

General Chemistry

A Mastery-Oriented Curriculum

Third Edition

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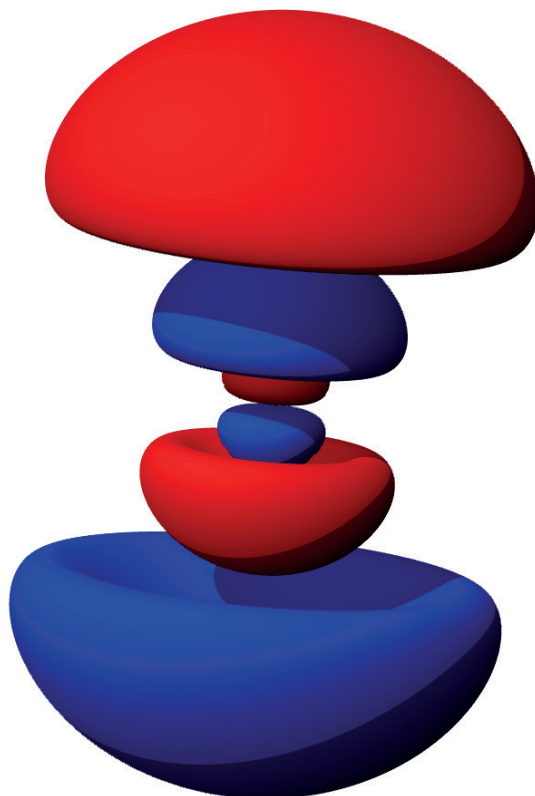
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Introduction

What is Chemistry All About?



This computer model depicts three of the *orbitals* available in atoms for holding electrons. Shown are one of the three *p* orbitals in each of three different *subshells*. The inner pair of orbitals can hold two of the highest energy electrons for elements 5 through 10 in the Periodic Table of the Elements. The middle pair is available to hold two of the highest energy electrons for elements 13 through 18 in the periodic table, and the outer pair can hold two of the highest energy electrons belonging to elements 31 through 36 in the periodic table.

In this Introduction, we touch briefly on electron orbitals. We treat the subject in more depth in Chapter 3.

Objectives for the Introduction

After studying this chapter and completing the exercises, you should be able to do each of the following tasks, using supporting terms and principles as necessary.

1. Briefly explain how electrons, electrical forces, minimizing energy, whole number ratios, and modeling can each be thought of as central to understanding what chemistry is all about.
2. State and explain examples illustrating a system moving to a lower energy state and a system experiencing an increase in entropy.
3. Briefly explain hydrogen bonding and why it plays such a large role in mixtures containing water.
4. Explain the relationship between energy and atomic orbitals.

1.1 A Few Major Themes

Chemistry is the study of the elements, how they combine to form mixtures and compounds, the properties of these substances, and the processes involved. One of the astonishing things about the physical world is that as complex as the details are, we can understand a lot about how it works in terms of just a few basic principles from physics. This striking situation is a direct result of the fact that nature is governed by an orderly, mathematical set of physical laws—the laws set in place by God according to his wisdom when he created the universe.

The existence of creation and of the laws of physics are two obvious clues to God's role in creating the universe: the universe is here because God made it, and it is governed in an orderly, mathematical way because it was God's pleasure to make it so. A third clue is that *we can understand it*.

Studying chemistry involves learning a great deal of terminology, and exploring quite a few different types of processes. The amount of information involved can be daunting! But one way to help organize all this information is to be alert to a few fundamental principles that turn up time and again. In this introductory chapter, we take a brief look at a few of these principles. As you read through the chapters ahead, you will see again and again that we can understand a lot about topics such as molecular structure, solubility, and chemical reactions in terms of a few basic concepts.

1.1.1 Chemistry Is All About Electrons

You probably already know that atoms consist of a tiny nucleus containing particles called protons and neutrons, and that the nucleus is surrounded by cloud-like regions containing the atoms' electrons. The protons and electrons carry electrical charge—protons are positively charged and electrons are negatively charged. The protons in an atom stay permanently in the atom's nucleus,¹ but atoms lose or gain electrons by interacting with other atoms. An atom that gains or loses one or more electrons is called an *ion*. Ions are charged particles. Gaining an electron means gaining negative charge, resulting in a negatively charged ion. Losing electrons means losing negative charge, and ending up with more protons (positive charges) than electrons. This results in a positively charged ion.

As it turns out, a lot of chemistry can be understood in terms of the atoms' electrons—where they are, how many there are, whether an atom has ionized by gaining or losing electrons, whether an atom is sharing electrons with another atom, and so on.

1 Except in the case of radioactive elements.

The cloud-like regions containing an atom's electrons are called *orbitals*, and electrons reside in different orbitals according to how much energy they have. The arrangement of the orbitals is the same for all atoms, although the specific energies associated with each orbital vary from atom to atom, depending on the size of the nucleus and how many electrons an atom has. The orbitals in atoms are grouped into different energy groupings called *shells*. There are seven main shells containing the orbitals with the electrons of all the elements discovered so far. There are additional shells above these that high-energy electrons can move into when they absorb more energy.

In each shell, there is a specific number of orbitals, and each orbital and set of orbitals holds a specific number of electrons. One of the essential facts about atomic behavior is that atoms seek to gain, lose, or share electrons until they have just the right number of electrons so that they have only full shells, without any extra electrons and without electrons missing from any orbitals in the full shells. If only the first shell is full, an atom has two electrons. If the first two shells are full, 10 electrons. If the first three are full, 28 electrons, and so on. Significantly, these numbers relate to the numbers of elements in the rows of the Periodic Table of the Elements. With this one fact, we can understand a great deal about how atoms of one element bond with atoms of other elements to form compounds.

The position of the electrons within an atom also has a lot to do with how an atom behaves. One aspect of atoms that affects the position of electrons is the shapes of the different orbitals. Some orbitals are spherically shaped, some are shaped in pairs of protruding lobes often described as “dumbbells,” and some are shaped as rings. There are other more complex shapes as well. Since all electrons repel each other due to their negative electrical charge, electrons located in the lobes of dumbbell-shaped orbitals push away from each other, resulting in molecules with very particular shapes. Examples are the water, ammonia, and methane molecules illustrated in Figure I.1.

Electron position is also affected by the fact that within molecules some atoms attract electrons more strongly than others, an effect denoted by a value called the *electronegativity* of the atom. We discuss this in more detail later, but I will mention an important example here to illustrate this point. The electronegativity values for oxygen and hydrogen are 3.44 and 2.20, respectively. This means the oxygen atoms in water molecules attract electrons more strongly than the hydrogen atoms do. As a result, the four bonding electrons in the molecule crowd over toward the oxygen atom, making the oxygen region of the water molecule more electrically negative and the hydrogen regions more electrically positive. These differences make the water molecule electrically imbalanced—or *polar*, as we say—negative on one side and positive at the ends on the other side, as illustrated in Figure I.2. In this diagram, the arrows point from the positive regions of the molecule toward the negative region of the molecule.

The shapes and polarizations affect atomic behavior because of electrical attractions and repulsions, the basic theme we discuss in the next section.

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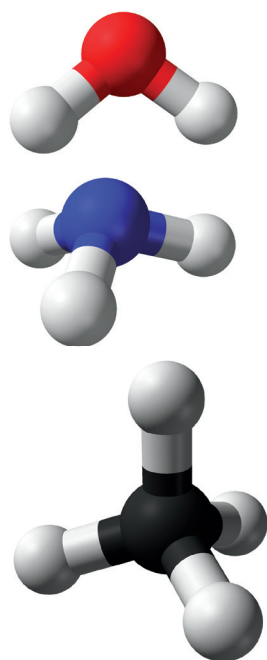


Figure I.1. Representations of the H_2O water molecule (top), the NH_3 ammonia molecule (middle), and the CH_4 methane molecule (bottom).

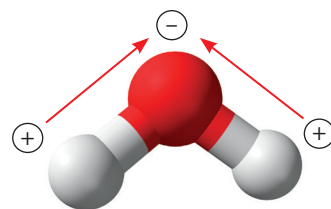


Figure I.2. The higher electronegativity of oxygen atoms compared to hydrogen atoms results in the polar water molecule.

1.1.2 Chemistry Is All About Electrical Forces

Consider again the two types of electrical charge, the positive protons and the negative electrons. You know that like charges repel each other (such as two positive charges) and opposite charges attract each other (positive and negative).

These electrical attractions and repulsions are highly important for chemistry because atoms and molecules are as prickly as porcupines with charges that repel or attract other charges. Some of these attractions and repulsions are stable and long-lasting, like the attractions between ions that hold together the atoms in a crystal of table salt. Other electrical interactions are sort of semi-stable, you might say, given the fact that molecules are moving around all the time. The world-class example of this is *hydrogen bonding*, which we examine in detail later. Hydrogen bonding takes us back to the polar water molecule described in the previous section. Since water molecules are polar, the positive regions of one water molecule are attracted toward the negative regions of other water molecules, as illustrated in Figure I.3.

The importance of hydrogen bonding cannot be overstated. Water is everywhere, and thus so is hydrogen bonding. Hydrogen bonding explains why so many things dissolve in water. It explains why water travels upwards against the force of gravity when soaking into the fibers of a towel. And it explains why water gets less dense right before it freezes (which in turn explains why ice floats). Figure I.4 is a model of how the water molecules are arranged in ice. The dashed lines in the figure indicate the hydrogen bonds between water molecules. The result of these bonds is the three-dimensional, hexagonal structure of ice.

There are several other ways electrical forces between atoms and molecules are made manifest. In general, these different attractions and repulsions are called *intermolecular forces*. There is an electron cloud around every atom (except in the case of a hydrogen atom that has lost its only electron due to ionization). There is also an electron cloud around and between the atoms of every molecule. As the electrons swarm around in these clouds, there are moments when some regions in the molecule are more negatively charged because of electrons crowding together. There are other moments when regions are more positively charged because electrons have temporarily moved away and the positive charge on the protons in atomic nuclei are dominant in the area. These electron movements and crowding go on all the time and at extremely high speeds, giving rise to ever-changing patterns of intermolecular forces. Intermolecular forces are all caused by electrical attractions.

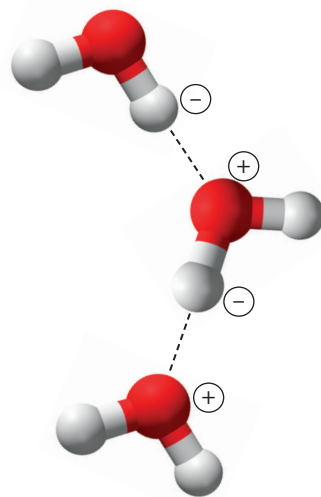


Figure I.3. Hydrogen bonding in water molecules. Dashed lines indicate hydrogen bonds between water molecules.

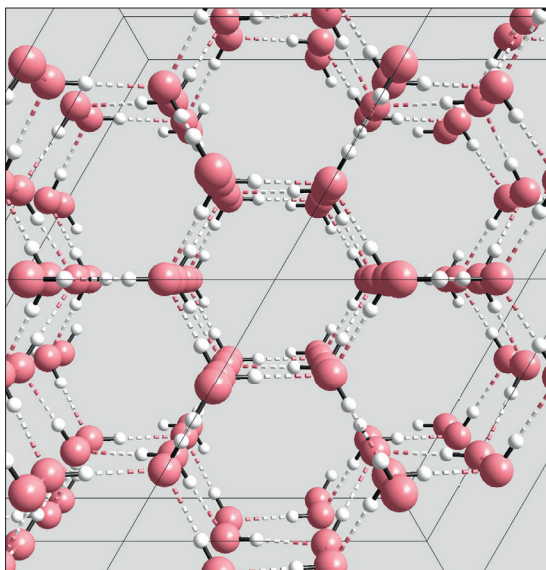
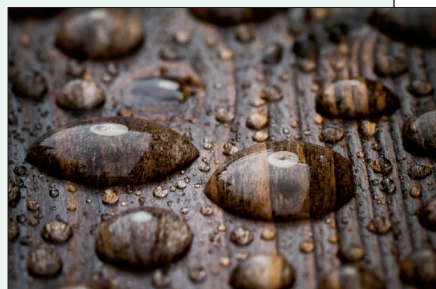
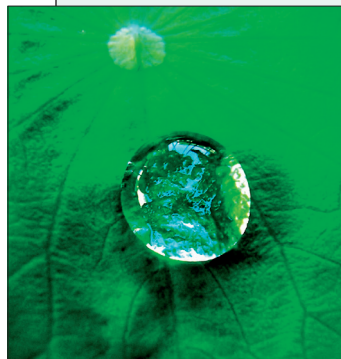


Figure I.4. The crystal structure in ordinary water ice. In this model, oxygen atoms are red and hydrogen atoms are white. Hydrogen bonds are shown as dashed lines. Thick black lines indicate the bonding of hydrogen to oxygen inside individual water molecules.

Hmm... Interesting.**Why water forms beads**

As described in this chapter, the polarity of water molecules makes them cling to one another. When nonpolar molecules are in contact with water, the water molecules are attracted to each other but not to the nonpolar molecules. The molecules in waxy leaves and oil-based wood finishing products are nonpolar. When water molecules rest on a surface of nonpolar molecules, they cling to each other but not to the surface, and the result is the formation of water drops. Small drops are nearly spherical because this shape minimizes the energy between the molecules. Larger drops flatten out due to their greater weight.



Nonpolar molecules do not dissolve in water. The attractions between the polar water molecules squeeze out the nonpolar molecules, causing the two substances to separate. This is why oil and vinegar separate—oil molecules are nonpolar and vinegar is mostly water.

1.1.3 Chemistry Is All About Minimizing Energy

One of the primary drivers causing atoms to do what they do is the natural tendency of all things to minimize the energy associated with the state they are in. Minimizing energy is a concept that explains a great deal of chemical behavior. Here, we look at several examples of objects in different energy states. Then we apply the concept of minimizing energy to phenomena we see occurring in chemistry.

To begin, in a previous science course you may have studied different forms of potential energy. For example, gravitational potential energy is the energy an object has after being lifted up in a gravitational field. Figure 1.5 shows a ball up on the side of the hill. The ball is trapped in a small valley or depression. The ball is located up above the ground, so it has gravitational potential energy. The ball always acts to reduce its potential energy if given a chance. If a tunnel opens up to a lower energy state, the ball goes there, releasing potential energy into some other form of energy (such as kinetic energy) as it goes. Another way for the ball to release potential energy and move to a lower state is for someone to hit it or kick it so that it has enough kinetic energy to get over the small hill where it is trapped. The point is that given the chance, the ball releases potential energy and moves to a lower energy state.

Another example of this idea is shown in Figure 1.6. A cone held on its point has potential energy that is released if the cone is released and allowed to fall. In this case, the cone doesn't even need any kind of push or kick; it spontaneously moves to the lower energy state (laying down on its side) if released.

As a third example, consider the act of stretching a rubber band. To stretch out a rubber band, you have to supply energy.

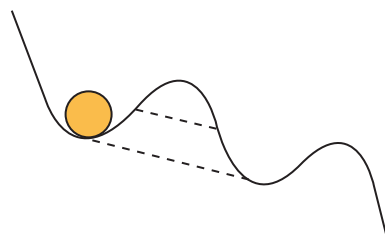
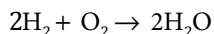


Figure 1.5. The ball is trapped in the lowest-energy region in its vicinity. However, if a tunnel to a lower-energy region opens up, the ball goes there.

That is, you have to do mechanical work on the rubber band. If you release the stretched rubber band, it spontaneously contracts back to its lower energy (unstretched) state.

Let's now apply the idea of minimizing energy to a chemical reaction you may already be familiar with: the combustion of hydrogen to produce water. This reaction is represented by the following chemical equation:



The left side of this equation indicates that the *reactants*—the substances taking part in the reaction—are molecules of hydrogen (H_2) and oxygen (O_2). Each of these gases exists as *diatomic* molecules, meaning that each hydrogen molecule is a pair of hydrogen atoms bonded together, and each oxygen molecule is a pair of oxygen atoms bonded together, as illustrated in Figure I.7. At room temperature, these gas molecules zoom around inside their container, colliding with one another several billion times per second, but otherwise nothing else happens.

In terms of the energy of these molecules, they are in a situation similar to the ball in Figure I.5: there are lower energy states the molecules can go to, releasing energy in the process, but they can't get there without a boost of energy to get the process started. Now, if a spark or flame is introduced to this gas mixture, the heat from the spark or flame excites the nearby molecules, causing them to move much faster and slam into each other with enough energy to break the bonds holding the molecules together. The result—which only lasts for a tiny fraction of a second—is a soup of unbonded gas atoms.

At this point, we have a situation similar to the stretched rubber band the instant after being released, before it has had a chance to shrink. Electrical attractions between the protons and electrons in the isolated atoms of oxygen and hydrogen draw the atoms toward each other at an extremely high rate. Consider the collapse of the rubber band after it is released. It collapses to its unstretched state—a lower energy state—releasing energy in the process. The energy released might result in a snap (kinetic energy) that stings your hand and a sound wave (kinetic energy in moving air molecules) producing a snapping sound. Just as the relaxed rubber band is at a lower energy state and releases energy to get there, the hydrogen and oxygen atoms collapse together to the lowest energy state they can find, which is to form water molecules (H_2O). As they do so, they release a lot of energy in the form of light and heat and all this happens in an instant.

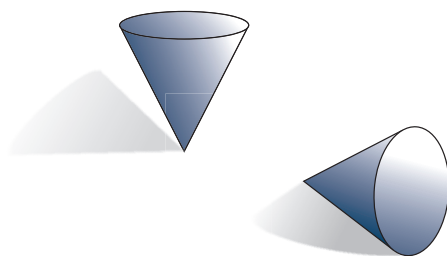


Figure I.6. The cone on the left is in a higher energy state. If released, it falls to the lower energy state shown on the right.

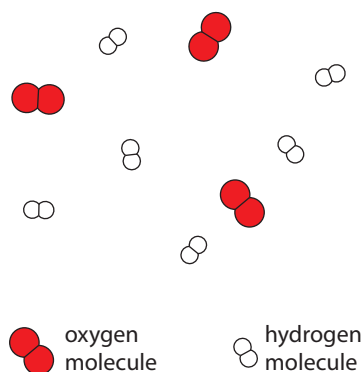


Figure I.7. Diatomic oxygen and hydrogen gas molecules.



Figure I.8. Energy released as heat and light as hydrogen and oxygen atoms combine to form water molecules.

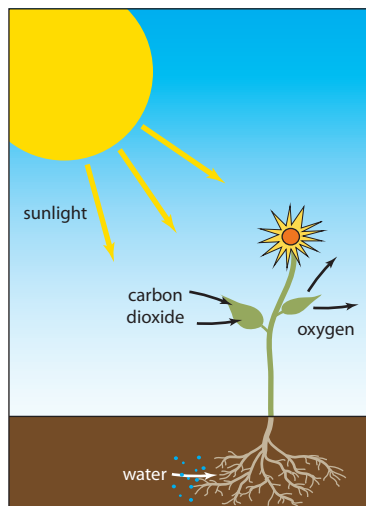
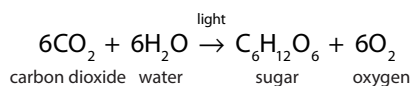


Figure I.9. The photosynthesis reaction is endothermic, as light from the sun is used by plants to convert water and carbon dioxide into sugar and oxygen.

This is the explosion of hydrogen, captured in the photograph of Figure I.8.

This release of heat indicates the reaction is *exothermic*—the reaction releases energy. When considering the way energy relates to various chemical processes, I have found it very helpful to remember the rubber band and to compare it in my mind to the way positive and negative ions are attracted to each other. If separated positive and negative ions are released and allowed to fly together, energy is released—the light and heat of the exothermic reaction—as the ions move to a lower energy state. To separate them, one has to pull them apart by putting in energy (doing work on them) and thus moving them to a higher energy state—just like stretching the rubber band. This is an *endothermic* process, where energy is being absorbed by the ions. The most well-known example of an endothermic chemical reaction is the photosynthesis reaction that occurs in plants, depicted in Figure I.9.

And I can't help pointing out in passing the exquisite elegance of the photosynthesis reaction, a process both simple and incredibly complex that happens automatically and continuously all over the world. Consider the care with which God placed oxygen-breathing creatures like ourselves on a planet covered with oxygen-producing vegeta-

tion. Of course, every school kid learns about photosynthesis, but do we also learn that the delicate balance displayed everywhere in the environment around us in creation is a tremendous gift? I encourage you, as a young student made in God's image, to give thanks and worship to our loving Creator for this most wonderful gift!

There are two more important concepts about the role of energy in chemistry to note here. The first involves a quantity called *entropy*. Entropy is a term that originated in the field of thermodynamics. Entropy is a measure of the *disorder* present in a system, and the second law of thermodynamics states that left to themselves, physical processes go in a direction that increases the entropy (disorder) in the system. As an example, consider a glass of water you may be holding in your hand versus a broken glass on the floor with water splashed everywhere. While the glass is intact with the water contained in it, the system of glass and water is in an orderly state. When you release the glass, disorder increases—the glass breaks and the water goes everywhere on the floor. If you leave the mess like this, the disorder continues to increase: the water evaporates and the water molecules are not even together any more at all. Instead, they are randomly distributed around in the atmosphere. And with time, the chunks of glass get trampled and broken more and more until the remnants of the glass are completely gone. You will never see this process occur in reverse!

For some chemical processes, the minimizing of energy and the increase of entropy both pull in the same direction. In other cases, they try to pull the system in opposite directions. This sets up a sort of tug of war, and the process goes in the most favorable direction. We discuss this in more detail later.

Finally, some detail is in order regarding the boundaries surrounding energy minimization. Try this little thought experiment: imagine a hydrogen ion, which is simply a proton with its positive charge. Nearby is a negatively charged electron, as illustrated in Figure I.10. Since these particles have opposite charges, they are strongly attracted to each other, and since the

proton's mass is 1,836 times greater than the electron's mass, the electron dashes toward the proton while the proton essentially stays put, waiting for the electron to arrive. You might expect that the electron would crash right into the proton, bringing the potential energy between them right down to zero. But this is not what happens.

In 1905, Albert Einstein theorized that energy is *quantized*—it comes in discrete chunks or packets. Since 1905, a host of scientists have explored the quantization of energy, confirming Einstein's proposal over and over and giving birth to the now well-developed theory of quantum mechanics. What quantum mechanics suggests for our proton-electron scenario is that an electron in an orbital of an atom cannot possess just any old amount of energy; it can only possess particular values of energy. In the context of dropping into one of the orbitals surrounding the proton, the electron can only possess an amount of energy corresponding to the one of the energies of the proton's orbitals. The bottom line is that instead of crashing into the proton and sticking to it like cat hair sticking to your pants, the electron instead pops into the lowest energy orbital available around the proton and stays there, captive, buzzing around furiously like a bee in a bottle. (But though this analogy may be suggestive, it is strictly metaphorical. Electrons are not at all like bees. For one thing, they don't have wings. And they don't make honey, either.)

1.1.4 Chemistry Is All About Whole-Number Ratios of Atoms

It is strange to think that even as recently as the beginning of the 20th century there was no consensus among scientists as to whether atoms even existed. In 1803, English scientist John Dalton put forward the first detailed, scientific atomic theory. Dalton proposed that all material substances are composed of atoms, and that the way different compounds are formed is by atoms combining together. Since various substances are composed of discrete, individual particles and not just a continuum of matter, there is always a whole number of each type of atom in the substance.

Although many scientists throughout the 19th century refused to accept the existence of atoms, we now agree that Dalton was correct. Compounds do form with whole-number ratios of the atoms involved. (Back then, those who accepted the existence of atoms were called “atomists.” Today, everyone is an atomist, so we don't need a name for this view any more.) As an example, sulfuric acid, H_2SO_4 , has two hydrogen atoms, one sulfur atom, and four oxygen atoms in every molecule, so the ratio of oxygen atoms to hydrogen atoms in the molecule is 2 to 1. The

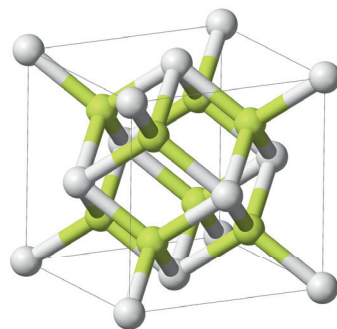


Figure 1.10. Oppositely charged particles strongly attracted to each other.

Figure 1.11. The ratio of fluorine atoms (greenish-yellow) to calcium atoms (white) in fluorite is 2 to 1. To see this, note that there are 8



fluorine atoms completely within this crystalline cell. Each of the 8 calcium atoms at the corners is shared by 8 cells—the one shown and 7 other surrounding cells. Eight calcium atoms each shared 8 ways contributes a net of 1 calcium atom to the cell. Then there are 6 calcium atoms on the faces of the cell, each shared by the cell shown and the adjacent cell. Six atoms each shared 2 ways contributes a net of 3 atoms to the cell. In total then, this cell claims 8 fluorine atoms and 4 calcium atoms, a ratio of 2 to 1.



ratio of oxygen atoms to sulfur atoms is 4 to 1. Of course, there are strange exceptions to every rule, including this one. Nevertheless, it is correct to say that just about every compound, regardless of how the atoms are structured, consists of atoms of different elements joined together in predictable whole-number ratios. Figure I.11 contains a photo of the mineral fluorite along with a computer model of the crystal structure of fluorite, or calcium fluoride, which has the formula CaF_2 . In this crystal structure, the ratio of fluorine atoms to calcium atoms is 2 to 1. The caption in the figure explains this, using the computer model of the crystal structure as an aid.

The fact that atoms combine in whole-number ratios is a powerful computational tool. When we get into the math behind chemical reactions (*stoichiometry*, as it is called), we will appeal often to the whole-number ratios of atoms involved in order to compute how much of one compound reacts with a given quantity of another compound.

1.1.5 Chemistry Is All About Modeling

Chemical reactions are happening around us all the time. Just pour a can of soft drink into a glass and watch the carbonic acid (H_2CO_3) in the can convert into the carbon dioxide bubbles and water (CO_2 and H_2O). Light up the gas grill and watch propane (C_3H_8) reacting with the oxygen (O_2) in the air to produce carbon dioxide (CO_2) and water (H_2O). Heat up a pan of cake batter in an oven and the rather complicated molecules in the batter react and change into different complicated molecules in a cake.

In these examples, even though we see bulk materials going into a chemical reaction (the *reactants*) and resulting from the chemical reaction (the *products*), we are not able to see the actual atoms and molecules as they zoom around, combining with and separating from each other. Understanding the behavior of things we cannot see is tricky business.

As mentioned in the previous section, even in recent scientific history the existence of atoms was debated for a hundred years. The issue was finally resolved with experiments in the early 20th century that gave more and more support to the theory that material substances were composed of atoms. We certainly know a lot more today about atoms and their internal structure than we did just a few decades ago. We can even put this knowledge to use in designing amazing new engineering materials, specialty drugs, and even chemical delivery systems to get the drugs into our bodies. But there remains much we do not understand about atoms.

It is helpful to think of science as the process of building “mental models” of the natural world. These mental models are called *theories*. The information we use to build our mental models—scientific facts—comes from experiments, observations, and inferences from these.

Since chemistry deals so much with atoms and molecules, which we can't see, we are almost completely dependent on inferences to develop atomic models describing how the atomic world works. Knowing that the gunpowder in a firecracker explodes when ignited doesn't require a model. It is obvious to all of us that gunpowder is explosive. But *why* is it explosive? What are the rules governing how the atoms in those compounds behave? Understanding why gunpowder explodes does require a model. And the models we work with in chemistry come at us from two different directions.

First, there is the information we gather from experiments. Chemical experimentation has been going on for hundreds of years. In the early days of the scientific revolution, scientists were amazed to discover quantitative laws such as Dalton's whole number ratios and the inverse relationship between the pressure and volume of a gas, a relationship known as Boyle's Law. Second, there is the theoretical modeling that occurs when scientists attempt to apply physical principles from quantum mechanics, thermodynamics, and statistical mechanics to the solution of chemical problems. The shapes and sizes of the atomic orbitals, which we address in Chapter 3, are an example of this type of theoretical modeling.

The theoretical models developed by scientists are the basis for our entire understanding of how the natural world functions. Successful theories are those that account for the facts we know and lead to new hypotheses (predictions) that can be put to the test. It is helpful to think about the relationship between facts, theories, hypotheses, and experiments as illustrated in Figure I.12. This diagram illustrates what I call the *Cycle of Scientific Enterprise*. It is important for every student to develop a correct understanding of the kind of knowledge scientific study provides for us. The *goal* of science is to uncover the truth about how nature works, but scientific theories are always works in progress. Even our best theories are provisional and subject to change. For this reason, science is not in the business of making truth claims about scientific knowledge. Science is in the business of modeling how nature works with theories based on research.

As our theories develop over time, our hope is that they get closer and closer to the truth—the amazing and profound truth about mysteries such as what protons and electrons are, why they have the properties they have, and how the two most successful theories of the 20th century—quantum mechanics and general relativity—can be reconciled with each other. But the truth about nature is always out in front of us somewhere, always outside our grasp. To know the truth about nature, we would have to understand nature as God understands it. We are nowhere close to that.

Here are some definitions to keep in mind as you consider the models we discuss in future chapters.

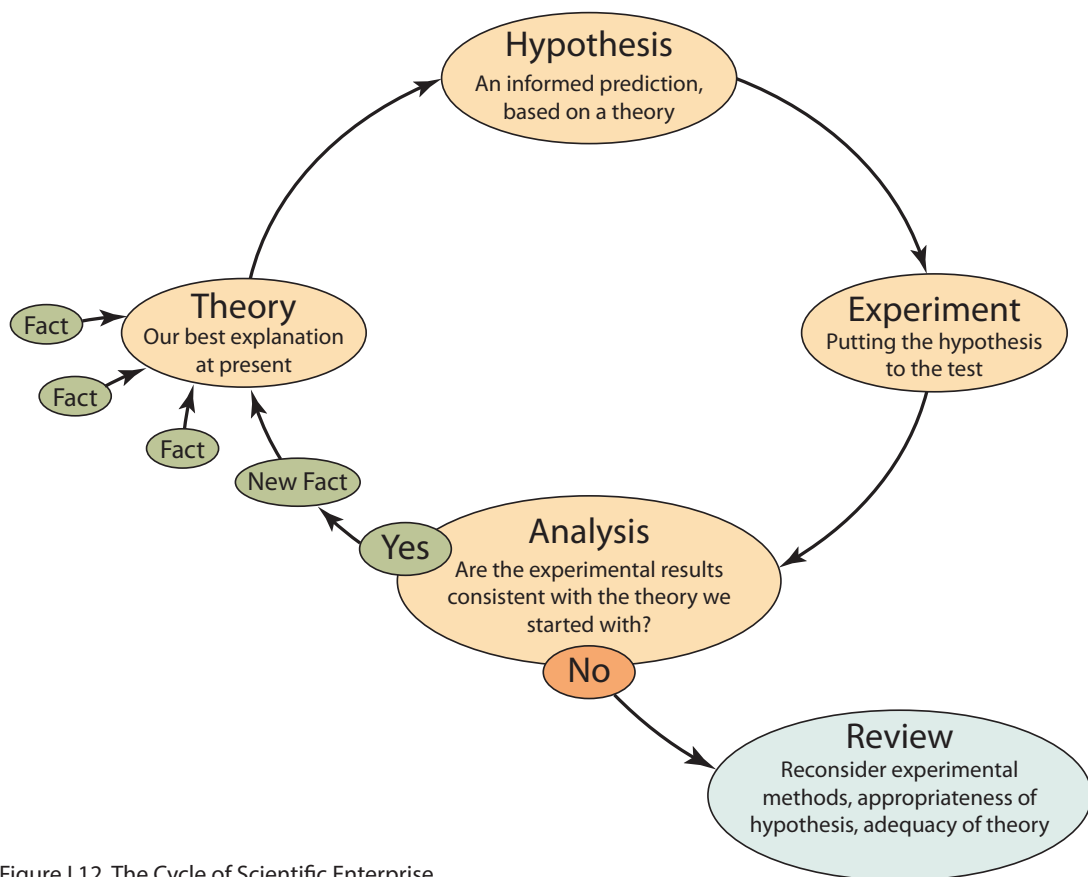


Figure I.12. The Cycle of Scientific Enterprise.

- Fact** A proposition based on a large amount of scientific data that is correct so far as we know. Facts are discovered by experiment, observation, and inferences from experiments and observations. Facts can and do change as new scientific knowledge—new data—is acquired. Since facts are always subject to change, we generally avoid terms like *true* or *proven*. Instead, we say a fact is correct so far as we know.
- Theory** A mental model that accounts for the data (facts) in a certain field of research, and attempts to relate them together, interpret them, and explain them. Scientific theories are successful if they repeatedly allow scientists to form new hypotheses that can be put to experimental test. *Successful theories are the glory and goal of science.* Nevertheless, theories, like facts, are provisional and subject to change. Indeed, theories are almost constantly evolving as research continues. And as with facts, when referring to theories we avoid terms like *true* or *proven*. Instead, we speak in terms of how successful theories have been in generating hypotheses that are confirmed by experiments. A widely accepted scientific theory should be understood as our best explanation at present—our best *model* of how nature works.²
- Hypothesis** An informed prediction about what will happen in certain circumstances. Every hypothesis is based on a particular theory. It is hypotheses that are tested in scientific experiments.
- Experiment** A test designed to confirm or disconfirm a particular hypothesis. If a hypothesis is confirmed through experiment, and if other scientists are able to validate the confirmation by replicating the experiment, then the new facts gained from the experimental results become additional support for the theory the hypothesis came from.

Chemistry is a subject loaded with facts and heavily based on theories—*models*—that we know are incomplete descriptions of nature. That is why the research continues, as our models (hopefully) get nearer and nearer to the truth.

1.2 Conclusion

The goal of this introductory chapter is simply to alert you to some of the key concepts undergirding our understanding of chemical processes. Over and over in the coming chapters, you will find that thinking about the content in terms of one or more of these central ideas will help you develop a better grasp of the material.

Introduction Study Questions

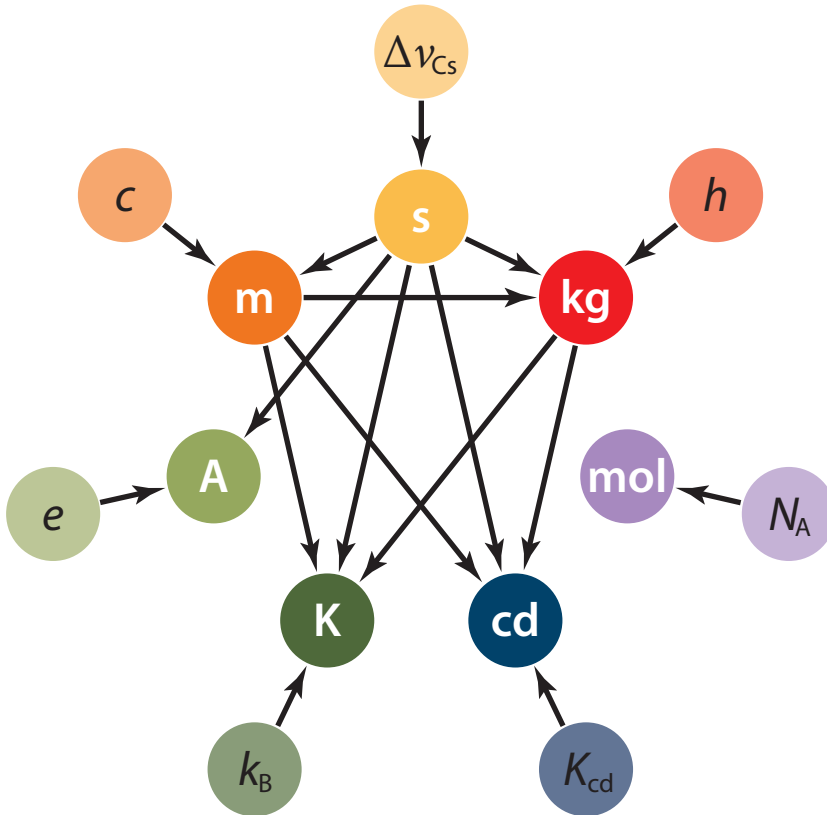
1. Write five brief paragraphs summarizing the main ideas behind the titles of Sections I.1.1 through I.1.5.
2. Describe two examples, other than those in this Introduction, of a system of some kind spontaneously (without help) moving from a higher energy state to a lower energy state. (Hint: If energy is being released, it means the entities involved in the process are moving to a lower energy state.)

² Note that the term *law* is simply an obsolete term for what we call a theory. For historical reasons, the term is still in use.

3. Describe two examples, other than those in this Introduction, of a system that will move to a lower energy state if allowed to, but which needs an initial boost of energy to get started (like the ball in Figure I.5 being kicked and then having enough energy to get out of the valley).
4. Describe two examples of processes in which entropy *decreases*. In each case, describe what source of energy and/or intelligence must be present for the decrease in entropy to occur. Here is an example to assist your thinking: an oxygen tank contains pressurized oxygen gas. The oxygen in this tank is more ordered than the oxygen in air because it has been separated from the air; there is a boundary (the tank) between the oxygen and the air. And if the valve on the tank is opened, the oxygen flows out into the air to increase the entropy (disorder). What we will never see: opening the tank valve and oxygen atoms from the atmosphere spontaneously flow into the tank. But the oxygen is put into the tank somehow, and the process that put it there decreases the entropy of that oxygen.
5. What is the ratio of nitrogen atoms to hydrogen atoms in ammonia molecules? What is the ratio of hydrogen atoms to carbon atoms in propane molecules?
6. Why are water molecules polar and what is the significance of this fact?
7. If oppositely charged objects attract, why can't a free electron and a free proton collide into one another and stick together because of their opposite charges?
8. A hydronium ion is a water molecule that has gained an extra proton. (A proton is identical to a hydrogen ion.) Hydronium ions form spontaneously in water, and are formed in greater quantities any time an acid is poured into water. What is the ratio of hydrogen atoms to oxygen atoms in hydronium ions?
9. What is hydrogen bonding?
10. Distinguish between endothermic and exothermic processes.
11. In a previous course, you may have learned about the "gold foil experiment" conducted by Ernest Rutherford in 1909. (I describe this experiment in Chapter 2.) This experiment led Rutherford to propose that the positive charge in atoms is concentrated in a tiny nucleus in the center of the atom. Think about this experiment and explain why Rutherford had to depend on inference as he interpreted his experimental data.
12. Why doesn't oil dissolve in water?
13. Distinguish between facts, theories, and hypotheses.
14. Explain why it is scientifically inappropriate to say, "no theory is true until it is proven."

Chapter 1

Measurements



The SI unit system—or metric system—underwent a serious overhaul in 2019. Prior to the make-over, the kilogram was still defined by a physical object kept in a vault in France. Now, however, the seven base units in the SI system, represented by the circles in the inner ring in the graphic above, are defined in terms of physical constants (the outer ring) and each other. For example, the meter (m) is defined as the distance light travels in $1/299,792,458$ seconds. That big number in the denominator is the speed of light in meters per second. So, the definition of the meter involves the speed of light (c) and the definition for the second (s). The arrows in the graphic indicate which units and constants affect others. As you see, the definition of the second affects every other unit definition except one.

Objectives for Chapter 1

After studying this chapter and completing the exercises, you should be able to do each of the following tasks, using supporting terms and principles as necessary.

SECTION 1.1

1. Define *matter* and *mass*.
2. Describe the advantages the SI system has over the USCS system for scientific work.
3. State the SI base units for length, mass, and time.
4. For the SI system of units, define the terms *base unit* and *derived unit*.
5. State several examples of base units and derived units in the SI system of units.
6. Use the metric prefixes listed in Table 1.4 from memory with various units of measure to perform unit conversions and solve problems.

SECTION 1.2

7. Given appropriate unit conversion factors, convert the units of measure for given quantities to different units of measure.
8. Given appropriate conversion factors, convert USCS units to SI units and vice versa.
9. Convert units of temperature measurements between °F, °C, and K.

SECTION 1.3

10. Define *accuracy* and *precision*.
11. Name several possibilities for sources of error in experimental measurements.
12. Explain how the significant digits in a measurement relate to the precision of the measurement.
13. Use significant digits correctly to record measurements from digital and analog measurement instruments.
14. Use significant digits correctly to perform computations, including multiplication, division, addition, subtraction, and combinations of these operations.

SECTION 1.4

15. Convert numerical values from standard notation into scientific notation and vice versa.
16. Use scientific notation when recording measurements and performing computations.
17. Calculate the percent difference between an experimental value and a theoretical value or accepted value.

1.1 Science and Measurements

1.1.1 No Measurements, No Science

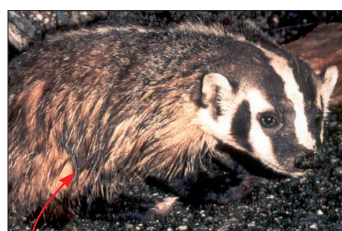
One of the things that distinguishes scientific research from other fields of study is the central role played in science by *measurement*. In every branch of science, researchers study the natural world, and they do it by making measurements. The measurements we make in science are the data we use to quantify the facts we have and to test new hypotheses. These data—our measurements—answer questions such as What is its volume?, How fast is it moving?, What is its mass?, How much time does it take?, What is its diameter?, What is its frequency and wavelength?, When will it occur again?, and many others. Without measurements, modern science would not exist.

Units of measure are crucial in science. Since science is so deeply involved in making measurements, you will work with measurements a lot in this course. The value of a measurement is always accompanied by the units of measure—a measurement without its units of measure is a meaningless number. For this reason, your answers to computations in scientific calculations must always show the units. In this chapter, we discuss units of measure at some length. The material in the chapter is very important—for the rest of this book, we will be engaged with calculations involving units of measure.

However, before we launch into our study of units, there is a topic of great importance we need to nail down—*mass*. Students often do not fully understand what mass is, so we will start by reviewing the topic.

1.1.2 Matter, Volume, and Mass

The best way to understand mass is to begin with *matter* and its properties. The term *matter* refers to anything composed of atoms or parts of atoms. Note that there are many things that are not material, that is, they are not matter, as illustrated in Figure 1.1. Your thoughts, your soul, and your favorite song are not matter. You can write down your thoughts in ink, which is matter, and your song can be recorded onto a CD, which is matter. But ideas and souls are not material and are not made of what we call matter. Another part of this world that is not matter



A badger is matter.

A person's thoughts are not matter.

I love science!



Ideas are not matter.

$$E = mc^2$$



Potassium permanganate is matter.



Light is not matter.

Figure 1.1. Some things are matter and some things are not.

is *electromagnetic radiation*—light, radio waves, X-rays, and all other forms of electromagnetic radiation. Light is pure energy; it is not matter and it has no mass.

We get into electromagnetic radiation a bit in Chapter 3. For now, we are going to focus on matter. A lot of what we discuss in this text is about different properties of matter. There are many different properties to discuss, but here we focus on just two properties that all matter possesses: all matter takes up space and all matter has inertia. Describing and comparing these two properties helps make clear what we mean by the term *mass*.

All matter takes up space. Even individual atoms and protons inside of atoms take up space. Now, how do we *quantify* how much space an object takes up? That is, how do we put a numerical measurement to it? The answer is, of course, by specifying its *volume*. Volume is the name of the variable we use to quantify how much space an object takes up. There are many different units of measure we use to specify an object's volume. Examples are gallons, liters, cubic meters, and pints. When we say that the volume of an object is 338 cubic centimeters, what we mean is that if we could hollow the object out and fill it up with little cubes, each with a volume of one cubic centimeter, 338 of them are needed to fill up the hollowed object.

All matter has inertia. The effect of this property is that objects resist being accelerated. The more inertia an object has, the more difficult it is to accelerate the object. For example, if the inertia of an object is small, as with, say, a golf ball, the object is easy to accelerate. Golf balls are easy to throw, and if you hit one with a golf club it accelerates at a high rate to a very high speed. But if the amount of inertia an object has is large, as with say, a grand piano, the object is difficult to accelerate. Just try throwing a grand piano or hitting one with a golf club and you will see that it doesn't accelerate at all. This is because the piano has a great deal more inertia than a golf ball.

As with the property of taking up space, we need to quantify the property of inertia. The way we do this is with the variable we call *mass*. The mass of an object is a numerical measurement specifying the amount of inertia the object has. Since inertia is a property of matter, and since all matter is composed of atoms, it should be pretty obvious that the more atoms there are packed into an object, the more mass it has. And since the different types of atoms themselves have different masses, an object made of more massive atoms has more mass than an object made of an equal number of less massive atoms.

The main unit of measure we use to specify an object's mass is the *kilogram*. There are other units such as the gram and the microgram. The kilogram (kg) is one of the base units in the metric system, our topic in Section 1.1.4. On the earth, an object weighing 2.2 pounds (lb) has a mass of one kilogram. To give you an idea of what a kilogram mass feels like in your hand on the earth, the lantern battery pictured in Figure 1.2 weighs 2.2 lb, and thus has a mass of 1 kg.

We have established that the mass of an object is a measure of its inertia, which in turn depends on how many atoms it is composed of and how massive those atoms are. The implication of this is that an object's mass does not depend on where it is. A golf ball on the earth has the same mass as a golf ball at the bottom of the ocean, on the moon, or in outer space. Even where there is no gravity, the mass of the golf ball is the same. This is what distinguishes the *mass* of an object from its *weight*.

Weight is caused by the force of gravity acting on an object composed of matter (which we often simply refer to as *a mass*). The weight of an object depends on where it is. An object—or mass—on the moon weighs only about 1/6 its weight on earth, and in outer space, where there is no gravity, a mass has no weight at all. But the mass of an object



Figure 1.2. The mass of this battery is about one kilogram.

does not depend on where it is. This is because an object's mass is based on the matter the object is made of. The lantern battery in the figure has a certain weight on the earth (2.2 lb). In outer space, it weighs nothing and floats right in front of you. But if you try to throw the battery, the force you feel on your hand is the same on the earth or in space. That's because the force you feel depends on the object's mass.

Here is a summary using slightly different terminology that may help even more. Inertia is a *quality* of all matter; mass is the *quantity* of a specific portion of matter. Inertia is a quality or property all matter possesses. Mass is a quantitative variable, and it specifies an amount of matter, a quantity of matter.

1.1.3 The US Customary System

The two major systems of units students should know about are the *International System of Units*, known as the *SI system* or the metric system, and the *U.S. Customary System*, or USCS. You have probably studied these systems before and should be already familiar with some of the SI units and prefixes. In this course, we do not make much use of the measurement system you are most familiar with—the USCS. For scientific work, the entire international scientific community uses the SI system. But here I address the USCS briefly before moving on.

Americans are generally comfortable with measurements in feet, miles, gallons, inches, and degrees Fahrenheit because they grow up using this system and are very familiar with it. But in fact, the USCS is rather cumbersome. One problem is that there are many different units of

measure for every kind of physical quantity. Just for measuring length or distance, for example, we have the inch, foot, yard, and mile. The USCS is also full of random numbers such as 3, 12, and 5,280. A third problem is that there is no inherent connection between units for different types of quantities. Gallons have nothing whatsoever to do with feet, and quarts have nothing to do with miles.

The USCS may be familiar ground, and it may even feel patriotic to prefer it, but it is not the system of measurement scientists use. Scientists everywhere use the SI, and it is to that system we now turn.

Unit	Symbol	Quantity
meter	m	length
kilogram	kg	mass
second	s	time
ampere	A	electric current
kelvin	K	temperature
candela	Cd	luminous intensity
mole	mol	amount of substance

Table 1.1. The seven base units in the SI unit system.

1.1.4 The SI Unit System

In contrast to the USCS, the SI system is simple and has many advantages. There is usually only one basic unit for each kind of quantity, such as the meter for measuring length. Instead of having many unrelated units of measure for measuring quantities of different sizes, prefixes based on powers of ten are used on all the units to accommodate various sizes of quantities. And

units for different types of quantities relate to one another in some way. Unlike the gallon and the foot, which have nothing to do with each other, the cubic meter is 1,000,000 cubic centimeters. For all these reasons, the USCS is not used much at all in scientific work. The SI system is the international standard.

There are seven *base units* in the SI System, listed in Table 1.1. All other SI units of measure,

Unit	Symbol	Quantity
joule	J	energy
newton	N	force
cubic meter	m ³	volume
watt	W	power
pascal	Pa	pressure

Table 1.2. Some SI System derived units.

such as the joule (J) for measuring quantities of energy and the newton (N) for measuring amounts of force, are based on these seven base units. Units based on combinations of the seven base units are called *derived units*. A few common derived units are listed in Table 1.2.

You are already familiar with the SI unit for time: the second. You may or may not be familiar with some of the other base units, so here are some facts and photos to help familiarize you with these. A meter is just a few inches longer than a yard (3 feet). Figure 1.3 shows a wooden measuring rule one meter long, commonly called a meter stick, along with a metal yardstick for comparison.

On earth, a mass of one kilogram weighs about 2.2 pounds. The six-volt lantern battery shown in Figure 1.2 weighs just under 2.2 pounds, so the mass of the battery is just about one kilogram, as I mentioned before.

I can't show you a picture of one ampere of electric current, but it may be helpful to know that a standard electrical receptacle (or "outlet") such as the one shown in Figure 1.4 is rated to carry 15 amperes of current. (However, the largest continuous current that the receptacle is allowed to supply is 80% of its rating, or 12 amperes. This is why vacuum cleaners are often advertised as having 12-amp motors. That's the upper limit of the current available to run them.)

Regarding temperature units, the Celsius scale is generally used for making scientific temperature measurements, but the Kelvin scale must be used for nearly all scientific calculations involving temperature. You must become familiar with both scales.

On the Celsius scale, water freezes at 0°C and boils at 100°C . Since Celsius temperature measurements can take on negative values, the Celsius scale (like the Fahrenheit scale) is not an absolute temperature scale. The Kelvin scale is an absolute scale, with 0 K (0 kelvins) being equal to absolute zero, theoretically the lower limit of possible temperatures. Note that the term "degrees" is not used when stating or writing values in kelvins.

A temperature change of one kelvin is the same as a temperature change of one degree Celsius, and both are almost double the change that a change of one degree Fahrenheit is. For reference, room temperature on the three scales is 72°F , 22.2°C , and 295.4 K .

All seven of the SI base units are defined in terms of physical constants, each one of which is defined with its own exact value. For example, the speed of light in a vacuum is defined to be 299,792,458 meters per second. The meter is defined from this: the distance light travels in $1/299,792,458$ seconds. (And the second has its own definition.) Formerly, units such as the kilogram and the meter were defined by man-made physical objects (artifacts), such as the one shown in Figure 1.5. The bar shown in the figure was the standard in the U.S. for the meter from 1893 to 1960. But this method of definition was not at all convenient.



Figure 1.3. A meter stick (left), with a yardstick for comparison.



Figure 1.4. A standard receptacle in American homes is rated for a current of 15 amperes.



Figure 1.5. Standard meter bar number 27, owned by the U.S. and used as the standard meter from 1893 to 1960.

Multiples	Prefix	deca-	hecto-	kilo-	mega-	giga-	tera-	peta-	exa-	zetta-	yotta-
	Symbol	da	h	k	M	G	T	P	E	Z	Y
	Factor	10	10 ²	10 ³	10 ⁶	10 ⁹	10 ¹²	10 ¹⁵	10 ¹⁸	10 ²¹	10 ²⁴
Fractions	Prefix	deci-	centi-	milli-	micro-	nano-	pico-	femto-	atto-	zetto-	yocto-
	Symbol	d	c	m	μ	n	p	f	a	z	y
	Factor	1/10	1/10 ²	1/10 ³	1/10 ⁶	1/10 ⁹	1/10 ¹²	1/10 ¹⁵	1/10 ¹⁸	1/10 ²¹	1/10 ²⁴

Table 1.3. The SI System prefixes.

In 1960, the definition of the meter was changed so that the meter was equal to a certain number of wavelengths of a certain color of light emitted by a certain isotope of the element krypton. In 1983, the definition of the meter was changed again to its present definition in terms of the speed of light and the second. The kilogram has a similar history, and its definition in terms of an artifact only ended in 2019.

The SI system includes not only the base and derived units, but all the other units that can be formed by adding metric prefixes to these units. We address the prefixes in the next section. But first I will mention a particular subset of the SI system known as the MKS system. MKS stands for meter-kilogram-second. The MKS system uses only the base and derived units without the prefixes (except for the kilogram, the only base unit with a prefix). The nice thing about the MKS system is that any calculation performed with MKS units produces a result in MKS units. For this reason, the MKS system is used almost exclusively in physics. However, in chemistry, it is common to use SI units that are not MKS units. Some commonly used non-MKS units are the gram (g), the centimeter (cm), the cubic centimeter (cm³), the liter (L), and the milliliter (mL). The liter is not actually an official SI unit, but it is used all the time in chemistry anyway.

1.1.5 Metric Prefixes

In the system of units commonly used in the U.S., different units are used for different sizes of objects. For example, for short distances we might use the inch or the foot, whereas for longer distances we switch to the mile. For the small volumes used in cooking, we use the fluid ounce (or pint, quart, teaspoon, tablespoon, etc.), but for larger volumes like the gasoline in the gas tank of a car, we switch to the gallon. (That's six different volume units I just listed!)

The SI System is much simpler. Each type of quantity—such as length or volume—has one main unit of measure. Instead of using several different units for different sizes of quantities, the SI System uses multipliers on the units to multiply them for large quantities, or to scale them down for smaller quantities. We call these multipliers the *metric prefixes*. The complete list of the 20 metric prefixes is in Table 1.3. You do not need to memorize all these; some are rarely used. But you do need to memorize some of them. I recommend that all science students commit to

memory the prefixes listed in Table 1.4.

Table 1.5 shows a few representative examples of how to use the prefixes to represent multiples (quantities larger than the SI base unit)

Fractions			Multiples		
Prefix	Symbol	Factor	Prefix	Symbol	Factor
centi-	c	1/10 ²	kilo-	k	10 ³
milli-	m	1/10 ³	mega-	M	10 ⁶
micro-	μ	1/10 ⁶	giga-	G	10 ⁹
nano-	n	1/10 ⁹	tera-	T	10 ¹²
pico-	p	1/10 ¹²			

Table 1.4. SI System prefixes to commit to memory.

	Prefix	Symbol	Meaning	Examples of usage
Multiples	kilo-	k	1,000	One kilojoule is 1,000 joules. There are 1,000 joules in one kilojoule, so $1,000 \text{ J} = 1 \text{ kJ}$.
	mega-	M	1,000,000	One megawatt is 1,000,000 watts. There are 1,000,000 watts in one megawatt, so $1,000,000 \text{ W} = 1 \text{ MW}$.
Fractions	centi-	c	1/100	One centimeter is 1/100 of a meter. There are 100 centimeters in one meter, so $100 \text{ cm} = 1 \text{ m}$.
	milli-	m	1/1,000	One milligram is 1/1,000 of a gram. There are 1,000 milligrams in one gram, so $1,000 \text{ mg} = 1 \text{ g}$.
	micro-	μ	1/1,000,000	One microliter is 1/1,000,000 of a liter. There are 1,000,000 microliters in one liter, so $1,000,000 \mu\text{L} = 1 \text{ L}$.

Table 1.5. Examples of correct usage of metric prefixes.

and fractions (quantities smaller than the SI base unit). As illustrations, let's look at a couple of these more closely.

The prefix *kilo-* is a *multiple*, and it means 1,000. One kilogram is 1,000 grams, one kilometer is 1,000 meters, and so on. Figure 1.6 illustrates with my favorite dairy product—chocolate chip cookie dough ice cream. One gram is only a fraction of a taste (and contains only two chocolate chips and no cookie dough). A kilogram of ice cream is 1,000 grams of ice cream, equivalent to two large bowls of ice cream.

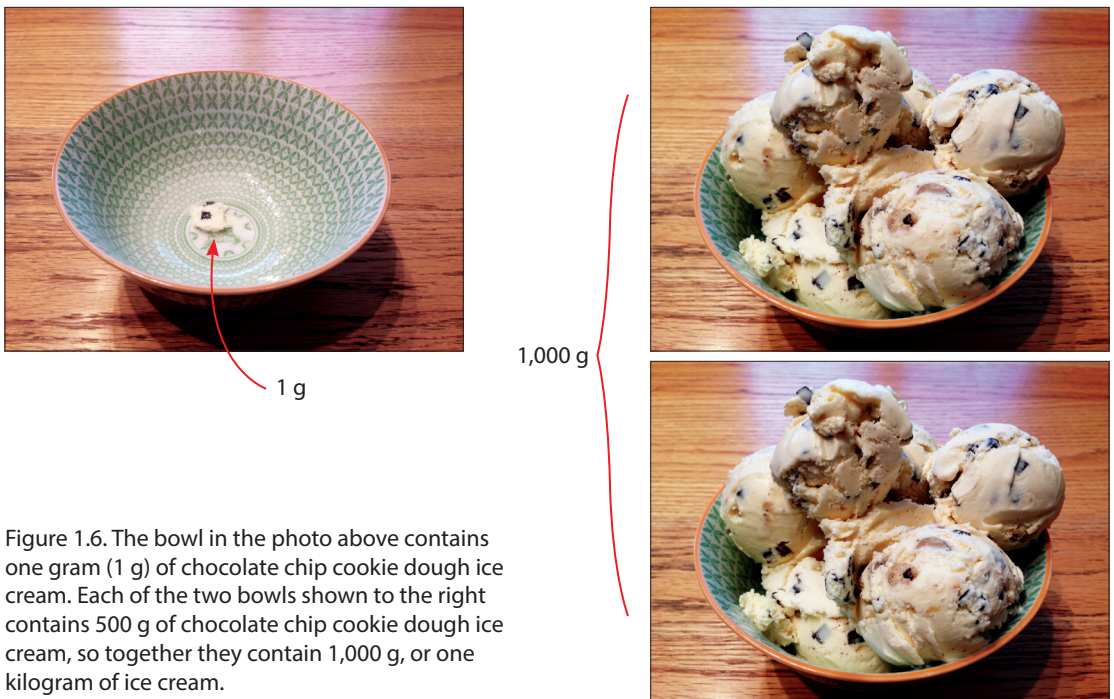


Figure 1.6. The bowl in the photo above contains one gram (1 g) of chocolate chip cookie dough ice cream. Each of the two bowls shown to the right contains 500 g of chocolate chip cookie dough ice cream, so together they contain 1,000 g, or one kilogram of ice cream.

The prefix *milli-* is a *fraction*, and it means one thousandth. One millimeter is one thousandth of a meter, and so on. The wooden rule in Figure 1.3 is one meter in length. A millimeter is one thousandth of this length, equal to the width of the line in Figure 1.7.

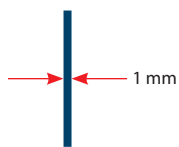


Figure 1.7. One millimeter, which is one thousandth of a meter.

We conclude this introduction to metric prefixes with a few brief notes. First, when using the prefixes for quantities of mass, prefixes are never added to the kilogram. Prefixes are only added to the gram, even though the kilogram is the base unit in the SI system, not the gram. Second, note that when writing the symbols for metric prefixes, the case of the letter matters: *kilo-* always takes a lower-case *k*, *mega-* always takes an upper-case *M*, and so on. Third, one of the prefix symbols is not an English letter. The prefix μ for *micro-* is the lower-case Greek letter *mu*, the *m* in the Greek alphabet. Finally, pay close attention to the difference between multiplier prefixes and fraction prefixes. Learning to use the fraction prefixes properly is the most challenging part of mastering the SI System of units, and using them incorrectly in unit conversion factors (our next topic) is a common student error.

1.2 Converting Units of Measure

1.2.1 Basic Principles of Unit Conversion Factors

For scientists and engineers, one of the most commonly used skills is re-expressing quantities into equivalent quantities with different units of measure. These calculations are called *unit conversions*. Mastery of this skill is essential for all students studying science. In this section, I describe what unit conversion factors are and how they are used.

Let's begin with the basic principles of how unit conversions work. First, we all know that multiplying any value by *unity* (one) leaves its value unchanged. Second, we also know that in any fraction, if the numerator and denominator are equivalent, the value of the fraction is unity (one). For example, the expression “12 bricks over 12 bricks” is equal to one:

$$\frac{12 \text{ bricks}}{12 \text{ bricks}} = 1$$

This is because the numerator and denominator are equivalent, and any time this is the case, the value of the fraction is one, or unity.

A *unit conversion factor* is simply a fractional expression in which the numerator and denominator are equivalent ways of writing the same physical quantity with different units of measure. This means a conversion factor is just a special way of writing unity (one).

The third basic principle is that when multiplying fractions, factors that appear in both the numerator and denominator may be “cancelled out.” So when performing ordinary unit conversions, what we are doing is repeatedly multiplying a given quantity by unity so that cancellations alter the units of measure until they are expressed the way we wish. Since all we are doing is multiplying by one, the value of our original quantity is unchanged; it simply looks different because it is expressed with different units of measure.

There are many different units of measure and there are many different conversion factors used for performing unit conversions. Table A.3 in Appendix A lists a number of important ones.

Let me elaborate a bit more on the idea of unity I mentioned above, using one common conversion factor as an example. American school kids all learn that there are 5,280 feet in one mile, which means $5,280 \text{ ft} = 1 \text{ mi}$. One mile and 5,280 feet are equivalent ways of writing the same length. If we place these two expressions into a fraction, the numerator and denominator

are equivalent, so the value of the fraction is unity, regardless of the way we write it. The equation $5,280 \text{ ft} = 1 \text{ mi}$ can be written as a conversion factor two different ways, and the fraction equals unity either way:

$$\frac{5280 \text{ ft}}{1 \text{ mi}} = \frac{1 \text{ mi}}{5280 \text{ ft}} = 1 \quad (1.1)$$

Now, I need to make an important clarification about the use of the equal sign in the expressions I just wrote. In mathematics, the equal sign means *identity*. An expression such as 3 miles = 3 miles is a mathematical *identity*; the first expression, “3 miles” is *identical* to the second expression, also “3 miles.” But when we are dealing with converting units of measure from one set of units to a different set of units, we don’t use the equal sign to mean identity. Obviously, the expression

$$\frac{5280 \text{ ft}}{1 \text{ mi}}$$

is not *identical* to the expression

$$\frac{1 \text{ mi}}{5280 \text{ ft}}$$

One of these has units of ft/mi and the other has units of mi/ft. But even though these expressions are not identical, they are *equivalent*. When we are dealing with converting units of measure, this is the sense in which we interpret the equal sign. We are using it to mean *equivalent*. This is why I can write Equation (1.1) using equal signs. The three terms in Equation (1.1) are not identical, but they are equivalent.

Suppose you have a measurement such as 43,000 feet that you wish to re-express in miles. To convert the units from feet to miles, first write down the quantity you are given, with its units of measure:

$$43,000 \text{ ft}$$

Next, select a unit conversion factor containing the units you presently have and the ones you want to convert to. (This is not always possible. Sometimes more than one conversion factor is required, as Example 1.1 below illustrates.) In this case, those units are feet and miles, and the conversion factors containing these units are the two written in Equation (1.1). As I explain below, the one we need for converting 43,000 ft into miles is the second one. So to perform the conversion, you multiply your given quantity by the conversion factor. Then you cancel any units that appear both in the numerator and the denominator, as follows:

$$43,000 \text{ ft} \cdot \frac{1 \text{ mi}}{5280 \text{ ft}} = 8.1 \text{ mi}$$

There are two important comments to make here. First, since any conversion factor can be written two ways (depending on which quantity is placed in the numerator), how do we know which way to write the conversion factor? Well, we know from algebra that when we have quantities in the numerator of a fraction that are multiplied (factors), and quantities in the denominator of the fraction that are multiplied (factors), any quantities that appear in both the numerator and denominator can be cancelled out. In the example above, we want to cancel out the “feet” in the given quantity (which is in the numerator), so the conversion factor needs to be written with feet in the denominator and miles in the numerator.

Second, if you perform the calculation above ($43,000 \div 5,280$), the result that appears on your calculator screen is 8.143939394. So why didn't I write down all those digits in my result? Why did I round my answer off to simply 8.1 miles? The answer to that question has to do with the significant digits in the value 43,000 ft that we started with. We address the issue of significant digits later in this chapter, but in the examples that follow I always write the results with the correct number of significant digits for the values involved in the problem.

The following example illustrates the use of conversion factors based on metric prefixes. This example also illustrates how to perform a conversion when more than one conversion factor is required.

▼ Example 1.1

Convert the value 2,953,000 μg into kilograms.

Referring to Tables 1.4 and 1.5, you see that the symbol μ means *micro-*, which means one millionth. Thus, μg means millionths of a gram. We use this information to make conversion factors. Since it takes 1,000,000 millionths of a gram to make one gram, $1,000,000 \mu\text{g} = 1 \text{ g}$, and thus

$$\frac{1,000,000 \mu\text{g}}{1 \text{ g}} = \frac{1 \text{ g}}{1,000,000 \mu\text{g}} = 1$$

Note that converting from μg to g only gets us part of the way toward the solution. We need another conversion factor to get from g to kg. Looking again at Tables 1.4 and 1.5, we see that there are 1,000 grams in one kilogram, or $1,000 \text{ g} = 1 \text{ kg}$. From this we can make additional conversion factors:

$$\frac{1000 \text{ g}}{1 \text{ kg}} = \frac{1 \text{ kg}}{1000 \text{ g}} = 1$$

Now to perform the conversion, first convert micrograms into grams. Then convert grams into kilograms. You can do this two-step conversion at the same time by simply multiplying both conversion factors at the same time as follows:

$$2,953,000 \mu\text{g} \cdot \frac{1 \text{ g}}{1,000,000 \mu\text{g}} \cdot \frac{1 \text{ kg}}{1000 \text{ g}} = 0.002953 \text{ kg}$$

The μg in the given quantity cancels with the μg in the denominator of the first conversion factor. The g in the first conversion factor cancels with the g in the denominator of the second conversion factor. The units we are left with are the kg in the numerator of the second conversion factor. The kg did not cancel out with anything, so these are the units of our result.



1.2.2 Tips for Converting Units of Measure

There are several important points you must remember in order to perform unit conversions correctly. I illustrate them below with examples. You should rework each of the examples on your own paper as practice to make sure you can do them correctly. The conversion factors used in the examples below are all listed in Table A.3 in Appendix A.

Point 1 *Never use slant bars in your unit fractions. Use only horizontal bars.*

In printed materials, we often see values written with a slant fraction bar in the units, as in the value 35 m/s. Although writing the units this way is fine for a printed document, you should not write values this way when you are performing unit conversions. This is because it is easy to get confused and not notice that one of the units is in the denominator in such an expression (s, or seconds, in my example), and the conversion factors used must take this into account.

▼ **Example 1.2**

Convert 57.66 mi/hr into m/s.

Writing the given quantity with a horizontal bar makes it clear that the “hours” is in the denominator. This helps you write the hours-to-seconds factor correctly.

To perform this conversion, we must convert the miles in the given quantity into meters, and we must convert the hours into seconds. From Table A.3, we select the two conversion factors we need. Then we multiply the given quantity by them to convert the mi/hr into m/s. When doing the multiplying, we write all the unit fractions with horizontal bars.

$$57.66 \frac{\text{mi}}{\text{hr}} \cdot \frac{1609 \text{ m}}{1 \text{ mi}} \cdot \frac{1 \text{ hr}}{3600 \text{ s}} = 25.77 \frac{\text{m}}{\text{s}}$$

As you see, the miles cancel and the hours cancel, leaving meters in the numerator and seconds in the denominator. Now that you have your result, you may write it as 25.77 m/s if you wish, but do not use slant fraction bars in the units when you are working out the unit conversion.



Point 2 *The term “per,” abbreviated p, implies a fraction.*

Some units of measure are commonly written with a “p” for “per,” such as mph for miles per hour, or gps for gallons per second. Change these expressions to fractions with horizontal bars when you work out the unit conversion.

▼ **Example 1.3**

Convert 472.2 gps to L/hr.

When you write down the given quantity, change the gps to gal/s and write these units with a horizontal bar:

$$472.2 \frac{\text{gal}}{\text{s}} \cdot \frac{3.785 \text{ L}}{1 \text{ gal}} \cdot \frac{3600 \text{ s}}{\text{hr}} = 6,436,000 \frac{\text{L}}{\text{hr}}$$



Point 3 *Use the $\boxed{\times}$ and $\boxed{\div}$ keys correctly when entering values into your calculator.*

When dealing with several numerator terms and several denominator terms, multiply all the numerator terms together first, hitting the $\boxed{\times}$ key between each, then hit the $\boxed{\div}$ key and enter

all the denominator terms, hitting the [=] key between each. This way you do not need to write down intermediate results, and you do not need to use any parentheses.

▼ Example 1.4

Convert 43.2 mm/hr into km/yr.

The setup with all the conversion factors is as follows:

$$43.2 \frac{\text{mm}}{\text{hr}} \cdot \frac{1 \text{ m}}{1000 \text{ mm}} \cdot \frac{1 \text{ km}}{1000 \text{ m}} \cdot \frac{24 \text{ hr}}{1 \text{ dy}} \cdot \frac{365 \text{ dy}}{1 \text{ yr}} = 0.378 \frac{\text{km}}{\text{yr}}$$

To execute this calculation in your calculator, you enter the values and operations in this sequence:

$$43.2 \times 24 \times 365 \div 1000 \div 1000 =$$

If you do so, you get 0.37843200. (Again, significant digits rules require us to round to 0.378.)



Point 4

When converting units for area and volume such as cm^2 or m^3 , you must use the appropriate length conversion factor twice for areas and three times for volumes.

The units “ cm^2 ” for an area mean the same thing as “ $\text{cm} \times \text{cm}$.” Likewise, “ m^3 ” means “ $\text{m} \times \text{m} \times \text{m}$.” So when you use a length conversion factor such as $100 \text{ cm} = 1 \text{ m}$ or $1 \text{ in} = 2.54 \text{ cm}$, you must use it twice to get squared units (areas) or three times to get cubed units (volumes).

▼ Example 1.5

Convert $3,550 \text{ cm}^3$ to m^3 .

$$3550 \text{ cm}^3 \cdot \frac{1 \text{ m}}{100 \text{ cm}} \cdot \frac{1 \text{ m}}{100 \text{ cm}} \cdot \frac{1 \text{ m}}{100 \text{ cm}} = 0.00355 \text{ m}^3$$

Notice in this example that the unit cm occurs three times in the denominator, giving us cm^3 when they are all multiplied together. This cm^3 term in the denominator cancels with the cm^3 term in the numerator. And since the m unit occurs three times in the numerator, they multiply together to give us m^3 for the units in our result.



The issue of needing to repeat conversion factors only arises when you are using a unit raised to a power, such as a when a length unit is used to represent an area or a volume. When using a conversion factor such as $3.785 \text{ L} = 1 \text{ gal}$, the units of measure are written using units that are strictly volumetric (liters and gallons), and are not obtained from lengths the way they are with in^2 , ft^2 , cm^3 , and m^3 . Another common unit that uses an exponent is acceleration, which has units of m/s^2 in the MKS unit system.

▼ Example 1.6

Convert 5.85 mi/hr^2 into MKS units.

The MKS unit for length is meters (m), so we must convert miles to meters. The MKS units for time is seconds (s), so we must convert hr^2 into s^2 .

$$5.85 \frac{\text{mi}}{\text{hr}^2} \cdot \frac{1609 \text{ m}}{1 \text{ mi}} \cdot \frac{1 \text{ hr}}{3600 \text{ s}} \cdot \frac{1 \text{ hr}}{3600 \text{ s}} = 0.000726 \frac{\text{m}}{\text{s}^2}$$

With this example, you see that since the “hours” unit is squared in the given quantity, the conversion factor converting the hours to seconds must appear twice in the conversion calculation. The “miles” unit in the given quantity has no exponent, so the conversion factor used to convert miles to meters only appears once in the calculation.



1.2.3 Converting Temperature Units

Converting temperature values from one scale to another requires the use of equations rather than conversion factors. This is due to the fact that the Fahrenheit and Celsius scales are not absolute temperature scales. If all temperature scales were absolute scales like the Kelvin scale is, temperature conversions could be performed with conversion factors just as other conversions are. To convert a temperature in degrees Fahrenheit (T_F) into degrees Celsius (T_C), use this equation:

$$T_C = \frac{5}{9}(T_F - 32^\circ)$$

Using a bit of algebra, we can work this around to give us an equation that can be used to convert Celsius temperatures to Fahrenheit values:

$$T_F = \frac{9}{5}T_C + 32^\circ$$

To convert a temperature in degrees Celsius into kelvins (T_K), use this equation:

$$T_K = T_C + 273.15$$

Again, some algebra gives us the equation the other way around.

$$T_C = T_K - 273.15$$

All four of the temperature conversion equations above are exact. This is important to know later when we discuss significant digits. These equations are listed in Table A.3 in Appendix A.

▼ Example 1.7

The normal temperature of the human body is 98.6°F . Express this value in degrees Celsius and kelvins.

Since the given value is in degrees Fahrenheit, write down the equation that converts values from $^\circ\text{F}$ to $^\circ\text{C}$.

$$T_C = \frac{5}{9}(T_F - 32^\circ)$$

Now insert the Fahrenheit value and calculate the Celsius value.

$$T_C = \frac{5}{9}(98.6^\circ - 32^\circ) = 37.0^\circ\text{C}$$

Now we are able to use the Celsius value to compute the Kelvin value.

$$T_K = T_C + 273.15 = 37.0 + 273.15 = 310.2 \text{ K}$$

The reason the answer is 310.2 K instead of 310.15 K is due to the significant digits rule for addition. Again, the topic of significant digits is coming up next.



▼ Example 1.8

The melting point of aluminum is 933.5 K. Express this temperature in degrees Celsius and degrees Fahrenheit.

Write down the equation that converts Kelvin values to Celsius values.

$$T_C = T_K - 273.15$$

From this we calculate the Celsius value as

$$T_C = T_K - 273.15 = 933.5 - 273.15 = 660.3^\circ\text{C}$$

Next, write down the equation for converting a Celsius value to a Fahrenheit value.

$$T_F = \frac{9}{5}T_C + 32^\circ$$

Insert the Celsius value into this equation and calculate the Fahrenheit value.

$$T_F = \frac{9}{5}T_C + 32^\circ = \frac{9}{5} \cdot 660.3^\circ\text{C} + 32^\circ = 1220.5^\circ\text{F}$$



1.3 Accuracy and Precision

1.3.1 Distinguishing Between Accuracy and Precision

The terms *accuracy* and *precision* refer to the practical limitations inherent in making measurements. Science is all about investigating nature, and to do that we must make measurements.

Accuracy relates to error—that is, to the lack of it. Error is the difference between a measured value and the true value. The lower the error is in a measurement, the better the accuracy. Error arises from many different sources, including human mistakes, malfunctioning equipment, incorrectly calibrated instruments, vibrations, changes in temperature or humidity, or unknown causes that are influencing a measurement without the knowledge of the experimenter. All measurements contain error, because (alas!) perfection is simply not a thing we have access to in this world.

Precision refers to the resolution or degree of “fine-ness” in a measurement. The limit to the precision that can be obtained in a measurement is ultimately dependent on the instrument being used to make the measurement. If you want greater precision, you must use a more precise instrument. The degree of precision in every measurement is signified by the measurement value itself because the precision is a built-in part of the measurement. *The precision of a measurement is indicated by the number of significant digits (or significant figures) included in the measurement value when the measurement is written down* (see below).

Here is an example that illustrates the idea of precision and also helps distinguish between precision and accuracy. The photograph in Figure 1.8 shows a machinist’s rule and an architect’s scale set one above the other. Since the marks on the two scales line up consistently, *these two scales are equally accurate*. But the machinist’s rule (on top) is more precise. The architect’s scale is marked in 1/16-inch increments, but the machinist’s rule is marked in 1/64-inch increments. Thus, *the machinist’s rule is more precise*.

It is important that you are able to distinguish between accuracy and precision. Here is another example to help illustrate the difference. Let’s say Shana and Marius each buy digital thermometers for their homes. The thermometer Shana buys costs \$10 and measures to the nearest 1°F. Marius pays \$40 and gets one that reads to the near-

est 0.1°F. Shana reads the directions and properly installs the sensor for her new thermometer in the shade. Marius doesn’t read the directions and mounts his sensor in the direct sunlight, which causes a significant error in the thermometer reading when the sun is shining on it; thus Marius’ measurements are not accurate. The result is that Shana has lower-precision, higher-accuracy measurements!

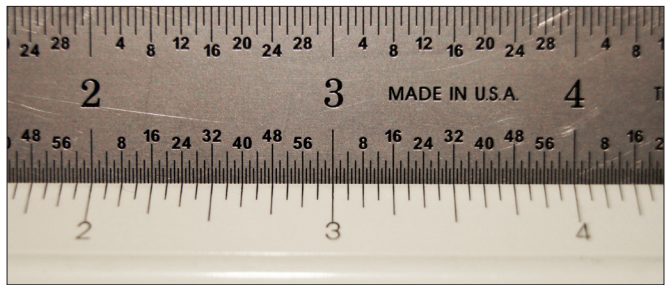


Figure 1.8. The accuracy of these two scales is the same, but the machinist’s rule (above) is more precise than the architect’s scale (below).

1.3.2 Significant Digits

The precision in any measurement is indicated by the number of *significant digits* it contains. Thus, the number of digits we write in any measurement we deal with in science is very important. The number of digits is meaningful because it shows the precision inherent in the instrument used to make the measurement.

Let’s say you are working a computational exercise in a science book. The problem tells you that a person drives a distance of 110 miles at an average speed of 55 miles per hour and wants you to calculate how long the trip takes. The correct answer to this problem *is different* from the correct answer to a similar problem with given values of 110.0 miles and 55.0 miles per hour. And if the given values are 110.0 miles and 55.00 miles per hour, the correct answer is different yet again. Mathematically, of course, all three answers are the same. If you drive 110 miles at 55 miles per hour, the trip takes two hours. But scientifically, the correct answers to these three problems are different: 2.0 hours, 2.00 hours, and 2.000 hours, respectively. The difference between these cases is in the precision indicated by the given data, which are *measurements*. (Even though this is just a made-up problem in a book and not an actual measurement someone made in an experiment, the given data are still measurements. There is no way to talk about distances or speeds without talking about measurements, even if the measurements are only imaginary or hypothetical.)

So when you perform a calculation with physical quantities (measurements), you can't simply write down all the digits shown by your calculator. The precision inherent in the measurements used in a computation governs the precision in any result you might calculate from those measurements. And since the precision in a measurement is indicated by the number of significant digits, data and calculations must be written with the correct numbers of significant digits. To do this, you need to know how to count significant digits, and you must use the correct number of significant digits in all your calculations and experimental data.

Correctly counting significant digits involves four different cases:

1. Rules for determining how many significant digits there are in a given measurement.
2. Rules for writing down the correct number of significant digits in a measurement you are making and recording.
3. Rules for computations you perform with measurements—multiplication and division.
4. Rules for computations you perform with measurements—addition and subtraction.

We address each of these cases below, in order.

Case 1

We begin with the rule for determining how many significant digits there are in a given measurement value. The rule is as follows:

- *The number of significant digits in a number is found by counting all the digits from left to right beginning with the first nonzero digit on the left. When no decimal is present, trailing zeros are not considered significant.*

Let's apply this rule to the following values to see how it works.

15,679 This value has five significant digits.

21.0005 This value has six significant digits.

37,000 This value has only two significant digits because when there is no decimal, trailing zeros are not significant. Notice that the word *significant* here is a reference to the precision of the measurement, which in this case is rounded to the nearest thousand. The zeros in this value are certainly *important*, but they are not *significant* in the context of precision.

0.0105 This value has three significant digits because we start counting digits with the first nonzero digit on the left.

0.001350 This value has four significant digits. Trailing zeros count when there is a decimal.

The significant digit rules enable us to tell the difference between two measurements such as 13.05 m and 13.0500 m. Again, these values are obviously equivalent *mathematically*. But they are different in what they tell us about the process of how the measurements were made—and science deals in measurements. The first measurement has four significant digits. The second measurement is more precise—it has six significant digits and was made with a more precise instrument.

Now, just in case you are bothered by the zeros at the end of 37,000 that are not significant, here is one more way to think about significant digits that may help. The precision in a measurement depends on the instrument used to make the measurement. If we express the measure-

ment in different units, this cannot change the precision of the value. A measurement of 37,000 grams is equivalent to 37 kilograms, as shown in the following calculation:

$$37,000 \text{ g} \cdot \frac{1 \text{ kg}}{1000 \text{ g}} = 37 \text{ kg}$$

Whether we express this value in grams or kilograms, it still has two significant digits.

Case 2

The second case addresses the rules that apply when you are recording a measurement yourself, rather than reading a measurement someone else has made. When you make measurements yourself, as when conducting the laboratory experiments in this course, you must know the rules for which digits are significant in the reading you are making on the measurement instrument. The rule for making measurements depends on whether the instrument you are using is a digital instrument or an analog instrument. Here are the rules for these two possibilities:

- *Rule 1 for digital instruments* *For the digital instruments commonly found in introductory science labs, assume all the digits in the reading are significant except leading zeros.*
- *Rule 2 for analog instruments* *The significant digits in a measurement include all the digits known with certainty, plus one digit at the end that is estimated between the finest marks on the scale of your instrument.*

The first of these rules is illustrated in Figure 1.9. The reading on the left has leading zeros, which do not count as significant. Thus, the first reading has three significant digits. The second reading also has three significant digits. The third reading has five significant digits.

The fourth reading also has five significant digits because with a digital display the only zeros that don't count are the leading zeros. Trailing zeros are significant with a digital instrument. *However*, when you write this measurement down, you must write it in a way that shows those zeros to be significant. The way to do this is by using scientific notation. When a value is written in scientific notation, *the digits that are written down in front of the power of 10* (the stem, also called the mantissa) *are the significant digits*. Thus, the right-hand value in Figure 1.9 must be written as 4.2000×10^4 . We address scientific notation in more detail in the next section.

Dealing with digital instruments is actually more involved than the simple rule above implies, but the issues involved go way beyond what we can deal with in introductory science classes. So, simply make your readings and assume that all the digits in the reading except leading zeros are significant.

Now let's look at some examples illustrating the rule for analog instruments. Figure 1.10 shows a machinist's rule being used to measure the length in millimeters (mm) of a brass block. We know the first two digits of the length with certainty; the block is clearly between 31 mm and 32 mm long. We have to estimate the third significant digit. The scale on the rule is marked in increments of 0.5 mm. Comparing the edge of the block with these marks, I estimate the next digit to be a 6, giving a measurement of 31.6 mm. Others might estimate the last digit to be a 5 or a 7; these small differences in the last digit are unavoidable because the last digit is estimated. Whatever you estimate the last digit to be, two digits of this measurement are known with certainty, the third digit is estimated, and the measurement has three significant digits.



Figure 1.9. With digital instruments, all digits are significant except leading zeros. Thus, the numbers of significant digits in these readings are, from left to right, three, three, five, and five.

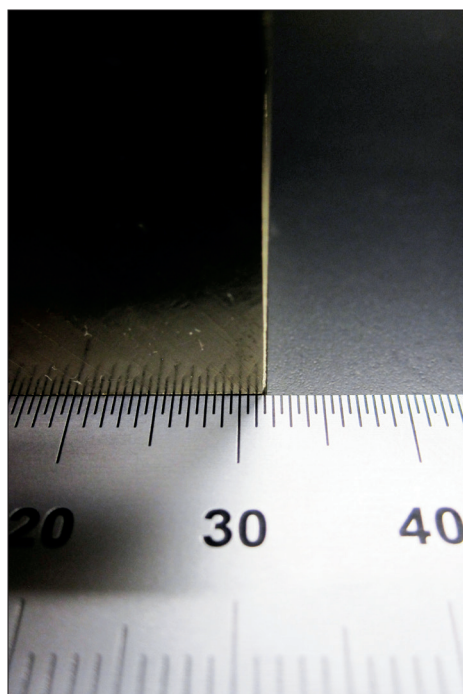


Figure 1.10. Reading the significant digits with a machinist's rule.

The photograph in Figure 1.11 shows a liquid volume measurement in milliliters (mL) being made with an article of apparatus called a *buret*. Notice in this figure that when measuring liquid volume the surface of the liquid curls up at the edge of the cylinder, forming a bowl-shaped surface on the liquid. This curved surface is called a *meniscus*. For most liquids, liquid measurement readings are taken at the bottom of the meniscus. Liquid mercury is the major exception, because the meniscus in liquid mercury is inverted—liquid mercury curves down at the edges. In that case, the measurement is read at the top of the meniscus. But that is an unusual case. For most liquids, the reading is made at the bottom of the meniscus.

For the buret in the figure, you can see that the scale is marked in increments of 0.1 milliliters (mL). This means we are to estimate to the nearest 0.01 mL. To one person, it may look like the bottom of the meniscus (where the black curve touches the bottom of the silver bowl) is just above 2.2 mL, so that person would call this measurement 2.19 mL. To someone else, it may seem that the bottom of the meniscus is right on 2.2, in which case that person would call the reading 2.20 mL. Either way, the reading has three significant digits and the last digit is estimated to be either 9 or 0.

The third example involves a liquid volume measurement with an article of apparatus called a *graduated cylinder*. The scales on small graduated cylinders like this one are marked in increments of 1 mL. In the photo of Figure 1.12, the entire meniscus appears silvery in color with a black curve at the bottom.

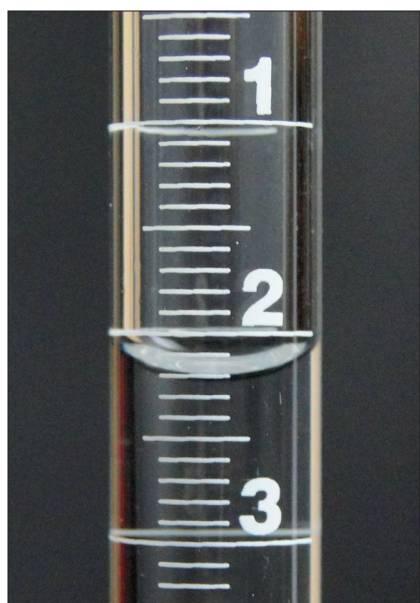


Figure 1.11. Reading the significant digits on a buret.

For the liquid shown in the figure, we know the first two digits of the volume measurement with certainty because the reading at the bottom of the meniscus is clearly between 82 mL and 83 mL. We have to estimate the third digit, and I estimate the edge of the meniscus to be at 60% of the distance between 82 and 83, giving a reading of 82.6 mL. Others may prefer a different value for that third digit.



Figure 1.12. Reading the significant digits on a graduated cylinder.

It is important for you to keep the significant digits rules in mind when you are making measurements and entering data for your lab reports. The data in your lab journal and the values you use in your calculations and report should correctly reflect the use of the significant digits rules as they apply to the actual instruments you use to make your measurements.

Case 3

The third and fourth cases of rules for significant digits apply to the calculations you perform with measurements. In Case 3, we deal with multiplication and division. The main idea behind the rule for multiplying and dividing is that the precision you report in your result cannot be higher than the precision that is in the measurements you start with. The precision in a measurement depends on the instrument used to make the measurement, nothing else. Multiplying and dividing things cannot increase that precision, and thus your results can be no more precise than the measurements used in the calculations. In fact, your result can be no more precise than the *least precise value* used in the calculation. The least precise value is, so to speak, the “weak link” in the chain, and a chain is no stronger than its weakest link.

Here are the two rules for using significant digits in calculations involving multiplication and division:

- *Rule 1* When multiplying or dividing, count the significant digits in each of the values you are using in the calculation, including any conversion factors involved. (However, note: Conversion factors that are exact are not considered.) Determine how many significant digits there are in the least precise of these values. The result of your calculation must have this same number of significant digits.
- *Rule 2* When performing a multi-step calculation, keep at least one extra digit during intermediate calculations, and round off to the final number of significant digits you need at the very end. This practice ensures that small round-off errors don't accumulate during a multi-step calculation. This extra digit rule also applies to unit conversions performed as part of the computation.

I illustrate the two rules above, along with some more unit conversions, in the following example problem and calculation.

▼ Example 1.9

At a chemical research lab, a stream of a reactant solution is flowing into a reaction vessel at a rate of 56.75 μL per second. A volume of 1.0 ft^3 of this solution is required in the vessel for the reaction. Determine the amount of time needed for the required volume to be collected. State your result in hours.

First note that the value of the flow rate has four significant digits, and the required volume has two significant digits. The two-digit value is the least precise of these, so our result must be rounded to two significant digits. But to avoid rounding error, we must work with values having at least three significant digits (one more than we need) until the very end.

One of the volumes in this problem is in μL and the other is in ft^3 . I begin by converting the required volume from ft^3 to μL so our volumes all have the same units. We have no conversion factor that goes directly from ft^3 to μL , so we must use a chain of conversion factors that we know or have available. Since we are dealing with relatively small volumes based on length units, the main conversion from USCS units to SI units is the inch to centimeter factor of 1 in = 2.54 cm. This factor is exact and should be committed to memory. If we were starting with gallons instead of ft^3 , we might use the conversion 1 gal = 3.785 L, although this factor is not exact.

We first convert ft^3 to in^3 , then from in^3 to cm^3 , then from cm^3 to L, and finally from L to μL .

$$1.0 \text{ ft}^3 \cdot \frac{12 \text{ in}}{1 \text{ ft}} \cdot \frac{12 \text{ in}}{1 \text{ ft}} \cdot \frac{12 \text{ in}}{1 \text{ ft}} \cdot \frac{2.54 \text{ cm}}{1 \text{ in}} \cdot \frac{2.54 \text{ cm}}{1 \text{ in}} \cdot \frac{2.54 \text{ cm}}{1 \text{ in}} \cdot \frac{1 \text{ L}}{1000 \text{ cm}^3} \cdot \frac{1 \times 10^6 \mu\text{L}}{1 \text{ L}} = 28,300,000 \mu\text{L}$$

This result has three significant digits. We are keeping one extra digit during the intermediate calculations. It is not incorrect to write down all the digits your calculator shows. But it *is* pointless. A person who writes all the digits regardless of whether they are needed simply shows that he or she doesn't understand significant digits. Those extra digits are meaningless.

Notice that all the conversion factors used in the calculation above are exact; none of them are approximations. Since they are all exact, they play no role in limiting the significant digits in our result. Note that if you must use a conversion factor that is approximate, you should make sure the precision of the value in the conversion factor is at least as high as the precision in your data (two significant digits, in this case). That way your conversion factor does not limit the precision of your result. If, for some reason, you do not have a conversion factor with as many significant digits as your least precise measurement, then the precision of your result must match the precision of the conversion factor. The least precise value in the entire calculation always governs the precision in your result.

Now we compute the time required by dividing the required volume by the flow rate.

$$t = \frac{28,300,000 \mu\text{L}}{56.75 \frac{\mu\text{L}}{\text{s}}} = 499,000 \text{ s}$$

This value also has three significant digits—one more than we need. We now convert this value from seconds to hours as the problem statement requires.

$$499,000 \text{ s} \cdot \frac{1 \text{ hr}}{3600 \text{ s}} = 139 \text{ hr}$$

Finally, we need to round the result to the required two significant digits. The second non-zero digit from the left (3) is in the tens place, so we round to the nearest ten.

$$t = 140 \text{ hr}$$



Case 4

The fourth case of rules for significant digits also applies to the calculations you perform with measurements. In Case 4, we deal with addition and subtraction.

The rule for addition and subtraction is completely different from the rule for multiplication and division. When performing addition, it is not the number of significant digits that governs the precision of the result. Instead, it is the *place value of the last digit that is farthest to the left in the numbers being added* that governs the precision of the result. This rule is quite wordy and is best illustrated by an example. Consider the following addition problem:

$$\begin{array}{r} 13.65 \\ 1.9017 \\ + 1,387.069 \\ \hline 1,402.62 \end{array}$$

Of the three values being added, 13.65 has digits out to the hundredths place, 1.9017 has digits out to the ten thousandths place, and 1,387.069 has digits out to the thousandths place. Looking at the final digits of these three, you can see that the final digit farthest to the left is the 5 in 13.65, which is in the hundredths place. This is the digit that governs the final digit of the result. There can be no digits to the right of the hundredths place in the result. The justification for this rule is that one of our measurements is precise only to the nearest hundredth, even though the other two are precise to the nearest thousandth or ten thousandth. We are going to add these values together, and one of them is precise only to the nearest hundredth. It makes no sense to have a result that is precise to a place more precise than that, so hundredths are the limit of the precision in the result.

Correctly performing addition problems in science (where nearly everything is a measurement) requires that you determine the place value governing the precision of your result, perform the addition, then round the result. In the above example, the sum is 1,4602.6207. Rounding this value to the hundredths place gives 1,4602.62.

Going back to Example 1.7, we saw the following equation for converting a temperature from kelvins to degrees Celsius:

$$T_k = T_c + 273.15 = 37.0 + 273.15 = 310.2 \text{ K}$$

The two values we are adding are 37.0, which has digits out to the tenths place, and 273.15, which has digits out to the hundredths place. The final digit in 37.0 is the one farther to the left, so it governs the final digit we can have in the sum. The final digit in 37.0 is in the tenths place, so the final digit in the sum must also be in the tenths place. Adding the values gives 310.15 K. Rounding this to the tenths place gives us 310.2 K for our result.

1.4 Other Important Math Skills

1.4.1 Scientific Notation

No doubt you have studied *scientific notation* in your math classes. However, beginning in high school, scientific notation is used all the time in scientific study. Knowing how to use scientific notation correctly—including the use of the special key found on scientific calculators for working with values in scientific notation—is very important.

Mathematical Principles

Scientific notation is a way of expressing very large or very small numbers without all the zeros, unless the zeros are *significant*. This is of enormous benefit when one is dealing with a value such as 0.0000000000001 cm (the approximate diameter of an atomic nucleus). The basic idea will be clear from a few examples.

Let's say we have the value 3,750,000. This number is the same as 3.75 million, which can be written as $3.75 \times 1,000,000$. Now, 1,000,000 itself can be written as 10^6 (which means one followed by six zeros), so our original number can be expressed equivalently as 3.75×10^6 . This expression is in scientific notation. The number in front, the stem, is always written with one digit followed by the decimal and the other digits. The multiplied 10 raised to a power has the effect of moving the decimal over as many places as necessary to recreate our original number.

As a second example, the current population of earth is about 7,200,000,000, or 7.2 billion. One billion has nine zeros, so it can be written as 10^9 . So we can express the population of earth in scientific notation as 7.2×10^9 .

When dealing with extremely small numbers such as 0.000000016, the process is the same, except the power on the 10 is negative. The easiest way to think of it is to visually count how many places the decimal in the value has to be moved over to get 1.6. To get 1.6, the decimal has to be moved to the right 8 places, so we write our original value in scientific notation as 1.6×10^{-8} .

Using Scientific Notation with a Scientific Calculator

All scientific calculators have a key for entering values in scientific notation. This key is labeled $\boxed{\text{EE}}$ or $\boxed{\text{EXP}}$ on most calculators, but others use a different label.¹ It is very common for those new to scientific calculators to use this key incorrectly and obtain incorrect results. So read carefully as I outline the procedure.

The whole point of using the $\boxed{\text{EE}}$ key is to make keying in the value as quick and error free as possible. *When using the scientific notation key to enter a value, you do not press the $\boxed{\times}$ key, nor do you enter the 10.* The scientific calculator is designed to reduce all this key entry, and the potential for error, by use of the scientific notation key. You only enter the stem of the value and the power on the ten and let the calculator do the rest.

Here's how. To enter a value, simply enter the digits and decimal in the stem of the number. Then hit the $\boxed{\text{EE}}$ key, and then enter the power on the ten. The value is then in the calculator and you may do with it whatever you need to. As an example, to multiply the value 7.2×10^9 by 25 using a standard scientific calculator, the sequence of key strokes is as follows:

7.2 $\boxed{\text{EE}}$ 9 $\boxed{\times}$ 25 $\boxed{=}$

Notice that between the stem and the power, the only key pushed is the $\boxed{\text{EE}}$ key.

When entering values in scientific notation with negative powers on the 10, the $\boxed{+/-}$ key is used before the power to make the power negative. Thus, to divide 1.6×10^{-8} by 36.17, the sequence of key strokes is:

1.6 $\boxed{\text{EE}}$ $\boxed{+/-}$ 8 $\boxed{\div}$ 36.17 $\boxed{=}$

Again, neither the 10 nor the \times sign that comes before it are keyed in. The $\boxed{\text{EE}}$ key has these built in.

Students sometimes wonder why it is *incorrect* to use the $\boxed{10^x}$ key for scientific notation. To calculate 7.2×10^9 times 25, they are tempted to enter the following:

7.2 $\boxed{\times}$ $\boxed{10^x}$ 9 $\boxed{\times}$ 25 $\boxed{=}$

The problem with this approach is that sometimes it works and sometimes it doesn't, and calculator users need to use key entries that *always* work. The scientific notation key ($\boxed{\text{EE}}$) keeps all the parts of a value in scientific notation together as one number. That is, when the $\boxed{\text{EE}}$ key is used, a value such as 7.2×10^9 is not two separate numbers to the calculator; it is a single numerical value. But when the $\boxed{\times}$ key is manually inserted, the calculator treats the numbers separated by the $\boxed{\times}$ key as two separate values, and this can cause the calculator to use a different order of operations than you intend. For example, using the $\boxed{10^x}$ key causes the calculator to render an incorrect answer for a calculation such as this:

$$\frac{3.0 \times 10^6}{1.5 \times 10^6}$$

1 One infuriating model uses the extremely unfortunate label $\boxed{\times 10^x}$ which looks a *lot* like $\boxed{10^x}$, a different key with a completely different function.

The denominator of this expression is exactly half the numerator, so the value of this fraction is obviously 2.0. But when using the 10^x key, the 1.5 and the 10^6 in the denominator are separated and treated as separate values. The calculator then performs the following calculation:

$$\frac{3.0 \times 10^6}{1.5} \times 10^6$$

This comes out to 2,000,000,000,000 (2×10^{12}), which is not the same as 2.0!

The bottom line is that the \boxed{EE} key, however it may be labeled, is the correct key to use for scientific notation.

Finally, when writing a result in scientific notation, it is not acceptable to write it using the EE notation your calculator uses. For example, your calculator might display a result as 3.14EE8 or 3.14E8, but you must write this as 3.14×10^8 .

1.4.2 Calculating Percent Difference

One of the conventional calculations in science experiments is the so-called “experimental error.” Experimental error is typically defined as the difference between a predicted value and an experimental value, expressed as a percentage of the predicted value, or

$$\text{experimental error} = \frac{|\text{predicted or accepted value} - \text{experimental value}|}{\text{predicted or accepted value}} \times 100\%$$

Although the term “experimental error” is widely used, it is a poor choice of words. When there is a mismatch between theory and experiment, the experiment may not be the source of the error. Often, it is the theory that is found wanting—this is how science advances.

I now prefer to use the phrase *percent difference* to describe the value computed by the above equation. When quantitative results are compared to quantitative predictions or accepted values, students should compute the percent difference as

$$\text{percent difference} = \frac{|\text{predicted or accepted value} - \text{experimental value}|}{\text{predicted or accepted value}} \times 100\%$$

Chapter 1 Exercises

For all exercises, note that physical constants and unit conversion factors are found in Tables A.2 and A.3 of Appendix A.

SECTION 1.1

1. Write a paragraph distinguishing between matter and mass.
2. Distinguish between base units and derived units in the SI system of units and give three examples of each.
3. Describe the advantages the SI system has over the USCS system for scientific work.
4. Why does the SI system use prefixes on the units of measure?

Chapter 1

5. Re-express the quantities in the following table using only a single numerical digit followed by an SI unit symbol, with a metric prefix if necessary. Example: 5 thousand liters = 5 kL

- | | | |
|----------------------|------------------------------|----------------------------|
| a. 8 pascals | b. 5 hundredths of a meter | c. 3 million amperes |
| d. 2 thousand meters | e. 4 thousandths of a second | f. 6 thousand newtons |
| g. 8 thousand grams | h. 7 millionths of a liter | i. 1 thousandth of a joule |

6. Re-write the quantities in the following table by writing out the unit names without symbols. Example: 5 km = 5 kilometers

- | | | |
|----------------------|-------------|--------------|
| a. 14 m ³ | b. 164.1 kg | c. 250 MPa |
| d. 16.533 ms | e. 160 kA | f. 19.55 cL |
| g. 31.11 μJ | h. 2300 K | i. 13.0 mmol |

SECTION 1.2

7. Why must equations be used instead of conversion factors for most temperature unit conversions?

8. Perform the USCS unit conversions required in the following table. (Note: The answers in the back of the book are given with the correct number of significant digits.)

Convert This Quantity	Into These Units	Convert This Quantity	Into These Units
a. 12.55 ft	yd	b. 0.44556 mi	ft
c. 147.55 in	ft	d. 55.08 gal	ft ³
e. 934 ft ³	in ³	f. 739.22 ft ³ /s	gal/hr
g. 12.4 yr	hr	h. 51,083 in	mi
i. 14,560.77 gal/hr	qt/s	j. 15.90 mi/dy	in/hr

9. Perform the SI/metric unit conversions required in the following table.

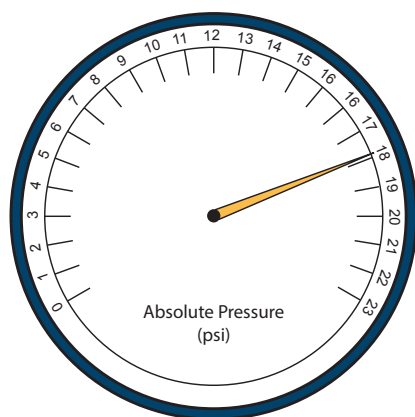
Convert This Quantity	Into These Units	Convert This Quantity	Into These Units
a. 35.4 mm	m	b. 76.991 mL	μL
c. 34.44 cm ³	L	d. 6.33 g/cm ²	kg/m ²
e. 9.35 m/s ²	mm/ms ²	f. 542.2 mJ/s	J/s
g. 56.6 μs	ms	h. 44.19 mL	cm ³
i. 532 nm	μm	j. 96,963,000 mL/ms	m ³ /s
k. 295.6 cL	μL	l. 0.007873 m ³	mL
m. 8,750 mm ²	m ²	n. 87.1 cm/s ²	m/s ²
o. 15.75 kg/m ³	g/cm ³	p. 0.875 km	m
q. 16,056 MPa	kPa	r. 7,845 μA	mA

10. Reproduce the following table on your own paper and fill in the empty cells.

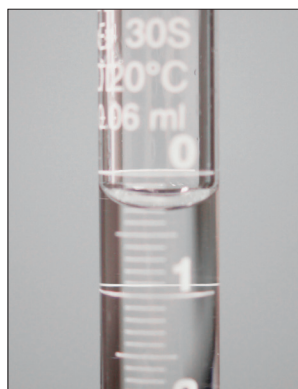
	°F	°C	K		°F	°C	K
a.	431.1			b.		-56.1	
c.			16.0	d.	0.0 (exact)		
e.		-77.0		f.			4,002
g.	-32.0			h.		65.25	
i.		1,958		j.			998.0

SECTION 1.3

- Distinguish between accuracy and precision.
- Describe the measurements you would obtain from an instrument that was very precise but not very accurate.
- Which is more important on the speedometer of a car—accuracy or precision?
- Explain why accuracy is important on a heart rate monitor but precision is not.
- Sometimes we want high accuracy in a measurement, but are not too concerned about high precision. Sometimes we want both high accuracy and high precision. Explain why no one wants low accuracy and high precision.
- On the package of a digital stopwatch I once purchased was the phrase: “1/100th second accuracy.” The stopwatch readings in seconds contained two decimal places, but the values the stopwatch actually displayed were spaced 0.03 seconds apart. Thus, it could read 12.31 s, 12.34 s, 12.37 s, etc. Comment on the accuracy and precision of this stopwatch with respect to the claim on the package.
- Using the correct number of significant digits and the correct units of measure, record the measurements represented by the following instruments.



(a)



(b)

(c) **01320**
gallons

(d) **010**
wind speed
(mph)

SECTION 1.4

- Using the correct number of significant digits, compute the percent difference for the experimental results in each of the following cases:

Chapter 1

- a. A scientist measures the masses of three compounds resulting from a certain chemical reaction. Her measurements are 0.234 g, 1.678 g, and 4.446 g. Her calculations predict that the reaction results in masses of 0.239 g, 1.688 g, and 4.678 g, respectively. Determine the percent difference for each of the three compounds.
 - b. A student measures the density of aluminum and finds it to be 2.81 g/cm^3 . The accepted density value for this alloy is 2.72 g/cm^3 .
 - c. According to the Periodic Table of the Elements, the atomic mass of carbon is 12.011 g/mol. A calculation from experimental data results in a figure of 12.0117 g/mol.
 - d. The predicted yields for the products of certain chemical reaction are 23.4 kg of compound A and 2.21 kg of compound B. Careful measurements of the masses of the compounds produced indicate masses of 21.610 kg for compound A and 1.995 kg of compound B.
19. Perform each of the unit conversions indicated in the table below. Express each result using the correct number of significant digits. Where possible and appropriate, express your result in both standard notation and scientific notation. (Note: By possible, I refer to the fact that sometimes a result can only be expressed with the correct number of significant digits if it is written in scientific notation, such as a value of 100 with two or three significant digits. By appropriate, I refer to the fact that it is silly to write a value with a very large number of zeros. Such values should always be expressed in scientific notation. It is also silly to use scientific notation to express a value such as 3 or 4.1. Such values should only be expressed in standard notation.)

Convert This Quantity	Into These Units
a. 1,737 km (radius of the earth's moon)	ft
b. 2.20 g (mass of a single peanut m&m)	kg
c. 591 mL (volume of a typical water bottle)	μL
d. $7 \times 10^8 \text{ m}$ (radius of the sun)	mi
e. $1.616 \times 10^{-35} \text{ m}$ (Planck length, a fundamentally small length)	ft
f. 750 cm^3 (size of the engine in my old motorcycle)	m^3
g. $2.9979 \times 10^8 \text{ m/s}$ (speed of light in a vacuum)	mi/hr
h. 168 hr (one week)	s
i. $5,570 \text{ kg/m}^3$ (average density of the earth)	g/cm^3
j. 45 gps (gal/sec, flow rate of Mississippi River at the source)	m^3/min
k. $600,000 \text{ ft}^3/\text{s}$ (flow rate of Mississippi River at New Orleans)	L/hr
l. 5,200 mL (volume of blood in a typical man's body)	m^3
m. $5.65 \times 10^2 \text{ mm}^2$ (area of a postage stamp)	in^2
n. 32.16 ft/s^2 (acceleration of gravity, or one "g")	m/s^2
o. $10.6 \mu\text{m}$ (wavelength of light from a CO_2 laser)	in
p. 1.1056 g/mL (density of heavy water)	kg/m^3

Convert This Quantity	Into These Units
q. 13.6 g/cm ³ (density of liquid mercury metal)	mg/m ³
r. 93,000,000 mi (distance from earth to the sun)	cm
s. 65 mph (typical highway speed limit)	m/s
t. 633 nm (wavelength of light from a red laser)	in
u. 5.015% of the speed of light (see item g, or Table A.2)	mph
v. 6.01 kJ/mol (molar heat of fusion of water)	J/mol
w. 32.1 bar (pressure in saltwater at 318 m, free diving record depth)	psi
x. 0.116 nm (radius of a sodium atom)	cm
y. 6.54×10^{-24} cm ³ (volume of a sodium atom)	in ³
z. 0.385 J/(g·K) (specific heat capacity of copper)	J/(mg·K)
aa. 370 mL (volume of a soft drink can)	ft ³
ab. 268,581 mi ² (land area of Texas)	mm ²
ac. 50,200 mi ² /yr (current rate of global deforestation) ²	ft ² /s

² As a science educator who believes that we should take care of our planet, I note that items (ab) and (ac) above indicate that every 5.4 years, we lose an area of forest the size of Texas. Texas is a big place. Think about it.

Chapter 2

Atoms and Substances



Some substances pose a minimal threat to human health and safety. Others need to be labeled and classified so we know about the risks they pose. If you examine the packaging on a shipment of chemicals or look on the back of a tank truck on the highway, you may find the Fire Diamond, a symbolic representation of the various hazards associated with the substance inside. The numerical values in each of the three colored zones range from 0 to 4, with 4 representing the most extreme hazard. The blue region pertains to health, the red to flammability, the yellow to instability or reactivity, and the white to special notices. In the sample symbol above, the numerical codes specify the following:

Blue: Health 3—Short exposure could cause serious temporary or moderate residual injury. (Example: chlorine)

Red: Flammability 2—Must be moderately heated before ignition can occur (Example: diesel fuel)

Yellow: Instability/reactivity 1—Normally stable, but can become unstable at elevated temperature and pressure (Example: alcohol)

White: Special notices “W bar”—Reacts with water in an unusual or dangerous manner (Example: sodium)

Objectives for Chapter 2

After studying this chapter and completing the exercises, you should be able to do each of the following tasks, using supporting terms and principles as necessary.

SECTION 2.1

1. Define and describe *atom* and *molecule*.
2. State the five points of John Dalton's atomic model.
3. Write brief descriptions of J.J. Thomson's cathode ray tube experiment, Robert Millikan's oil drop experiment, and Ernest Rutherford's gold foil experiment. In your descriptions, include definitions for the terms *cathode ray* and *alpha particle*.
4. Describe the atomic models proposed by J.J. Thomson and Ernest Rutherford.

SECTION 2.2

5. Define *pure substance*, *element*, *compound*, *mixture*, *heterogeneous mixture*, and *homogeneous mixture* and give several examples of each.
6. Define *suspension* and *colloidal dispersion* and give several examples of each.
7. Explain the Tyndall effect and how it can be used to identify a mixture as colloidal.
8. Describe the two basic types of structures atoms form when bonding together.
9. Distinguish between compounds and mixtures.
10. Define and distinguish between *physical properties*, *chemical properties*, *physical changes*, and *chemical changes*. Give several examples of each.
11. Define and give examples of the terms *malleable* and *ductile*.

SECTION 2.3

12. Define *isotope*, *nuclide*, and *atomic mass*.
13. Given isotope mass and abundance data, calculate the atomic mass of an element.
14. Given the periodic table, determine the number of protons, electrons, and neutrons in the atoms of a given nuclide.
15. Define the *unified atomic mass unit*, u.

SECTION 2.4

16. Use the density equation to calculate the density, volume, or mass of a substance.
17. Define the *mole*.
18. State the Avogadro constant to four digits of precision.
19. Define *molar mass*, *formula mass*, and *molecular mass*.
20. Calculate the molar mass, formula mass, or molecular mass of a compound or molecule.
21. Calculate the mass in grams of a given mole quantity of a compound or molecule, or vice versa.
22. Calculate the number of atoms or molecules in a given quantity of substance.
23. Calculate the gram masses of an atom or molecule of a given pure substance.

2.1 Atoms and Molecules

2.1.1 Atomic Facts

We begin this chapter with a summary of the basic facts about atoms and molecules. Much of this information you probably already know.

As we saw in the previous chapter, all matter is made of atoms, the smallest basic units matter is composed of. An atom of a given element is the smallest unit of matter that possesses all the properties of that element.

Atoms are almost entirely empty space. Each atom has an incredibly tiny nucleus in the center containing all the atom's protons and neutrons. Since the protons and the neutrons are in the nucleus, they are collectively called *nucleons*. The masses of protons and neutrons are very nearly the same, although the neutron mass is slightly greater. Each proton and neutron has nearly 2,000 times the mass of an electron, so the nucleus of an atom contains practically all the atom's mass. Outside the nucleus is a weird sort of cloud surrounding the nucleus containing the atom's electrons.

We address the details about electrons in the next chapter, but here is a brief preview. The electron cloud consists of different *orbitals* where the electrons are contained. Electrons are sorted into the atomic orbitals according to the amount of energy they have. For an electron to be in a specific orbital means the electron has a certain amount of energy—no more, no less.

I wrote above that atoms are almost entirely empty space because the nucleus is incredibly small compared to the overall size of the atom with its electron cloud. It's quite easy for us to pass over that remark without pausing to consider what it means. To help visualize the meaning, consider the athletic stadium pictured in Figure 2.1. Using this stadium as an enlarged atomic model, the electrons in their orbitals would be zipping around in the region where the red seating sections are in the stadium. Each electron in this enormous atomic model is far smaller than the period at the end of this sentence.

The atomic nucleus containing the protons and neutrons is located at the center of the playing field, and is the size of a pinhead. And what fills all the vast space inside the atom? Nothing, not even air, since air, of course, is also made of atoms. The inside of an atom is empty space.

Returning to our discussion of atomic facts, one of the fundamental physical properties of the subatomic particles is *electric charge*. Neutrons have no electric charge. They are electrically neutral, hence their name. Protons and electrons each contain exactly the same amount of charge, but the

charge on protons is positive and the charge on electrons is negative. If an atom or molecule has no net electric charge, it contains equal numbers of protons and electrons.

Atoms are significantly smaller than the wavelengths of light, which means light does not reflect off atoms and there is no way to see them. The same is true of *molecules*. Molecules are clusters of atoms chemically bonded together. When atoms of different elements are bonded together in a molecule they form a compound, which we discuss later in this chapter. But sometimes atoms of the same element bond together in molecules, as illustrated in Figure 2.2. Oxygen and chlorine are two of the *diatomic gases* that form molecules consisting of a pair of atoms



Figure 2.1. The head of a pin at center field in a stadium is analogous to the nucleus in the center of an atom.

chemically bound together. Hydrogen, nitrogen, and fluorine also exist naturally as diatomic gases.

2.1.2 The History of Atomic Models

The story of atomic theory starts back with the ancient Greeks. As we look at how the contemporary model of the atom developed, we also hit on some of the great milestones in the history of chemistry and physics along the way.

In the 5th century BC, the Greek philosopher Democritus proposed that everything was made of tiny, indivisible particles. Our word atom comes from the Greek word *atomos*, meaning “indivisible.” Democritus’ idea was that the properties of substances were due to characteristics of the atoms they are made from. So atoms of metals were supposedly hard and strong, atoms of water were assumed to be wet and slippery, and so on. At this same time, there were various views about what the most basic substances—that is, the elements—were. One of the most common views was that there are four elements—earth, air, water, and fire—and that everything is composed of these.

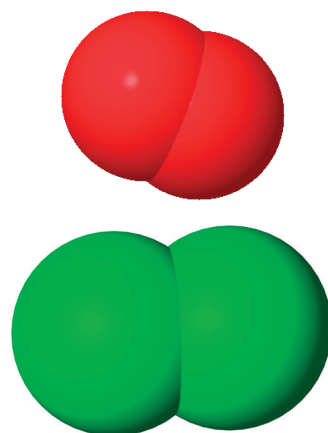


Figure 2.2. Space-filling models of the diatomic oxygen (top) and chlorine molecules.

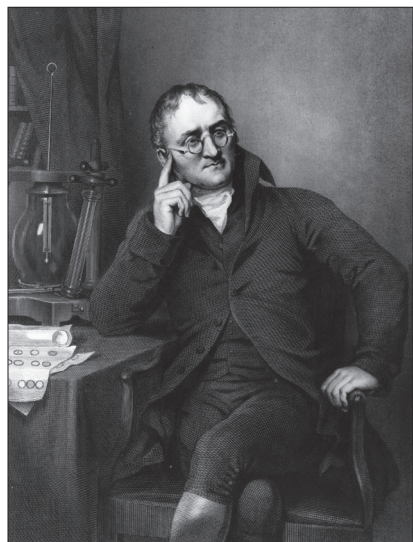


Figure 2.3. English scientist John Dalton (1766–1844).

Not much real chemistry went on for a very long time.

During the medieval period, of course, there were the alchemists, who sought to transform lead and other materials into gold. But this cannot be done by the methods available to them, so their efforts were not successful.

But in the 17th century, things started changing as scientists became interested in experimental research. The goal of the scientists described here was to figure out what the fundamental constituents of matter are. This meant figuring out how atoms are put together, what the basic elements are, and understanding what is going on when various chemical reactions take place. The nature of earth, air, fire, and water was under intense scrutiny over the next 200 years.

In 1803, English scientist John Dalton (Figure 2.3) produced the first scientific model of the atom. Dalton’s atomic model is based on five main points, listed in Table 2.1.

The impressive thing about Dalton’s atomic theory is that even today the last three of these points are regarded as correct, and the first two are at least partially correct. On the first point, it is still scientifically factual that all substances

are made of atoms, but we now know that atoms are not indivisible. This is now obvious, since atoms themselves are composed of protons, neutrons, and electrons. The second point is correct in every respect but one. Except for the number of neutrons in the nucleus, every atom of a given element is identical. However, we now know that

1. All substances are composed of tiny, indivisible substances called atoms.
2. All atoms of the same substance are identical.
3. Atoms of different elements have different weights.
4. Atoms combine in whole-number ratios to form compounds.
5. Atoms are neither created nor destroyed in chemical reactions.

Table 2.1. The five tenets of Dalton’s 1803 atomic model.

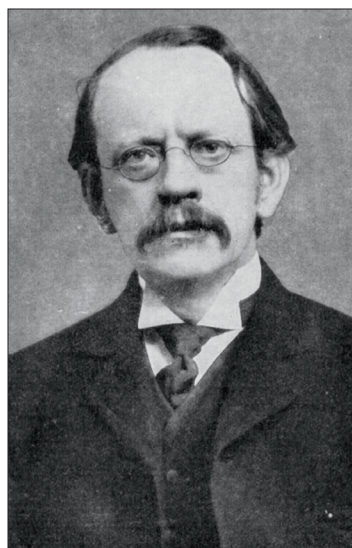


Figure 2.4. English scientist Joseph John (J.J.) Thomson (1856–1940).

atoms of the same element can vary in the number of neutrons they have in the nucleus. These varieties of nuclei are called *isotopes*, a topic we return to soon.

After Dalton, the next breakthrough in our understanding of atomic structure came from English scientist J.J. Thomson (Figure 2.4). Thomson worked at the Cavendish Laboratory in Cambridge, England. In 1897, he conducted a series of landmark experiments that revealed the existence of electrons. Because of his work, he won the Nobel Prize in Physics in 1906. A photograph of the *cathode ray tube* Thomson used for his work is shown in Figure 2.5.

Thomson placed electrodes from a high-voltage electrical source inside a very elegantly made, sealed-glass vacuum tube. This apparatus can generate a so-called *cathode ray* from the negative electrode (1), called the *cathode*, to the positive one, called the *anode* (2). A cathode ray is simply a beam of electrons, but this was not known at the time. The anode inside Thomson's vacuum tube had a hole in it for some of the electrons to escape through, which created a beam of cathode rays heading toward the other end of the tube (5).

Thomson placed the electrodes of another voltage source in-

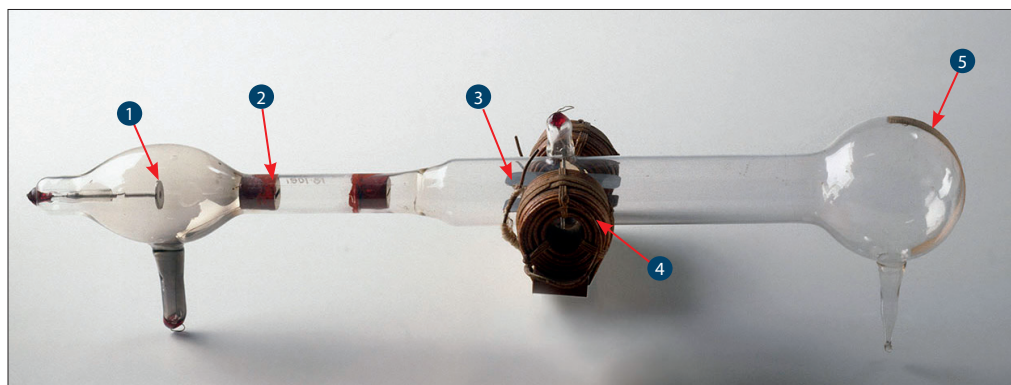


Figure 2.5. J.J. Thomson's cathode ray tube.

side the tube (3), above and below the cathode ray, and discovered that the beam of electrons deflected when this voltage was turned on. He also placed magnetic coils on the sides of the tube (4) and discovered that the electrons also deflected as they passed through the magnetic field produced by the coils. The deflection of the beam toward the positive electrode led Thomson to theorize that the beam was composed of negatively charged particles, which he called “corpuscles.” (The name *electron* was first used a few years later by a different scientist.) By trying out many different arrangements of cathode ray tubes, Thomson confirmed that the ray was negatively charged. Then using the scale

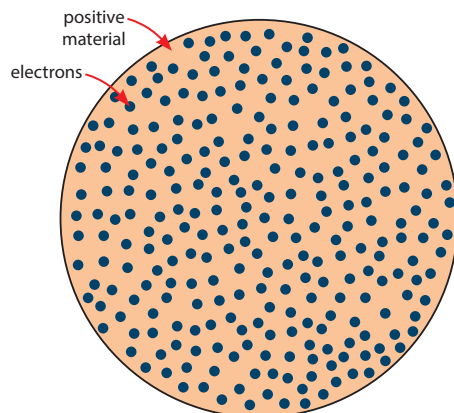


Figure 2.6. Thomson's plum pudding model.

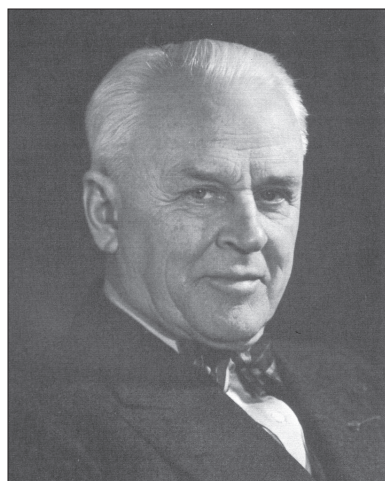


Figure 2.7. American scientist Robert Millikan (1868–1953).

on the end of the tube to measure the deflection angle (5), he was able to determine the charge-to-mass ratio of the individual electrons he had discovered. This value is 1.8×10^{11} C/kg, where C stands for coulomb, the SI unit of electric charge.

Thomson went on to theorize that electrons came from inside atoms. He developed a new atomic model that envisions atoms as tiny clouds of massless, positive charge sprinkled with thousands of the negatively charged electrons, as illustrated in Figure 2.6. Thomson's model is usually called the *plum pudding model*.

In 1911, American scientist Robert Millikan (Figure 2.7) devised his famous *oil drop experiment*, an extraordinary procedure that allowed him to determine the charge on individual electrons, 1.6×10^{-19} C. Once this value was known, Millikan used Thomson's charge-to-mass ratio and calculated the mass of the electron, 9.1×10^{-31} kg. Millikan's apparatus is pictured in Figure 2.8, and his sketch of the system is shown in Figure 2.9.

Inside a heavy metal drum (1) about the size of a 5-gallon bucket, Millikan placed a pair of horizontal metal plates (2) connected to an adjustable high-voltage source. The upper plate had a hole in the center and was connected to the positive voltage, the lower plate to the negative. He used an atomizer spray pump (3) to spray in a fine mist of watchmaker's oil above the positive plate. Some of the oil droplets would fall through the hole in the upper plate and move into the region between the plates. Connected through the side of the drum between the two plates was a telescope eyepiece (4) and lamp (5) so that Millikan could see the oil droplets between the plates.

The process of squirting in the oil droplets with the atomizer sprayer caused some of the droplets to acquire a charge of static electricity. This means the droplets had excess electrons on them and carried a net negative charge. They picked up these extra electrons by friction as the droplets squirted through the rubber sprayer tube. As Millikan looked at an oil droplet through the eye-

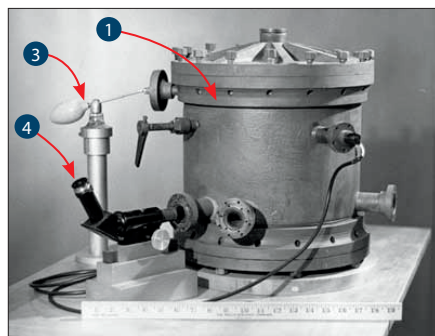


Figure 2.8. Apparatus for the oil drop experiment.

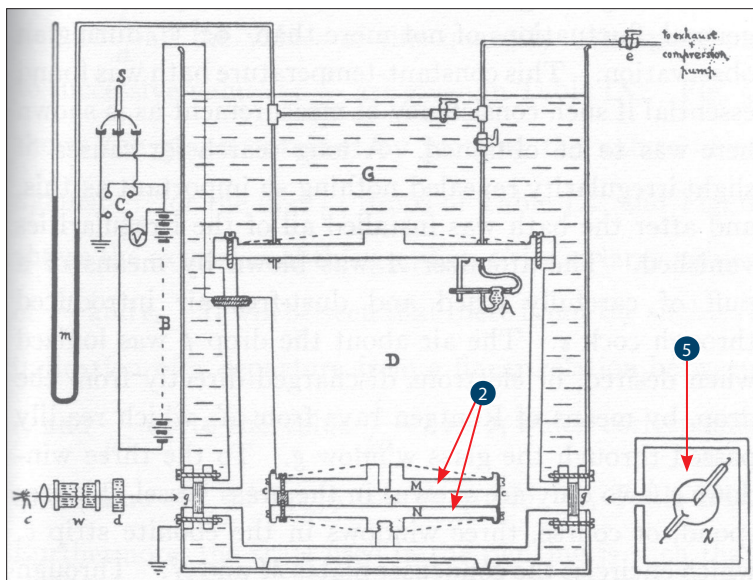


Figure 2.9. Millikan's schematic of his oil drop apparatus.



Figure 2.10. New Zealander and physicist Ernest Rutherford (1871–1937).

piece and adjusted the voltage between the plates, he could make the charged oil droplet hover when the voltage was just right. Millikan took into account the weight of the droplets and the viscosity of the air as the droplets fell and was able to determine that every droplet had a charge on it that was a multiple of 1.6×10^{-19} C. From this he deduced that this must be the charge on a single electron, which it is. Millikan won the Nobel Prize in Physics in 1923 for this work.

The last famous experiment in this basic history of atomic models was initiated in 1909 by one of Thomson's students at Cambridge, New Zealander Lord Ernest Rutherford (Figure 2.10). Rutherford was already famous when this experiment occurred, having just won the Nobel Prize in Chemistry the previous year. Rutherford's *gold foil experiment* resulted in the discovery of the atomic nucleus. To understand this experiment you need to know that an alpha particle, or α -particle (using the Greek letter alpha, α) is a particle composed of two protons and two neutrons. Alpha particles are naturally emitted by some radioactive materials in a process called *nuclear decay*.

Rutherford created a beam of α -particles by placing some radioactive material (radium bromide) inside a lead box with a hole in one end. The α -particles from the decaying radium atoms streamed out the hole at very high speed (15,000,000 m/s!). Rutherford aimed the α -particles at an extremely thin sheet of gold foil only a few hundred atoms thick. Surrounding the gold foil was a ring-shaped screen coated with a material that glows when hit by α -particles. Rutherford could then determine where the α -particles went after encountering the gold foil.

When Rutherford began taking data, he had Thomson's plum pudding model in mind and was expecting results consistent with that atomic model. This scenario is depicted in the upper part of Figure 2.11. The atom's positive charge, shown in the light orange color, is spread throughout the atom and the negative electrons are embedded in the positive material. Rutherford expected the massive and positively charged α -particles to blow right through the gold foil.

What Rutherford found was astonishing. Most of the α -particles passed straight through the foil and struck the screen on the other side, just as Rutherford expected. However, occasionally an α -particle (one particle out of every several thousand) deflected with a small angle. And sometimes the deflected particles bounced almost straight back. This situation is depicted in the lower part of Figure 2.11.

The astonished Rutherford commented that it was like firing a huge artillery shell into a piece of tissue paper and having it bounce back and hit you! Rutherford's work led to his new proposal in 1911 for a model of the atom. Rutherford's model included the key points listed in Table 2.2.

In 1917, Rutherford became the first to "split the atom." In this experiment he used α -particles again, this time strik-

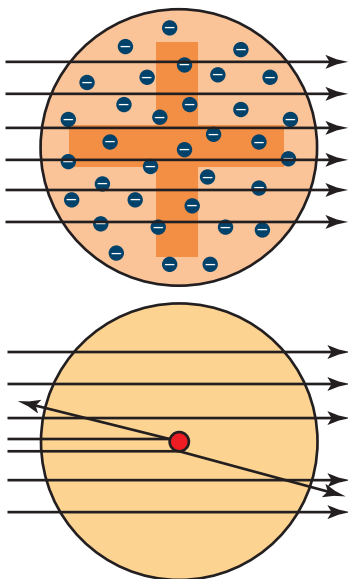


Figure 2.11. Alpha-particle pathways through the gold atoms expected from Thomson's plum pudding model (top) and the pathways Rutherford observed (bottom).

ing nitrogen atoms. His work led to the discovery of the positively-charged particles in the atomic nucleus, which he named protons.

It took another twenty years before James Chadwick (Figure 2.12), another Englishman, discovered the neutron. Before World War I, Chadwick studied under Rutherford (at Cambridge, of course). Then the war began. Not only did the war interrupt the progress of the research in general, but Chadwick was a prisoner of war in Germany. Working back in England after the war, he discovered the neutron in 1932 and received the Nobel Prize in Physics for his discovery in 1935.

Chadwick's discovery of the neutron enabled physicists to fill in a lot of blanks in their understanding of the basic structure of atoms. But years before Chadwick made his discovery, Rutherford's atomic model was already being taken to another level through the work of Niels Bohr. We explore Bohr's atomic model, and the quantum model to which it led, in the next chapter.

2.2 Types of Substances

A *substance* is anything that contains matter. There are several major classifications of substances, but as shown in Figure 2.13, they all fall into two major categories, *pure substances* and *mixtures*.

2.2.1 Pure Substances: Elements and Compounds

There are two kinds of pure substances, *elements* and *compounds*. We will discuss elements first. In previous science classes you may have seen or studied the Periodic Table of the Elements, which lists all the known elements. This famous table plays a major role in the study of chemistry and is shown in Figure 2.14 on the next page. We dive into the periodic table in detail in Chapter 4, but I bring up the table here to assist in our discussion of elements.

Elements

The characteristic that defines each element in the periodic table is the number of protons the element has in each of its atoms, a number called the *atomic number*. The elements are ordered in the periodic table by atomic number. For example, carbon is element number six in the periodic table. This means that an atom of carbon has six

1. The positive charge in atoms is concentrated in a tiny region in the center of the atom, which Rutherford called the *nucleus*.
2. Atoms are mostly empty space.
3. The electrons, which contain the atoms' negative charge, are outside the nucleus.

Table 2.2. The main ideas in Ernest Rutherford's 1911 atomic model.



Figure 2.12. English physicist James Chadwick (1891–1974).

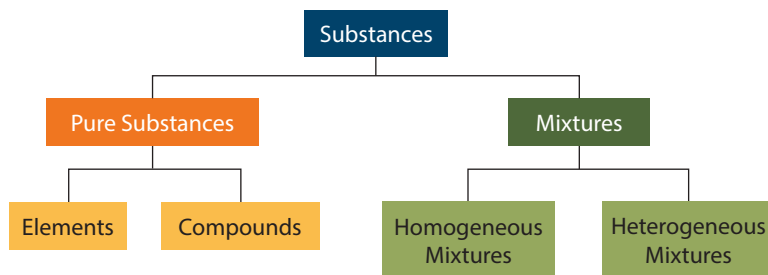


Figure 2.13. Classifications of substances.

Figure 2.14. The Periodic Table of the Elements.

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	
1A	2A					7B	8B		1B	2B		3A	4A	5A	6A	7A	8A	
1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	
H Hydrogen 1.0079	Li Lithium 6.941	Na Sodium 22.9898	K Potassium 39.098	Rb Rubidium 85.468	Cs Cesium 132.905	Fr Francium 223.0197	Be Beryllium 9.0122	Mg Magnesium 24.3050	Ca Calcium 40.078	Sr Strontium 87.62	Ba Barium 137.327	B Boron 10.811	C Carbon 12.011	N Nitrogen 14.0067	O Oxygen 15.9994	F Fluorine 18.9984	He Helium 4.0026	
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
Potassium 39.098	Calcium 40.078	Scandium 44.9559	Titanium 47.88	Vanadium 50.9415	Chromium 51.9961	Manganese 54.9380	Iron 55.847	Cobalt 58.9332	Nickel 58.6934	Copper 63.546	Zinc 65.39	Gallium 69.723	Germanium 72.61	Arsenic 74.9216	Selenium 78.96	Bromine 79.904	Krypton 83.80	
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54	
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
Rubidium 85.468	Strontium 87.62	Yttrium 88.9059	Zirconium 91.224	Niobium 92.9064	Molybdenum 95.94	Technetium 98.9072	Ruthenium 101.07	Rhodium 101.07	Palladium 106.42	Silver 107.8682	Cadmium 112.411	Indium 114.82	Tin 118.710	Antimony 121.76	Tellurium 127.60	Iodine 126.9045	Xenon 131.29	
55	56	57	71	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86
Cs	Ba	Lu	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Cesium 132.905	Barium 137.327	Lutetium 174.967	Lanthanum 138.9055	Hafnium 178.49	Tantalum 180.9479	Tungsten 183.85	Rhenium 186.207	Osmium 190.2	Iridium 192.22	Platinum 195.08	Gold 196.9665	Mercury 200.59	Thallium 204.3833	Lead 207.2	Bismuth 208.9804	Polonium 208.9824	Astatine 209.9871	Radon 222.0176
87	88	103	104	105	106	107	108	109	110	111	112	113	114	115	116	117	118	
Fr	Ra	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og	
Francium 223.0197	Radium 226.0254	Lawrencium 262.11	Rutherfordium 261.11	Dubnium 262.114	Seaborgium 263.118	Bohrium 262.12	Hassium (265)	Mitneerium (266)	Darmstadtium (281)	Roentgenium (281)	Copernicium (285)	Nihonium (284)	Flerovium (289)	Moscovium (288)	Livermorium (293)	Tennesse (294)	Ognesson (294)	

■ liquid at room temperature
■ radioactive

57	58	59	60	61	62	63	64	65	66	67	68	69	70
La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb
Lanthanum 138.9055	Cerium 140.115	Praseodymium 140.9077	Neodymium 144.24	Promethium 144.9127	Samarium 150.36	Europium 151.965	Gadolinium 157.25	Terbium 158.9253	Dysprosium 162.50	Holmium 164.9303	Erbium 167.26	Thulium 168.9342	Ytterbium 173.04
89	90	91	92	93	94	95	96	97	98	99	100	101	102
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No
Actinium 227.0278	Thorium 232.0381	Protactinium 231.0359	Uranium 238.0289	Neptunium 237.0482	Plutonium 244.0642	Ameicium 243.0614	Curium 247.0703	Berkelium 247.0703	Californium 251.0796	Einsteinium 252.083	Fermium 257.0951	Mendelevium 258.10	Nobelium 259.1009

Note: The upper set of group numbers has been recommended by the International Union of Pure and Applied Chemistry (IUPAC) and is now in wide use. The lower set of numbers is still in common use in America.

protons—all carbon atoms have six protons. If an atom does not have six protons, it is not a carbon atom, and if an atom does have six protons, it is a carbon atom. An element is therefore a type of atom, classified according to the number of protons the atom has. A lump of elemental carbon—which could be graphite, diamond, coal, or several other varieties of pure carbon—is any lump of atoms that contain only six protons apiece. Oxygen (element 8) is another example of an element. Pure oxygen is a gas (ordinarily) that contains only atoms with eight protons each, because oxygen is element number eight. Other examples of elements you have heard of are iron, gold, silver, neon, copper, nitrogen, lead, and many others.

For every element, there is a *chemical symbol* that is used in the periodic table and in the chemical formulas for compounds. For a few elements, a single upper-case letter is used, such as N (7) for nitrogen and C for carbon. But for most elements, an upper-case letter is followed by one lower-case letter, such as Na (11) for sodium and Mg (12) for magnesium. The three-letter symbols for elements 113, 115, 117, and 118 are placeholders until official names and two-letter symbols are selected by the appropriate governing officials.

Some of the chemical symbols are based on the Latin names of elements, such as Ag for silver, from its Latin name *argentum*. Other examples are Au for gold, from the Latin *aurum*, and Pb for lead, from the Latin *plumbum* (everyone's favorite Latin name).

The most common representations of the periodic table show four pieces of information for each element, indicated in Figure 2.15. At the top of the cell is the atomic number, symbolically represented as *Z*. Again, this number indicates the number of protons in each atom of the element and is the number used to order the elements in the periodic table. Below the atomic number are the element's chemical symbol and name. At the bottom of the cell is the atomic mass. We look at the atomic mass more closely in Section 2.3.

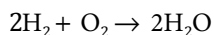
17	← atomic number
Cl	← chemical symbol
Chlorine	← element name
35.4527	← atomic mass

Figure 2.15. The basic information in each cell of the periodic table.

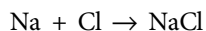
Compounds

As the name implies, a compound is formed when two or more different elements are chemically bonded together. This bonding is always the result of a chemical reaction. A chemical reaction is any process in which connecting bonds between atoms are formed or broken. Once bonded together chemically, the elements in a compound can only be separated by chemical means. In other words, it takes a different chemical reaction to break atoms apart.

The elements or compounds that go into a chemical reaction are called the *reactants*. The compounds formed by the reaction are called the *products*. The physical and chemical properties of a compound are completely different from the properties of any of the elements in the compound. For example, consider oxygen, hydrogen, and water. Hydrogen and oxygen react with a boom (Figure I.8) to form water, according to this chemical equation:



Oxygen is an invisible gas that we breathe in the air and that supports combustion. Hydrogen is an invisible, flammable gas. Water is composed of oxygen atoms bonded to hydrogen atoms, but one cannot breathe water, nor does water combust or support combustion. Or consider the sodium and chlorine in sodium chloride. These elements combine according to this chemical equation:



We all require salt in our diets, and we find it tasty. But both sodium and chlorine, the two elements of which sodium chloride is composed, are deadly dangerous in their pure, elemental forms. Sodium is a shiny, peach-colored metal, pictured in Figure 2.16. Sodium slices just like cheddar cheese and the photo shows me slicing a chunk of sodium open to show its beautiful, pinkish shiny color. Chlorine is a greenish gas, pictured in a glass bottle in Figure 2.17. As you can see, the properties of these substances are nothing at all like the properties of the salt they form when they react together. We discuss physical and chemical properties in more detail a bit later.



Figure 2.16. Sodium metal.



Figure 2.17. Chlorine gas.

When atoms bond together to form a compound, the atoms in the compound can be arranged in either of two basic types of structures. In many cases, the atoms join together to form molecules. In every molecule of a given substance, the atoms bond together in the same whole-number ratio—a perfect example of the important general principle that we first encountered in the Introduction. In the Introduction are shown computer models of water (H_2O), ammonia (NH_3), and methane (CH_4) molecules. A few other well-known molecular substances are carbon dioxide (CO_2), propane (C_3H_8), and ozone (O_3), all represented by space-filling models in Figure 2.18. The standard color coding used in computer models is white for hydrogen, black or charcoal gray for carbon, and red for oxygen. It is important to note that in any chemical formula for a molecular substance, the formula indicates the number of each type of atom in the molecule. Propane, C_3H_8 , has three carbon atoms and eight hydrogen atoms in each molecule. The subscripts are only shown when the quantity of atoms of an element is greater than one. The formula for sucrose—table sugar—is $\text{C}_{12}\text{H}_{22}\text{O}_{11}$. Each molecule of sucrose contains 12 carbon atoms, 22 hydrogen atoms, and 11 oxygen atoms.

The other common way atoms combine is by forming a continuous geometric arrangement. These compounds are called *crystals*, and the structure the atoms in the compound make when they join together is called a *crystal lattice*. The number of different arrangements atoms can make in a lattice is endless, and these arrangements are responsible for many of the unusual properties crystals possess. But what all lattices have in common is the regular arrangement of the atoms into

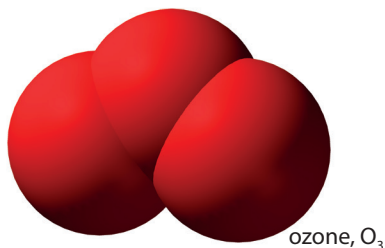
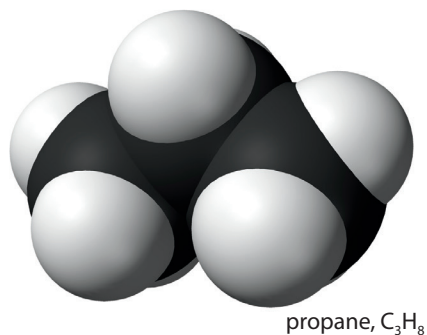
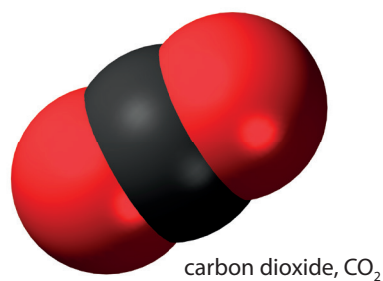


Figure 2.18. More common molecular substances. Hydrogen atoms are white, carbon atoms are black, and oxygen atoms are red.

repeating, geometrical patterns. A space-filling model of the very simple crystal structure for sodium chloride, NaCl (table salt), is shown in Figure 2.19. The formula tells us that the ratio of sodium atoms to chlorine atoms in the crystal is 1 : 1. The model shows that the atoms are bonded together in a simple alternating arrangement.

We have seen several space-filling models so far. Another common type of computer model is the ball-and-stick model. Figure 2.20 shows a ball-and-stick model of the somewhat more complex crystal structure of copper(II) chloride, CuCl_2 . As the formula indicates, in the lattice structure there are two chlorine atoms for each copper atom.

2.2.2 Mixtures

So far we have been discussing one major category of substances—pure substances. Elements and compounds are pure substances. The other major category is *mixtures*. Anytime substances are mixed together without a chemical reaction occurring, a mixture is formed. Remember—if a chemical reaction occurs, compounds are formed, not mixtures. If you toss vegetables in a salad, you've made a mixture. If you put sugar in your tea or milk in your coffee, you've made a mixture. If you mix up a batch of chocolate chip cookie dough, a bowl of party mix, or the batter for a vanilla cake, you've made a mixture.

In contrast to compounds, when a mixture is formed the individual substances in the mixture retain their physical and chemical properties. If you mix salt in water, you've made a mixture. The salt is still there and tastes salty. The water is still there too, and tastes watery. Also in contrast to compounds, the substances in a mixture can be separated by physical means such as filtering, boiling, freezing, or settling.

Again, in compounds, the original properties of the elements in the compound are chemically changed into the properties of the new chemical substance that is formed by means of a chemical reaction. Further, the elements in a compound cannot be separated by physical means. They can only be separated by the same means that brought them together in the first place—a chemical reaction. The distinguishing features of mixtures and compounds are summarized in Table 2.3. There are two classes of mixtures—*homogeneous mixtures* and *heterogeneous mixtures*. We examine these next.

Homogeneous Mixtures

Homogeneous mixtures have uniform composition down to, but not including, the groups of atoms at the atomic level. The individual particles of the different substances in a homogeneous mixture cannot be seen with the eye, not even with the most powerful microscope. Particles at the atomic level—atoms, molecules, ions—are too small to reflect visible light, and so cannot be seen in the ordinary way, regardless of the magnification.

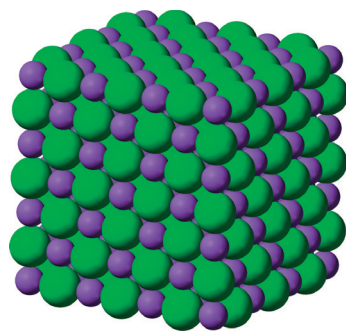


Figure 2.19. A space-filling model of the sodium chloride lattice structure. Sodium atoms are purple and chlorine atoms are green.

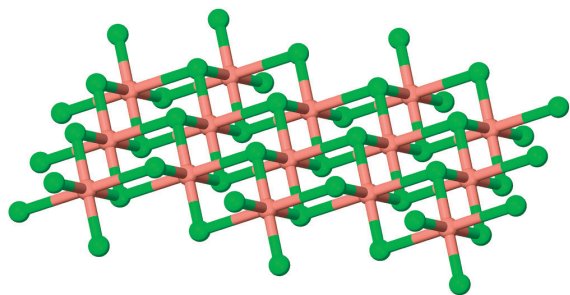


Figure 2.20. A ball-and-stick model of the copper(II) chloride crystal structure. Copper atoms are copper colored and chlorine atoms are green.

Mixtures	Compounds
<ul style="list-style-type: none"> • Formed when substances combine without a chemical reaction occurring. • The individual substances in the mixture retain their physical and chemical properties. • The substances in mixtures can be separated by physical means such as filtering, boiling, freezing, or settling. 	<ul style="list-style-type: none"> • Elements combine chemically to form a new substance—a compound. • The physical and chemical properties of the compound are completely different from those of the reactants that formed the compound. • The elements in a compound can only be separated by chemical means.

Table 2.3. Summary of the distinctions between mixtures and compounds.

The implication of this definition is that homogeneous mixtures are identical with *solutions*, mixtures in which one pure substance (or more than one) is dissolved in another pure substance. Solutions are so important in chemistry that we devote an entire chapter to the topic later on. For now, we move on to the other types of mixtures.

Heterogeneous Mixtures

In contrast to homogeneous mixtures, in a heterogeneous mixture there are lumps of different substances mixed together. You might be able to see the lumps with the naked eye, as with the mixture of spices in meat seasoning. Or the different substances may be visible only under a microscope, such as microscopic particulates in well water. Either way, if the different substances can be seen, the mixture is heterogeneous. In addition to the examples of salads and so on I mentioned above, there are two classes of heterogeneous mixtures that we encounter every day: *suspensions* and *colloidal dispersions*.

Suspensions

A *suspension* is formed when particles of size approximately 1 micrometer ($1\ \mu\text{m}$) or larger are dispersed in a fluid (liquid or gas) medium. Particles this large do not remain in suspension indefinitely; they eventually settle out due to gravity.

Figure 2.21 shows an example of a suspension. At room temperature, corn starch is not soluble in water; that is, it does not dissolve in water. But if corn starch is stirred into water, a suspension is formed, as shown in the upper photo in the figure. After a few minutes, it is evident that the starch does not dissolve in the water, and after a few hours the starch particles are all at the bottom of the container, as shown in the lower photo.

Muddy water is another example of a suspension. If left to stand, the particles of sand, soil, and other organic matter eventually settle out, leaving the water quite transparent.

Colloidal Dispersions

Another class of heterogeneous mixtures is the *colloidal dispersions*, usually referred to simply as *colloids*. A colloid is formed when microscopic particles, ranging in

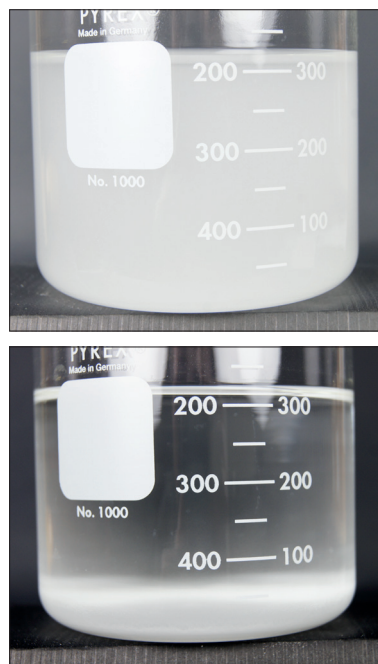


Figure 2.21. Corn starch mixed in water forms a suspension (top). The starch particles eventually settle out of the liquid and fall to the bottom of the beaker (bottom).

		Dispersing Medium State			
		Solid	Liquid	Gas	
Dispersed Particle State	Solid	solid sol cobalt glass, cranberry glass	sol blood, paint	solid emulsion butter, cheese	solid aerosol smoke, fine airborne dust
	Liquid	gel gelatin, hair gel	liquid emulsion mayonnaise, milk		liquid aerosol fog, hair spray, insect repellent
	Gas	solid foam Styrofoam, foam rubber, marshmallow	foam shaving cream, whipped cream		(none)

Table 2.4. Types of colloids with examples.

size from 1 nm to 1,000 nm, are dispersed throughout a dispersing medium. Unlike the particles in suspensions, these particles do not settle out. Forces from *Brownian motion* (see the box on the next page) keep the particles dispersed in the dispersing medium.

Table 2.4 lists the different types of colloids, distinguished by the states of the dispersed particles and the dispersing medium.¹ Since most of these substances are probably familiar to you, although perhaps not by name, I will just point out a few things about the information in the table.

You may or may not have seen items made of cobalt glass or cranberry glass, both examples of *solid sols* and shown in Figure 2.22. Glass is silicon dioxide, SiO_2 , just like sand. Cobalt glass is silicon dioxide with cobalt salts such as cobalt(II) aluminate, CoAl_2O_4 , dispersed within the molten glass. Cranberry glass contains particles of gold(III) oxide in the molten glass. Once the glass hardens, the colloid that remains consists of solid particles dispersed within a solid medium.

Two different terms are used for colloids consisting of solid particles dispersed within a liquid medium, *sol* and *solid emulsion*. As the examples in the table show, these two classes of colloids are rather different.

In a *gel*, the solid particles link together in long molecules known as *polymers*, and these linked solids form a network or matrix throughout the substance. Finally, there are no colloids consisting of gas particles dispersed in other gas particles. This is because when no reaction occurs, all gases are *miscible* in each other. That is, the particles of the gases mix together in all proportions. Thus, mixed gases form solutions, not colloids.

We conclude this section with two more interesting facts about colloids. First, although the dispersed particles in a colloid are smaller than the particles in a suspension, they are still not



Figure 2.22. Cobalt glass (left) and cranberry glass (right), also known as “gold ruby.”

1 Sources differ to a frustrating degree on the definitions of some of the terms and on the identifications of some of the substances. If you go digging around in different texts or online sources, prepare to get confused.

Hmm... Interesting.**Brownian motion**

In 1827, Scottish botanist Robert Brown (1773–1858) was using a microscope to examine tiny particles from pollen grains. The particles were suspended in water, and Brown noticed that they jittered back and forth and moved through the water in an erratic and apparently random path. The motion of those particles is now known as *Brownian motion*, although no one knew what caused it until 1905. That was the year Albert Einstein published a paper demonstrating that the motion was caused by a particle being bombarded on all sides by moving water molecules. On this topic, still pictures just don't cut it. There are lots of animations online, but nothing beats the real thing. Go to commons.wikimedia.org, search under Brownian motion, and watch the clip with the file name *Brownianmotion beads in water spim video.gif* (sic).

small enough to pass through a filtration medium. In a true solution, the individual particles are ions or molecules, and their diameters are down in the sub-one nanometer range. Particles this small can pass right through a filtration medium such as filter paper. (So don't bother trying to separate salt from water by passing saltwater through a filter.) Colloidal particles are much larger than this and cannot pass through a filter. This is why people can wear dust masks to filter out pollens and dust (but not water vapor) from breathing air.

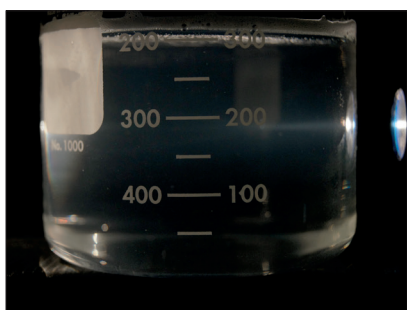
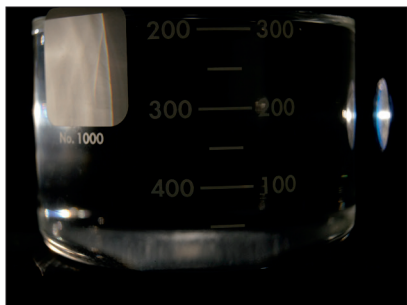


Figure 2.23. A salt water solution (top) allows the flashlight beam to pass through without scattering the light. A colloidal dispersion (bottom) scatters the light, making the beam visible, illuminating the surrounding liquid and the markings on the beaker from behind.

For colloids in fluid media, there is a test that distinguishes between solutions and colloids. Figure 2.23 illustrates the so-called *Tyndall effect* in a colloid. In the upper photograph, a flashlight beam is passed from right to left through a solution of saltwater. The light from the beam cannot be seen passing through the solution because the particles in the solution are so small they do not scatter light. In the lower photo, the beam is passed through a mixture of gelatin and water. The particles of gelatin form a colloidal dispersion, and are large enough that they do scatter the light. This is demonstrated by the fact that the beam is visible, by the illumination of the liquid, and by the brightness of the white markings on the beaker. The two photos were taken under the same lighting conditions. The light in the room was very dim except for the flashlight.

The Tyndall effect is what makes the beams from a car's headlights visible in foggy conditions. Fog is a colloid of liquid particles of water in air (a liquid aerosol), and the liquid particles scatter the light. And by the way, true water vapor—which is what steam is—is an invisible gas, just as the humidity in the air is invisible. If you can see the moisture, you are seeing light scattered by liquid water particles in air, and you are looking at a colloid, a mist of tiny condensed water droplets, not at steam.

2.2.3 Physical and Chemical Properties

All substances have certain properties. We divide up the different properties substances can possess into two broad classes. Some properties have to do with the physical characteristics of the substance, such as color, shape, size, phase, boiling point, texture, thermal conductivity, electrical conductivity, opacity, and density. These properties are called *physical properties*. Below is a list of example statements about the physical properties of substances. Consider how each one relates to the definition of physical properties just given.

- Iron is gray in color.
- Copper has a high electrical conductivity.
- Mica is shiny.
- Glass is smooth, but has sharp edges.
- Aluminum has a high thermal conductivity.
- Ethyl alcohol is transparent and colorless.
- Chlorine is greenish-yellow gas at atmospheric pressure and room temperature.
- Helium is a gas at atmospheric pressure and room temperature.
- At standard pressure, water freezes at 0°C.
- Milk is opaque.
- Oil is slippery.
- Clay brick is ochre in color.
- At 4.0°C, the density of water is 1.0 g/cm³.
- Gold is malleable and ductile.
- Cast iron is not malleable.
- Glass is not malleable.
- Play-Doh is malleable, but not ductile.
- Jello is not ductile.

You may have noticed a couple of unfamiliar terms in the above list. The terms *malleable* and *ductile* are used to describe two important properties possessed by many metals. A substance is malleable if it can be hammered into different shapes, or hammered flat into sheets. A substance is ductile if it can be *drawn* into a wire. Wire drawing is a process of making wire by pulling the metal through a small hole in a metal block called a *die*. Usually the metal is already formed into a wire of larger diameter. The end of this larger diameter wire is hammered down or filed to get it through the hole in the die, and then a machine pulls the wire through the die to make the new, smaller diameter wire. Substances that can be drawn through a die like this without simply snapping are said to be ductile.

Notice from the above examples that a good student of science should be careful when describing physical properties. We need to make sure our statements are accurate in cases where temperature or pressure affect the property in question. For example, it is inaccurate to say that H₂O is a liquid. A more accurate statement is to say that one of the physical properties of water is that it is a liquid at temperatures between 0°C and 100°C. The statement is made even more accurate by specifying that the preceding sentence is correct at atmospheric pressure because at other pressures the boiling and freezing points of water are different.

The second broad class of properties has to do with the kinds of chemical bonds a substance forms, that is, the chemical reactions a substance does or does not participate in. These properties are called *chemical properties*. We have not yet studied chemical reactions, so you may not know that much about them. However, there are two common chemical reactions that you are quite familiar with—burning and rusting. Both of these are chemical reactions in which a substance combines with oxygen, and both are examples of a type of reaction called *oxidation*. Fiery explosions are simply combustions that happen very rapidly. But whether the combustion hap-

pens slowly, as with a log on a fire, or rapidly, as with a firecracker, combustion is a chemical reaction with oxygen. Substances that react with oxygen in this way are said to be *flammable* or *combustible*. (Oddly, *inflammable* means the same thing!)

Iron oxidizes to form compounds known as *oxides*. There are several different forms of iron oxide, colored red, yellow, brown, and black. Other metals oxidize as well. When copper oxidizes, it can form two different oxides, one red and one black. This is why copper objects exposed to the air turns dark brown or black. (Over a longer period of time the copper oxide forms other compounds, such as copper carbonate, which give the copper its pretty blue-green color. The Statue of Liberty is made of copper, and has been there for a long time. It is essentially covered with a layer of copper carbonate and other copper compounds.) Aluminum also oxidizes. Aluminum oxide is dark gray, and anyone who has done a lot of hand work with aluminum parts has noticed his or her hands blackened by the particles of aluminum oxide building up. Figure 2.24 shows a few different oxides.

Here are some examples describing chemical properties of substances:

- Hydrogen is combustible.
- Aluminum oxidizes to form aluminum oxide.
- Water is not flammable.
- Platinum does not oxidize. (This is why it is so valuable. It stays shiny as a pure element.)
- Baking soda (sodium bicarbonate, NaHCO_3) reacts with vinegar (acetic acid, CH_3COOH).
- Iron oxidizes to form iron oxide, or rust.
- Sodium reacts violently with water.
- Hydrogen reacts with a number of different polyatomic ions to form acids.
- Dynamite is explosive.
- Sodium hydroxide reacts with aluminum.
- Sulfuric acid reacts with many metals.

We discuss the physical and chemical properties of substances further in the coming chapters.

The two broad classes of properties we have been discussing, physical properties and chemical properties, are related to two broad classes of changes that substances undergo. If a substance experiences a change with respect to one of its physical properties, we call this a *physical change*. When a physical change occurs, the substance is still the same substance, it just looks or feels different. If a chemical reaction occurs to a substance, this is called a *chemical change*. Chemical properties basically describe the kinds of chemical changes a substance can undergo. When a chemical change occurs, the original substances that go into the reaction—the reactants—are converted into new substances—the products—with totally different physical and chemical properties. When asked to describe a given change as physical or chemical, ask yourself if the substance is still the same substance, or if it has actually gone through a chemical reaction to become a different substance.

Table 2.5 lists some examples of physical and chemical changes, with comments explaining why the type of change is physical or chemical.



Figure 2.24. Oxides: red iron oxide (top), yellow iron oxide (center), and black copper oxide (bottom).

Process	Change	Comments
glass breaking	physical	The broken glass is still glass, it just changes shape.
firecracker exploding	chemical	This explosion is a combustion. All combustions are chemical reactions. The substances in the firecracker react to form new substances such as ash and various gases.
mercury boiling	physical	The mercury is still mercury, it just changes from the liquid state to the vapor state.
copper turning dark brown or black	chemical	This occurs because the copper is oxidizing and forming copper oxide, a new substance. This is a chemical reaction.
iron pipes corroding	chemical	Corrosion is a chemical reaction. In this case, the iron reacts with the substances surrounding the pipes to form a new substance.
water evaporating	physical	The substance is still H ₂ O, it simply changes state from liquid to vapor.
mixing cake batter	physical	The eggs and flour and so on form a mixture, but no chemical change (reaction) occurs.
baking cookies	chemical	The heat causes a chemical reaction to occur in the dough. The substance is no longer cookie dough. It is cookie.
spilled pancake batter drying out	physical	No chemical reaction occurs. Dried batter is still batter. (If you want a pancake you have to cook it, which causes a chemical change.)
molten lead hardening	physical	The lead is still lead, it just changes state from liquid to solid.
balloon popping	physical	The balloon material is still the same material, it is just in shreds now. The air inside the balloon is at a lower pressure and is not contained in the balloon any longer, but it is still air.

Table 2.5. Examples of physical and chemical changes.

2.3 Isotopes and Atomic Masses

2.3.1 Isotopes

As you know, the atomic number (Z) of an element designates the number of protons in the nucleus of an atom of that element. For a given element, the atomic number is fixed: if an atom has a different number of protons, it is an atom of a different element. But the number of neutrons that may be present in the nucleus is not always the same for atoms of a given element. For most elements, there are variations in the number of neutrons that can be present in the nucleus. These varieties are called *isotopes*. For most elements, there is one isotope that is the most abundant in nature and several other isotopes that are also present but in smaller quantities. The general term for any isotope of any element is *nuclide*.

Isotopes are designated by writing the name of the element followed by the number of nucleons (protons and neutrons) in the isotope. The number of nucleons in a nucleus is called

the *mass number*. For example, the most common isotope of carbon is carbon-12, accounting for about 98.9% of all the naturally occurring carbon. In the nucleus of an atom of carbon-12 there are six protons and six neutrons. There are two other naturally occurring carbon isotopes. Carbon-13 with seven neutrons accounts for about 1.1% of natural carbon. Atoms of carbon-14, of which only a trace exists in nature, have eight neutrons in the nucleus.

2.3.2 The Unified Atomic Mass Unit

The mass of a single atom is an extremely small number. But so much of our work in chemistry depends on atomic masses that scientists having been using units of *relative* atomic mass for a long time—all the way back to John Dalton, before actual masses of atoms were even known. Prior to the discovery of isotopes in 1912, the so-called *atomic mass unit* (amu) was defined as 1/16 the mass of an oxygen atom. After the discovery of isotopes, physicists defined the amu as 1/16 the mass of an atom of oxygen-16, but the definition used by chemists was 1/16 the average mass of naturally occurring oxygen, which is composed of several isotopes. To eliminate the confusion resulting from these conflicting definitions, the new *unified atomic mass unit* (u) was adopted in 1961 to replace them. Many texts continue to use the amu as a unit, but they define it as the u is defined. Strictly speaking, the amu is an obsolete unit that has been replaced by the u, now also called the dalton (Da). The u and the Da are alternative names (and symbols) for the same unit. The use of the dalton has increased in recent years, particularly in molecular biology.

The unified atomic mass unit, u, is defined as exactly 1/12 the mass of an atom of carbon-12. Table 2.6 lists a few nuclides and their atomic masses using the u as a unit of mass. All the elements listed exist as other isotopes in addition to those shown, but as you see from the percentage abundances, the ones shown are the major ones for the elements represented in the table.

2.3.3 Atomic Masses

In addition to the atomic number, the Periodic Table of the Elements lists the atomic mass in unified atomic mass units (u) for each element. But since there are multiple isotopes for just about every element, *the atomic mass values in the periodic table represent the weighted average of the masses of naturally occurring isotopes.*

An example of a weighted average is the average age of the students in the junior class at your school. Let's say there are 47 juniors, 40 of whom are 16 years old and 7 of whom are 17 years old at the beginning of the school year. To determine the average age of these students, let's first determine the proportion of the students at each age.

$$\frac{40}{47} = 0.851 \text{ (85.1\%)}$$

$$\frac{7}{47} = 0.149 \text{ (14.9\%)}$$

Z	isotope	mass (u)	abundance (%)
1	hydrogen-1	1.0078	99.9885
1	hydrogen-2	2.0141	0.0115
6	carbon-12	12.0000	98.93
6	carbon-13	13.0034	1.078
14	silicon-28	27.9769	92.223
14	silicon-29	28.9765	4.685
14	silicon-30	29.9738	3.092
17	chlorine-35	34.9689	75.76
17	chlorine-37	36.9659	24.24
20	calcium-40	39.9626	96.941
20	calcium-42	41.9586	0.647
26	iron-54	53.9396	5.845
26	iron-56	55.9349	91.754
26	iron-57	56.9354	2.119
26	iron-58	57.9333	0.282
29	copper-63	62.9296	69.15
29	copper-65	64.9278	30.85
92	uranium-235	235.0439	0.7204
92	uranium-238	238.0508	99.2742

Table 2.6. Major isotopes for a few elements.

To calculate the average age, we first multiply each student age by the proportion of students of that age to find the contribution to the average from each age group. Then we add the contributions together to find the weighted average age for the junior class.

$$\begin{array}{r} 16 \text{ years} \cdot 0.851 = 13.6 \text{ years} \\ + 17 \text{ years} \cdot 0.149 = 2.53 \text{ years} \\ \hline = 16.1 \text{ years} \end{array}$$

We perform a similar calculation when computing the average atomic mass of an element from the masses of its isotopes, as shown in the following example.

▼ Example 2.1

Given the isotope masses and abundances for copper-63 and copper-65 in Table 2.6, determine the atomic mass for naturally occurring copper.

We need to multiply each isotope's mass by its abundance to get the isotope's contribution to the average mass of the element, which is the atomic mass. Then we add together the contributions from each isotope. The data from the table are:

copper-63: mass = 62.9296 u, abundance = 69.15%

copper-65: mass = 64.9278 u, abundance = 30.85%

$$\begin{array}{r} 62.9296 \text{ u} \cdot 0.6915 = 43.51 \text{ u} \\ + 64.9278 \text{ u} \cdot 0.3085 = 20.03 \text{ u} \\ \hline = 63.55 \text{ u} \end{array}$$

Compare this value to the value shown in the periodic table in Figure 2.14 or inside the back cover of the text.



The unified atomic mass unit, u, is defined as 1/12 the mass of an atom of carbon-12. Although the value of this mass is quite close to the masses of the proton and neutron, it is not exact because of the mass of the electrons in atoms of carbon-12, and also because of the mass-energy involved in binding the nucleus of the atom together. (The mass of nucleons bound together in a nucleus does not equal the sum of their individual masses.) Table 2.7 shows the masses of the three basic subatomic particles in unified atomic mass units.

Still, the proton and neutron masses are very close to unity (one) and the electron mass is extremely small. This means that for elements with a very large abundance of one isotope we can use the atomic mass and atomic number in the periodic table to determine the numbers of protons and neutrons in the nucleus of the most common isotope. For example, from Table 2.6, the mass of uranium-238 is very close to 238 u. Since an atom of uranium-238 has 92 protons, the balance of the mass is essentially all neutrons. Thus, there are $238 - 92 = 146$ neutrons in uranium-238.

particle	mass
proton	1.007277 u
neutron	1.008665 u
electron	0.0005486 u

Table 2.7. Masses in u of the three basic subatomic particles.

2.4 Density and Quantity of Substances

2.4.1 Density

Density is a physical property of substances. Density is a measure of how much matter is packed into a given volume for different substances. No doubt you are already familiar with the concept of density. You know that if you hold equally sized balloons in each hand, one filled with water and one filled with air, the water balloon weighs more because water is denser than air. You know that for equal weights of sand and Styrofoam packing peanuts the volume of the packing material is much larger because the packing material is much less dense. And you probably also know that objects less dense than water float, while objects denser than water sink.

The equation for density is

$$\rho = \frac{m}{V} \quad (2.1)$$

where the Greek letter ρ (spelled rho and pronounced “row,” which rhymes with snow) is the density in kg/m^3 , m is the mass in kg, and V is the volume in m^3 . These are the variables and units in the MKS unit system. However, since laboratory work typically involves only small quantities of substances, it is more common in chemistry for densities to be expressed in g/cm^3 (for solids) or g/mL (for liquids). In the examples that follow, I illustrate the use of g/cm^3 and kg/m^3 . Since $1 \text{ mL} = 1 \text{ cm}^3$ (see Table A.3 in Appendix A), calculations of densities in g/mL are essentially the same as those solving for densities in g/cm^3 . If all you are doing is using the density equation, then any of these units of measure is fine. One final item for you to note is that the density of water at room temperature is

$$\rho_w = 0.998 \frac{\text{g}}{\text{cm}^3} \quad (22.0^\circ\text{C})$$

This value is useful to know because water comes up in many different applications. The densities at other temperatures, along with other properties of water, are listed in Table A.5 of Appendix A.

▼ Example 2.2

The density of germanium is 5.323 g/cm^3 . A small sample of germanium has a mass of 17.615 g . Determine the volume of this sample.

Begin by writing the given information.

$$\rho = 5.323 \frac{\text{g}}{\text{cm}^3}$$

$$m = 17.615 \text{ g}$$

$$V = ?$$

Now write Equation (2.1) and solve for the volume.

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

$$\rho \cdot V = m$$

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

Next, insert the values and compute the result.

$$V = \frac{m}{\rho} = \frac{17.615 \text{ g}}{5.323 \frac{\text{g}}{\text{cm}^3}} = 3.309 \text{ cm}^3$$

This value has four significant digits, as it should based on the given information.



▼ Example 2.3

Determine the density of a block of plastic that has a mass of 1,860 g and dimensions 4.0 in × 2.5 in × 9.50 in. State your result in kg/m³.

To solve this problem, we use the given dimensions to calculate the volume of the block. Then we use Equation (2.1) to calculate the density.

Always begin your problem solutions by writing down the given information and performing any necessary unit conversions. Since the units of measure required for the result are kg/m³, we convert the mass to kilograms and the lengths to meters. When solving problems requiring unit conversions like this, write down the given information on separate lines down the left side of your page. Then perform the unit conversions by multiplying the conversion factors out to the right.

From the given information, our result must have two significant digits. This means we must perform the unit conversions and volume calculation with three significant digits (one more than we need), and round to two digits when we get our final result.

$$m = 1860 \text{ g} \cdot \frac{1 \text{ kg}}{1000 \text{ g}} = 1.86 \text{ kg}$$

$$l = 4.0 \text{ in} \cdot \frac{2.54 \text{ cm}}{1 \text{ in}} \cdot \frac{1 \text{ m}}{100 \text{ cm}} = 0.102 \text{ m}$$

$$w = 2.5 \text{ in} \cdot \frac{2.54 \text{ cm}}{1 \text{ in}} \cdot \frac{1 \text{ m}}{100 \text{ cm}} = 0.0635 \text{ m}$$

$$h = 9.5 \text{ in} \cdot \frac{2.54 \text{ cm}}{1 \text{ in}} \cdot \frac{1 \text{ m}}{100 \text{ cm}} = 0.241 \text{ m}$$

Now that the units are squared away, let's determine the volume of the block.

$$V = l \cdot w \cdot h = 0.102 \text{ m} \cdot 0.0635 \text{ m} \cdot 0.241 \text{ m} = 0.00156 \text{ m}^3$$

Finally, using Equation (2.1) the density is

$$\rho = \frac{m}{V} = \frac{1.86 \text{ kg}}{0.00156 \text{ m}^3} = 1192 \frac{\text{kg}}{\text{m}^3}$$

Rounding to two significant digits we have

$$1200 \frac{\text{kg}}{\text{m}^3}$$



2.4.2 The Mole and the Avogadro Constant

When solving problems in chemistry, we are generally working with chemical reactions in which huge numbers of atoms are involved, including all the naturally occurring isotopes, so performing reaction calculations with the masses of individual atoms is not practical. However, the average mass of a given multiple of some kind of atom is simply that multiple times the atomic mass. The mass of one million atoms of aluminum is 1,000,000 times the atomic mass of aluminum.

In chemistry, the standard bulk quantity of substance used in calculations is the *mole* (mol), the SI base unit for quantity of substance (Table 1.1). The mole is a particular number of particles of a substance, just as the terms *dozen*, *score*, and *gross* refer to specific numbers of things (12, 20, and 144, respectively). A mole is exactly $6.02214076 \times 10^{23}$ particles of a substance.

This value is known today as the *Avogadro constant*, N_A . More formally, the Avogadro constant is defined as exactly:

$$N_A = 6.02214076 \times 10^{23} \text{ mol}^{-1} \quad (2.2)$$

Usually, we just round this value to $6.022 \times 10^{23} \text{ mol}^{-1}$. In the next section, I'll describe why this value is what it is, instead of being a more convenient round number. For the moment, let's focus on what it means. Now, don't freak out over the unit of measure. Allow me to explain. Raising a unit of measure to the power -1 is mathematically equivalent to placing the unit in a denominator because $x^{-1} = \frac{1}{x}$. In other words, Equation (2.2) is the same thing as saying "6.02214076 $\times 10^{23}$ per mole." To make things even clearer, it's okay to say it this way: " N_A is about 6.022×10^{23} particles per mole." This is the way I like to think of it when performing unit conversions, as we do quite a lot in coming chapters. Without the units of measure, the value $6.02214076 \times 10^{23}$ is called *Avogadro's number*. With the units, it is called the *Avogadro constant*. Using this terminology, the mole can be defined this way:

A mole is the amount of a pure substance (element or compound) that contains Avogadro's number of particles of the substance.

Let's now consider what we mean when we refer to *particles* of a substance. For substances that exist as molecules, the particles are the molecules. Examples of these are the molecular substances we encountered back in Figure 2.18. For substances that exist as individual atoms, the particles are the individual atoms. Metals are like this, since a pure metal is composed of individual atoms of the same element joined together in a crystal lattice. The noble gases are also like this. The noble gases are located in the far right-hand column of the Periodic Table of the Elements. As I discuss more in coming chapters, atoms of noble gases are almost completely

unreactive—they don't bond with other atoms at all. At ordinary temperature and pressure, the noble gases are gases composed of individual atoms.

For crystalline compounds, the “particles” in a mole of the substance are the *formula units* in the crystal lattice. A formula unit is one set of the atoms represented by the chemical formula of the compound. For example, the chemical formula for calcium carbonate is CaCO_3 . One formula unit of calcium carbonate includes one calcium atom, one carbon atom, and three oxygen atoms.

The value of the Avogadro constant was determined approximately by French Physicist Jean Perrin (Figure 2.25) in the early 20th century. Perrin determined the value of the constant through several different experimental methods. In the 19th century, many scientists did not yet accept the existence of atoms as a scientific fact and Perrin's research put the atomic nature of matter beyond dispute. For this work, he received the Nobel Prize in Physics in 1926. Perrin proposed naming the constant after Amedeo Avogadro, a 19th-century Italian scientist who was the first to propose that the volume of a gas at a given temperature and pressure is proportional to the number of particles of the gas (atoms or molecules), regardless of the identity of the gas. In fact, at 0°C and atmospheric pressure, one mole of any gas occupies a volume of 22.4 L.

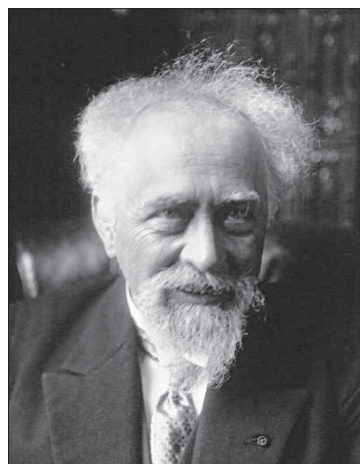


Figure 2.25. French physicist Jean Perrin (1870–1942).

2.4.3 Molar Mass and Formula Mass

Since 2019, the value of the Avogadro constant in Equation (2.2) is exact by definition. But the number has the value it does because it was originally chosen so that the average atomic mass in u of a molecule of a compound, as computed from the mass values in the periodic table, would be numerically equivalent to the mass of one mole of the compound in grams per mole. (We address these calculations below.) Now, recall that the definition of the unified atomic mass unit (or dalton) is such that an atom of carbon-12 has a mass of exactly 12 u. According to the original definition of Avogadro's number, there were also exactly 12 grams of carbon-12 in one mole of carbon-12. So according to these definitions, an atom of carbon-12 has a mass of exactly 12 u, and a mole of carbon-12 had a mass of exactly 12 grams. This quantity, the mass of one mole of a substance, is called the *molar mass*. Because of the way the molar and atomic masses were defined, the molar mass for an atom was numerically equivalent to the atomic mass. As a result of the 2019 redefinition of Avogadro's number, the atomic mass in u and the molecular mass in g/mol are no longer exactly equivalent. However, they are extremely close and may still be treated as equal for practical purposes. (The difference is a factor of only about 4×10^{-10} .)

Even though the exact equivalence ended in 2019, these are still very handy definitions! For example, from the periodic table we find that the average mass of one atom of silicon ($Z = 14$) is 28.0855 u. This also tells us that the mass of one mole of silicon is 28.0855 g, so the molar mass of silicon is 28.0855 g/mol. Likewise, from the periodic table we find that the average mass of one atom of copper ($Z = 29$) is 63.546 u. This also tells us that the mass of one mole of copper is 63.546 g, so the molar mass of copper is 63.546 g/mol. For the elements that exist as single atoms, the molar mass in g/mol and the atomic mass in u are numerically equivalent (almost).

From the periodic table, we can also determine the molar mass of compounds—the mass of one mole of the compound. We simply add up the molar masses for the elements in the chemical formula, taking into account any subscripts present in the formula, and we have the molar mass for the compound in g/mol. If we add up the element atomic masses in unified atomic mass

Quantity	Units	Definition
molar mass	g/mol	The mass of one mole of a substance, approximately equal to the sum of the atomic masses of the elements in a chemical formula, taking into account the subscripts indicating atomic ratios in the compound.
formula mass	u	The mass of one formula unit of a substance. Numerically nearly equivalent to the molar mass.
molecular mass	u	The average mass of a single molecule of a molecular substance. Numerically equivalent to the formula mass. (May also be converted to grams and expressed in grams, see Section 2.4.4.)

Table 2.8. Definitions and units for molar mass, formula mass, and molecular mass.

units, we obtain what is called the *formula mass* of the compound in u (or Da). If the compound is molecular, then the formula mass may also be referred to as the *molecular mass*, the average mass of a single molecule of the substance.

The details of these three different mass terms are summarized in Table 2.8.

▼ Example 2.4

Determine the molar mass and formula mass for water, H₂O. Note that since water is composed of molecules, the formula mass may also be called the molecular mass.

From the periodic table, the atomic masses of hydrogen (H) and oxygen (O) are:

H: 1.0079 u

O: 15.9994 u

Written as molar masses, these are:

H: $1.0079 \frac{\text{g}}{\text{mol}}$

O: $15.9994 \frac{\text{g}}{\text{mol}}$

The formula for water, H₂O, tells us there are two hydrogen atoms and one oxygen atom in each molecule of water, so we multiply these numbers by the molar masses and add them up to get the molar mass of H₂O.

$$\left(2 \times 1.0079 \frac{\text{g}}{\text{mol}}\right) + \left(1 \times 15.9994 \frac{\text{g}}{\text{mol}}\right) = 18.0152 \frac{\text{g}}{\text{mol}}$$

The calculation of the formula mass is identical, except we use units of u instead of g/mol. From the periodic table, the atomic masses of hydrogen (H) and oxygen (O) are:

H: 1.0079 u

O: 15.9994 u

There are two hydrogen atoms and one oxygen atom in each molecule, so we multiply these numbers by the element atomic masses and add them up to get the formula mass of H₂O.

$$(2 \times 1.0079 \text{ u}) + (1 \times 15.9994 \text{ u}) = 18.0152 \text{ u}$$

Thus, the formula mass for water is 18.0152 u. This value is also the molecular mass of water.



▼ Example 2.5

Nitrogen is one of several elements that exist in nature as diatomic gases. (The others include hydrogen, oxygen, fluorine, and chlorine.) Determine the molar mass of nitrogen gas, N_2 .

From the periodic table, the atomic mass of nitrogen (N), written as a molar mass for individual nitrogen atoms, is:

$$\text{N: } 14.0067 \frac{\text{g}}{\text{mol}}$$

There are two nitrogen atoms in each molecule of N_2 , so we multiply the molar mass of N by two to get the molar mass of N_2 .

$$\left(2 \times 14.0067 \frac{\text{g}}{\text{mol}} \right) = 28.0134 \frac{\text{g}}{\text{mol}}$$



▼ Example 2.6

Determine the mass in grams of 2.5 mol sodium bicarbonate, NaHCO_3 (baking soda).

In any problem like this, we first find the molar mass of the given compound. Then we simply use that molar mass to compute the mass of the given quantity. From the periodic table, the molar masses of the elements in the compound are:

$$\text{Na: } 22.9898 \frac{\text{g}}{\text{mol}}$$

$$\text{H: } 1.0079 \frac{\text{g}}{\text{mol}}$$

$$\text{C: } 12.011 \frac{\text{g}}{\text{mol}}$$

$$\text{O: } 15.9994 \frac{\text{g}}{\text{mol}}$$

The oxygen appears three times in the formula, so its mass must be multiplied by three and added to the others.

$$22.9898 \frac{\text{g}}{\text{mol}} + 1.0079 \frac{\text{g}}{\text{mol}} + 12.011 \frac{\text{g}}{\text{mol}} + \left(3 \times 15.9994 \frac{\text{g}}{\text{mol}} \right) = 84.007 \frac{\text{g}}{\text{mol}}$$

This value is the molar mass of NaHCO_3 . To find the mass of 2.5 mol, we multiply:

$$2.5 \text{ mol} \cdot 84.007 \frac{\text{g}}{\text{mol}} = 210 \text{ g}$$

Notice that the significant digits in the molar mass are obtained by the addition rule. The atomic mass of carbon, 12.011 u, only goes to three decimal places, so the molar mass of the compound goes to three decimal places. The final result is obtained by multiplying the compound's molar mass by the quantity 2.5 mol, which only has two significant digits. Thus, the result must have two significant digits as well.



▼ Example 2.7

A scientist measures out 125 g of potassium chloride (KCl). How many moles of KCl does this quantity represent?

First, determine the molar mass of KCl. From the periodic table:

$$\text{K: } 39.098 \frac{\text{g}}{\text{mol}}$$

$$\text{Cl: } 35.4527 \frac{\text{g}}{\text{mol}}$$

The formula includes one atom of each, so we add them to obtain the molar mass:

$$39.098 \frac{\text{g}}{\text{mol}} + 35.4527 \frac{\text{g}}{\text{mol}} = 74.551 \frac{\text{g}}{\text{mol}}$$

Beginning now, *always think of the molar mass of any substance as a conversion factor* that can be written right side up or upside down to convert grams to moles or vice versa. For KCl, 74.551 g is equivalent to 1 mol, so these quantities can be written as conversion factors, like this:

$$\frac{74.551 \text{ g}}{1 \text{ mol}} = \frac{1 \text{ mol}}{74.551 \text{ g}}$$

This makes the last step of this problem easy. Just select the way of writing the molar mass conversion factor that cancels out the given units (g) and gives the units required (mol). This is nothing but a unit conversion.

$$125 \text{ g} \cdot \frac{1 \text{ mol}}{74.551 \text{ g}} = 1.68 \text{ mol}$$



The photograph in Figure 2.26 shows one mole of each of four substances. The first is one mole of copper, equal to 63.5 g. The second is a 250-mL beaker containing one mole of water. As you can see, this is not much water—only 18 mL. At the upper right is a weigh tray



Figure 2.26. Clockwise from left are shown 1 mole of copper, 1 mole of water, 1 mole of table salt, and 1 mole of baking soda.

containing one mole of sodium chloride, 40.0 g. (This is just under 1/4 cup.) Finally, one mole of baking soda, 84.1 g. (This is right at 1/3 cup.)

▼ Example 2.8

Calculate the number of water molecules in a 1.00-liter bottle of water.

The logic of this problem, in reverse, is as follows: To calculate a number of molecules, we must use the Avogadro constant. To use the Avogadro constant, we need to know the number of moles of water we have. To determine the number of moles, we need to know both the molar mass and the mass of the water. To determine the mass from a volume, we use the density equation.

So we begin with the given information and the density equation to determine the mass of water we have. The given information and unit conversions are as follows:

$$V = 1.00 \text{ L} \cdot \frac{1000 \text{ cm}^3}{1 \text{ L}} = 1.00 \times 10^3 \text{ cm}^3$$

$$\rho = 0.998 \frac{\text{g}}{\text{cm}^3}$$

$$m = ?$$

Now we write down Equation (2.1) and solve for the mass:

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

$$m = \rho \cdot V = 0.998 \frac{\text{g}}{\text{cm}^3} \cdot 1.00 \times 10^3 \text{ cm}^3 = 998 \text{ g}$$

Next, we need the molar mass of water. We calculated this in Example 2.4 and obtained 18.0152 g/mol. We use this molar mass as a conversion factor to convert the mass of water into a number of moles of water:

$$998 \text{ g} \cdot \frac{1 \text{ mol}}{18.0152 \text{ g}} = 55.40 \text{ mol}$$

This intermediate result has four significant digits—one more than we need in the final result. Finally, with the number of moles in hand we use the Avogadro constant to determine how many particles of water this is, which is identical to the number of water molecules.

$$55.40 \text{ mol} \cdot \frac{6.022 \times 10^{23} \text{ particles}}{\text{mol}} = 3.34 \times 10^{25} \text{ particles}$$



2.4.4 Gram Masses of Atoms and Molecules

The atomic mass from the periodic table and the Avogadro constant can be used to calculate the mass in grams of an individual atom. Recall that the atomic mass value in the periodic table gives both the average atomic mass in u, and the molar mass in g/mol. Knowing the molar mass in g/mol we can simply divide by the number of atoms there are in one mole to find the mass of

one atom in grams. Although this kind of calculation is quite simple, I have found that it is *very* easy for students to get confused and not be able to determine whether one should multiply or divide or what. So here's a problem solving tip: let the units of measure help you figure out what to do. If you include the units of measure in your work and pay attention to how the units cancel out or don't cancel out, these calculations are pretty straightforward. *Keep this principle firmly in mind throughout your study of chemistry!* Units of measure are not an annoying burden; they are the student's friend.

▼ Example 2.9

Determine the average mass in grams of an atom of boron.

From the periodic table, we find that the molar mass of boron is 10.811 g/mol. One mole consists of Avogadro's number of atoms of boron, so if we divide the molar mass by the Avogadro constant, we have the mass of a single atom of boron. Let's begin by setting up the division I just described, and then use the old invert-and-multiply trick for fraction division to help with the unit cancellations.

$$\begin{aligned} \frac{10.811 \frac{\text{g}}{\text{mol}}}{6.0221 \times 10^{23} \frac{\text{particles}}{\text{mol}}} &= 10.811 \frac{\text{g}}{\text{mol}} \cdot \frac{1}{6.0221 \times 10^{23}} \frac{\text{mol}}{\text{particles}} \\ &= \frac{10.811}{6.0221 \times 10^{23}} \frac{\text{g}}{\text{particle}} = 1.7952 \times 10^{-23} \frac{\text{g}}{\text{particle}} \end{aligned}$$

So the average mass of one boron atom is 1.7952×10^{-23} g. Note that I use five digits in the value of the Avogadro constant to preserve the precision we have in the molar mass.



For molecular substances, the molar mass is used to compute the molecular mass in grams—the average mass of one molecule. This is done by first computing the molar mass of the compound, just as we did before. Then we simply divide by the Avogadro constant to obtain the mass of a single molecule.

Like the atomic mass, the molecular mass is an average mass, since the atomic masses used in calculating the molar mass are all based on the average mass of different isotopes with their abundances taken into account. The molecular mass for a *specific* molecule must be calculated using the specific masses of the nuclides in the molecule.

▼ Example 2.10

Determine the mass in grams of one molecule of carbon tetrachloride, CCl_4 .

From the periodic table we find that the molar masses of carbon and chlorine are 12.011 g/mol and 35.4527 g/mol, respectively. From this we calculate the molar mass of CCl_4 :

$$\left(1 \times 12.011 \frac{\text{g}}{\text{mol}}\right) + \left(4 \times 35.4527 \frac{\text{g}}{\text{mol}}\right) = 153.822 \frac{\text{g}}{\text{mol}}$$

With this molar mass we can use the Avogadro constant to get the molecular mass in grams. This time, instead of writing the Avogadro constant in the denominator of a big fraction, I simply treat it as a conversion factor and write it in the equation such that the mole units cancel

out. (This is the way I always perform such calculations.) I also use six digits in the Avogadro constant to preserve the precision we have in the molar mass.

$$153.822 \frac{\text{g}}{\text{mol}} \cdot \frac{1 \text{ mol}}{6.02214 \times 10^{23} \text{ particles}} = 2.55427 \times 10^{-22} \frac{\text{g}}{\text{particle}}$$



Chapter 2 Exercises

SECTION 2.1

1. Write paragraphs describing the experiments performed by J.J. Thomson, Robert Millikan, and Ernest Rutherford.
2. Describe the main points or features in the atomic models proposed by John Dalton, J.J. Thomson, and Ernest Rutherford.
3. Explain why Ernest Rutherford found the reflection of alpha particles off gold foil so astonishing.

SECTION 2.2

4. Write paragraphs distinguishing between these pairs of terms:
 - a. compounds and elements
 - b. mixtures and compounds
 - c. heterogeneous mixtures and homogeneous mixtures
 - d. suspensions and colloids
5. Classify each of the following as element, compound, homogeneous mixture, or heterogeneous mixture.

a. water	b. cesium chloride	c. pond water	d. methane
e. a soft drink	f. nitric acid	g. black coffee	h. argon
i. air	j. hydrogen nitrate	k. exhaust fumes	l. quartz
m. brass	n. hydrogen gas	o. hydrogen cyanide	p. mouthwash
q. platinum	r. dirt	s. radon	t. a smoothie
6. Explain why salt water and sugar water are homogeneous mixtures while automotive paint, which contains invisible particulates, is not.
7. Write a paragraph describing the two basic types of structures atoms can take when bonding together.
8. Select three pure substances not mentioned in the chapter. For each substance, list at least eight physical properties and three chemical properties.
9. Explain why colloids reflect light. What is this effect called?

10. Identify each of the following as a physical change or a chemical change. For each, explain your choice.
- | | | |
|------------------------|-----------------------------|-----------------------------------|
| a. an avalanche | b. a cigar burning | c. spilling a glass of milk |
| d. digesting your food | e. swatting a fly | f. stirring cream into coffee |
| g. firing a pop gun | h. firing a real gun | i. boiling mercury |
| j. welding steel | k. filling a helium balloon | l. allowing molten iron to harden |
| m. frying chicken | n. snow melting | o. a car exhaust pipe rusting |
| p. paint “drying” | q. wood rotting | r. a ball rolling down a hill |

SECTION 2.3

11. How is the unified atomic mass unit, u , defined?
12. Referring to Table 2.6, calculate the atomic mass for silicon, calcium, and uranium. Compare your results to the values shown in the periodic table.
13. Which two nuclides in Table 2.6 have 20 neutrons?
14. In Table 2.6, how many protons, neutrons, and electrons are there in the heaviest nuclide listed? How many protons, neutrons, and electrons are there in the lightest nuclide listed?
15. As mentioned in the chapter, the sum of the masses of the particles in an atom does not equal the mass of the atom. Some of the mass of the individual particles is converted to energy, and the atom weighs less than the sum of the weights of its parts. How much mass is converted into energy when the individual protons, neutrons, and electrons are assembled to form an atom of uranium-238?

SECTION 2.4

16. What is the density of carbon dioxide gas if 0.196 g of the gas occupies a volume of 100.1 mL?
17. Oil floats because its density is less than that of water. Determine the volume of 550 g of a particular oil with a density of 955 kg/m^3 . State your answer in mL.
18. A factory orders 15.7 kg of germanium. The density of germanium is 5.32 g/cm^3 . Calculate the volume of this material and state your answer both in m^3 and cm^3 .
19. A graduated cylinder contains 23.35 mL of water. An irregularly shaped stone is placed into the cylinder, raising the volume to 27.79 mL. If the mass of the stone is 32.1 g, what is the density of the stone?
20. A standard 55-gallon drum is 34.5 inches tall and 24 inches in diameter. Consider a 55-gallon drum filled with kerosene. Using the dimensions in inches to calculate the volume, determine the mass of kerosene that fills this drum, given that the density of kerosene is 810 kg/m^3 .
21. Iron has a density of $7,830 \text{ kg/m}^3$. An iron block is 2.1 cm by 3.5 cm at the base and has a mass of 94.5 g. How tall is the block?
22. A student measures out 22.5 mL of mercury and finds the mass to be 306 g. Determine the density of mercury, and state your answer in kg/m^3 .
23. A large contemporary water tower holds over 3 million gallons of water. Determine the mass in kilograms of 3.0×10^6 gallons of water.
24. The famous Kon-Tiki was a raft sailed by Norwegian explorers in 1947 from South America to the Polynesian islands in the South Pacific. (Thor Heyerdahl’s book about it is a great read.) The trip took three and a half months and covered 4,300 miles in the Pacific Ocean. The crew of six men built the main section of the raft out of 9 massive balsa trees, each 2.0 ft

in diameter. Assume for simplicity that the balsa trees had a typical balsa wood density of 160 kg/m^3 , and the logs were each 45 ft in length. Determine the total mass in kilograms of these logs used to build the raft.

25. How is the mole defined?

26. Determine the number of atoms in each of the following.

- a. 73.2 g Cu b. 1.35 mol Na c. 1.5000 kg W

27. Determine the mass in grams for each of the following.

- a. 6.022×10^{23} atoms K b. 100 atoms Au c. 0.00100 mol Xe
d. 2.0 mol Li e. 4.2120 mol Br f. 7.422×10^{22} atoms Pt

28. Determine the number of moles present in each of the following.

- a. 25 g $\text{Ca}(\text{OH})_2$ b. 286.25 g $\text{Al}_2(\text{CrO}_4)_3$ c. 2.111 kg KCl
d. 47.50 g LiClO_3 e. 10.0 g O_2 f. 1.00 mg $\text{C}_{14}\text{H}_{18}\text{N}_2\text{O}_5$

29. Calculate the molar mass for each of the following compounds or molecules.

- a. ammonia, NH_3 b. carbon dioxide, CO_2 c. chlorine gas, Cl_2
d. copper(II) sulfate, CuSO_4 e. calcium nitrite, $\text{Ca}(\text{NO}_2)_2$ f. sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$
g. ethanol, $\text{C}_2\text{H}_5\text{OH}$ h. propane, C_3H_8 i. glass, SiO_2

30. Determine the formula masses for these compounds:

- a. MgCl_2 b. $\text{Ca}(\text{NO}_3)_2$ c. $(\text{SO}_4)^{2-}$ (The 2- indicates this is an ion with an electrical charge of -2. The charge does not affect your calculation.)
d. CuSO_4 e. BF_3 f. CCl_4

31. Determine the mass in grams of 2.25 mol silver nitrate, AgNO_3 .

32. Given 2.25 kg CCl_4 , answer these questions:

- a. How many moles CCl_4 are present?
b. How many carbon atoms are present?
c. Approximately how many carbon-13 atoms are present?

33. Given 1.00 gal H_2O at 4°C , answer the questions below. (Hint: You must use the appropriate volume conversion and the density of water to determine the mass of 1.00 gal H_2O . See the information in Tables A.3 and A.5 in Appendix A.)

- a. How many moles H_2O are present?
b. How many hydrogen atoms are present?
c. Approximately how many deuterium (hydrogen-2) atoms are present?