## Introduction

Chemistry is defined as the study of matter and its properties. Matter is defined as everything that has mass and occupies space. Although these definitions are acceptable, they do not explain why one needs to know chemistry. The answer to that query is that the world in which we live is a chemical world. Your own body is a complex chemical factory that uses chemical processes to change the food you eat and the air you breathe into bones, muscle, blood, and tissue and even into the energy that you use in your daily living. When illness prevents some part of these processes from functioning correctly, the doctor may prescribe as a medicine a chemical compound, either isolated from nature or prepared in a chemical laboratory by a chemist.

The world around us is also a vast chemical laboratory. The daily news is filled with reports of acid rain, toxic wastes, the risks associated with nuclear power plants, and the derailment of trains carrying substances such as vinyl chloride, sulfuric acid, and ammonia. Not all chemical news is of disasters. The daily news also carries stories (often in smaller headlines) of new drugs that cure old diseases; of fertilizers, insecticides, and herbicides designed by chemists to allow the farmers to feed our growing populations, and of other new products to make our lives more pleasant. The packages we buy at the grocery store list their contents, including what chemicals the package contains, such as preservatives, and the nutritional content in terms of vitamins, minerals, fats, carbohydrates, and proteins.

Everyday life is besieged with chemicals. In beginning the study of chemistry, it is unwise to start with topics as complex as the latest miracle drugs. We will begin with the composition of matter and the different kinds of matter. We can then talk about the properties of the different types of matter and the changes that each can undergo. You will learn that each of these changes is accompanied by an energy change and learn the significance of these energy changes.

## The Kinds of Matter

### Pure Substances

A pure substance consists of a single kind of matter. It always has the same composition and the same set of properties. For example, baking soda is a single kind of matter, known chemically as sodium hydrogen carbonate. A sample of pure baking soda, regardless of its source or size, will be a white solid containing 57.1% sodium, 1.2% hydrogen, 14.3% carbon, and 27.4% oxygen. The sample will dissolve in water. When heated to 270°C the sample will decompose, giving off carbon dioxide and water vapor and leaving a residue of sodium carbonate. Thus, by definition, baking soda is a pure substance because it has a constant composition and a unique set of properties, some of which we have listed. The properties we have described hold true for any sample of baking soda. These properties are the kinds in which we are interested.

A note about the term pure; in this text, the word pure means a single substance, not a mixture of substances. As used by the U.S. Food and Drug Administration (USFDA), the term pure means "fit for human consumption." Milk, whether whole, 2% fat, or skim, may be pure (fit for human consumption) by public health standards, but it is not pure in the chemical sense. Milk is a mixture of a great many substances, including water, butterfat, proteins, and sugars.

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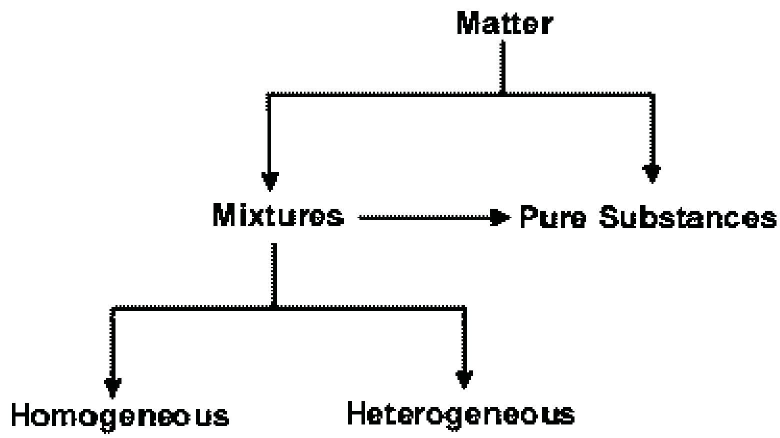
**Mixtures**

A mixture consists of two or more pure substances. Most of the matter we see around us is composed of mixtures. Seawater contains dissolved salts; river water contains suspended mud; hard water contains salts of calcium, magnesium, and iron. Both seawater and river water also contain dissolved oxygen, without which fish and other aquatic life could not survive.

Unlike the constant composition of a simple substance, the composition of a mixture can be changed. The properties of the mixture depend on the percentage of each pure substance in it. Steel is an example of a mixture. All steel starts with the pure substance iron. Refiners then add varying percentages of carbon, nickel, chromium, vanadium, or other substances to obtain steels of a desired hardness, tensile strength, corrosion resistance, and so on. The properties of a particular type of steel depend not only on which substances are mixed with the iron but also on the relative percentage of each. One type of chromium-nickel steel contains 0.6% chromium and 1.25% nickel. Its surface is easily hardened, a property that makes it valuable in the manufacture of automobile gears, pistons, and transmissions. The stainless steel used in the manufacture of surgical instruments, food-processing equipment, and kitchenware is also a mixture of iron, chromium, and nickel; it contains 18% chromium and 8% nickel. Steel with this composition can be polished to a very smooth surface and is very resistant to rusting.

You can often tell from the appearance of a sample whether it is a mixture. For example, if river water is clouded with mud or silt particles, you know it is a mixture. If a layer of brown haze lies over a city, you know the atmosphere is mixed with pollutants. However, the appearance of a sample is not always sufficient evidence by which to judge its composition. A sample of matter may look pure without being so. For instance, air looks like a pure substance but it is actually a mixture of oxygen, nitrogen, and other gases.

Rubbing alcohol is a clear, colorless liquid that looks pure but is actually a mixture of isopropyl alcohol and water, both of which are clear, colorless liquids. As another example, you cannot look at a piece of metal and know whether it is pure iron or a mixture of iron with some other substance such as chromium or nickel. Figure 1.2 shows the relationships between different kinds of matter.




**The Properties of Matter**

Each kind of matter possesses a number of properties by which it can be identified. In Section 1.2A , we listed some of the properties by which the pure substance baking soda can be identified. These properties fall into two large categories (1) physical properties, those that can be observed without changing the composition of the sample, and (2) chemical properties, those whose observation involves a change in composition.

Baking soda dissolves readily in water. If water is evaporated from a solution of baking soda, the baking soda is recovered unchanged; thus solubility is a physical property. The decomposition of baking soda on heating is a chemical property. You can observe the decomposition of baking soda, but, after you make this observation, you no longer have baking soda. Instead you have carbon dioxide, water, and sodium carbonate. A physical change alters only physical properties, such as size and shape.

This discussion of properties points to another difference between pure substances and mixtures. A mixture can be separated into its components by differences in their physical properties. A mixture of salt and sand can be separated because salt dissolves in water but sand does not. If we add water to a salt-sand mixture, the salt will dissolve, leaving the sand at the bottom of the container. If we pour off the water, the sand will remain. If we boil off the water from the salt solution, we will get the salt by itself. We have separated the two components of the mixture by a difference in their ability to dissolve in water. Solubility is a physical property. Pure substances, on the other hand, can be separated into their components only by chemical changes. When added to water, the pure substance sodium bicarbonate does not separate into sodium, hydrogen, carbon, and oxygen, although these components of sodium bicarbonate differ greatly in their solubilities in water.

One of the important physical properties of a substance is its physical state at room temperature. The three physical states of matter are solid, liquid and gas. Most kinds of matter can exist in all three states. You are familiar with water as a solid (ice), a liquid, and a gas (steam). You have seen wax as a solid at room temperature and a liquid when heated. You have probably seen carbon dioxide as a solid (dry ice) and been aware of it as a colorless gas at higher temperatures. The temperatures at which a given kind of matter changes from a solid to a liquid (its melting point) or from a liquid to a gas (its boiling point) are physical properties. For example, the melting point of ice (0°C) and the boiling point of water (100°C) are physical properties of the substance water.

Like pure substances, mixtures can exist in the three physical states of solid, liquid and gas. Air is a gaseous mixture of approximately 78% nitrogen, 21% oxygen, and varying percentages of several other gases. Rubbing alcohol is a liquid mixture of approximately 70% isopropyl alcohol and 30% water. Steel is a solid mixture of iron and other pure substances.

**The Law of Conservation of Mass**

The Law of Conservation of Mass states that matter can be changed from one form into another, mixtures can be separated or made, and pure substances can be decomposed, but the total amount of mass remains constant. We can state this important law in another way. The total mass of the universe is constant within measurable limits; whenever matter undergoes a change, the total mass of the products of the change is, within measurable limits, the same as the total mass of the reactants.

The formulation of this law near the end of the eighteenth century marked the beginning of modern chemistry. By that time many elements had been isolated and identified, most notably oxygen, nitrogen, and hydrogen. It was also known that, when a pure metal was heated in air, it became what was then called a calx (which we now call an oxide) and that this change was accompanied by an increase in mass. The reverse of this reaction was also known: Many calxes on heating lost mass and returned to pure metals. Many imaginative explanations of these mass changes were proposed. Antoine Lavoisier (1743-1794), a French nobleman later guillotined in the revolution, was an amateur chemist with a remarkably analytical mind. He considered the properties of metals and then carried out a series of experiments designed to allow him to measure not just the mass of the metal and the calx but also the mass of the air surrounding the reaction. His results showed that the mass gained by the metal in forming the calx was equal to the mass lost by the surrounding air.

With this simple experiment, in which accurate measurement was critical to the correct interpretation of the results, Lavoisier established the Law of Conservation of Mass, and chemistry became an exact science, one based on careful measurement. For his pioneering work in the establishment of that law and his analytical approach to experimentation, Lavoisier has been called the father of modern chemistry.

**The Law of Conservation of Energy**

A study of the properties of matter must include a study of energy. Energy, defined as the capacity to do work, has many forms. Potential energy is stored energy; it may be due to composition (the composition of a battery determines the energy it can release), to position (a rock at the top of a cliff will release energy if it falls to lower ground), or to condition (a hot stone will release heat energy if it is moved to a cooler place). Kinetic energy is energy of motion. You are undoubtedly aware that the faster a car is moving, the more damage it does in crashing into an object. Because it is moving faster, it has more kinetic energy and has a greater capacity to do work (in this case, damage).

One of the characteristics of energy is that one form of energy can be converted to another. When wood is burned, some of its potential energy is changed to radiant energy (heat and light). Some is changed to kinetic energy as the water and carbon dioxide formed move away from the burning log. Some remains as potential energy in the composition of the water and carbon dioxide produced by the burning. Throughout all these changes, the total amount of energy remains constant. All changes must obey the Law of Conservation of Energy, which states that energy can neither be created nor destroyed. An alternative statement is that the total amount of energy in the universe remains constant.

The Law of Conservation of Mass and the Law of Conservation of Energy are interrelated principles. Mass can be changed into energy and energy into mass according to the equation:

This relationship allows us to state the two laws as a single law, called the Law of Conservation of Mass/Energy: Energy and mass can be interconverted, but together they are conserved. This law was first stated by Albert Einstein (1879-1955). In most changes, the amount of matter converted to energy is much too small to be detected by even the most sensitive apparatus, and we can say "in this change both mass and energy are separately conserved." It is nevertheless important to be aware of this relationship between mass and energy, because nuclear energy is obtained by just such a conversion of mass to energy.

$$E = mc^2$$

where E = energy change

m = mass change (in grams)

c = speed of light ($3.00 \times 10^8$ m/sec, or 186,000 mi/sec)

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**Systems of Measurement**

Measurements in the scientific world, and increasingly, in the nonscientific world are made in SI

(Systéme International) units. The system was established in order to allow comparison of measurements made in one country with those made in another. SI units and their relative values were adopted by an international association of scientists meeting in Paris in 1960. Table 2.1 lists the basic SI units and derived units. Notice that metric units are part of this system. The system still in common, nonscientific use in the United States is called the English system, even though England, like most other developed countries, now uses metric units. Anyone using units from both the English and SI systems needs to be aware of a few simple relationships between the two systems.

| Property being measured | Basic SI Unit | Derived units | Relationship to English Unit |
|---|---|---|---|
| length | meter (m) | kilometer (km)<br>1 km = 1000 m<br>centimeter (cm)<br>1 cm = 0.01 m | 1 m = 39.37 in<br>1.61 km = 1 mi<br>2.54 cm = 1 in |
| mass | kilogram (kg) | gram (g)<br>1 g = 0.001 kg | 1 kg = 2.204 lb<br>453.6 g = 1 lb |
| volume | cubic meter ($m^3$) | liter (L)<br>1 L = 0.001 $m^3$<br>cubic centimeter ($cm^3$ or cc)<br>1 $cm^3$ = 0.001 L<br>milliliter (mL)<br>1 mL = 1 $cm^3$ | 1 L = 1.057 qt<br>946 mL = 1.0 qt |
| temperature | Kelvin (K) | Celsius (C)<br>K = °C + 273.15 | Fahrenheit (F)<br>$°C = \frac{°F - 32}{1.8}$<br>$= \frac{5}{9}(°F - 32)$ |
| energy | Joule (J) | calorie (cal)<br>1 cal = 4.184 J<br>kilocalorie (kcal) | |

Two features of the SI system make it easy to use. First, it is a base-10 system; that is, the various units of a particular dimension vary by multiples of ten. Once a base unit is defined, units larger and smaller than the base unit are indicated by prefixes added to the name of the base unit. Table 2.2 (next page) lists some of these SI prefixes,along with the abbreviation for each and the numerical factor relating it to the base unit.

| Prefix | Symbol | Base unit multiplied by |
|---|---|---|
| mega- | M | 1,000,000, or $10^6$* |
| kilo- | k | 1,000, or $10^3$ |
| deci- | d | 0.1, or $10^{-1}$ |
| centi- | c | 0.01, or $10^{-2}$ |
| milli- | m | 0.001, or $10^{-3}$ |
| micro- | μ | 0.000001, or $10^{-6}$ |
| nano- | n | 0.000000001, or $10^{-9}$ |
| pico- | p | 0.000000000001, or $10^{-12}$ |

The second feature that increases the usefulness of the SI system is the direct relationship between the base units of different dimensions. For example, the unit of volume (cubic meter) is the cube of the unit of length (meter). We shall see later how the unit of mass is related to the unit of volume.

The base unit of length in the SI system is the meter (m). The meter, approximately 10% longer than a yard, is equivalent to 39.37 inches, or 1.094 yards.

The base unit of volume in the SI system is the cubic meter (m3). Other commonly used units of volume are the liter (L), the cubic centimeter (cm3 or cc), and the milliliter (mL).

One liter has a volume equal to 0.001 m3. The nearest unit of comparable volume in the English system is the quart (1.000 L = 1.057 qt). Note particularly that the volume of 1 cm3 is the same as the volume of 1 mL.

The standard of mass in the SI system is the kilogram (kg). A safe in Sèvres, France, holds a metal cylinder with a mass of exactly 1 kg. The mass of that cylinder is the same as the mass of 1000 mL (1 L) of water at 4°C, thereby relating mass to volume. Notice that the base unit of mass in the SI system is the gram (g), even though the standard of mass in this system is the kilogram.

**Conversion Factors**

Measurements made during a chemical experiment are often used to calculate another property. Frequently it is necessary to change measurements from one unit to another - inches to feet, meters to centimeters, or hours to seconds. A relationship between two units that measure the same quantity is a conversion factor. For example, the conversion factor between feet and yards is:

3 ft = 1 yd

A conversion factor relates two measurements of the same sample. The measurements may be of the same property (in 3 ft = 1 yd, both measurements are of length) or of different properties of the same sample. In saying that 3 mL alcohol weigh 2.4 g, we are considering two different properties of the same sample - volume and mass. Together these measurements express a conversion factor, for they refer to the same sample and show a relationship between its volume and its mass.

Conversion factors are so named because they offer a way of converting a measurement made in one dimension to another dimension. They do not change the original property, only how it is measured. Table 2.1 listed many conversion factors within the metric system and between the metric and English systems.

Conversion factors that define relationships, such as 3 ft = 1 yd or 1 L = 1000 mL, are said to be infinitely significant. This statement means that the number of figures in these factors does not affect the number of significant figures in the answer to the problem.

**Density**

We have said that chemists determine the properties of matter, particularly those properties that help identify the composition of a sample. We can measure the mass and volume of a sample, as was done for several samples of iron with results shown in Table 2.7. However, neither their masses nor their volumes show that all the samples are iron, but all the samples do have the same ratio of mass to volume, as is shown in the far right column. This ratio is called density.

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

All samples of the same kind of matter under the same conditions have the same density. Density is a physical property that characterizes and identifies a particular kind of matter (see Figure 2.6). Table 2.8 lists the densities of some common solids and liquids under normal conditions. The densities of solids are usually given in grams per cubic centimeter (g/cm3), the densities of gases in grams per liter (g/L), and the densities of liquids in grams per milliliter (g/mL). Recall from Table 2.1 that 1 mL = 1 cm3. Using these units, the density of water is given as 1.000 g/mL at 4°C. Based on the information in Table 2.8, we can make some basic observations. The densities of most metals are greater than that of water. The densities of liquids vary; some are less dense than water, whereas others are more dense. For example, the density of gasoline is about 30% less than that of water, and the density of chloroform is about 50% greater.

Densities vary with temperature.

Density is a conversion factor that relates mass to volume. If you know two of the three quantities (mass, volume, and density), you can calculate the third.

**Specific Gravity**

Often, particularly in discussing fluids, specific gravity is reported rather than density. The specific gravity (sp gr) of a substance is the ratio of its density to that of a reference substance:

$$\text{Specific gravity} = \frac{\text{density of a substance}}{\text{density of a reference substance}}$$

Generally, water is the reference substance for comparing solids and liquids, and air is the reference substance for comparing gases. A value of specific gravity must state the temperature at which the densities were measured. Specific gravity has no units, because the density units cancel in its calculation. For example, we calculate the specific gravity of benzene at 20°C as follows:

$$\begin{aligned} \text{sp gr}^{20}_{4} &= \frac{\text{density of benzene at } 20\,^\circ\text{C}}{\text{density of water at } 4\,^\circ\text{C}} \\ &= \frac{0.8784 \text{ g/mL}}{1.0000 \text{ g/mL}} \\ &= 0.8784 \end{aligned}$$

**Energy Measurements**

Energy Measurements We have learned that chemistry is concerned with the properties of matter and with the energy changes that matter undergoes. We have discussed properties related to the mass and volume of a sample of matter. In this section we examine properties related to energy. Energy is measured either in joules (J) or in calories (cal), where the conversion factor relating the two units is:

4.184 J = 1 cal

The terms kilojoule (kJ), 1000 J, and kilocalorie (kcal), 1000 cal, are also commonly used. The large calorie (Calorie) used in nutrition is equal to one kilocalorie.

The amount of heat energy associated with a particular sample is dependent on its temperature, its mass, and its composition. Let us consider temperature before discussing its relationship to the energy of a sample.

Some measurement units are defined below:

**Ampere (amp):** A unit of electrical current or rate of flow of electrons. One volt across one ohm of resistance causes a current flow of one ampere. Amperes are used by utilities and electrical engineers to measure electrical flow.

**Joule (J):** A unit of electrical energy equal to the work done when a current of one ampere passes through a resistance of one ohm for one second (synonymous with watt-second).

**Ohm:** A measure of the electrical resistance of a material equal to the resistance of a circuit in which the potential difference of 1 volt produces a current of 1 ampere. Ohms are used by utilities and electrical engineers to measure the resistance of wires conducting electricity.

**Volt:** A unit of electrical force equal to the amount of electromotive force that will cause a steady current of one ampere to flow through a resistance of one ohm. High-voltage electricity moves faster than low-voltage electricity, as seen in the difference between high-voltage transmission lines used to move electricity quickly throughout a region and lower-voltage distribution lines used to move electricity directly to customers.

**Voltage:** The amount of electromotive force, measured in volts, that exists between two points. Voltage is used to describe the amount of power produced by a generator.

**Watt (W):** The rate of energy transfer equivalent to one ampere under an electrical pressure of one volt. One watt equals 1/746 horsepower, or one joule per second. It is the product of voltage and current (amperage). The term "watt" (in addition to the larger measurements of kilowatt and megawatt) is commonly used to describe the capacity of an electric generator. For example, a 1,000-watt photovoltaic system has the capacity to produce 1,000 watts of power at any given time, though it may not consistently produce this much.

**Conversion Factors**

**Temperature**

Temperature measures how hot or cold a sample is relative to something else, usually an arbitrary standard.

**Temperature scales**

Temperature is measured with a thermometer and is most commonly reported using one of three different scales: Fahrenheit (F), Celsius (C) (sometimes called centigrade), and Kelvin (sometimes called absolute).

The relationship between temperatures on these three scales is straightforward if you understand how a thermometer is constructed and calibrated. Two essential features of a thermometer are: (1) it contains a substance that expands when heated and contracts when cooled, and (2) it has some means to measure the expansion and contraction. In the thermometer with which you may be most familiar, the substance that expands and contracts is mercury. In order to measure its expansion or contraction, the mercury is confined within a small, thin-walled glass bulb connected to a very narrow or capillary tube. When the temperature increases, the mercury expands and its level in the capillary tube rises. This increase in height is proportional to the increase in temperature.

A thermometer is calibrated in the following manner. First, the mercury bulb of a new thermometer is immersed in a mixture of ice and water. When the height of the mercury in the column remains constant, a mark is made. This mark is one reference point. The ice-water mixture is then heated to boiling and kept at that temperature while the height of the mercury in the column rises to a new constant level. Another mark is made on the column at this level. This mark is a second reference point.

Further steps depend on whether this thermometer will measure temperature on the Celsius, Fahrenheit, or Kelvin scale. If the Celsius scale is to be used, the reference point for the ice-water mixture is labeled 0°C and that for boiling water is labeled 100°C. The distance between these two reference points is divided into 100 equal segments. If the thermometer is to measure temperature on the Fahrenheit scale, the reference point for the ice-water mixture is labeled 32°F and that for boiling water is labeled 212°F. The distance between these two points is divided into 180 equal segments. If the thermometer is to measure temperature on the Kelvin scale, the ice-water reference point is labeled 273.15 K, the boiling-water reference point is labeled 373.15 K, and the distance between these two marks is divided equally into 100 segments. Notice that K does not use a degree symbol. The symbol K means "degrees Kelvin." As you can see, the temperatures measured by any of these thermometers do not differ; the difference is in the units with which each temperature is reported.

**Conversions between the temperature scales**

A temperature reading on any one of the three scales can be converted to a reading on any other. First, consider a conversion from degrees Celsius to degrees Fahrenheit. Figure 2.7 shows that, between the temperature readings of the ice-water and boiling-water marks, there are 180 Fahrenheit degrees but only 100 Celsius degrees. This relationship can be written as a conversion factor:

$$180°F = 100°C \qquad \text{or} \qquad \frac{180°F}{100°C} = \frac{9°F}{5°C} = \frac{1.8°F}{1°C}$$

In other words, a temperature increase of 9 Fahrenheit degrees is equivalent to an increase of 5 Celsius degrees. Figure 2.7 also shows that the numerical values assigned to the two ice-water reference points differ by 32 degrees; a reading of 0° on the Celsius scale corresponds to a reading of 32° on the Fahrenheit scale. Combining these facts in an equation, we get:

$$°F = \tfrac{9}{5}(°C) + 32 \qquad \text{or} \qquad °F = 1.8(°C) + 32$$

This equation can be rearranged to give the Fahrenheit to Celsius conversion equation:

$$°C = \tfrac{9}{5}(°F) - 32 \qquad \text{or} \qquad °C = \frac{(°F - 32)}{1.8}$$

What is the relationship between the Celsius and Kelvin scales? Because each scale has exactly 100 divisions, or degrees, between the ice-water temperature and the boiling-water temperature, the temperature change represented by a Celsius degree is the same as that represented by a Kelvin. The scales differ in the readings at the ice-water reference point; the reading is 0° on the Celsius scale and 273.15 on the Kelvin scale. Therefore, to convert a Celsius temperature to a Kelvin temperature, simply add 273.15.

$$K = °C + 273.15 \qquad \text{or} \qquad °C = K - 273.15$$

Remember that K is not preceded by the degree symbol (°). The symbols for Fahrenheit and Celsius do require the degree symbol; for example we write 212°F and 100°C, but 373.15 K.

**Melting points and boiling points**

Among the data used to identify a substance are the temperatures at which it changes state. The melting point, (mp) of a substance is the temperature at which it changes from a solid to a liquid (or from a liquid to a solid, in which case it may be called the freezing point). The boiling point (bp) of a substance is the temperature at which under normal conditions the substance changes from a liquid to a gas.

**Specific Heat**

When energy in the form of heat is added to a sample, the resulting temperature change depends on the sample's mass and composition. We are aware of this dependence on composition when we notice that a piece of iron left in bright sunshine quickly becomes too hot to touch, whereas a sample of water with the same mass left in the same location for the same length of time becomes only pleasantly warm. The difference is due to the difference in composition and is expressed quantitatively in the specific heats of the two materials. The specific heat, of a substance is the amount of energy required to raise the temperature of a 1-g sample by 1 degree Celsius. Typically specific heat has units of joules per gram °C (J/g°C).

The specific heat of iron is 0.4525 J/g°C; that is, 0.4525 J is required to raise the temperature of 1 g iron by 1°C. The specific heat of water is 4.184 J/g°C, so 4.184 J are required to raise the temperature of 1 g water by 1°C. Each kind of matter has a unique specific heat.

Specific heat is a conversion factor that relates energy input to sample mass, composition, and temperature change.