Chapter 2

Atoms, Molecules, and Ions

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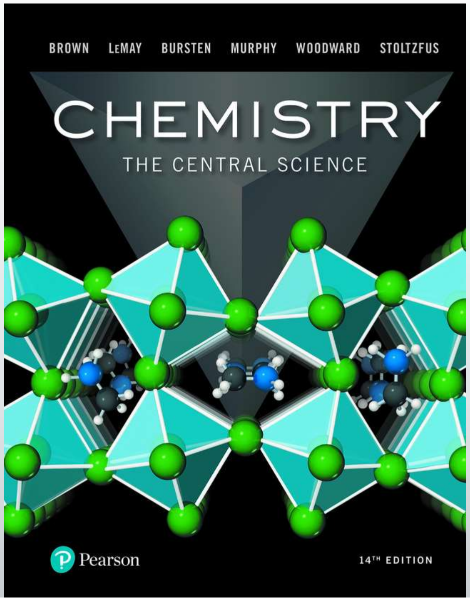


2.1 The Atomic Theory of Matter

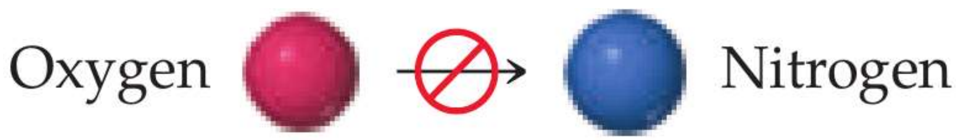
❖ Dalton's atomic theory:

1. Each element is composed of extremely small particles called atoms.

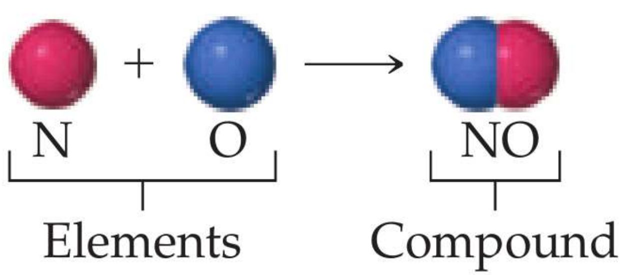
2. All atoms of a given element are identical, but the atoms of one element are different from the atoms of all other elements.

2.1 The Atomic Theory of Matter

3. Atoms of one element cannot be changed into atoms of a different element by chemical reactions; atoms are neither created nor destroyed in chemical reactions.



4. Compounds are formed when atoms of more than one element combine; a given compound always has the same relative number and kind of atoms.

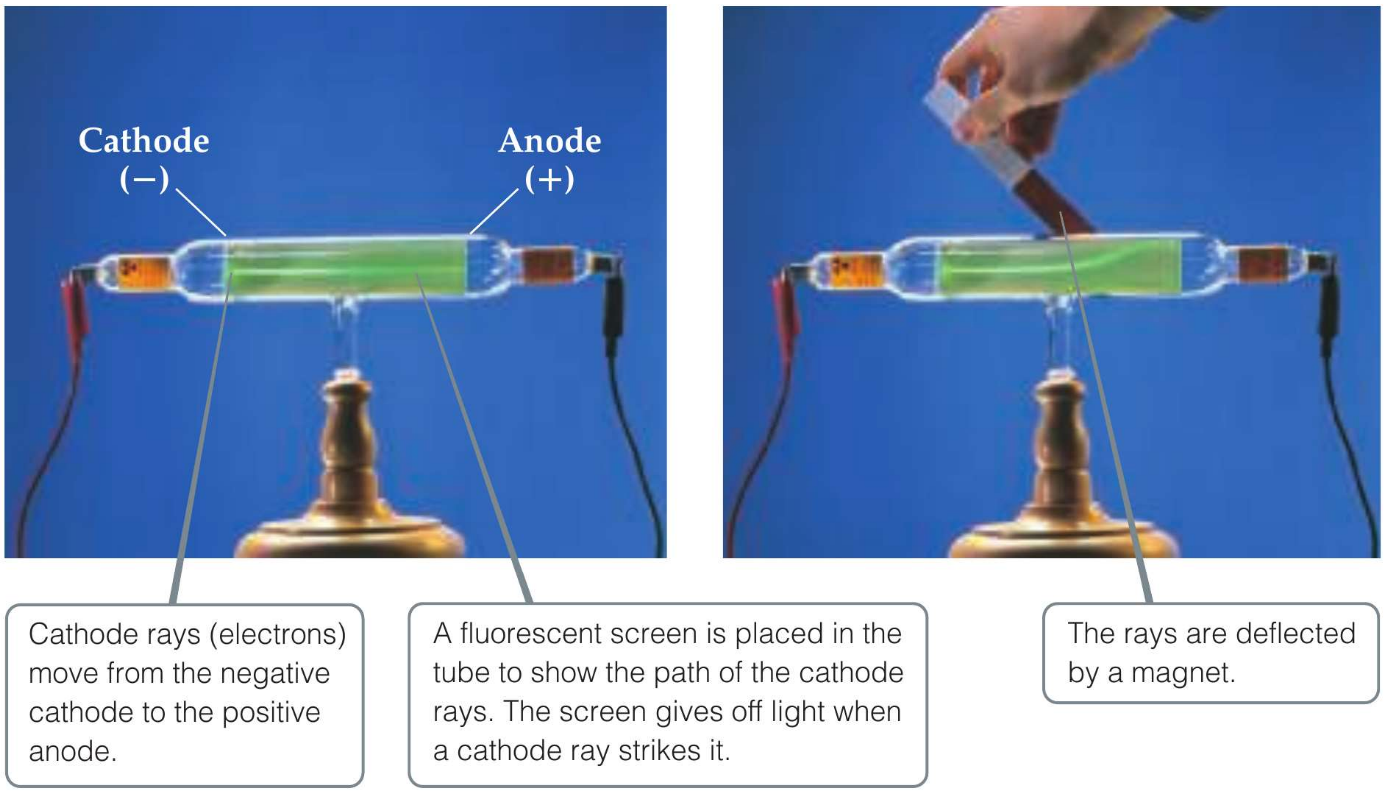


2.1 The Atomic Theory of Matter

- The law of constant composition, based on postulate 4: In a given compound, the relative numbers and kinds of atoms are constant.
- The law of conservation of mass, based on postulate 3: The total mass of materials present after a chemical reaction is the same as the total mass present before the reaction.
- The law of multiple proportions: If two elements A and B combine to form more than one compound, the masses of B that can combine with a given mass of A are in the ratio of small whole numbers.

2.2 The Discovery of Atomic Structure

Cathode Rays and electrons

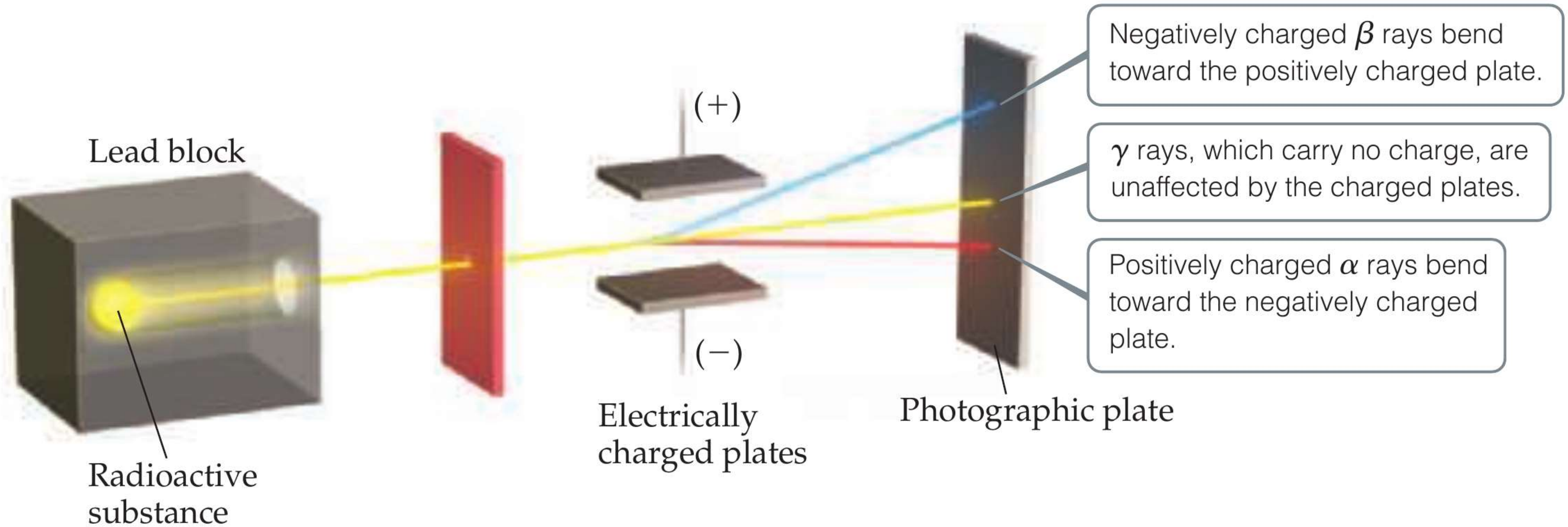


▲ **Figure 2.3** **Cathode-ray tube.**

2.2 The Discovery of Atomic Structure

Radioactivity

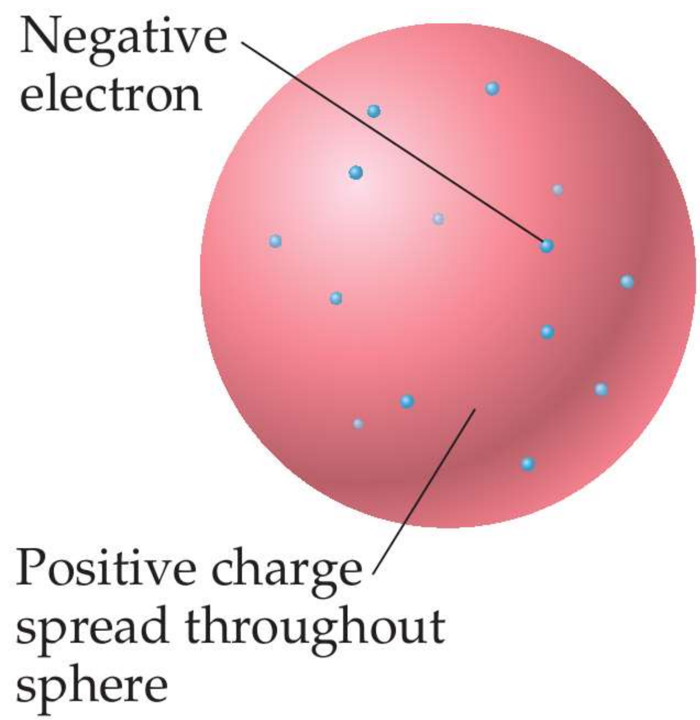
- When a compound, such as uranium, spontaneously emits high-energy radiation, this spontaneous emission of radiation is called radioactivity.
- A study of radioactivity, by Rutherford, revealed three types of radiation: alpha (α), beta (β), and gamma (γ).
- From this experiment, it was concluded that β particles are nothing more than high-speed electrons, which have a charge of 1–, whereas α particles have a charge of 2+ .
- Gamma radiation is high-energy electromagnetic; it does not consist of particles and it carries no charge.



▲ **Figure 2.7** **Behavior of alpha (α), beta (β), and gamma (γ) rays in an electric field.**

The Nuclear Model of the Atom

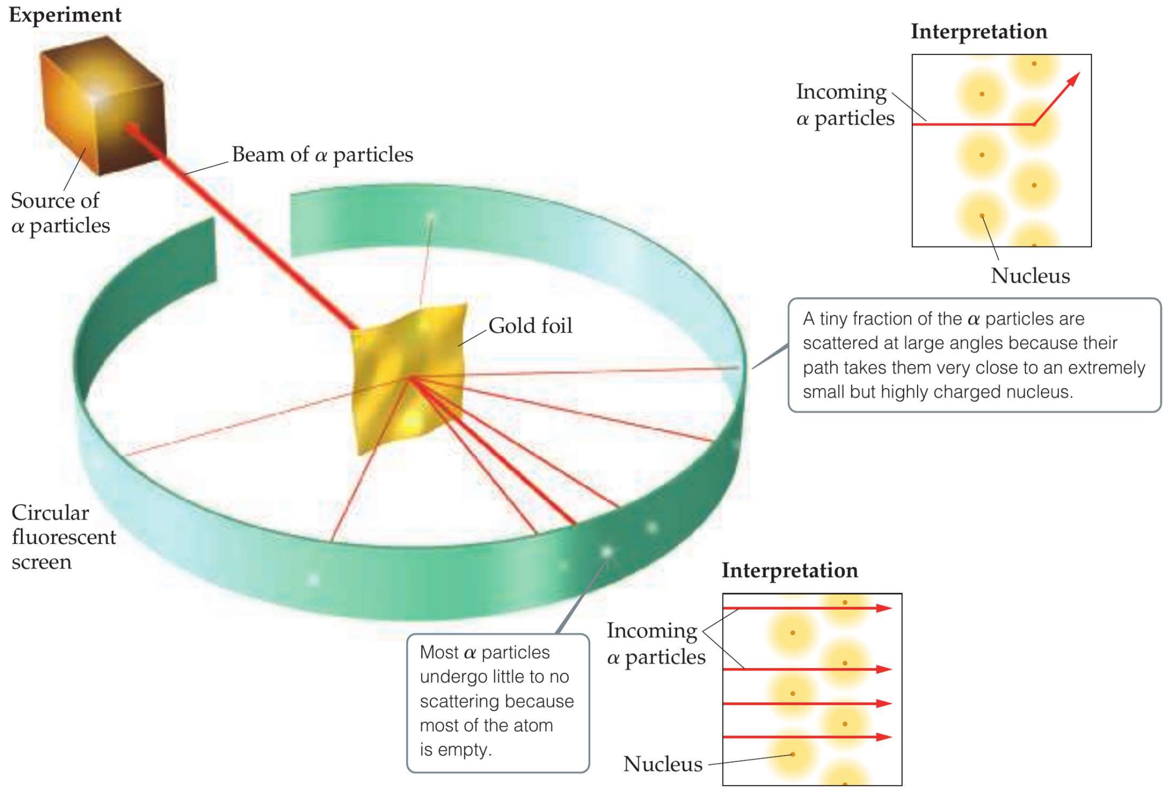
- Thomson proposed that the atom consists of a uniform positive sphere of matter in which the mass is evenly distributed and in which the electrons are embedded like raisins in a pudding or seeds in a watermelon.



▲ Figure 2.8 **J. J. Thomson's plum-pudding model of the atom.** Ernest Rutherford and Ernest Marsden proved this model wrong.

2.2 The Discovery of Atomic Structure

- Rutherford studied the angles at which α particles were deflected, or scattered, as they passed through a thin sheet of gold foil.

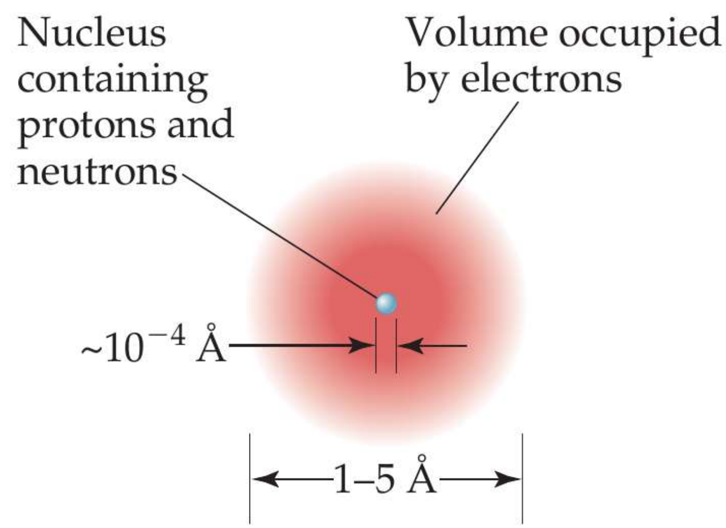


2.3 The Modern View of Atomic Structure

- The charge of an electron is -1.602×10^{-19} C, whereas the charge of a proton is opposite in sign but equal in magnitude to that of an electron $+1.602 \times 10^{-19}$ C.
- The quantity 1.602×10^{-19} C is called the electronic charge.
- For convenience, the charges of atomic and subatomic particles are usually expressed as multiples of this charge rather than as coulombs.
- Thus, the charge of an electron is 1– and that of a proton is 1+.
- Neutrons are electrically neutral.
- Every atom has an equal number of electrons and protons, so atoms have no net electrical charge.

2.3 The Modern View of Atomic Structure

- Protons and neutrons reside in the tiny nucleus of the atom.
- The vast majority of an atom's volume is the space in which the electrons reside.
- A convenient non-SI unit of length used for atomic dimensions is the angstrom (Å), where $1\ \text{Å} = 1 \times 10^{-10}\ \text{m}$.
- Electrons are attracted to the protons in the nucleus by the electrostatic force that exists between particles of opposite electrical charge.

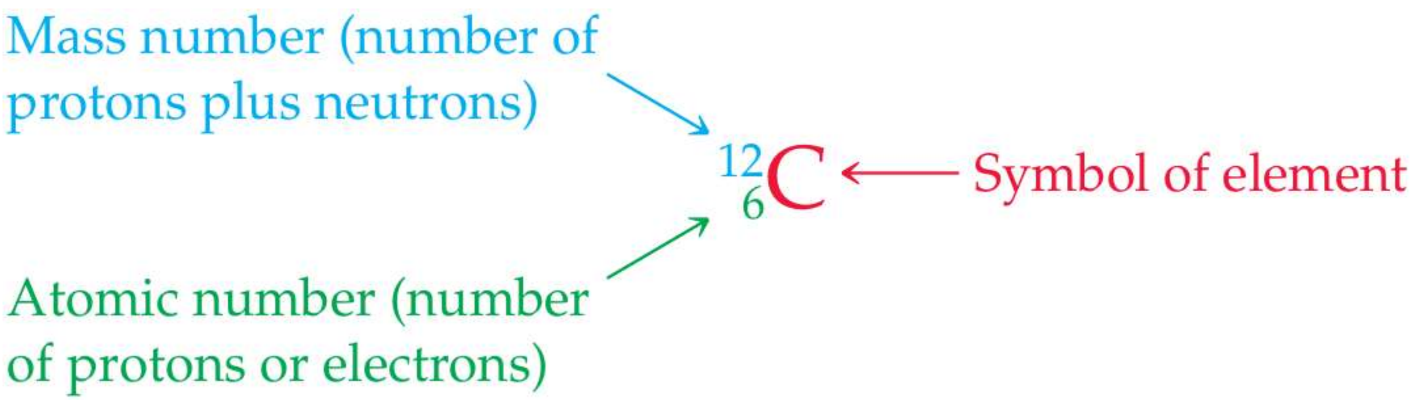


▲ Figure 2.10 The structure of the atom.

2.3 The Modern View of Atomic Structure

Atomic Numbers, Mass Numbers, and Isotopes

- The number of protons in an atom of any particular element is called that element's atomic number.
- The atomic number is indicated by the subscript; the superscript, called the mass number, is the number of protons plus neutrons in the atom:



2.3 The Modern View of Atomic Structure

- The symbol $^{12}_{6}C$ (read "carbon twelve," carbon-12) represents the carbon atom containing six protons and six neutrons.
- Atoms with identical atomic numbers but different mass numbers (that is, the same number of protons but different numbers of neutrons) are called isotopes of one another.

TABLE 2.2 Some Isotopes of Carbon[a]

Symbol	Number of Protons	Number of Electrons	Number of Neutrons
^{11}C	6	6	5
^{12}C	6	6	6
^{13}C	6	6	7
^{14}C	6	6	8

[a] Almost 99% of the carbon found in nature is ^{12}C.

Sample Exercise 2.1

The diameter of a U.S. dime is 17.9 mm, and the diameter of a silver atom is 2.88 Å . How many silver atoms could be arranged side by side across the diameter of a dime?

$$\text{No. of atoms} = 17.9\ \text{mm}\left(\frac{10^{-3}\ \text{m}}{1\ \text{mm}}\right)\left(\frac{1\ \text{Å}}{10^{-10}\ \text{m}}\right)\left(\frac{1\ \text{Ag atom}}{2.88\ \text{Å}}\right)$$

$$= 6.22 \times 10^{7}\ \text{Ag atoms}$$

Sample Exercise 2.2

How many protons, neutrons, and electrons are in an atom of

a. ^{197}Au

$\text{Atomic No.} = 79$ ⇒ $\text{No. of protons} = \text{No. of electrons} = 79$

$\text{No. of neutrons} = 197 - 79 = 118$

b. strontium-90?

$\text{Atomic No.} = 38$ ⇒ $\text{No. of protons} = \text{No. of electrons} = 38$

$\text{No. of neutrons} = 90 - 38 = 52$

Sample Exercise 2.3

Magnesium has three isotopes with mass numbers 24, 25, and 26.

a. Write the complete chemical symbol (superscript and subscript) for each.

Atomic No. = 12 → $^{24}_{12}Mg$ $^{25}_{12}Mg$ $^{26}_{12}Mg$

b. How many neutrons are in an atom of each isotope?

- The numbers of neutrons in an atom of each isotope are therefore 12, 13, and 14, respectively.

2.4 Atomic Weights

The Atomic Mass Scale

- The atomic mass unit is defined by assigning a mass of exactly 12 amu to a chemically unbound atom of the ^{12}C isotope of carbon.

Atomic Weight

- We can determine the **average atomic mass** of an element, usually called the element's atomic weight (*AW*), by summing over the masses of its isotopes multiplied by their relative abundances:

$$AW = \sum_i [(\text{isotope mass})_i (\text{fractional isotope abundance})_i]$$

Sample Exercise 2.4

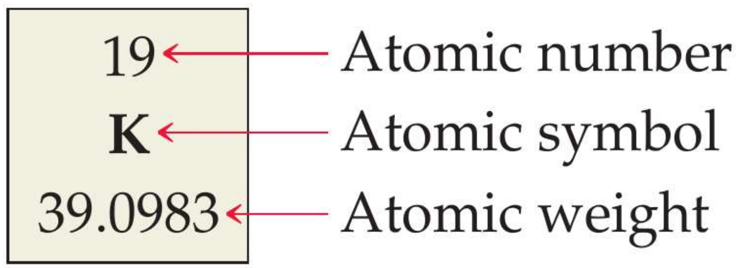
Naturally occurring chlorine is 75.78% ^{35}Cl (atomic mass 34.969 amu) and 24.22% ^{37}Cl (atomic mass 36.966 amu). Calculate the atomic weight of chlorine.

$$AW = \sum_i [(\text{isotope mass})_i (\text{fractional isotope abundance})_i]$$

$$= (34.969)(0.7578) + (36.966)(0.2422) = 35.45 \text{ amu}$$

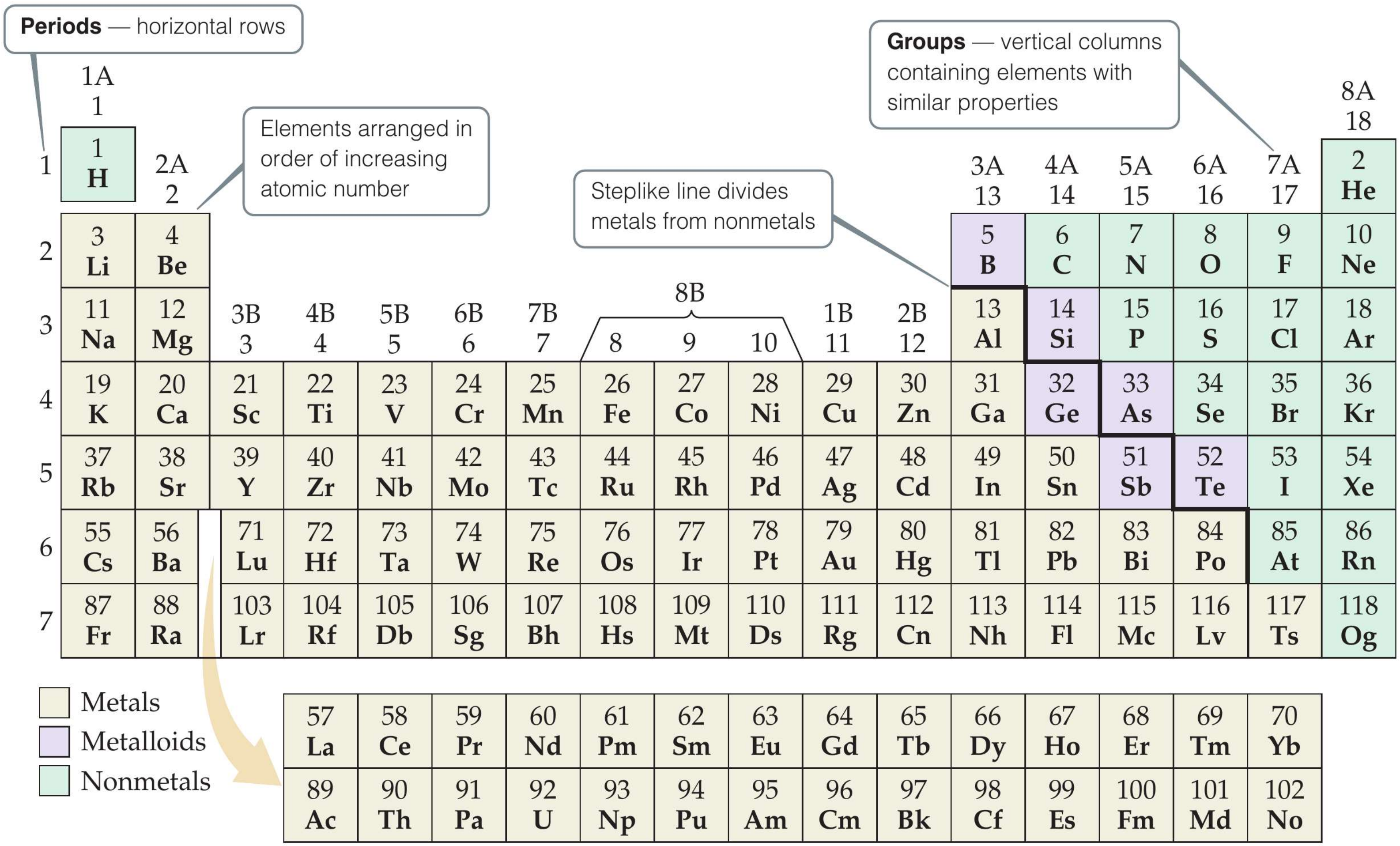
2.5 The Periodic Table

- The arrangement of elements in order of increasing atomic number, with elements having similar properties placed in vertical columns, is known as the periodic table.



- The horizontal rows of the periodic table are called periods.
- The vertical columns are groups.

2.5 The Periodic Table



2.5 The Periodic Table

TABLE 2.3 Names of Some Groups in the Periodic Table

Group	Name	Elements
1A	Alkali metals	Li, Na, K, Rb, Cs, Fr
2A	Alkaline earth metals	Be, Mg, Ca, Sr, Ba, Ra
6A	Chalcogens	O, S, Se, Te, Po
7A	Halogens	F, Cl, Br, I, At
8A	Noble gases	He, Ne, Ar, Kr, Xe, Rn

- The color code in the periodic table shows that, except for hydrogen, all the elements on the left and in the middle of the table are **metallic elements**, or **metals**.
- All the metallic elements share characteristic properties, such as luster and high electrical and heat conductivity, and all of them except mercury (Hg) are solid at room temperature.

2.5 The Periodic Table

- The metals are separated from the **nonmetallic elements**, or **nonmetals**, by a stepped line that runs from boron (B) to astatine (At).
- any of the elements that lie along the line that separates metals from nonmetals have properties that fall between those of metals and nonmetals.
- These elements are often referred to as **metalloids**.
- Only the noble-gas elements are normally found in nature as isolated atoms.

Sample Exercise 2.5

Which two of these elements would you expect to show the greatest similarity in chemical and physical properties: B, Ca, F, He, Mg, P?

- Ca and Mg are most alike because they are in the same group (2A, the alkaline earth metals).

2.6 Molecules and Molecular Compounds

Molecules and Chemical Formulas

- Most of the oxygen in air consists of molecules that contain two oxygen atoms, we represent this molecular oxygen by the **chemical formula** O_2.
- A molecule made up of two atoms is called a diatomic molecule.
- Compounds composed of molecules contain more than one type of atom and are called molecular compounds.

2.6 Molecules and Molecular Compounds

Molecular and empirical Formulas

- Chemical formulas that indicate the actual numbers of atoms in a molecule are called molecular formulas.
- Chemical formulas that give only the relative number of atoms of each type in a molecule are called empirical formulas.
- The subscripts in an empirical formula are always the smallest possible whole-number ratios.
- **Example**: the molecular formula for hydrogen peroxide is H_2O_2, whereas its empirical formula is HO.

Sample Exercise 2.6

Write the empirical formulas for

a. glucose, a substance also known as either blood sugar or dextrose—molecular formula $C_6H_{12}O_6$.

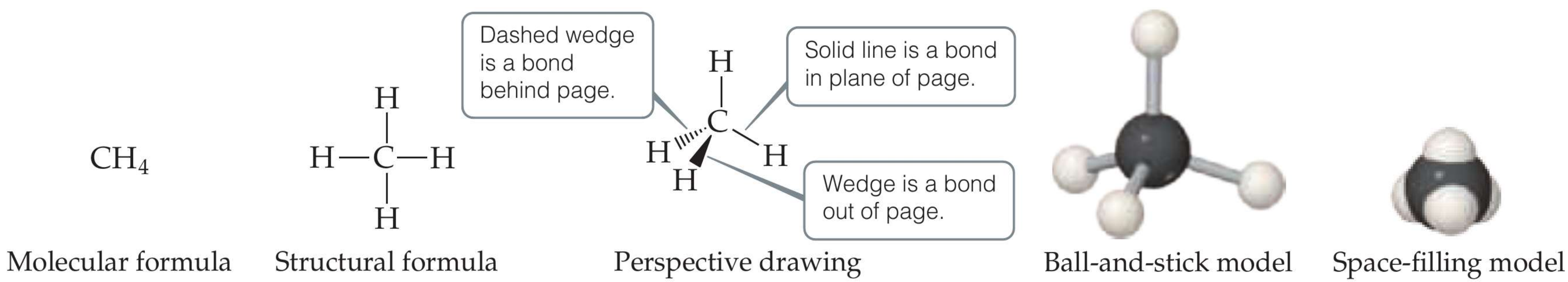
CH_2O

b. nitrous oxide, a substance used as an anesthetic and commonly called laughing gas—molecular formula N_2O.

N_2O

Picturing Molecules

- A **structural formula** show how the atoms of a substance are joined together.
- **Perspective drawings** use wedges and dashed lines to depict bonds that are not in the plane of the paper.



▲ **Figure 2.18 Different representations of the methane (CH_4) molecule.** Structural formulas, perspective drawings, ball-and-stick models, and space-filling models.

2.6 Molecules and Molecular Compounds

- **Ball-and-stick models** show atoms as spheres and bonds as sticks. This type of model has the advantage of accurately representing the angles at which the atoms are attached to one another in a molecule. The identities of the atoms are typically indicated by color.
- **Space-filling models** depict what a molecule would look like if the atoms were scaled up in size. These models show the relative sizes of the atoms, which help define their molecular geometry. The identities of the atoms are typically indicated by color.

2.7 Ions and Ionic Compounds

- If electrons are removed from or added to an atom, a charged particle called an ion is formed.
- An ion with a positive charge is a cation; a negatively charged ion is an anion.
- In general, metal atoms tend to lose electrons to form cations and nonmetal atoms tend to gain electrons to form anions.

Sample Exercise 2.7

Give the chemical symbol, including superscript indicating mass number, for

a. the ion with 22 protons, 26 neutrons, and 19 electrons.

$^{48}_{22}\mathrm{Ti}^{3+}$

b. the ion of sulfur that has 16 neutrons and 18 electrons.

$^{32}_{16}\mathrm{S}^{2-}$

2.7 Ions and Ionic Compounds

- In addition to simple ions such as Na^+ and Cl^-, there are polyatomic ions, such as NH_4^+ (ammonium ion) and SO_4^{2-} (sulfate ion), which consist of atoms joined as in a molecule, but carrying a net positive or negative charge.

Predicting Ionic Charges

- The noble gases are chemically nonreactive elements that form very few compounds.
- Many atoms gain or lose electrons to end up with the same number of electrons as the noble gas closest to them in the periodic table.

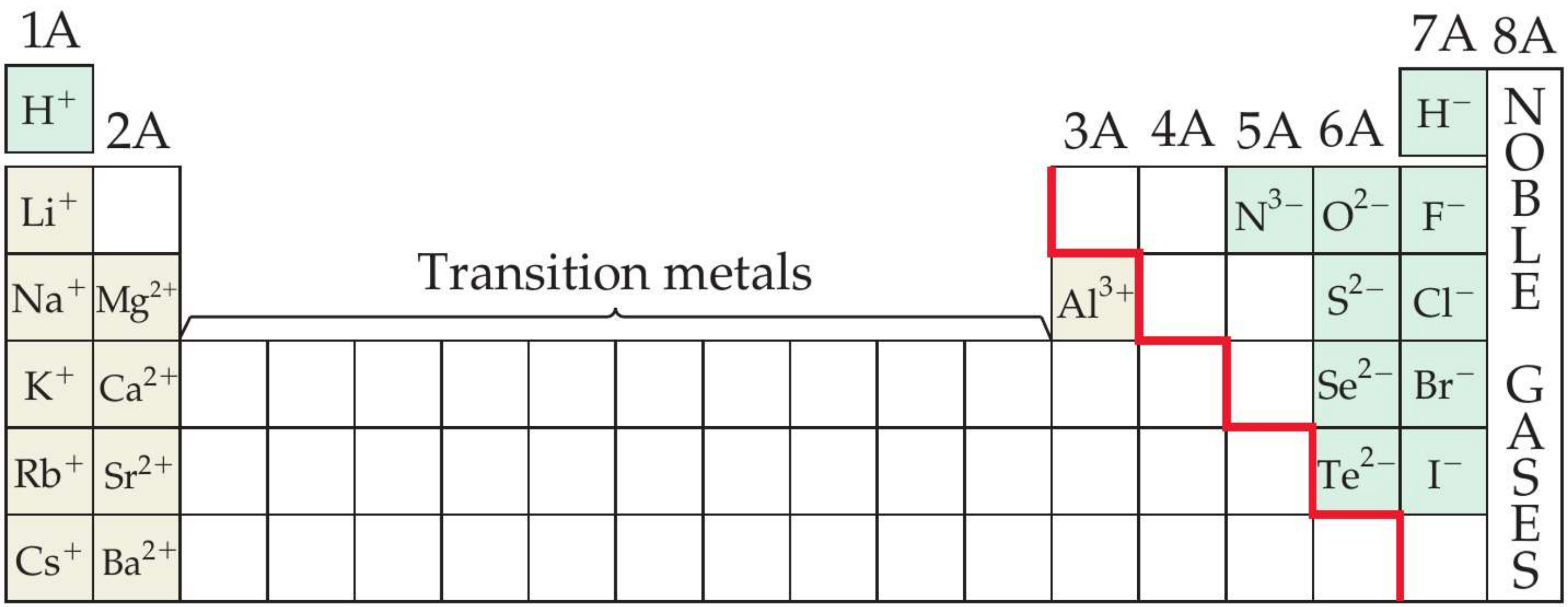
Sample Exercise 2.8

Predict the charge expected for the most stable ion of barium and the most stable ion of oxygen.

- Barium has atomic number 56. The nearest noble gas is xenon, atomic number 54. Barium can attain a stable arrangement of 54 electrons by losing two electrons, forming the Ba^{2+} cation.

- Oxygen has atomic number 8. The nearest noble gas is neon, atomic number 10. Oxygen can attain this stable electron arrangement by gaining two electrons, forming the O^{2-} anion.

2.7 Ions and Ionic Compounds



▲ **Figure 2.19 Predictable charges of some common ions.** Notice that the red stepped line that divides metals from nonmetals also separates cations from anions. Hydrogen forms both 1+ and 1− ions.

2.7 Ions and Ionic Compounds

Ionic Compounds

- An ionic compound is a compound made up of cations and anions.
- Ionic compounds are generally combinations of metals and nonmetals, as in NaCl.
- We can write the empirical formula for an ionic compound if we know the charges of the ions.
- Because chemical compounds are always electrically neutral, the ions in an ionic compound always occur in such a ratio that the total positive charge equals the total negative charge.
- If the charges are not equal, the charge on one ion (without its sign) will become the subscript on the other ion.

2.7 Ions and Ionic Compounds

$$Mg^{2+} \quad N^{3-} \longrightarrow Mg_3N_2$$

- There is one caveat to using this approach.
- So the empirical formula for the ionic compound formed between Ti^{4+} and O^{2-} is TiO_2 rather than Ti_2O_4.

Sample Exercise 2.9

Which of these compounds would you expect to be ionic: N_2O, Na_2O, $CaCl_2$, SF_4?

- We predict that Na_2O and $CaCl_2$ are ionic compounds because they are composed of a metal combined with a nonmetal.
- We predict that N_2O and SF_4 are molecular compounds because they are composed entirely of nonmetals.

Sample Exercise 2.10

Write the empirical formula of the compound formed by

a. Al^{3+} and Cl^{-} ions

$AlCl_3$

b. Al^{3+} and O^{2-} ions

Al_2O_3

c. Mg^{2+} and NO_3^{-} ions

$Mg(NO_3)_2$