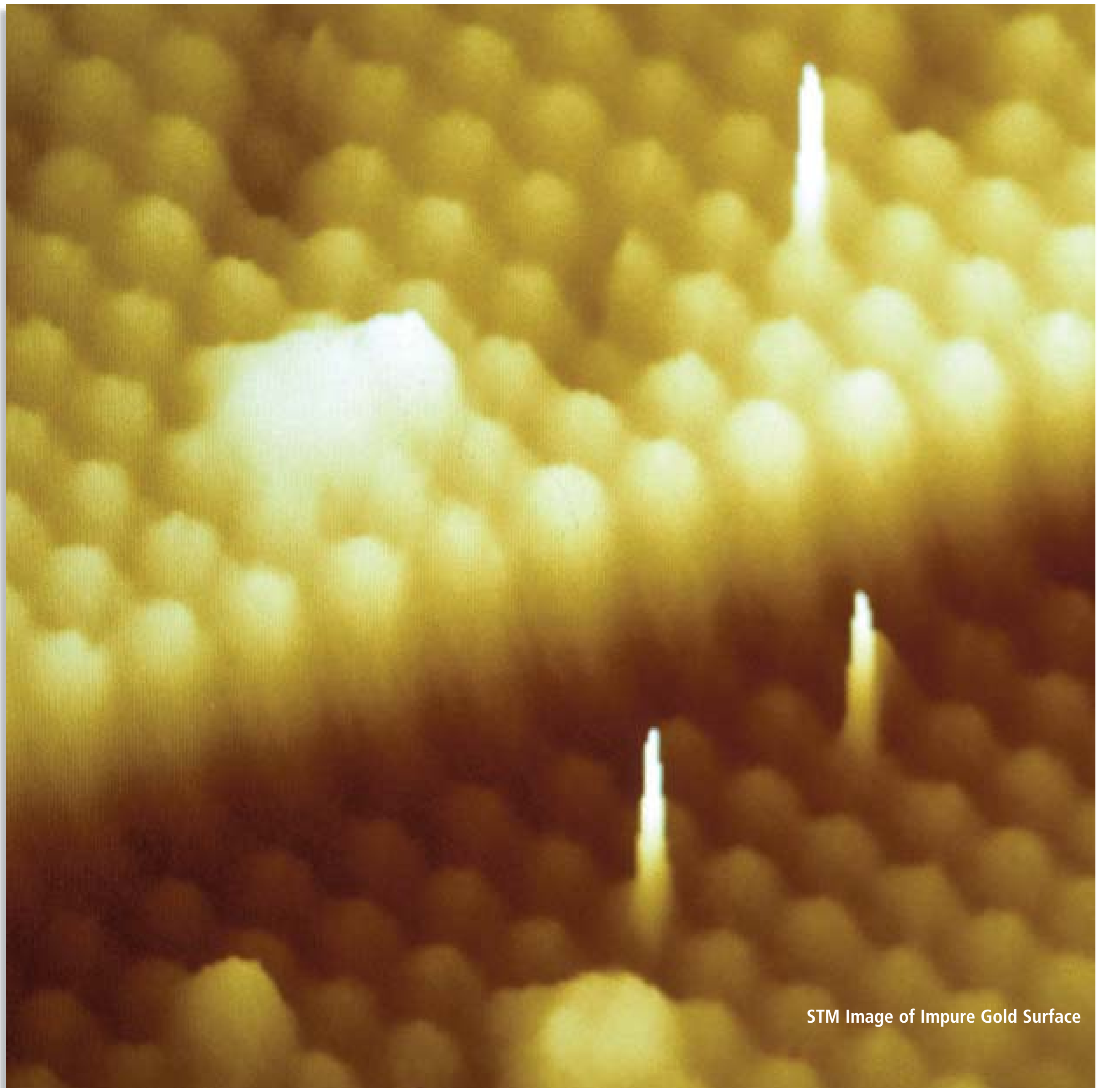
CHAPTER 3

Atoms: The Building Blocks of Matter

An atom is the smallest particle of an element that retains the chemical properties of that element.



STM Image of Impure Gold Surface

The Atom: From Philosophical Idea to Scientific Theory

SECTION 1

Objectives

- Explain the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.
- Summarize the five essential points of Dalton's atomic theory.
- Explain the relationship between Dalton's atomic theory and the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.

When you crush a lump of sugar, you can see that it is made up of many smaller particles of sugar. You may grind these particles into a very fine powder, but each tiny piece is still sugar. Now suppose you dissolve the sugar in water. The tiny particles seem to disappear completely. Even if you look at the sugar-water solution through a powerful microscope, you cannot see any sugar particles. Yet if you were to taste the solution, you'd know that the sugar is still there. Observations like these led early philosophers to ponder the fundamental nature of matter. Is it continuous and infinitely divisible, or is it divisible only until a basic, invisible particle that cannot be divided further is reached?

The particle theory of matter was supported as early as 400 B.C. by certain Greek thinkers, such as Democritus. He called nature's basic particle an *atom,* based on the Greek word meaning "indivisible." Aristotle was part of the generation that succeeded Democritus. His ideas had a lasting impact on Western civilization, and he did not believe in atoms. He thought that all matter was continuous, and his opinion was accepted for nearly 2000 years. Neither the view of Aristotle nor that of Democritus was supported by experimental evidence, so each remained speculation until the eighteenth century. Then scientists began to gather evidence favoring the atomic theory of matter.

extension

Historical Chemistry

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Foundations of Atomic Theory

Virtually all chemists in the late 1700s accepted the modern definition of an element as a substance that cannot be further broken down by ordinary chemical means. It was also clear that elements combine to form compounds that have different physical and chemical properties than those of the elements that form them. There was great controversy, however, as to whether elements always combine in the same ratio when forming a particular compound.

The transformation of a substance or substances into one or more new substances is known as a *chemical reaction*. In the 1790s, the study of matter was revolutionized by a new emphasis on the quantitative

FIGURE 1 Each of the salt crystals shown here contains exactly 39.34% sodium and 60.66% chlorine by mass.

analysis of chemical reactions. Aided by improved balances, investigators began to accurately measure the masses of the elements and compounds they were studying. This led to the discovery of several basic laws. One of these laws was the **law of conservation of mass,** *which states that mass is neither created nor destroyed during ordinary chemical reactions or physical changes*. This discovery was soon followed by the assertion that, regardless of where or how a pure chemical compound is prepared, it is composed of a fixed proportion of elements. For example, sodium chloride, also known as ordinary table salt, *always* consists of 39.34% by mass of the element sodium, Na, and 60.66% by mass of the element chlorine, Cl. *The fact that a chemical compound contains the same elements in exactly the same proportions by mass regardless of the size of the sample or source of the compound is known as the* **law of definite proportions.**

It was also known that two elements sometimes combine to form more than one compound. For example, the elements carbon and oxygen form two compounds, carbon dioxide and carbon monoxide. Consider samples of each of these compounds, each containing 1.00 g of carbon. In carbon dioxide, 2.66 g of oxygen combine with 1.00 g of carbon. In carbon monoxide, 1.33 g of oxygen combine with 1.00 g of carbon. The ratio of the masses of oxygen in these two compounds is 2.66 to 1.33, or 2 to 1. This illustrates the **law of multiple proportions:** *If two or more different compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element is always a ratio of small whole numbers.*

Dalton's Atomic Theory

In 1808, an English schoolteacher named John Dalton proposed an explanation for the law of conservation of mass, the law of definite proportions, and the law of multiple proportions. He reasoned that elements were composed of atoms and that only whole numbers of atoms can combine to form compounds. His theory can be summed up by the following statements.

1. All matter is composed of extremely small particles called atoms.
2. Atoms of a given element are identical in size, mass, and other properties; atoms of different elements differ in size, mass, and other properties.
3. Atoms cannot be subdivided, created, or destroyed.
4. Atoms of different elements combine in simple whole-number ratios to form chemical compounds.
5. In chemical reactions, atoms are combined, separated, or rearranged.

According to Dalton's atomic theory, the law of conservation of mass is explained by the fact that chemical reactions involve merely the combination, separation, or rearrangement of atoms and that during these processes atoms are not subdivided, created, or destroyed. This

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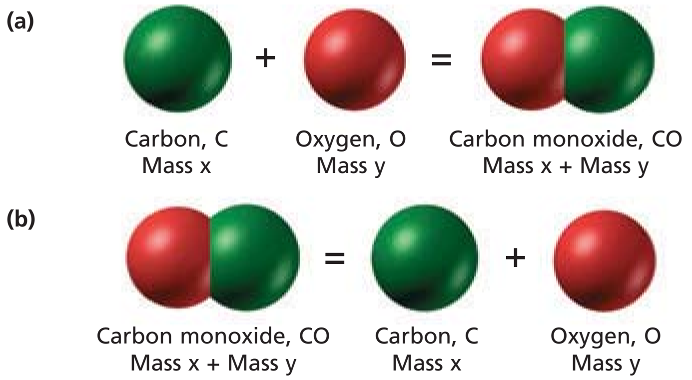


FIGURE 2 (a) An atom of carbon, C, and an atom of oxygen, O, can combine chemically to form a molecule of carbon monoxide, CO. The mass of the CO molecule is equal to the mass of the C atom plus the mass of the O atom. (b) The reverse holds true in a reaction in which a CO molecule is broken down into its elements.

idea is illustrated in **Figure 2** for the formation of carbon monoxide from carbon and oxygen.

The law of definite proportions, on the other hand, results from the fact that a given chemical compound is always composed of the same combination of atoms (see **Figure 3**). As for the law of multiple proportions, in the case of the carbon oxides, the 2-to-1 ratio of oxygen masses results because carbon dioxide always contains twice as many atoms of oxygen (per atom of carbon) as does carbon monoxide. This can also be seen in **Figure 3.**

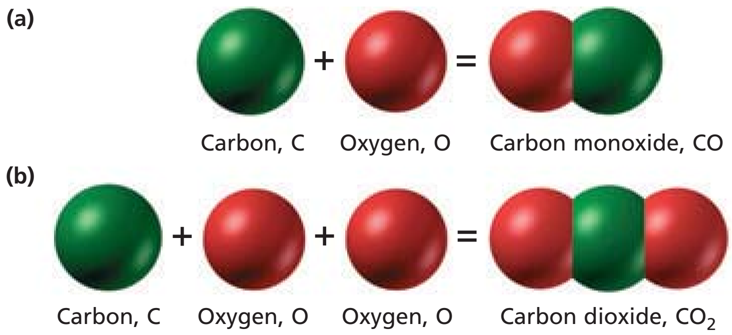


FIGURE 3 (a) CO molecules are always composed of one C atom and one O atom. (b) CO_2 molecules are always composed of one C atom and two O atoms. Note that a molecule of carbon dioxide contains twice as many oxygen atoms as does a molecule of carbon monoxide.

Modern Atomic Theory

By relating atoms to the measurable property of mass, Dalton turned Democritus's *idea* into a *scientific theory* that could be tested by experiment. But not all aspects of Dalton's atomic theory have proven to be correct. For example, today we know that atoms are divisible into even smaller particles (although the law of conservation of mass still holds true for chemical reactions). And, as you will see in Section 3, we know that a given element can have atoms with different masses. Atomic theory has not been discarded, however. Instead, it has been modified to explain the new observations. The important concepts that (1) all matter is composed of atoms and that (2) atoms of any one element differ in properties from atoms of another element remain unchanged.

CAREERS in Chemistry

Physical Chemist

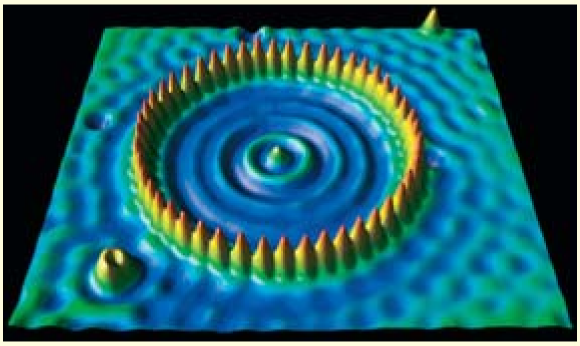
Physical chemists focus on understanding the physical properties of atoms and molecules. They are driven by a curiosity of what makes things work at the level of atoms, and they enjoy being challenged. In addition to chemistry, they study mathematics and physics extensively. Laboratory courses involving experience with electronics and optics are typically part of their training. Often, they enjoy working with instruments and computers. Physical chemists can be experimentalists or theoreticians. They use sophisticated instruments to make measurements, or high-powered computers to perform intensive calculations. The instruments used include lasers, electron microscopes, nuclear magnetic resonance spectrometers, mass spectrometers, and particle accelerators. Physical chemists work in industry, government laboratories, research institutes, and academic institutions. Because physical chemists work on a wide range of problems, taking courses in other science disciplines is important.

Scanning Tunneling Microscopy

For years, scientists have yearned for the ability to "see" individual atoms. Because atoms are so small, this had been nothing more than a dream. Now, the scanning tunneling microscope, STM, gives scientists the ability to look at individual atoms. It was invented in 1981 by Gerd Binnig and Heinrich Rohrer, scientists working for IBM in Zurich, Switzerland. They shared the 1986 Nobel Prize in physics for their discovery.

The basic principle of STM is based on the current that exists between a metallic needle that is sharpened to a single atom, the probe, and a conducting sample. As the probe passes above the surface of the sample at a distance of one or two atoms, electrons can "tunnel" from the needle tip to the sample's surface. The probe moves across, or "scans," the surface of the sample. When the probe comes close to the electrons of an individual atom, a signal is produced. A weaker signal is produced between atoms. These signals build a topographical (hill and valley) "map" of conducting and nonconducting regions. The resulting map shows the position and spacing of atoms.

Surface chemistry is a developing subdiscipline in physical chemistry, and STM is an important tool in the field. Scientists use STM to study surface reactions, such as those that take place in catalytic converters. Other areas of research in which STM is useful include semiconductors and microelectronics. Usually, STM is used with materials that conduct, but it has also been used to study biological molecules, such as DNA.



▲ *This STM image shows a "corral" of iron atoms on a copper surface.*

One innovative application of STM is the ability to position individual atoms. The figure shows the result of moving individual atoms. First, iron atoms were placed on a copper surface. Then, individual iron atoms were picked up by the probe and placed in position. The result is a "quantum corral" of 48 iron atoms on the surface of copper. The diameter of the corral is about 14 nm.

Questions

1. In addition to chemistry, what kinds of courses are important for a student interested in a physical chemistry career?
2. What part of an atom is detected by STM?

QuickLAB

Wear safety goggles and an apron.

Constructing a Model

Question

How can you construct a model of an unknown object by (1) making inferences about an object that is in a closed container and (2) touching the object without seeing it?

Procedure

Record all of your results in a data table.

1. Your teacher will provide you with a can that is covered by a sock sealed with tape. Without unsealing the container, try to determine the number of objects inside the can as well as the mass, shape, size, composition, and texture of each. To do this, you may carefully tilt or shake the can. Record your observations in a data table.
2. Remove the tape from the top of the sock. Do *not* look inside the can. Put one hand through the opening, and make the same observations as in step 1 by handling the objects. To make more-accurate estimations, practice estimating the sizes and masses of some known objects outside the can. Then compare your estimates of these objects with actual measurements using a metric ruler and a balance.

Discussion

1. Scientists often use more than one method to gather data. How was this illustrated in the investigation?
2. Of the observations you made, which were qualitative and which were quantitative?
3. Using the data you gathered, draw a model of the unknown object(s) and write a brief summary of your conclusions.

Materials

- can covered by a sock sealed with tape
- one or more objects that fit in the container
- metric ruler
- balance

SECTION REVIEW

1. List the five main points of Dalton's atomic theory.
2. What chemical laws can be explained by Dalton's theory?

Critical Thinking

3. **ANALYZING INFORMATION** Three compounds containing potassium and oxygen are compared. Analysis shows that for each 1.00 g of O, the compounds have 1.22 g, 2.44 g, and 4.89 g of K, respectively. Show how these data support the law of multiple proportions.

SECTION 2

The Structure of the Atom

Objectives

- Summarize the observed properties of cathode rays that led to the discovery of the electron.
- Summarize the experiment carried out by Rutherford and his co-workers that led to the discovery of the nucleus.
- List the properties of protons, neutrons, and electrons.
- Define *atom.*

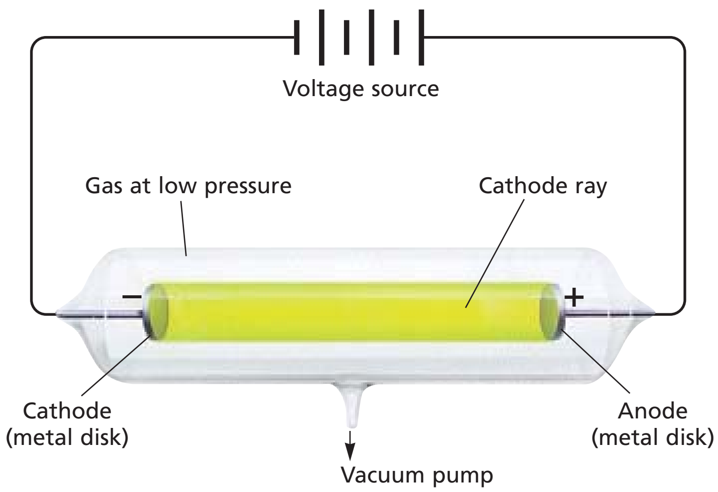
Although John Dalton thought atoms were indivisible, investigators in the late 1800s proved otherwise. As scientific advances allowed a deeper exploration of matter, it became clear that atoms are actually composed of several basic types of smaller particles and that the number and arrangement of these particles within an atom determine that atom's chemical properties. Today we define an **atom** *as the smallest particle of an element that retains the chemical properties of that element*.

All atoms consist of two regions. The *nucleus* is a very small region located at the center of an atom. In every atom, the nucleus is made up of at least one positively charged particle called a *proton* and usually one or more neutral particles called *neutrons.* Surrounding the nucleus is a region occupied by negatively charged particles called *electrons.* This region is very large compared with the size of the nucleus. Protons, neutrons, and electrons are often referred to as *subatomic particles*.

Discovery of the Electron

The first discovery of a subatomic particle resulted from investigations into the relationship between electricity and matter. In the late 1800s, many experiments were performed in which electric current was passed through various gases at low pressures. (Gases at atmospheric pressure don't conduct electricity well.) These experiments were carried out in glass tubes like the one shown in **Figure 4.** Such tubes are known as *cathode-ray tubes*.

FIGURE 4 A simple cathode-ray tube. Particles pass through the tube from the *cathode,* the metal disk connected to the negative terminal of the voltage source, to the *anode,* the metal disk connected to the positive terminal.



Cathode Rays and Electrons

Investigators noticed that when current was passed through a cathode-ray tube, the surface of the tube directly opposite the cathode glowed. They hypothesized that the glow was caused by a stream of particles, which they called a cathode ray. The ray traveled from the cathode to the anode when current was passed through the tube. Experiments devised to test this hypothesis revealed the following observations.

1. Cathode rays were deflected by a magnetic field in the same manner as a wire carrying electric current, which was known to have a negative charge (see **Figure 5**).
2. The rays were deflected away from a negatively charged object.

These observations led to the hypothesis that the particles that compose cathode rays are negatively charged. This hypothesis was strongly supported by a series of experiments carried out in 1897 by the English physicist Joseph John Thomson. In one investigation, he was able to measure the ratio of the charge of cathode-ray particles to their mass. He found that this ratio was always the same, regardless of the metal used to make the cathode or the nature of the gas inside the cathode-ray tube. Thomson concluded that all cathode rays are composed of identical negatively charged particles, which were named electrons.

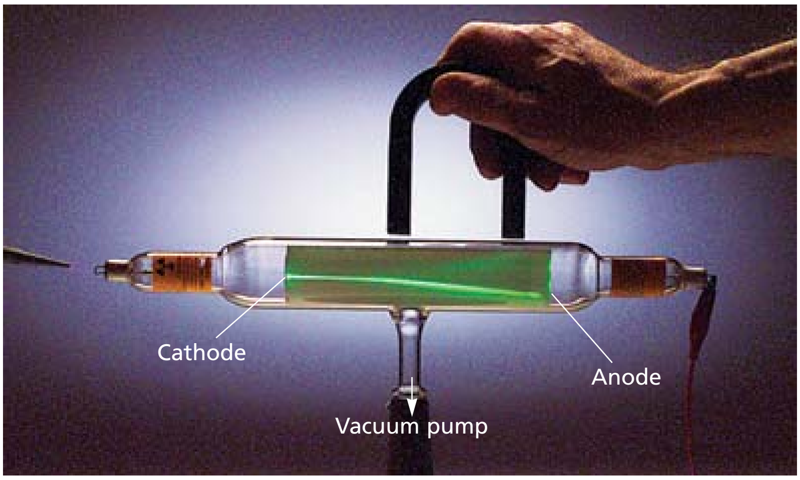


FIGURE 5 A magnet near the cathode-ray tube causes the beam to be deflected. The deflection indicates that the particles in the beam have a negative charge.

Charge and Mass of the Electron

Thomson's experiment revealed that the electron has a very large charge-to-mass ratio. Because cathode rays have identical properties regardless of the element used to produce them, it was concluded that electrons are present in atoms of all elements. Thus, cathode-ray experiments provided evidence that atoms are divisible and that one of the atom's basic constituents is the negatively charged electron. In 1909, experiments conducted by the American physicist Robert A. Millikan measured the charge of the electron. Scientists used this information and the charge-to-mass ratio of the electron to determine that the mass of the electron is about one two-thousandth the mass of the simplest type of hydrogen atom, which is the smallest atom known. More-accurate experiments conducted since then indicate that the electron has a mass of 9.109×10^{-31} kg, or 1/1837 the mass of the simplest type of hydrogen atom.

Based on what was learned about electrons, two other inferences were made about atomic structure.

1. Because atoms are electrically neutral, they must contain a positive charge to balance the negative electrons.
2. Because electrons have so much less mass than atoms, atoms must contain other particles that account for most of their mass.

Thomson proposed a model for the atom that is called the *plum pudding model* (after the English dessert). He believed that the negative electrons were spread evenly throughout the positive charge of the rest of the atom. This arrangement is similar to that of seeds in a watermelon: the seeds are spread throughout the fruit but do not contribute much to the overall mass. However, shortly thereafter, new experiments disproved this model.

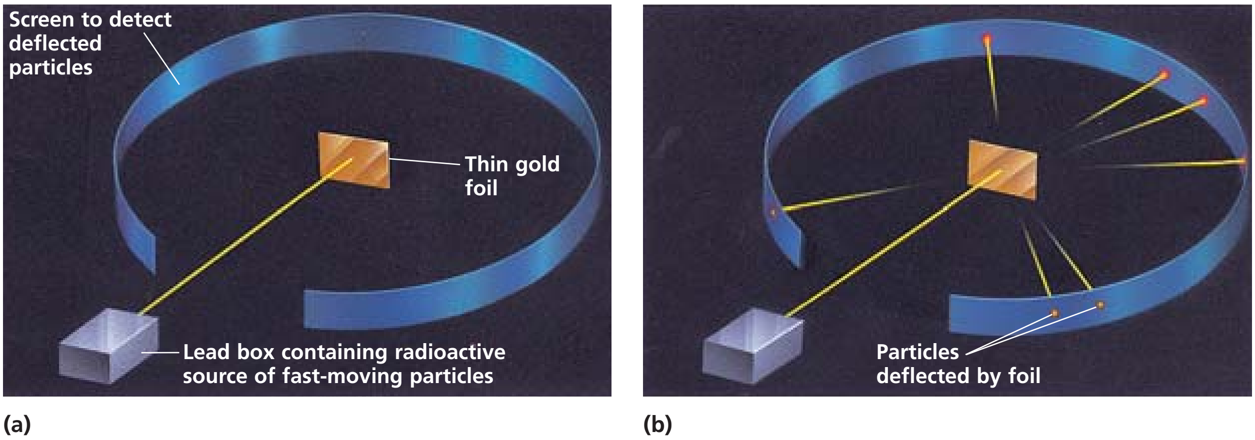
Discovery of the Atomic Nucleus

More detail of the atom's structure was provided in 1911 by New Zealander Ernest Rutherford and his associates Hans Geiger and Ernest Marsden. The scientists bombarded a thin piece of gold foil with fast-moving *alpha particles,* which are positively charged particles with about four times the mass of a hydrogen atom. Geiger and Marsden assumed that mass and charge were uniformly distributed throughout the atoms of the gold foil. They expected the alpha particles to pass through with only a slight deflection, and for the vast majority of the particles, this was the case. However, when the scientists checked for the possibility of wide-angle deflections, they were shocked to find that roughly 1 in 8000 of the alpha particles had actually been deflected back toward the source (see **Figure 6**). As Rutherford later exclaimed, it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue paper and it came back and hit you."

After thinking about the startling result for a few months, Rutherford finally came up with an explanation. He reasoned that the deflected alpha particles must have experienced some powerful force within the atom. And he figured that the source of this force must occupy a very small amount of space because so few of the total number of alpha particles had been affected by it. He concluded that the force must be caused by a very densely packed bundle of matter with a positive electric charge. Rutherford called this positive bundle of matter the nucleus (see **Figure 7**).

Rutherford had discovered that the volume of a nucleus was very small compared with the total volume of an atom. In fact, if the nucleus were the size of a marble, then the size of the atom would be about the size of a football field. But where were the electrons? This question was not answered until Rutherford's student, Niels Bohr, proposed a model in which electrons surrounded the positively charged nucleus as the planets surround the sun. Bohr's model will be discussed in Chapter 4.

FIGURE 6 (a) Geiger and Marsden bombarded a thin piece of gold foil with a narrow beam of alpha particles. (b) Some of the particles were deflected by the gold foil back toward their source.



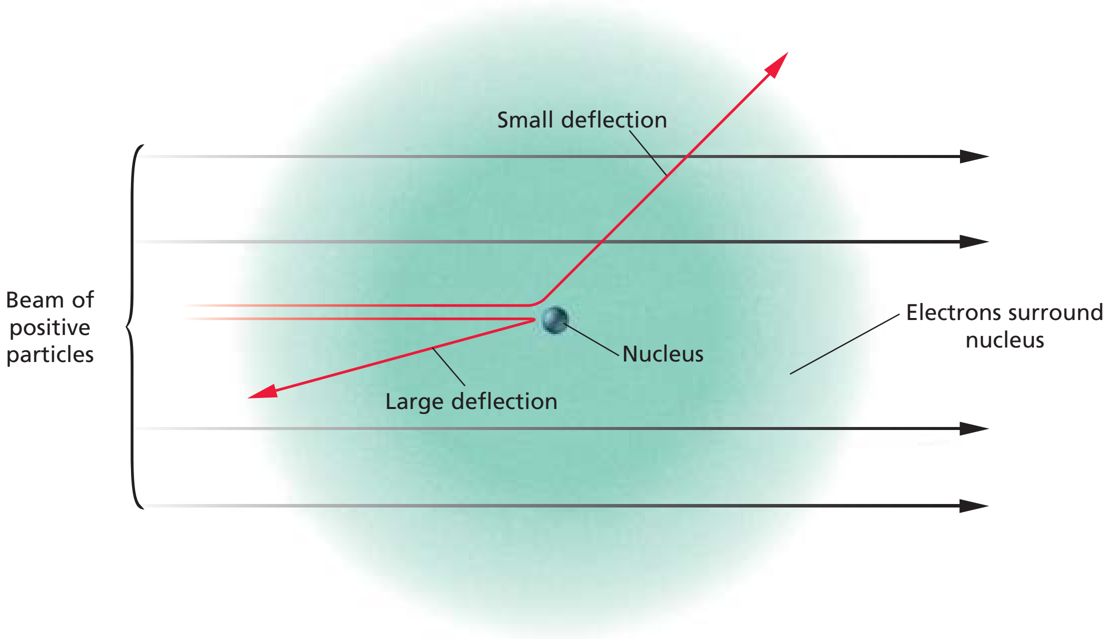


FIGURE 7 Rutherford reasoned that each atom in the gold foil contained a small, dense, positively charged nucleus surrounded by electrons. A small number of the alpha particles directed toward the foil were deflected by the tiny nucleus (red arrows). Most of the particles passed through undisturbed (black arrows).

Composition of the Atomic Nucleus

Except for the nucleus of the simplest type of hydrogen atom (discussed in the next section), all atomic nuclei are made of two kinds of particles, protons and neutrons. A proton has a positive charge equal in magnitude to the negative charge of an electron. Atoms are electrically neutral because they contain equal numbers of protons and electrons. A neutron is electrically neutral.

The simplest hydrogen atom consists of a single-proton nucleus with a single electron moving about it. A proton has a mass of 1.673×10^{-27} kg, which is 1836 times greater than the mass of an electron and 1836/1837, or virtually all, of the mass of the simplest hydrogen atom. All atoms besides the simplest hydrogen atom also have neutrons. The mass of a neutron is 1.675×10^{-27} kg—slightly larger than that of a proton.

The nuclei of atoms of different elements differ in their number of protons and therefore in the amount of positive charge they possess. Thus, the number of protons determines that atom's identity. Physicists have identified other subatomic particles, but particles other than electrons, protons, and neutrons have little effect on the chemical properties of matter. **Table 1** on the next page summarizes the properties of electrons, protons, and neutrons.

Forces in the Nucleus

Generally, particles that have the same electric charge repel one another. Therefore, we would expect a nucleus with more than one proton to be unstable. However, when two protons are extremely close to each other, there is a strong attraction between them. In fact, more than 100

TABLE 1 Properties of Subatomic Particles

Particle	Symbols	Relative electric charge	Mass number	Relative mass (amu*)	Actual mass (kg)
Electron	e^-, ${}^{0}_{-1}e$	–1	0	0.000 5486	9.109×10^{-31}
Proton	p^+, ${}^{1}_{1}H$	+1	1	1.007 276	1.673×10^{-27}
Neutron	n°, ${}^{1}_{0}n$	0	1	1.008 665	1.675×10^{-27}

*1 amu (atomic mass unit) = $1.660\ 540 \times 10^{-27}$ kg

protons can exist close together to help form a nucleus. A similar attraction exists when neutrons are very close to each other or when protons and neutrons are very close together. *These short-range proton-neutron, proton-proton, and neutron-neutron forces hold the nuclear particles together and are referred to as* **nuclear forces.**

The Sizes of Atoms

It is convenient to think of the region occupied by the electrons as an electron cloud—a cloud of negative charge. The radius of an atom is the distance from the center of the nucleus to the outer portion of this electron cloud. Because atomic radii are so small, they are expressed using a unit that is more convenient for the sizes of atoms. This unit is the picometer. The abbreviation for the picometer is pm (1 pm = 10^{-12} m = 10^{-10} cm). To get an idea of how small a picometer is, consider that 1 cm is the same fractional part of 10^3 km (about 600 mi) as 100 pm is of 1 cm. Atomic radii range from about 40 to 270 pm. By contrast, the nuclei of atoms have much smaller radii, about 0.001 pm. Nuclei also have incredibly high densities, about 2×10^8 metric tons/cm^3.

SECTION REVIEW

1. Define each of the following:

a. atom **c.** nucleus **e.** neutron

b. electron **d.** proton

2. Describe one conclusion made by each of the following scientists that led to the development of the current atomic theory:

a. Thomson **b.** Millikan **c.** Rutherford

3. Compare the three subatomic particles in terms of location in the atom, mass, and relative charge.

4. Why is the cathode-ray tube in **Figure 4** connected to a vacuum pump?

Critical Thinking

5. EVALUATING IDEAS Nuclear forces are said to hold protons and neutrons together. What is it about the composition of the nucleus that requires the concept of nuclear forces?

STAGES IN THE SCIENTIFIC METHOD

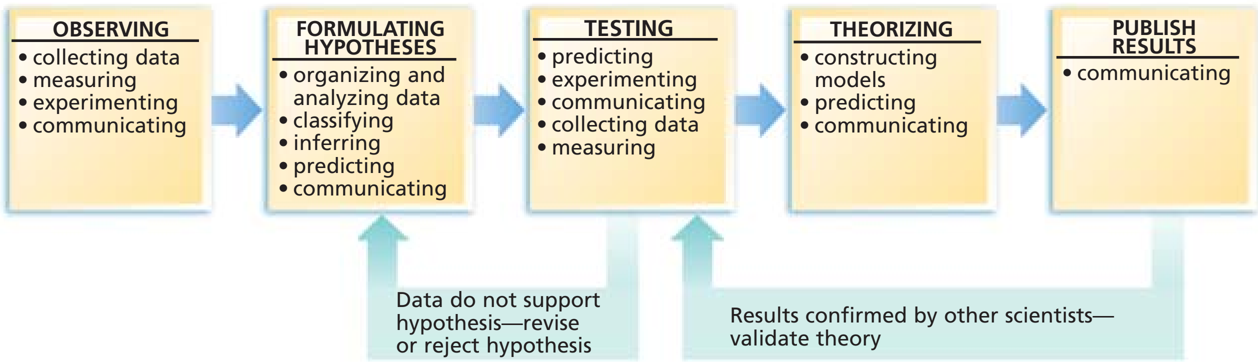


FIGURE 3 The scientific method is not a single, fixed process. Scientists may repeat steps many times before there is sufficient evidence to formulate a theory. You can see that each stage represents a number of different activities.

Theorizing

When the data from experiments show that the predictions of the hypothesis are successful, scientists typically try to explain the phenomena they are studying by constructing a model. *A* **model** *in science is more than a physical object; it is often an explanation of how phenomena occur and how data or events are related.* Models may be visual, verbal, or mathematical. One important model in chemistry is the atomic model of matter, which states that matter is composed of tiny particles called atoms.

If a model successfully explains many phenomena, it may become part of a theory. The atomic model is a part of the atomic theory, which you will study in Chapter 3. *A* **theory** *is a broad generalization that explains a body of facts or phenomena.* Theories are considered successful if they can predict the results of many new experiments. Examples of the important theories you will study in chemistry are kinetic-molecular theory and collision theory. **Figure 3** shows where theory fits in the scheme of the scientific method.

SECTION REVIEW

1. What is the scientific method?

2. Which of the following are quantitative?
a. the liquid floats on water
b. the metal is malleable
c. the liquid has a temperature of 55.6°C

3. How do hypotheses and theories differ?

4. How are models related to theories and hypotheses?

Critical Thinking

5. **INTERPRETING CONCEPTS** Suppose you had to test how well two types of soap work. Describe your experiment by using the terms *control* and *variable.*