g
- initial temperature = 274 K
- final temperature = 314 K
- quantity of energy absorbed (q) = 32 J

2 Plan your work.

- Determine ΔT by calculating the difference between the initial and final temperatures.
- Insert the values into the equation for calculating specific heat.

$$c_p = \frac{32\ \text{J}}{4.0\ \text{g} \times (314\ \text{K} - 274\ \text{K})}$$

3 Calculate.

$$c_p = \frac{32\ \text{J}}{4.0\ \text{g} \times (40\ \text{K})} = 0.20\ \text{J/g}\cdot\text{K}$$

4 Verify your results.

The units combine correctly to give the specific heat in J/g·K. The answer is correctly given to two significant figures.

PRACTICE HINT

The equation for specific heat can be rearranged to solve for one of the quantities, if the others are known. For example, to calculate the quantity of energy absorbed or released, rearrange the equation to get $q = c_p \times m \times \Delta T$.

PRACTICE

1. Calculate the specific heat of a substance if a 35 g sample absorbs 48 J as the temperature is raised from 293 K to 313 K.
2. The temperature of a piece of copper with a mass of 95.4 g increases from 298.0 K to 321.1 K when the metal absorbs 849 J of energy as heat. What is the specific heat of copper?
3. If 980 kJ of energy as heat are transferred to 6.2 L of water at 291 K, what will the final temperature of the water be? The specific heat of water is 4.18 J/g·K. Assume that 1.0 mL of water equals 1.0 g of water.
4. How much energy as heat must be transferred to raise the temperature of a 55 g sample of aluminum from 22.4°C to 94.6°C? The specific heat of aluminum is 0.897 J/g·K. Note that a temperature change of 1°C is the same as a temperature change of 1 K because the sizes of the degree divisions on both scales are equal.

PROBLEM SOLVING SKILL

Scientific Notation

Chemists often make measurements and perform calculations using very large or very small numbers. Very large and very small numbers are often written in *scientific notation.* To write a number in scientific notation, first know that every number expressed in scientific notation has two parts. The first part is a number that is between 1 and 10 but that has any number of digits after the decimal point. The second part consists of a power of 10. To write the first part of the number, move the decimal to the right or the left so that only one nonzero digit is to the left of the decimal. Write the second part of the value as an exponent. This part is determined by counting the number of decimal places the decimal point is moved. If the decimal is moved to the right, the exponent is negative. If the decimal is moved to the left, the exponent is positive. For example, 299 800 000 m/s is expressed as 2.998×10^8 m/s in scientific notation. When writing very large and very small numbers in scientific notation, use the correct number of significant figures.

3 SKILLS Toolkit

Using Scientific Notation

1. In scientific notation, exponents are count values.

2. In addition and subtraction problems, all values must have the same exponent before they can be added or subtracted.

- $6.2 \times 10^4 + 7.2 \times 10^3 = 62 \times 10^3 + 7.2 \times 10^3 = 69.2 \times 10^3 = 69 \times 10^3 = 6.9 \times 10^4$
- $4.5 \times 10^6 - 2.3 \times 10^5 = 45 \times 10^5 - 2.3 \times 10^5 = 42.7 \times 10^5 = 43 \times 10^5 = 4.3 \times 10^6$

3. In multiplication problems, the first factors of the numbers are multiplied and the exponents of 10 are added.

- $(3.1 \times 10^3)(5.01 \times 10^4) = (3.1 \times 5.01) \times 10^{4+3} = 16 \times 10^7 = 1.6 \times 10^8$

4. In division problems, the first factors of the numbers are divided and the exponent of 10 in the denominator is subtracted from the exponent of 10 in the numerator.

- $7.63 \times 10^3/8.6203 \times 10^4 = 7.63/8.6203 \times 10^{3-4} = 0.885 \times 10^{-1} = 8.85 \times 10^{-2}$

SKILLS Toolkit 4

Scientific Notation with Significant Figures

1. Use scientific notation to eliminate all placeholding zeros.

- $2400 \longrightarrow 2.4 \times 10^3$ (both zeros are not significant)
- $750\ 000. \longrightarrow 7.50000 \times 10^5$ (all zeros are significant)

2. Move the decimal in an answer so that only one digit is to the left, and change the exponent accordingly. The final value must contain the correct number of significant figures.

- $5.44 \times 10^7/8.1 \times 10^4 = 5.44/8.1 \times 10^{7-4} = 0.6716049383 \times 10^3 =$ 6.7×10^2 (adjusted to two significant figures)

3 Section Review

UNDERSTANDING KEY IDEAS

1. How does accuracy differ from precision?
2. Explain the advantage of using scientific notation.
3. When are zeros significant in a value?
4. Why are significant figures important when reporting measurements?
5. Explain how a series of measurements can be precise without being accurate.

PRACTICE PROBLEMS

6. Perform the following calculations, and express the answers using the correct number of significant figures.

   a. $0.8102 \text{ m} \times 3.44 \text{ m}$

   b. $\dfrac{94.20 \text{ g}}{3.167\ 22 \text{ mL}}$

   c. $32.89 \text{ g} + 14.21 \text{ g}$

   d. $34.09 \text{ L} - 1.230 \text{ L}$

7. Calculate the specific heat of a substance when 63 J of energy are transferred as heat to an 8.0 g sample to raise its temperature from 314 K to 340 K.
8. Express the following calculations in the proper number of significant figures. Use scientific notation where appropriate.

   a. 129 g/29.2 mL

   b. (1.551 mm)(3.260 mm)(4.9001 mm)

   c. 35 000 kJ/0.250 s

9. A clock gains 0.020 s/min. How many seconds will the clock gain in exactly six months, assuming 30 days are in each month? Express your answer in scientific notation.

CRITICAL THINKING

10. There are 12 eggs in a carton. How many significant figures does the value 12 have in this case?
11. If you measure the mass of a liquid as 11.50 g and its volume as 9.03 mL, how many significant figures should its density value have? Explain the reason for your answer.

2
He
Helium
4.002 602
$1s^2$

HELIUM

Element Spotlight

Where Is He?

Universe:
about 23% by mass

Earth's crust:
0.000001% by mass

Air:
0.0005% by mass

In Florida, divers on the Wakulla Springs project team breathed heliox at depths greater than 90 m.

Deep-sea Diving with Helium

Divers who breathe air while at great undersea depths run the risk of suffering from a condition known as nitrogen narcosis. Nitrogen narcosis can cause a diver to become disoriented and to exercise poor judgment, which leads to dangerous behavior. To avoid nitrogen narcosis, professional divers who work at depths of more than 60 m breathe heliox, a mixture of helium and oxygen, instead of air.

The greatest advantage of heliox is that it does not cause nitrogen narcosis. A disadvantage of heliox is that it removes body heat faster than air does. This effect makes a diver breathing heliox feel chilled sooner than a diver breathing air.

Breathing heliox also affects the voice. Helium is much less dense than nitrogen, so vocal cords vibrate faster in a heliox atmosphere. This raises the pitch of the diver's voice, and makes the diver's voice sound funny. Fortunately, this effect disappears when the diver surfaces and begins breathing air again.

Industrial Uses

- Helium is used as a lifting gas in balloons and dirigibles.
- Helium is used as an inert atmosphere for welding and for growing high-purity silicon crystals for semiconducting devices.
- Liquid helium is used as a coolant in superconductor research.

Real-World Connection Helium was discovered in the sun before it was found on Earth.

internet connect
www.scilinks.org
Topic: Helium
SciLinks code: HW4171
SCILINKS. Maintained by the National Science Teachers Association

A Brief History

1600 | 1700 | 1800 | 1900

1868: Pierre Janssen, studies the spectra of a solar eclipse and finds evidence of a new element. Edward Frankland, an English chemist, and Joseph Lockyer, an English astronomer, suggest the name helium.

1888: William Hillebrand discovers that an inert gas is produced when a uranium mineral is dissolved in sulfuric acid.

1894: Sir William Ramsay and Lord Rayleigh discover argon. They suspect that the gas Hillebrand found in 1888 was argon. They repeat his experiment and find that the gas is helium.

1908: Ernest Rutherford and Thomas Royds prove that alpha particles emitted during radioactive decay are helium nuclei.

Questions

1. Research the industrial, chemical, and commercial uses of helium.
2. Research properties of neon, argon, krypton, and xenon. How are these gases similar to helium? Are they used in a manner similar to helium?

CHAPTER HIGHLIGHTS 2

KEY TERMS

energy
physical change
chemical change
evaporation
endothermic
exothermic
law of conservation of energy
heat
kinetic energy
temperature
specific heat

scientific method
hypothesis
theory
law
law of conservation of mass

accuracy
precision
significant figure

KEY IDEAS

SECTION ONE Energy

- Energy is the capacity to do work.
- Changes in matter can be chemical or physical. However, only chemical changes produce new substances.
- Every change in matter involves a change in energy.
- Endothermic processes absorb energy. Exothermic processes release energy.
- Energy is always conserved.
- Heat is the energy transferred between objects that are at different temperatures. Temperature is a measure of the average random kinetic energy of the particles in an object.
- Specific heat is the relationship between energy transferred as heat to a substance and a substance's temperature change.

SECTION TWO Studying Matter and Energy

- The scientific method is a strategy for conducting research.
- A hypothesis is an explanation that is based on observations and that can be tested.
- A variable is a factor that can affect an experiment.
- A controlled experiment is an experiment in which variables are kept constant.
- A theory is a well-tested explanation of observations. A law is a statement or mathematical expression that describes the behavior of the world.

SECTION THREE Measurements and Calculations in Chemistry

- Accuracy is the extent to which a measurement approaches the true value of a quantity.
- Precision refers to how closely several measurements that are of the same quantity and that are made in the same way agree with one another.
- Significant figures are digits known with certainty as well as one estimated, or uncertain, digit.
- Numbers should be written in scientific notation.

KEY SKILLS

Rules for Determining Significant Figures
Skills Toolkit 1 p. 57

Rules for Using Significant Figures in Calculations
Skills Toolkit 2 p. 58
Sample Problem A p. 59

Calculating Specific Heat
Sample Problem B p. 61

Scientific Notation in Calculations
Skills Toolkit 3 p. 62

Scientific Notation with Significant Figures
Skills Toolkit 4 p. 63

2

Physical Setting/Chemistry
REGENTS EXAM PRACTICE

PART A

For each item, write on a separate piece of paper the number of the word, expression, or statement that best answers the item.

1. Temperature is defined as the
(1) quantity of heat present in a substance.
(2) average kinetic energy of the particles in a substance.
(3) energy that is transferred as heat between two substances.
(4) specific heat of a substance.

2. A transfer of energy is involved in
(1) physical changes only.
(2) chemical changes only.
(3) both physical changes and chemical changes.
(4) neither physical changes nor chemical changes.

3. Any change in matter in which energy is released
(1) is known as an exothermic process.
(2) is known as an endothermic process.
(3) does not illustrate the law of conservation of energy.
(4) occurs when a solid changes phase to become a liquid.

4. Which of the following always applies when energy is transferred as heat?
(1) The temperature changes.
(2) A chemical reaction occurs.
(3) The specific heat value of a substance increases.
(4) Energy is conserved.

5. Which of the following is not a form of energy?
(1) heat **(3)** chemical
(2) temperature **(4)** light

6. Which values represent the same temperature measurements?
(1) 110°C and 373 K
(2) –20°C and 253 K
(3) 0°C and 0 K
(4) 100°C and 100 K

7. Identify an example of an endothermic physical change.
(1) an explosion
(2) the formation of a solid when two liquids are mixed
(3) the melting of butter
(4) the condensation of a gas

8. The equation $E = mc^2$ shows that
(1) chemical reactions are either exothermic or endothermic.
(2) energy is not always involved when a chemical change occurs.
(3) matter and energy are related.
(4) exothermic reactions release energy.

9. Which transfer of energy is least likely to occur?
(1) from a system to its surroundings
(2) from a cooler object to a warmer object
(3) from a solid to a liquid
(4) from a gas to a liquid

Regents Test-Taking Tip

Carefully read questions that ask for the one choice that is not correct. Often these questions include words such as *is not, except,* or *all but*.

10. Every chemical change involves the
(1) formation of a different substance.
(2) vaporization of a liquid.
(3) separation of states of matter.
(4) release of energy.

11. A theory differs from a hypothesis in that a theory
(1) cannot be disproved.
(2) always leads to the formation of a law.
(3) has been subjected to experimental testing.
(4) describes past events.

12. Which of the following is not part of the scientific method?
(1) making measurements
(2) introducing bias
(3) proposing a hypothesis
(4) analyzing data

13. A control in an experiment
(1) is not needed when the scientist is testing a hypothesis.
(2) means that the scientist has everything under control.
(3) is required only if the hypothesis leads to the development of a theory.
(4) allows the scientist to identify the cause of the results in an experiment.

14. A variable in an experiment is
(1) a factor that can affect the results.
(2) the same as a hypothesis.
(3) needed only when a new procedure is being tried.
(4) never allowed to affect the outcome.

15. Which is characteristic of the scientific method?
(1) a set of steps that most scientists follow
(2) verifying that the hypothesis is valid
(3) a set of processes used to test a hypothesis
(4) a series of experiments where several variables are tested at the same time

16. Which of the following measurements contains three significant figures?
(1) 200 mL
(2) 0.20 mL
(3) 20.2 mL
(4) 200.0 mL

17. All measurements in science
(1) are expressed in scientific notation.
(2) include some degree of uncertainty.
(3) are both accurate and precise.
(4) include only those digits that are known with certainty.

18. The accuracy of a measurement
(1) is how close it is to the true value.
(2) does not depend on the instrument being used to measure the object.
(3) indicates that the measurement is also precise.
(4) is something that scientists rarely achieve.

19. A measurement of 23 456 mg converted to grams equals
(1) 2.3456 g.
(2) 23.456 g.
(3) 23.5 g
(4) 0.23456 g

20. Which of the following statements is true about significant figures?
(1) Zeros between nonzero digits are not significant.
(2) Zeros in front of nonzero digits are significant.
(3) Zeros both at the end of a number and to the right of a decimal point are significant.
(4) Nonzero digits are significant only if they are to the left of a decimal point.

PART B–1

For each item, write on a separate piece of paper the number of the word, expression, or statement that best answers the item.

21. Beaker A contains water at a temperature of 0°C. Beaker B contains the same volume of water at a temperature of 37°C. Which of the following statements is true about these beakers?

(1) The water molecules in beaker A have a greater average kinetic energy than those in beaker B.
(2) The water molecules in beaker A have a lower average kinetic energy than those in beaker B.
(3) The water molecules in both beakers have the same average kinetic energy.
(4) The water molecules in beaker A have no kinetic energy.

22. Which of the following statements is true about temperature values?

(1) The zero point on the Celsius scale is designated as absolute zero.
(2) The zero point on the Kelvin scale is designated as the freezing point of water.
(3) Values on the Celsius scale cannot be converted to values on the Fahrenheit scale.
(4) The zero point on the Kelvin scale is designated as absolute zero.

23. Which of the following statements can be considered a hypothesis?

(1) Sugar dissolves faster in warm water than it does in cold water.
(2) During a chemical reaction, energy is conserved.
(3) If I pay attention in class, I will succeed in this course.
(4) Matter and energy are related.

24. Which of the following statements can be classified as a law?

(1) The substance is silvery white, is fairly hard, and is a good conductor of electricity.
(2) Bases are substances that feel slippery in water.
(3) A system containing many particles will not change spontaneously from a disordered state to an ordered state.
(4) If salt is dissolved in water, the solution will conduct electricity.

25. In order to be valid, a hypothesis must be

(1) proven to be correct by the results of the experiment.
(2) testable.
(3) based on observations made in a laboratory or in the field.
(4) derived from unexpected observations made during an experiment.

26. What is a problem with having more than one variable in an experiment?

(1) The factor responsible for the results cannot be identified.
(2) A scientist will have trouble keeping track of everything.
(3) The results may take too long to obtain in order to be meaningful.
(4) Another scientist would not be able to follow the procedure and duplicate the experiment.

27. Which of the following statements contains an exact number?

(1) The circumference of Earth at the equator is 40000 km.
(2) The tank was filled with 54 L of gas.
(3) A nickel has a mass of 5 g.
(4) The accident injured 21 people.

PART B–2

Answer the following items.

28. How does Einstein's equation $E = mc^2$ seem to contradict both the law of conservation of energy and the law of conservation of mass?

29. Explain the relationship between models and theories.

30. Perform the following calculations and express your answers in the correct number of significant figures.
a. $(172.56/43.8) - 1.825$
b. $(9.8 \times 8.934) + 0.0048$
c. $1.36 \times 10^{-5} \times 5.02 \times 10^{-2}$

31. Calculate the specific heat of a substance if a 35 g sample absorbs 48 J as it is heated from 298 K to 313K. Be sure to include the correct units in your answer.

32. How much energy is needed to raise the temperature of 75 g of gold by 25°C? Refer to Table 1 on page 60.

33. Perform the following calculations and express the answers in scientific notation. Be sure to include units for each answer.
a. 37 000 000 m × 7 100 000 m
b. 0.000 312 cm / 486 cm
c. 4.6×10^4 cm × 7.5×10^3 cm
d. 8.3×10^6 kg / 2.5×10^9 cm^3

34. A 35 g sample of water is heated so that 420 J of energy is added. The initial temperature of the water sample is 10.0°C. Calculate the final temperature of the water. Refer to Table 1 on page 60.

35. Assume that students are testing the effects of five garden fertilizers by applying some of each fertilizer to five separate rows of radishes.
a. What variable is being tested?
b. Identify two factors that they should control.
c. Propose one set of observations that allows the students to conclude which garden fertilizer is best for growing radishes.

PART C

Answer the following items.

Items 36–38 are based on the following passage.

Secrets of the Cremona Violins

What are the most beautiful sounding of all violins? Most professionals will pick the instruments created in Cremona, Italy, following the Renaissance. At that time, Antonio Stradivari and other designers created instruments of extraordinary sound. The craftsmen were notoriously secretive about their techniques, but, based on 20 years of research, Dr. Joseph Nagyvary, a professor of biochemistry at Texas A & M University thinks he has discovered the key to the violins' sounds hidden in the chemistry of their materials.

According to Dr. Nagyvary, Stradivarius instruments are nearly free of the shrill, high-pitched noises produced by modern violin makers. Generally, violin makers attribute this to the design of the instrument, but Dr. Nagyvary traces it to a different source. In Stradivari's day, wood for the violin was transported by floating it down a river from the mountains in Venice, where it was stored in sea water. Dr. Nagyvary thought that the soaking process could have removed ingredients from the wood that made it inherently noisy.

Dr, Nagyvary found other clues to the sound of the violins in the work of the Renaissance alchemists. Aside from trying to turn lead into gold, alchemists made many useful experiments, including investigating different chemical means of preserving wood in musical instruments. Attempting to duplicate their techniques and to reproduce the effects of sea water, Dr. Nagyvary soaks all his wood in a "secret" solution. One of his favorite ingredients is a cherry-and-plum puree, which contains an enzyme called pectinase. The pectinase softens the wood, making it resonate more freely.

"The other key factor in a violin's sound," says Dr. Nagyvary, "is the finish, which is the filler and the varnish covering the instrument. Most modern finishes are made from rubbery materials, which limit the vibrations of the wood." Modern analysis has revealed the Cremona finish was different: it was a brittle mineral microcomposite of a very sophisticated nature. According to historical accounts, all violin makers, including Stradivari, procured their varnishes from the local drugstore chemist, and they didn't even know what they were using! Dr. Nagyvary thinks this unknown and unsung drugstore chemist could have been the major factor behind the masterpieces for which violin makes like Stradivari received exclusive credit.

Dr. Nagyvary and his coworkers have identified most of the key ingredients of the Cremona finish. Many new violins made from treated wood and replicated finish have been made, and their sounds have been analyzed by modern signal analyzers. These violins have been favorably compared with authentic Stradivari violins. However, some violin makers remain skeptical, claiming that it takes many years to reveal just how good a violin is. In the meantime, most everyone agrees that the art and science of violin making are still epitomized by the instruments of Cremona.

36. Identify the two variables that Dr. Nagyvary thinks might be responsible for the sounds produced by Cremona violins.

37. Dr. Nagyvary's work is an example of what is called "reverse engineering." What is meant by this term?

38. Why do some people refuse to accept the findings of Dr, Nagyvary's experiments?

FOCUS ON GRAPHING

Study the graph below, and answer the questions that follow.
For help in interpreting graphs, see Appendix B, "Study Skills for Chemistry."

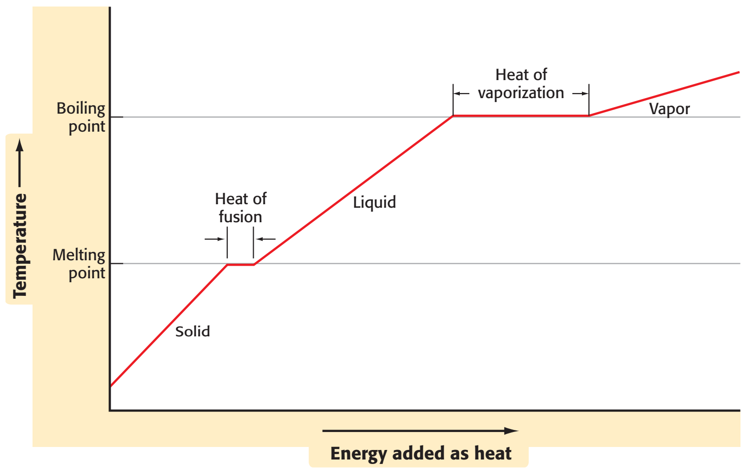
39. What is the value for the slope of the curve during the period in which the temperature is equal to the melting point temperature?

40. Is there another period in the graph where the slope equals the value in question 39?

41. Draw the cooling curve for water. Label the axes and the graph.

42. Suppose water could exist in four states of matter at some pressure. Draw what the heating curve for water would look like. Label the axes and the graph.

Heating Curve for H_2O



TECHNOLOGY AND LEARNING

43. Graphing Calculator

Graphing Celsius and Fahrenheit Temperatures

The graphing calculator can run a program that makes a graph of a given Fahrenheit temperature (on the x-axis) and the corresponding Celsius temperature (on the y-axis). You can use the TRACE button on the calculator to explore this graph and learn more about how the two temperature scales are related.

Go to Appendix C. If you are using a TI-83 Plus, you can download the program CELSIUS and run the application as directed. If you are using another calculator, your teacher will provide you with keystrokes and data sets to use. After the graph is displayed, press TRACE. An X-shaped cursor on the graph line indicates a specific point. At the bottom of the screen the values are shown for that point. The one labeled X= is the Fahrenheit temperature and the one labeled Y= is the Celsius temperature. Use the right and left arrow keys to move the cursor along the graph line to find the answers to these questions.

a. What is the Fahrenheit temperature when the Celsius temperature is zero? (This is where the graph line crosses the horizontal x-axis.) What is the significance of this temperature?

b. Human internal body temperature averages 98.6°F. What is the corresponding value on the Celsius scale?

c. Determine the Fahrenheit temperature in your classroom or outside, as given in a weather report. What is the corresponding Celsius temperature?

d. At what temperature are the Celsius and Fahrenheit temperatures the same?