Chemistry

Zumdahl, 7th edition

歐亞書局

CH2 Atoms, Molecules, and Ions

A worker in Thailand piles up salt crystals..

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Introduction

- A major goal of this chapter is to present the systems for naming chemical compounds to provide you with the vocabulary necessary to understand this book and to pursue your laboratory studies.

2.1 The Early History of Chemistry

- Chemistry has been important since ancient times.

• The Greeks were the first to try to explain why chemical changes occur. By about 400 B.C. they had proposed that all matter was composed of four fundamental substances: fire, earth, water, and air.

• The next 2000 years of chemical history were dominated by a pseudoscience called *alchemy (煉金術）*.

• The foundations of modern chemistry were laid in the sixteenth century with the development of systematic metallurgy (extraction of metals from ores) by a German, Georg Bauer(1494-1555), and the medicinal application of minerals by a Swiss alchemist/physician known as Paracelsus.

• The first “chemist” to perform truly quantitative experiments was Robert Boyle (1627-1691), who carefully measured the relationship between the pressure and volume of air.

- The phenomenon of combustion evoked intense interest in the seventeenth and eighteenth centuries.
- Oxygen gas, discovered by Joseph Priestley (1773-1804) an English clergyman and scientist (Fig. 2.1).

Figure 2.1

Figure 2.1

The Priestley Medal is the highest honor given by the American Chemical Society. It is named for Joseph Priestley, who was born in England on March 13, 1773. He performed many important scientific experiments, among them the discovery that a gas later identified as carbon dioxide could be dissolved in water to produce seltzer. Also, as a result of meeting Benjamin Franklin in London in 1776, Priestley became interested in electricity and was the first to observe that graphite was an electrical conductor. However, his greatest discovery occurred in 1774 when he isolated oxygen by heating mercuric oxide.

Because of his nonconformist political views, Priestley was forced to leave England. He died in the United States in 1804.

2.2 Fundamental Chemical Laws

- By the late eighteenth century, combustion had been studied extensively; the gases carbon dioxide, nitrogen, hydrogen, and oxygen had been discovered; and the list of elements continued to grow.
- It was Antoine Lavoisier (1743-1794), a French chemist (Fig. 2.2), who finally explained the true nature of combustion, thus clearing the way for the tremendous progress that was made near the end of the eighteenth century.

Figure 2.2

Figure 2.2

Antoine Lavoisier was born in Paris on August 26, 1743. Although Lavoisier's father wanted his son to follow him into the legal profession, young Lavoisier was fascinated by science. From the beginning of his scientific career, Lavoisier recognized the importance of accurate measurements. His careful weighings showed that mass is conserved in chemical reactions and that combustion involves reaction with oxygen. Also, he wrote the first modern chemistry textbook. It is not surprising that Lavoisier is often called the father of modern chemistry.

To help support his scientific work, Lavoisier invested in a private tax-collecting firm and married the daughter of one of the company executives. His connection to the tax collectors proved Fatal, for radical French revolutionaries demanded his execution, which occurred on the guillotine on May 8, 1794.

- His experiments, suggested that *mass is neither created nor destroyed*.
- Lavoisier's verification of this **law of conservation of mass** was the basis for the developments in chemistry in the nineteenth century.
- One of these chemists, a Frenchman, Joseph Proust (1754-1826), showed that *a given compound always contains exactly the same proportion of elements by mass*.

- The principle of the constant composition of compounds, original called “Proust’s law,” is now known as the **law of definite proportion**.
- Proust’s discovery stimulated John Dalton (1776-1844), an English schoolteacher (Fig. 2.3), to think about atoms as the particles that might compose elements.
- Dalton reasoned that if elements were composed of tiny individual particles, a given compound should always contain the same combination of these atoms.

	Mass of Oxygen That Combines with 1 g of Carbon
Compound I	1.33 g
Compound II	2.66 g

- Dalton noted that compound II contains twice as much oxygen per gram of carbon as compound I, a fact that could easily be explained in terms of atoms. Compound I might be CO, and compound II might be CO_2.

- This principle, which was found to apply to compounds of other elements as well, became known as the **law of multiple proportions**: When two elements form a series of compounds, the ratios of the masses of the second element that combine with I gram of the first element can always be reduced to small whole numbers.

Sample Exercise 2.1 Illustrating the Law of Multiple Proportions

The following data were collected for several compounds of nitrogen and oxygen:

	Mass of Nitrogen That Combines with 1 g of Oxygen
Compound A	1.750 g
Compound B	0.8750 g
Compound C	0.4375 g

Show how these data illustrate the law of multiple proportions.

Sample Exercise 2.1

Solution

For the law of multiple proportions to hold, the ratios of the masses of nitrogen combining with 1 gram of oxygen in each pair of compounds should be small whole numbers. We therefore compute the ratios as follows:

$$\frac{A}{B} = \frac{1.750}{0.875} = \frac{2}{1}$$

$$\frac{B}{C} = \frac{0.875}{0.4375} = \frac{2}{1}$$

$$\frac{A}{C} = \frac{1.750}{0.4375} = \frac{4}{1}$$

Sample Exercise 2.1

These results support the law of multiple proportions.

See Exercises 2.27 and 2.28

Figure 2.3

Figure 2.3

John Dalton (1766-1844), an Englishman, began teaching at a Quaker school when he was12. His fascination with science included an intense interest in meteorology, which led to an interest in the gases of the air and their ultimate components, atoms. Dalton is best known for his atomic theory, in which he postulated that the fundamental differences among atoms are their masses. He was the first to prepare a table of relative atomic weights.

Dalton was a humble man with several apparent handicaps: He was not articulate(咬字清楚） and he was color-blind, a terrible problem for a chemist. Despite these disadvantages, he helped to revolutionize the science of chemistry.

2.3 Dalton's Atomic Theory

- In 1808 Dalton published *A New System of Chemical Philosophy*, in which he presented his theory of atoms:

1. Each element is made up of tiny particles called atoms.
2. The atoms of a given element are identical; the atoms of different elements are different in some fundamental way or ways.
3. Chemical compounds are formed when atoms of different elements combine with each other. A given

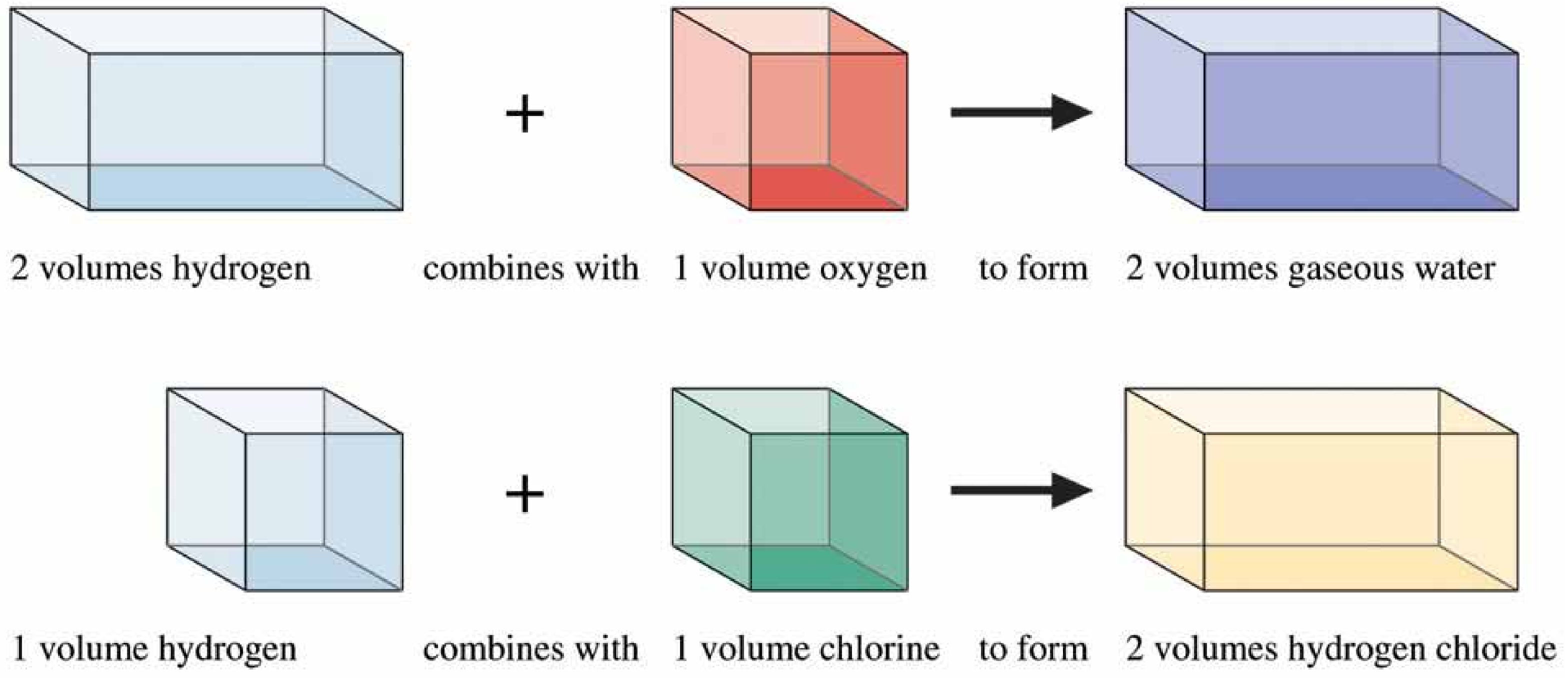
compound always has the same relative numbers and types of atoms.

4. Chemical reactions involve reorganization of the atoms—changes in the way they are bound together. The atoms themselves are not changed in a chemical reaction.

- Using similar reasoning for other compounds, Dalton prepared the first table of **atomic masses** (sometimes called **atomic weights** by chemists, since mass is often determined by comparison to a standard mass—a process called weighing).
- The keys to determining absolute formulas for compounds were provided in the experimental work of the French chemist Joseph Gay-Lussac (1778-1850) and by the hypothesis of an Italian chemist named Amadeo Avogadro (1776-1856).

- Gay-Lussac found that 2 volumes of hydrogen react with 1 volume of oxygen to form 2 volumes of gaseous water and that 1 volume of hydrogen reacts with 1 volume of chlorine to form 2 volumes of hydrogen chloride.
- These results are represented schematically in Fig. 2.4.

Figure 2.4



A representation of some of Gay-Lussac's experimental results on combining gas volumes.

- In 1811 Avogadro interpreted these results by proposing that *at the same temperature and pressure, equal volumes of different gases contain the same number of particles.*
- This assumption (called **Avogadro's hypothesis**) makes sense if the distances between the particles in a gas are very great compared with the sizes of the particles.

- If Avogadro’s hypothesis is correct, Gay-Lussac’s result,

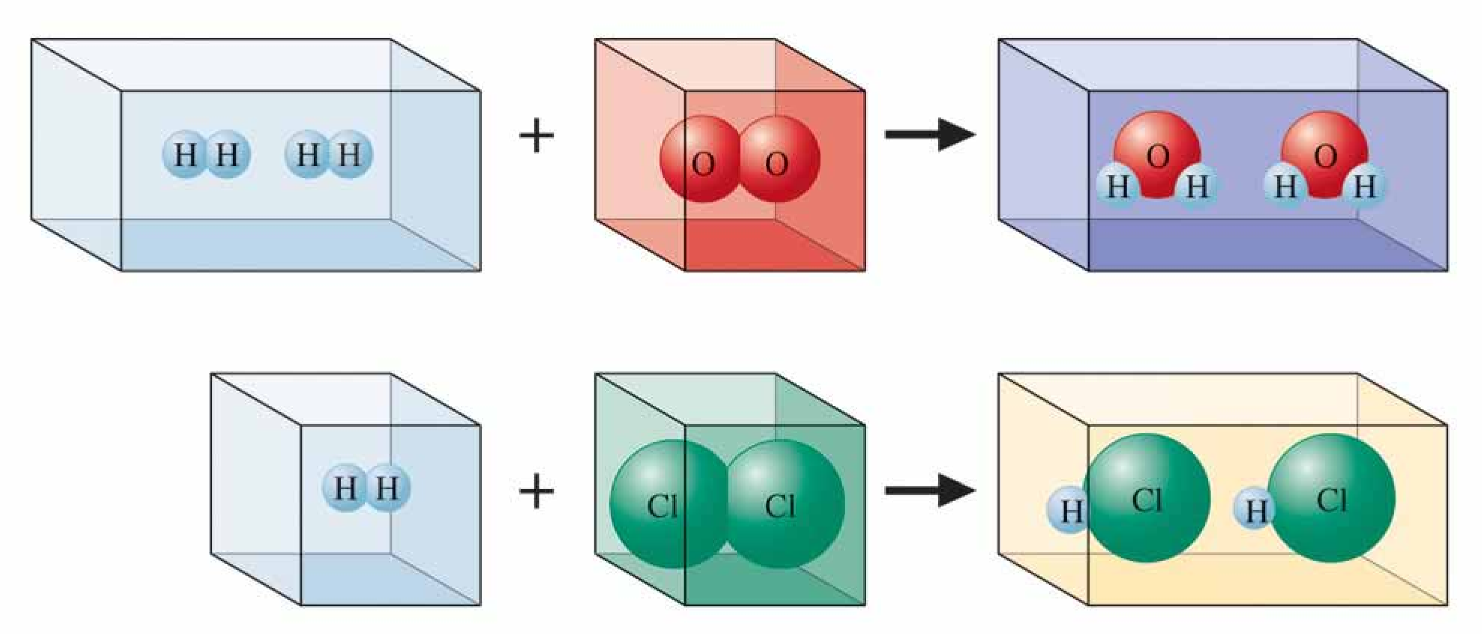
  2 volumes of hydrogen react with 1 volume of oxygen $\longrightarrow$ 2 volumes of water vapor

- can be expressed as follows:

  2 molecules of hydrogen react with 1 molecule of oxygen $\longrightarrow$ 2 molecules of water

- These observations can best be explained by assuming that gaseous hydrogen, oxygen, and chlorine are all composed of diatomic (two-atom) molecules: H_2, O_2, and Cl_2, respectively.
- Gay-Lussac's results can then be represented as shown in Fig. 2.5.

Figure 2.5



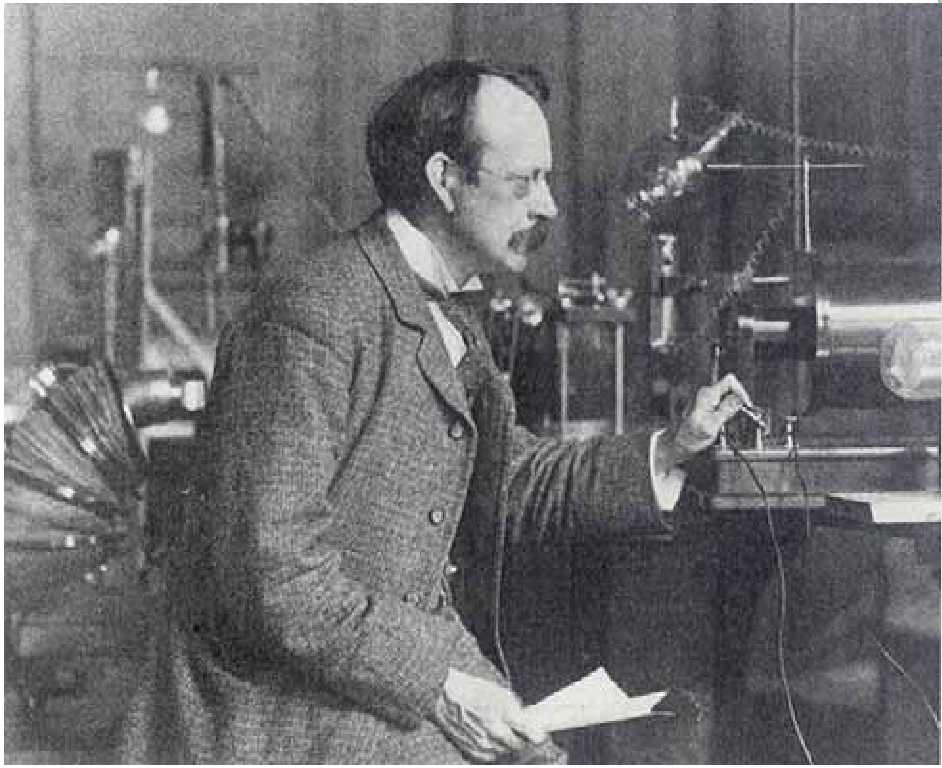
A representation of combining gases at the molecular level. The spheres represent atoms in the molecules.

2.4 Early Experiments to Characterize the Atom

The Electron

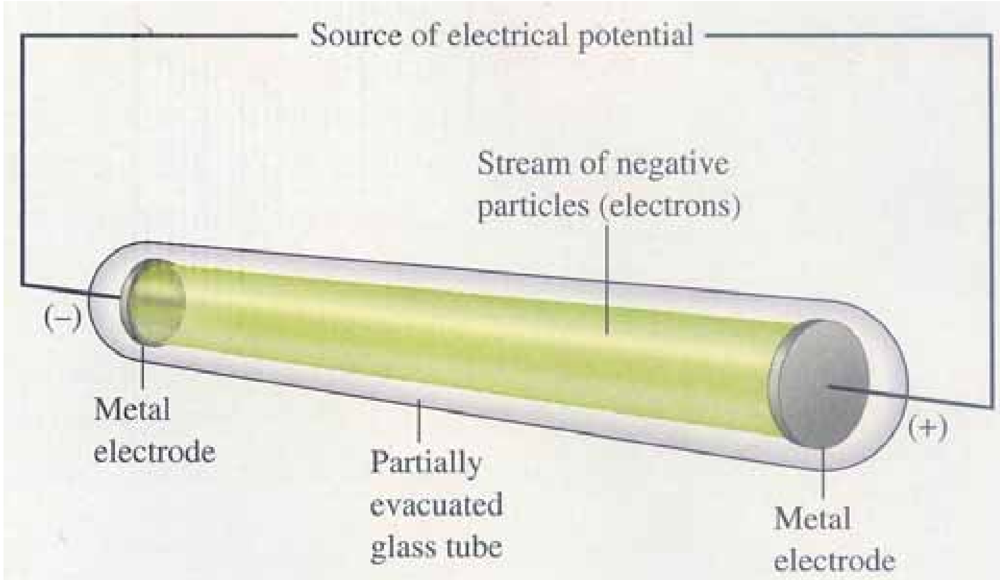
- The first important experiments that led to an understanding of the composition of the atom were done by the English physicist J.J. Thomson (Fig. 2.6), who studied electrical discharges in partially evacuated tubes called **cathode-ray tubes** (Fig. 2.7) during the period from 1898 to 1903.

Figure 2.6



J.J. Thomson (1856-1940) was an English physicist at Cambridge University. He received the Nobel Prize in physics in 1906.

Figure 2.7



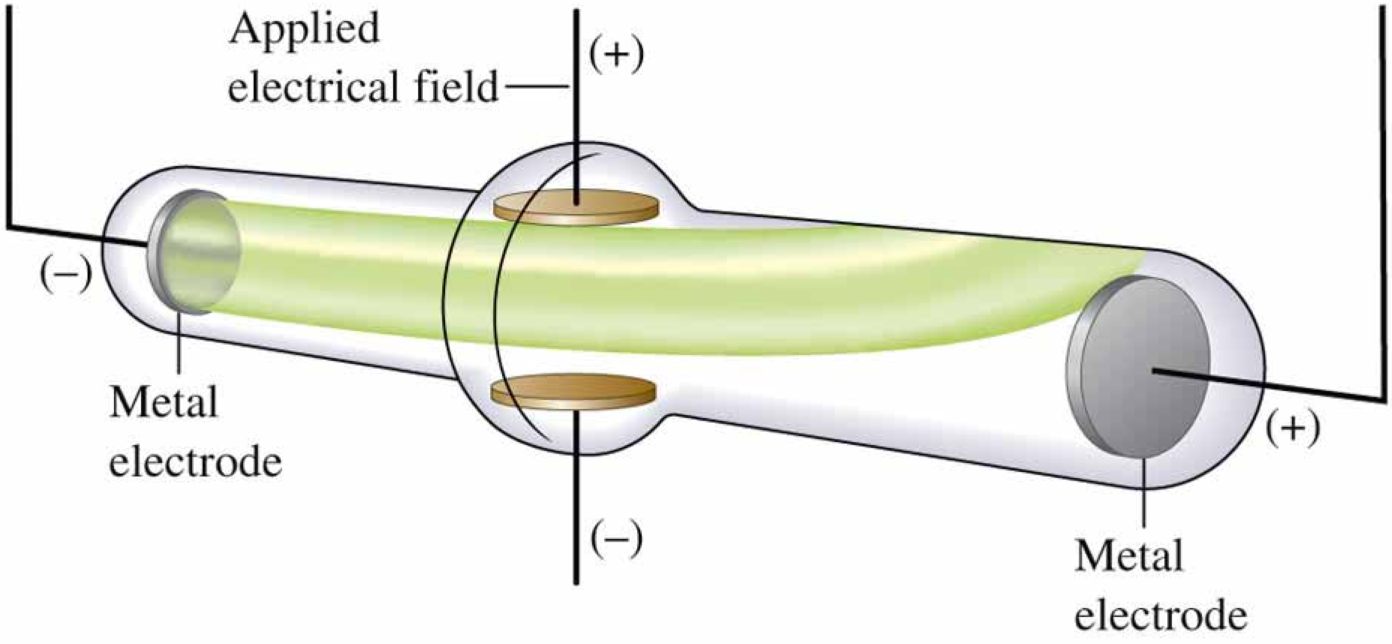
A cathode-ray tube. The fast-moving electrons excite the gas in the tube, causing a glow between the electrodes. The green color in the photo is due to the response of the screen (coated with zinc sulfide) to the electron beam.

- Because this ray was produced at the negative electrode and was repelled by the negative pole of an applied electric field (see Fig. 2.8), Thomson postulated that the ray was a stream of negatively charged particles, now called **electrons**.
- Thomson determined the *charge-to-mass ratio* of an electron:

$$\frac{e}{m} = -1.76 \times 10^{8}\, \mathrm{C/g}$$

- where e represents the charge on the electron in coulombs (C) and m represents the electron mass in grams.

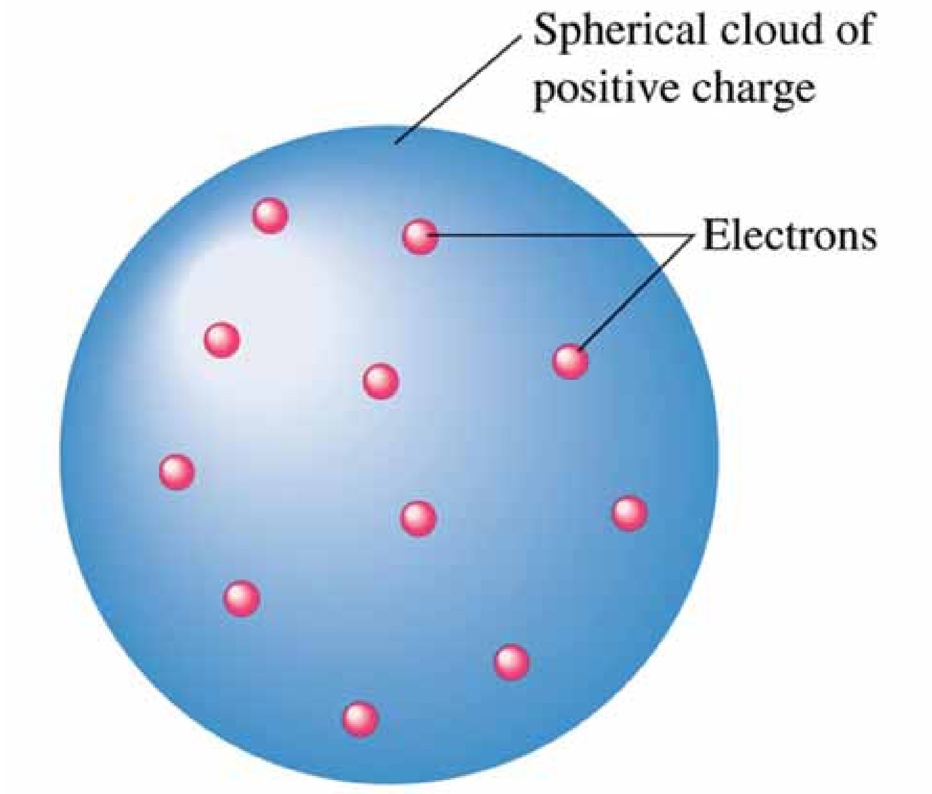
Figure 2.8



Deflection of cathode rays be an applied electric field.

- This model, shown in Fig. 2.9, is often called the plum pudding model because the electrons are like raisins dispersed in a pudding.
- In 1909 Robert Millikan (1868-1953), working at the University of Chicago, performed very clever experiments involving charged oil drops.
- These experiments allowed him to determine the magnitude of the electron charge (see. 2.10).

Figure 2.9



The plum pudding model of the atom.

Figure 2.10

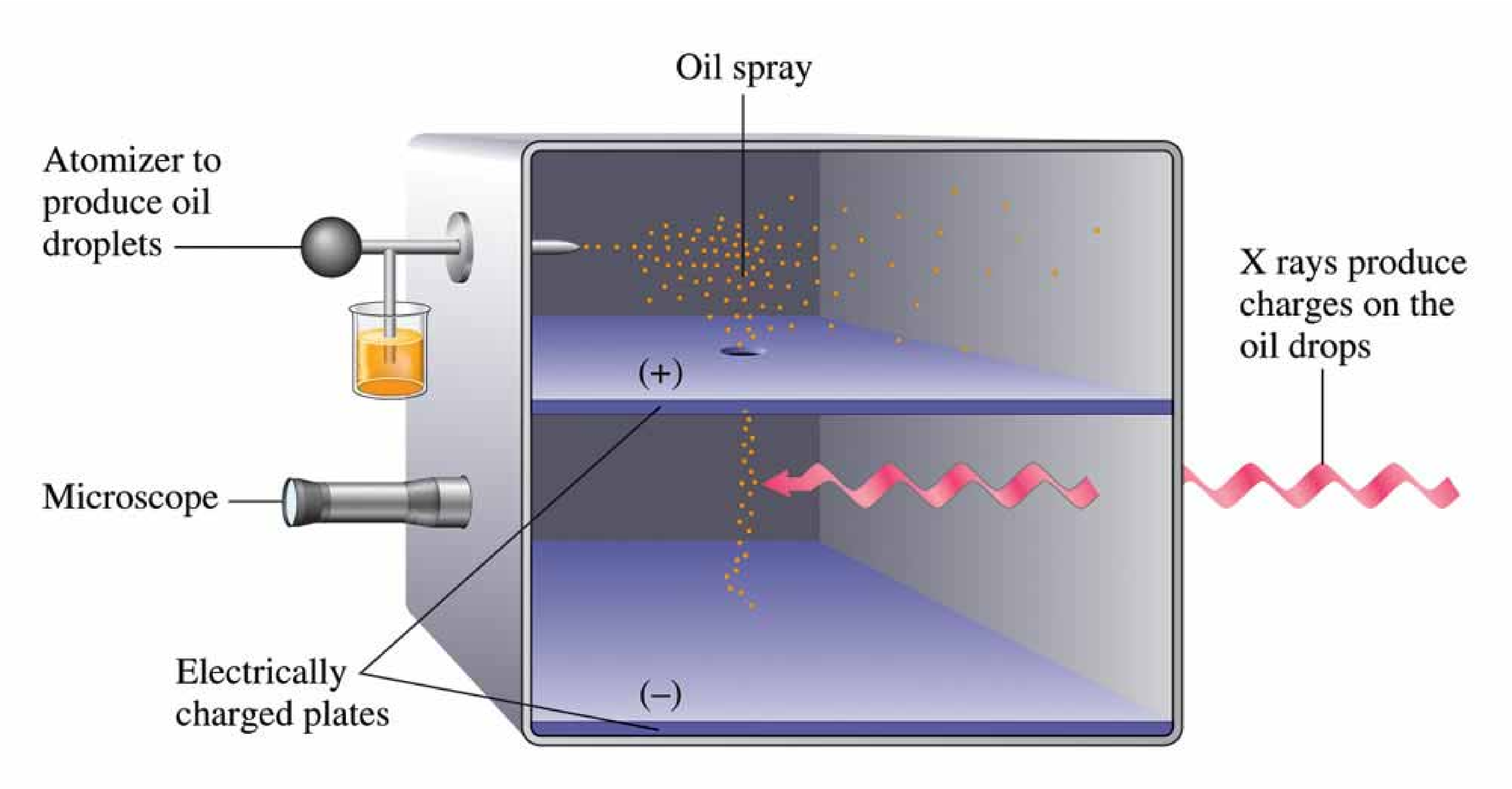
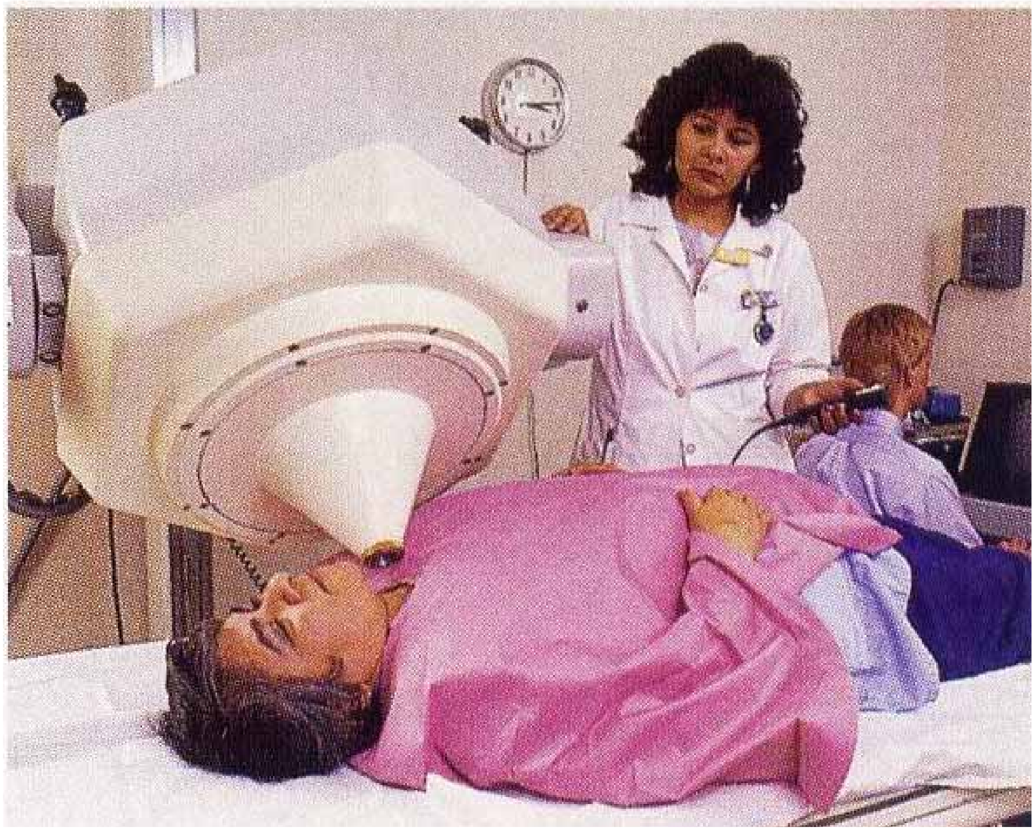

Oil spray
Atomizer to
produce oil
droplets
X rays produce
charges on the
oil drops
(+)
Microscope
(−)
Electrically
charged plates

Figure 2.10

A schematic representation of the apparatus Millikan used to determine the charge on the electron. The fall of charged oil droplets due to gravity can then be used to calculate the charge on the oil drop. Millikan's experiments showed that the charge on an oil drop is always a whole-number multiple of the electron charge.

A technician using a scanner to monitor the uptake of radioactive iodine in a patient's thyroid (甲狀腺).



Rodioactivity

- In the late nineteenth century scientists discovered that certain elements produce high-energy radiation.
- Studies in the early twentieth century demonstrated three types of radioactive emission: gamma (γ) rays, beta (β) particles, and alpha (α) particles.
- A γ ray is high-energy "light"; a β particle is a high-speed electron; and an α particle has a $2+$ charge, that is, a charge twice that of the electron and with the opposite sign.

The Nuclear Atom

- In 1911 Ernest Rutherford (Fig. 2.11), who performed many of the pioneering experiments to explore radioactivity, carried out an experiment to test Thomson's plum pudding model.
- The experiment involved direction a particles at a thin sheet of metal foil, as illustrated in Fig. 2.12.

Figure 2.11

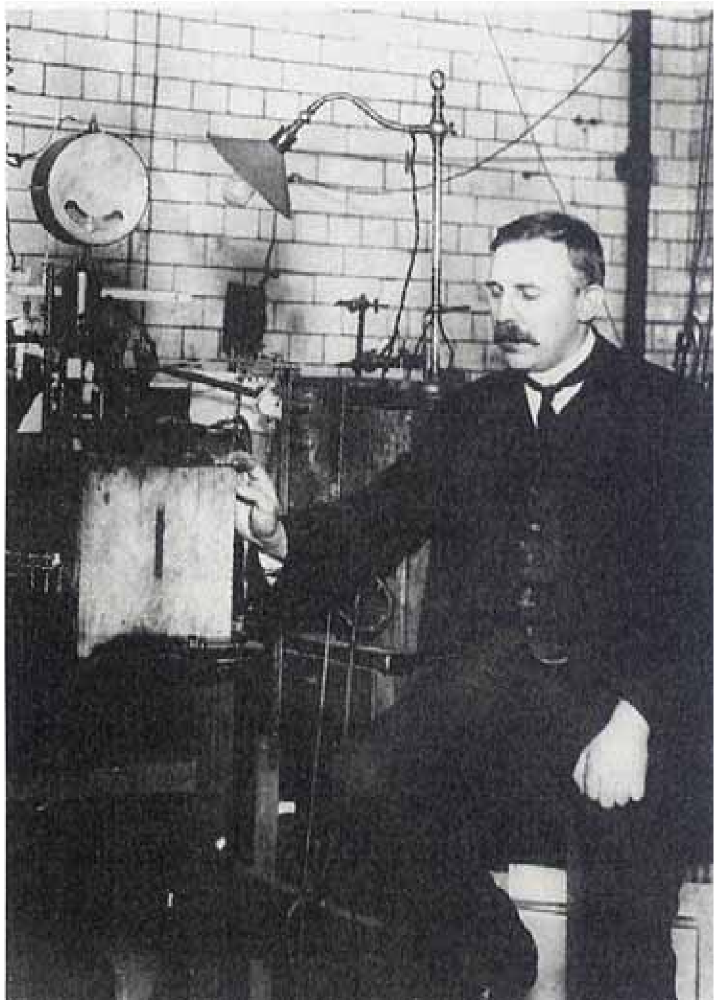
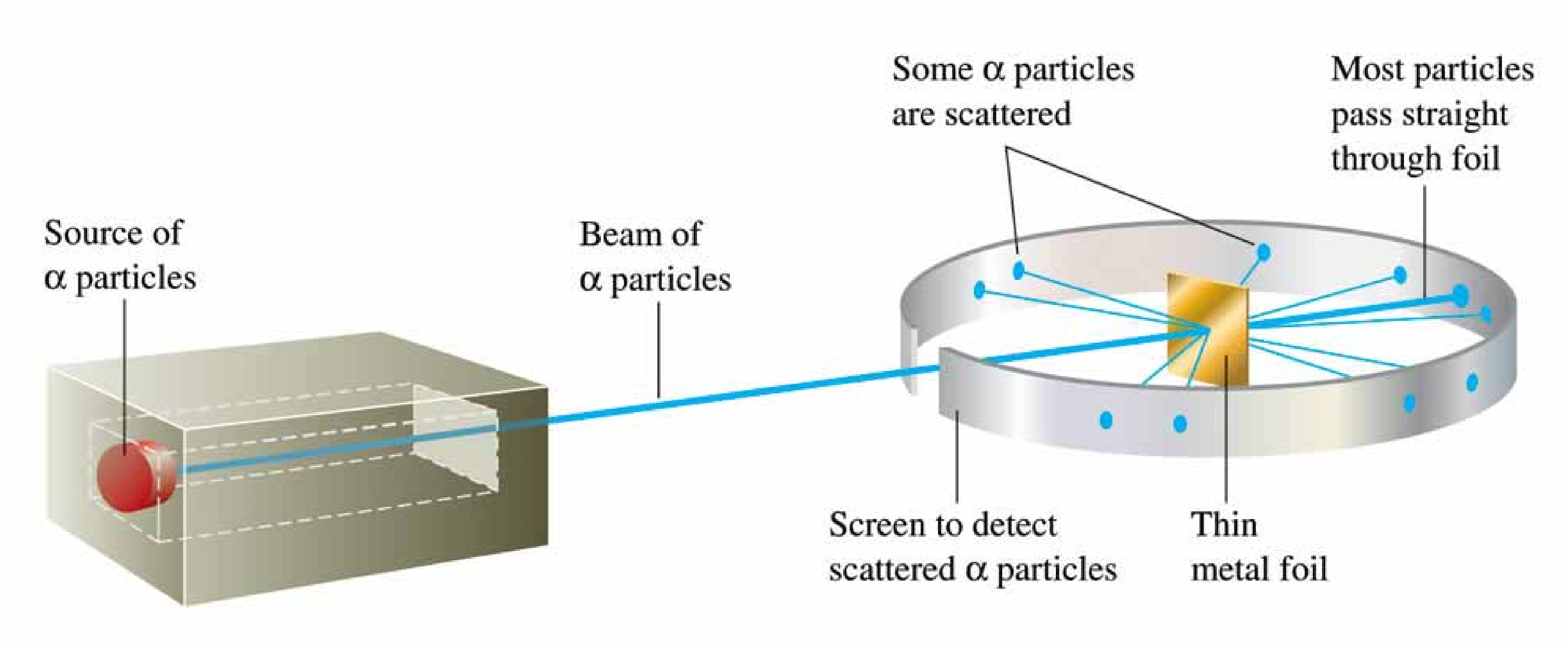


Figure 2.11

Ernest Rutherford (1871-1937) was born on a farm in New Zealand. In 1895 he placed second in a scholarship competition to attend Cambridge University but was awarded the scholarship when the winner decided to stay home and get married. As a scientist in England, Rutherford did much of the early work on characterizing radioactivity. He named the α and β particles and the γ ray and coined the term half-life to describe an important attribute of radioactive elements. His experiments on the behavior of α particles striking thin metal foils led him to postulate the nuclear atom. He also invented the name proton for the nucleus of the hydrogen atom. He received the Nobel Prize in chemistry in 1908.

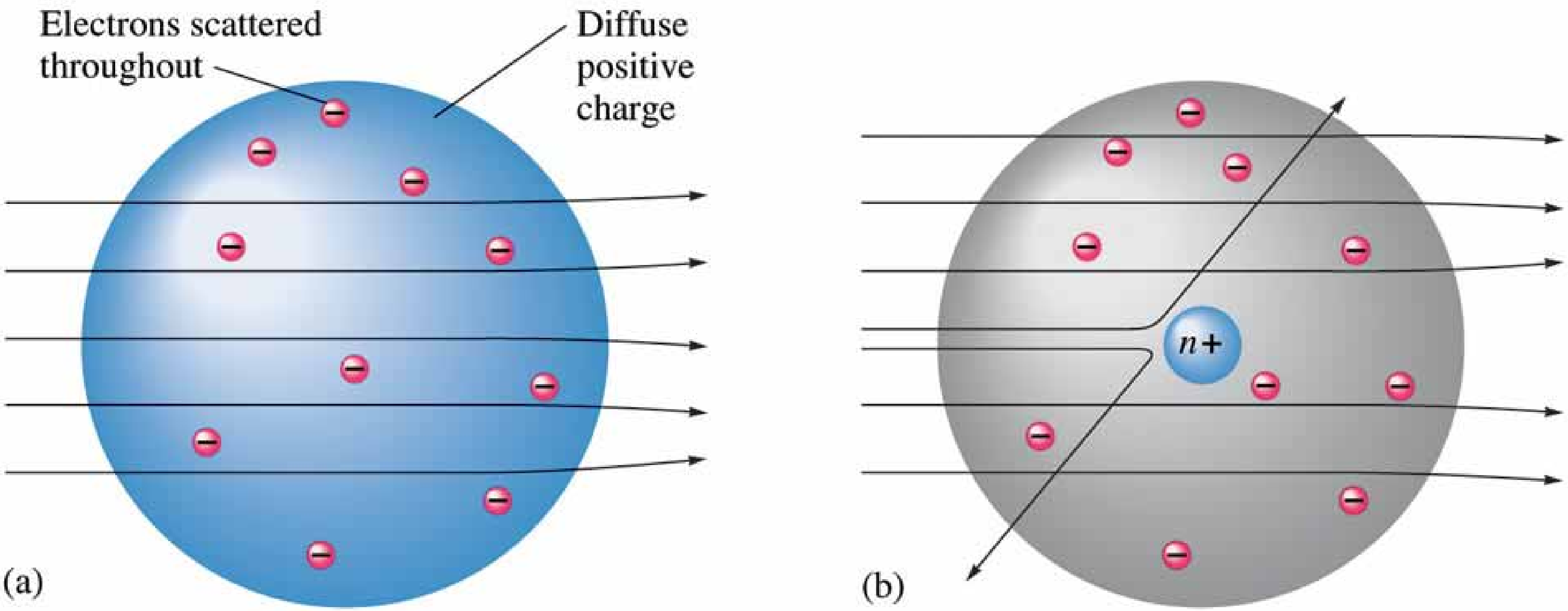
Figure 2.12



Rutherford's experiment on α-particle bombardment of metal foil.

- Rutherford reasoned that if Thomson's model were accurate, the massive a particles should crash through the thin foil like cannonballs through gauze, as shown in Fig. 2.13(a).
- Although most of the α particles passed straight through, many of the particles were deflected at large angles, as shown in Fig. 3.13(b), and some were reflected, never hitting the detector.

Figure 2.13



(a) The expected results of the metal foil experiment if Thomson's model were correct. (b) Actual results.

- In Rutherford's mind these results could be explained only in terms of a nuclear atom—an atom with a dense center of positive charge (the **nucleus**) with electrons moving around the nucleus at a distance that is large relative to the nuclear radius.

2.5 The Modern View of Atomic Structure: An Introduction

- The simplest view of the atom is that it consists of a tiny nucleus (with a diameter of about 10^{-13} cm) and electrons that move about the nucleus at an average distance of about 10^{-8} cm from it (see Fig. 2.14).
- The nucleus is assumed to contain **protons**, which have a positive charge equal in magnitude to the electron's negative charge, and **neutrons**, which have virtually the same mass as a proton but no charge.
- The masses and charges of the electron, proton, and neutron are shown in Table 2.1.

Figure 2.14

A nuclear atom viewed in cross section. Note that this drawing is not to scale.

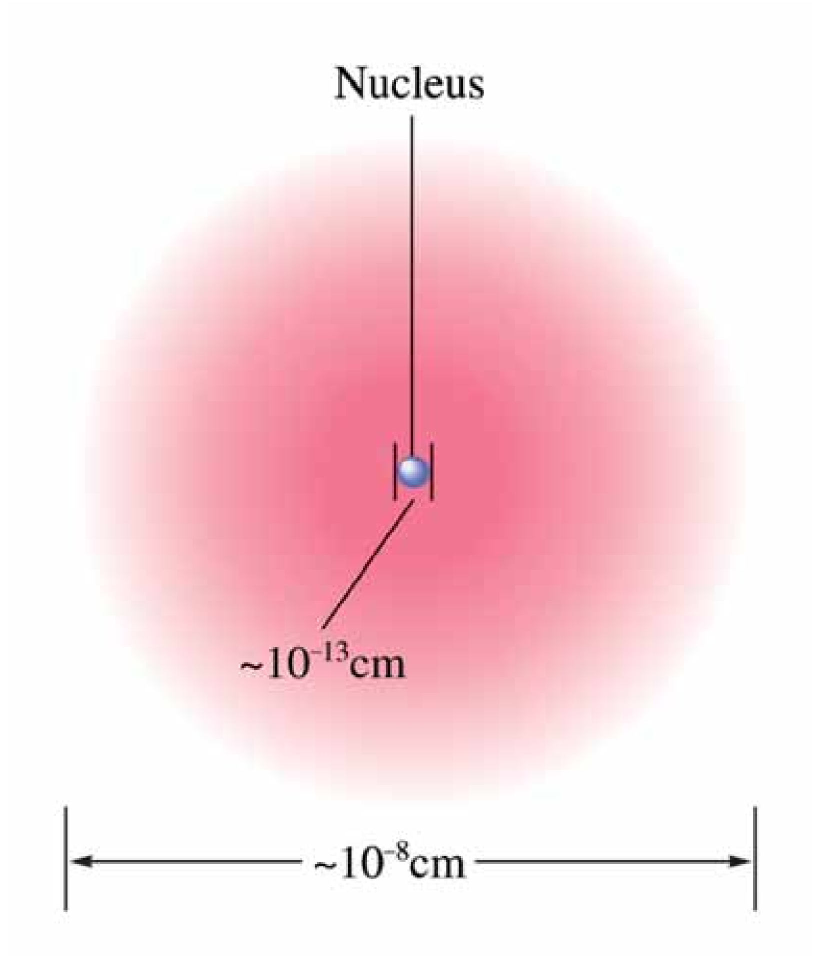


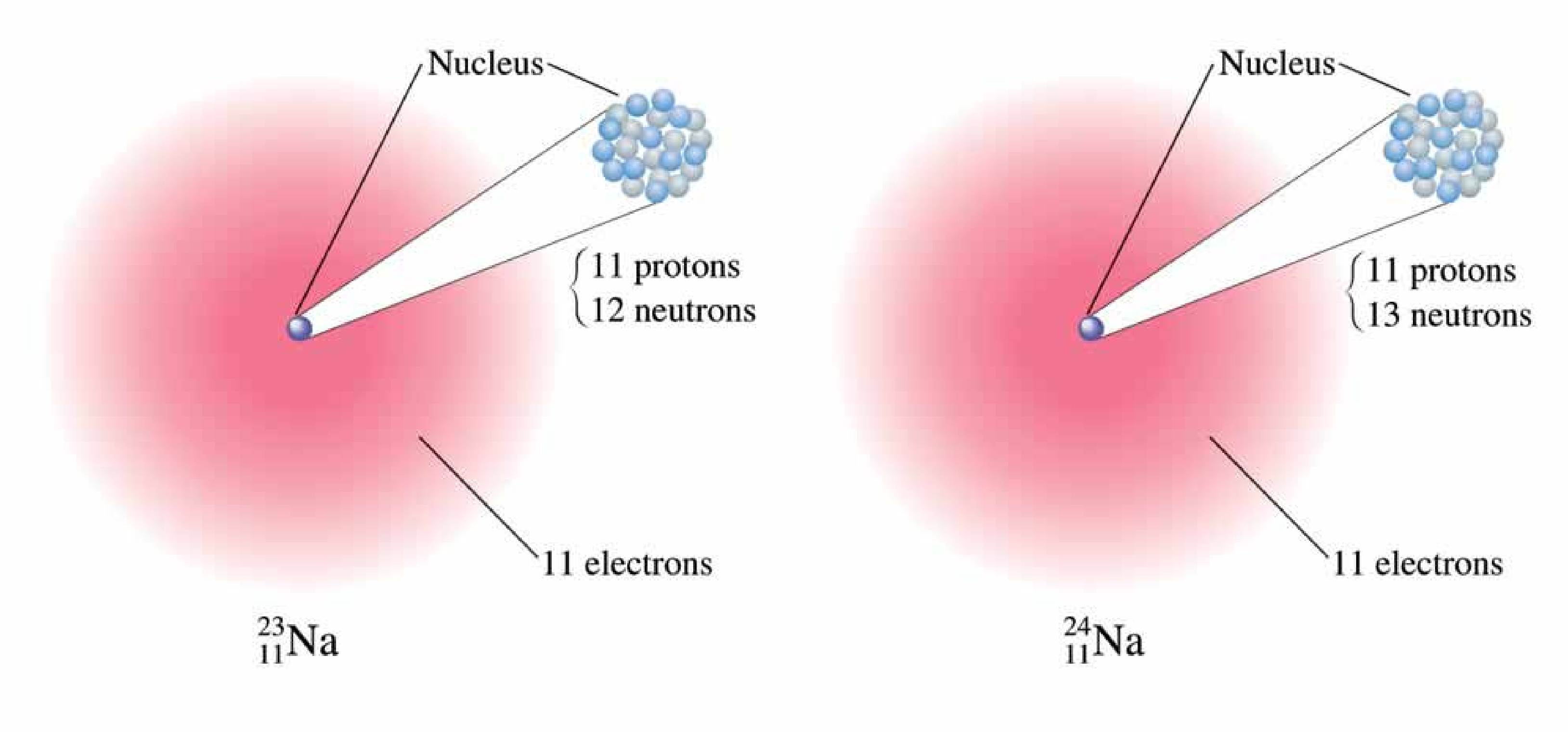
TABLE 2.1 The Mass and Charge of the Electron, Proton, and Neutron

Particle	Mass	Charge*
Electron	9.11×10^{-31} kg	1−
Proton	1.67×10^{-27} kg	1+
Neutron	1.67×10^{-27} kg	None

*The magnitude of the charge of the electron and the proton is 1.60×10^{-19} C.

- An important question to consider at this point is, "*If all atoms are composed of these same components, why do different atoms have different chemical properties*?"
- In Fig.2.15, these two atoms are **isotopes**, or *atoms with the same number of protons but different numbers of neutrons.*

Figure 2.15



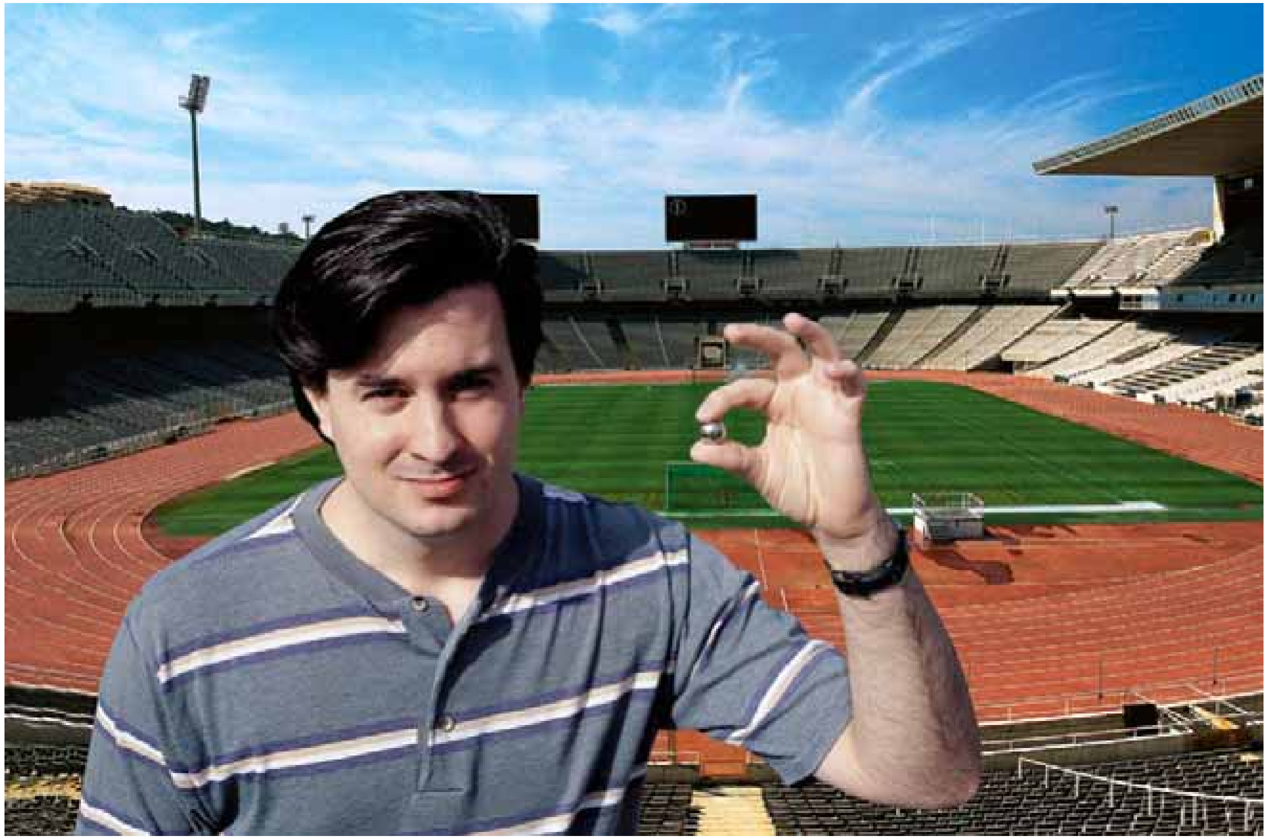
Two isotopes of sodium. Both have 11 protons and 11 electrons, but they differ in the number of neutrons in their nuclei.

- Note that the symbol for one particular type of sodium atom is written

Mass number → $^{23}_{11}Na$ ← Element symbol

Atomic number ↗

- where the atomic number Z (number of protons) is written as a subscript, and the mass number A (the total number of protons and neutrons) is written as a superscript. (The particular atom represented here is called “sodium twenty-three.”
- It has 11 electrons, 11 protons, and 12 neutrons.)



If the atomic nucleus were the size of this ball bearing, a typical atom would be the size of this stadium.

Sample Exercise 2.2

Writing the Symbols for Atoms

Write the symbol for the atom that has an atomic number of 9 and a mass number of 19. How many electrons and how many neutrons does this atom have?

Solution

The atomic number 9 means the atom has 9 protons. This element is called fluorine, symbolized by F. The atom is represented as

$$^{19}_{9}F$$

Sample Exercise 2.2

and is called "fluorine nineteen." Since the atom has 9 protons, it also must have 9 electrons to achieve electrical neutrality. The mass number gives the total number of protons and neutrons, which means that this atom has 10 neutrons.

See Exercises 2.43 though 2.46