Contents

Matter: Its Properties and Measurement

Properties of Matter

Dictionary definitions of **chemistry** usually include the terms *matter*, *composition*, and *properties*, as in the statement that "**chemistry** is the science that deals with the composition and properties of matter."

Matter: is anything that occupies space and displays the properties of mass and inertia. Every human being is a collection of matter. We all occupy space, and we describe our mass in terms of weight, a related property. All the objects that we see around us consist of matter. The gases of the atmosphere, even though they are invisible, are matter they occupy space and have mass. Sunlight is not matter; rather, it is a form of energy.

Composition: refers to the *parts* or *components* of a sample of matter and *their relative proportions*. Ordinary water is made up of two simpler substances hydrogen and oxygen present in certain fixed proportions. A chemist would say that the composition of water is 11.19% hydrogen and 88.81% oxygen by mass. Hydrogen peroxide, a substance used in bleaches and antiseptics, is also made up of hydrogen and oxygen, but it has a different composition. Hydrogen peroxide is 5.93% hydrogen and 94.07% oxygen by mass.

Properties: are those qualities or attributes that we can use to distinguish one sample of matter from others; and, as we consider next, the properties of matter are generally grouped into two broad categories: *physical* and *chemical* properties.

Physical Properties and Physical Changes

Physical property: is one that a sample of matter displays *without changing its composition*. Thus, we can distinguish between the reddish brown solid, copper, and the yellow solid, sulfur, by the physical property of color (Fig. 1).

Figure 1. Physical properties of sulfur and copper.

Sometimes a sample of matter undergoes a change in its physical appearance. In such a physical change, some of the physical properties of the sample may change, but its composition remains unchanged. When liquid water freezes into solid water (ice), it certainly looks different and, in many ways, it is different. Yet, the water remains 11.19% hydrogen and 88.81% oxygen by mass.

Chemical Properties and Chemical Changes

In a *chemical change*, or *chemical reaction*, one or more kinds of matter are converted to new kinds of matter with different compositions. The key to identifying chemical change, then, comes in observing a *change in composition*.

Example: The burning of paper involves a chemical change. Paper is a complex material, but its principal constituents are carbon, hydrogen, and oxygen. The chief products of the combustion are two gases, one consisting of carbon and oxygen (carbon dioxide) and the other consisting of hydrogen and oxygen (water, as steam).

Example: Zinc reacts with hydrochloric acid solution to produce hydrogen gas and a solution of zinc chloride in water (Fig. 2). This reaction is one of zinc s distinctive chemical properties.

Figure 2. A chemical property of zinc and gold: reaction with hydrochloric acid.

$$Zn + 2HCl\ (aq.) \longrightarrow ZnCl_2\ (aq.) + H_2\ (g)$$

Classification of Matter

Atom: Matter is made up of very tiny units called *atoms*.

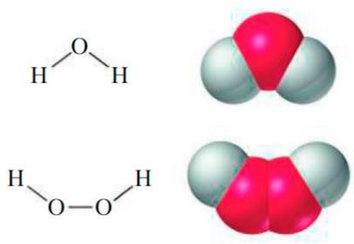
Element: Each different type of atom is the building block of a different chemical *element*.

Presently, the International Union of Pure and Applied Chemistry (IUPAC) recognizes 112 elements, and all matter is made up of just these types! The known elements range from common substances, such as carbon, iron, and silver, to uncommon ones, such as lutetium and thulium.

Compound: Chemical *compounds* are substances comprising atoms of two or more elements joined together.

Molecule: A *molecule* is the smallest entity having the same proportions of the constituent atoms as does the compound as a whole.

▷ **Example:** A molecule of water consists of three atoms: two hydrogen atoms joined to a single oxygen atom. A molecule of hydrogen peroxide has two hydrogen atoms and two oxygen atoms; the two oxygen atoms are joined together and one hydrogen atom is attached to each oxygen atom.



➘ The composition and properties of an element or a compound are uniform throughout a given sample and from one sample to another.

Substance: Elements and compounds are called *substances*. (In the chemical sense, the term substance should be used only for elements and compounds.)

➘ A mixture of substances can vary in composition and properties from one sample to another.

Homogeneous Mixture: One that is uniform in composition and properties throughout is said to be a *homogeneous mixture* or a *solution*.

Heterogeneous Mixture: In *heterogeneous mixtures*–sand and water, for example–the components separate into distinct regions. Thus, the composition and physical properties vary from one part of the mixture to another.

Separating Mixtures

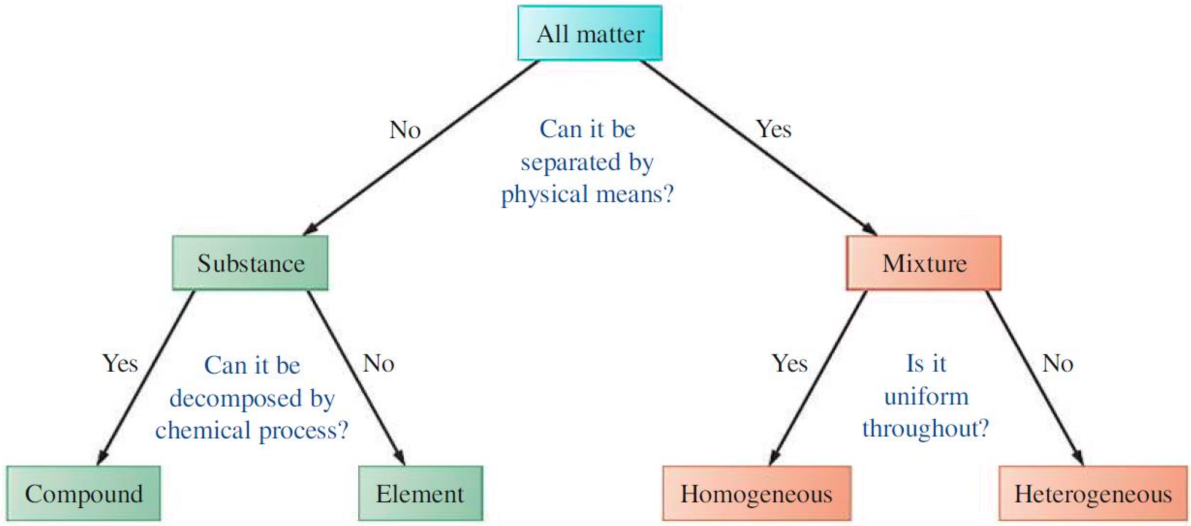


Figure 3. A classification scheme for matter.

➘ A mixture can be separated into its components by appropriate physical means (e.g. filtration, distillation, and chromatography).

➘ A chemical compound can be decomposed into its constituent elements by chemical changes. The decomposition of compounds into their constituent elements is a more difficult matter than the mere physical separation of mixtures.

▷ **Example:** For example, when heated, ammonium dichromate decomposes into the substances chromium(III) oxide, nitrogen, and water. This reaction, once used in movies to simulate a volcano, is illustrated in Fig. 4.

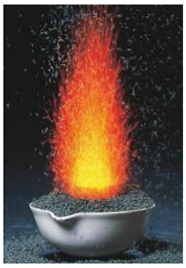


Figure 4. A chemical change: decomposition of ammonium dichromate

States of Matter

Matter is generally found in one of three *states*: solid, liquid, or gas (Fig. 5).

Solid: In a *solid*, atoms or molecules are in close contact, sometimes in a highly organized arrangement called a crystal.

Liquid: In a *liquid*, the atoms or molecules are usually separated by somewhat greater distances than in a solid. Movement of these atoms or molecules gives a liquid its most distinctive property the ability to flow, covering the bottom and assuming the shape of its container.

Gas: In a gas, distances between atoms or molecules are much greater than in a liquid. A gas always expands to fill its container. Depending on conditions, a substance may exist in only one state of matter, or it may be present in two or three states.

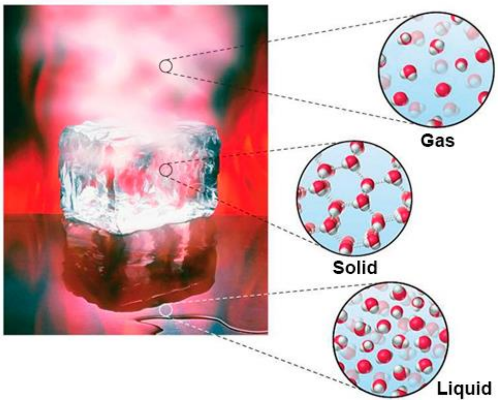


Figure 5. Macroscopic and microscopic views of matter

Measurement of Matter: SI (Metric) Units

Chemistry is a *quantitative* science, which means that in many cases we can measure a property of a substance and compare it with a standard having a known value of the property. We express the measurement as the product of a *number* and a *unit*.

The scientific system of measurement is called the *Système Internationale d'Unités* (International System of Units) and is abbreviated **SI**.

SI Base Quantities and Prefixes

Table 1. SI Base Quantities

Physical Quantity	**Unit**	**Symbol**
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	kelvin	K
Amount of substance	mole	mol
Electric current	ampere	A
Luminous intensity	candela	cd

TABLE 2. SI Prefixes

Multiple	**Prefix**
10^{18}	exa (E)
10^{15}	peta (P)
10^{12}	tera (T)
10^{9}	giga (G)
10^{6}	mega (M)
10^{3}	kilo (k)
10^{2}	hecto (h)
10^{1}	deka (da)
10^{-1}	deci (d)
10^{-2}	centi (c)
10^{-3}	milli (m)
10^{-6}	micro (a m)
10^{-9}	nano (n)
10^{-12}	pico (p)
10^{-15}	femto (f)
10^{-18}	atto (a)
10^{-21}	zepto (z)
10^{-24}	yocto (y)

Mass and Weight

Mass describes the quantity of matter in an object. In SI the standard of mass is 1 kilogram (kg), which is a fairly large unit for most applications in chemistry. More commonly we use the unit gram (g) (about the mass of three aspirin tablets).

Weight is the force of gravity on an object. It is directly proportional to mass, as shown in the following mathematical expressions.

$$W \propto m \text{ and } W = g \times m$$

An object has a fixed mass (*m*), which is independent of where or how the mass is measured. Its weight (*W*), however, may vary because the acceleration due to gravity (*g*) varies slightly from one point on Earth to another.

➘ The terms weight and mass are often used interchangeably, but only mass is a measure of the quantity of matter.

Temperature

There are two commonly used temperature scale.

Celsius: On the *Celsius* scale, the melting point of ice is 0 °C, the boiling point of water is 100 °C, and the interval between is divided into 100 equal parts called *Celsius* degrees.

Fahrenheit: On the *Fahrenheit* temperature scale, the melting point of ice is 32 °F, the boiling point of water is 212 °F, and the interval between is divided into 180 equal parts called Fahrenheit degrees.

Kelvin: The SI temperature scale, called the *Kelvin* scale, assigns a value of zero to the lowest possible temperature. The zero on the *Kelvin* scale is denoted 0 K and it comes at 273.15 °C.

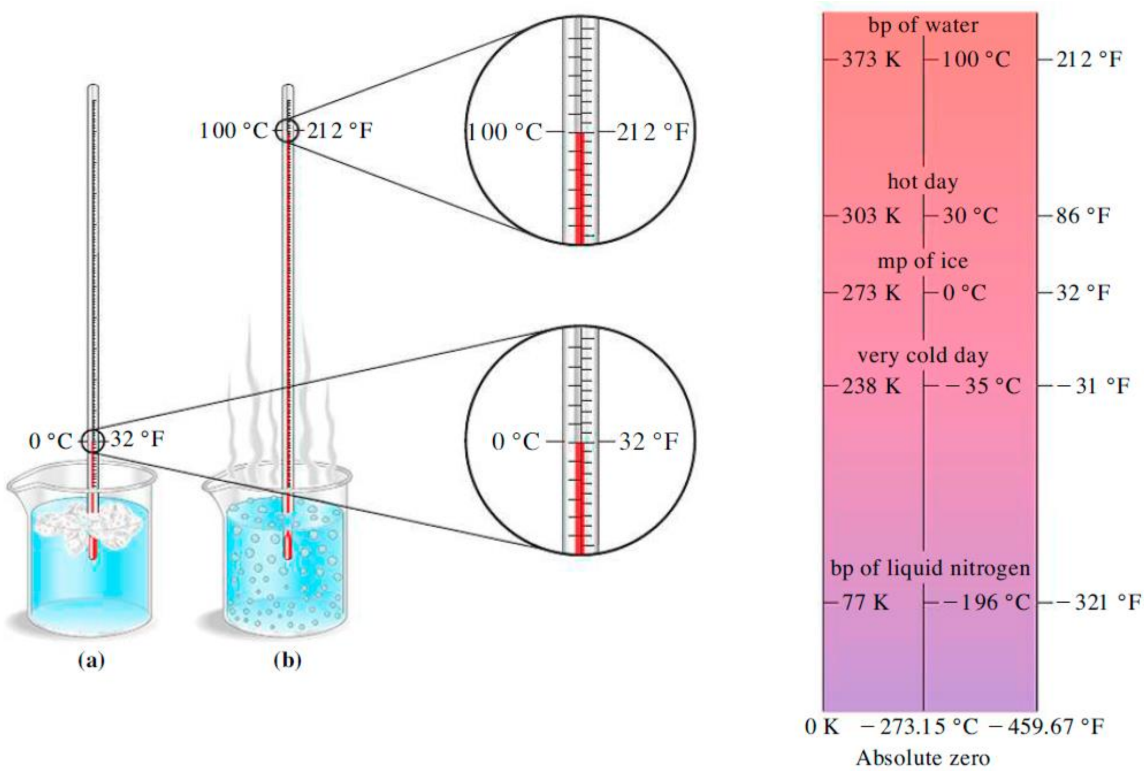


Figure 6. A comparison of temperature scales. (a) The melting point (mp) of ice. (b) The boiling point (bp) of water.

➘ Temperature conversions can be made in a straightforward way by using the algebraic equations shown below.

Kelvin from Celsius $T(K) = t(°C) + 273.15$

Fahrenheit from Celsius $T(°F) = \frac{9}{5}t(°C) + 32$

Celsius from Fahrenheit $t(°C) = \frac{5}{9}[t(°F) - 32]$

▷ **Example:** A recipe in an American cookbook calls for roasting a cut of meat at 350 °F. What is this temperature on the Celsius scale?

$$t(°\mathrm{C}) = \frac{5}{9}[t(°\mathrm{F}) - 32] = \frac{5}{9}[350 - 32] = 176.7\ °\mathrm{C}$$

Derived Units

Many measured properties are expressed as combinations of these fundamental, or base, quantities. We refer to the units of such properties as *derived* units.

Derived Unit	**Combination**
Velocity	m s^{-1}
Pressure (*pascal*)	kg m^{-1} s^{-2}
Energy (*joule*)	kg m^2 s^{-2}
Volume	m^3, cm^3, L (1000 cm^3), mL (1 cm^3)

Density and Percent Composition: Their Use in Problem Solving

In this section, we will introduce two quantities frequently required in problem solving: density and percent composition.

Density

Density is the ratio of mass to volume. Units of density: g/cm^3 or kg/m^3 or g/mL

$$\text{density } (d) = \frac{\text{mass } (m)}{\text{volume } (V)}$$

Mass and volume are both extensive properties. An *extensive property* is dependent on the quantity of matter observed. However, if we divide the mass of a substance by its volume, we obtain density, an intensive property. *An intensive property is independent of the amount of matter observed.* Thus, the density of pure water at 25 °C has a unique value, whether the sample fills a small beaker (small mass/small volume) or a swimming pool (large mass/large volume).

➘ There are several important consequences of the different densities of solids and liquids. A solid that is insoluble and floats on a liquid is *less* dense than the liquid, and it displaces a mass of liquid equal to its own mass. An insoluble solid that sinks to the bottom of a liquid is *more* dense than the liquid and displaces a volume of liquid equal to its own volume. Liquids that are immiscible in each other separate into distinct layers, with the *most* dense liquid at the bottom and the *least* dense liquid at the top.

Density in Conversion Pathways

If we measure the mass of an object and its volume, simple division gives us its density. Conversely, if we know the density of an object, we can use density as a conversion factor to determine the object s mass or volume.

▷ **Example:** For example, a cube of osmium 1.000 cm on edge weighs 22.59 g. The density of osmium (the densest of the elements) is 22.59 g / cm^3. What would be the mass of a cube of osmium that is 1.25 in. on edge (1 in. = 2.54 cm). To solve this problem, we begin by relating the volume of a cube to its length, that is, $V = l^3$. Then we can map out the *conversion pathway*:

$$\text{in. osmium} \longrightarrow \text{cm osmium} \longrightarrow \text{cm}^3\text{ osmium} \longrightarrow \text{g osmium}$$

(converts in. to cm) (converts cm to cm^3) (converts cm^3 to g osmium)

$$?\text{ g osmium} = \left[1.25\ \cancel{\text{in.}} \times \frac{2.54\ \cancel{\text{cm}}}{1\ \cancel{\text{in.}}}\right]^3 \times \frac{22.59\text{ g osmium}}{1\ \cancel{\text{cm}^3}} = 723\text{ g osmium}$$

▷ **Example:** At 25 °C the density of mercury, the only metal that is liquid at this temperature, is 13.5 g/mL. Suppose we want to know the volume, in mL, of 1.000 kg of mercury at 25 °C.

kg mercury → g mercury → mL mercury

$$?\text{ mL mercury} = 1.000\ \cancel{\text{kg}} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\text{ mL mercury}}{13.5\ \cancel{\text{g}}} = 74.1\text{ mL mercury}$$

Percent Composition as a Conversion Factor

Percent is the number of parts of a constituent in 100 parts of the whole. To say that a seawater sample contains 3.5% sodium chloride by mass means that there are 3.5 g of sodium chloride in every 100 g of the seawater. We make the statement in terms of grams because we are talking about percent *by mass*. We can express this percent by writing the following ratios:

▷ **Example:** A 75 g sample of sodium chloride (table salt) is to be produced by evaporating to dryness a quantity of seawater containing 3.5% sodium chloride by mass. What volume of seawater, in liters, must be taken for this purpose? Assume a density of 1.03 g/mL for seawater.

$$\begin{aligned} ?\ \text{L seawater} &= 75\ \cancel{\text{g sodium chloride}} \times \frac{100\ \cancel{\text{g seawater}}}{3.5\ \cancel{\text{g sodium chloride}}} \\ &\quad \times \frac{1\ \cancel{\text{mL seawater}}}{1.03\ \cancel{\text{g seawater}}} \times \frac{1\ \text{L seawater}}{1000\ \cancel{\text{mL seawater}}} \\ &= 2.1\ \text{L seawater} \end{aligned}$$