

AP Chemistry Summer Assignment

Any questions regarding the assignment below or for an online copy of the assignment:

Contact Mr. Cicale: jcicale@jerichoschools.org

Also, copies of all assignment materials will be posted or linked on the jerichoschools.org website.

1. Read the photocopied chemistry text chapters.
2. *Showing all work NEATLY and CLEARLY in a DETAILED, ACCURATE manner; ALL UNITS MUST BE WRITTEN and PROPERLY CANCELLED.*

Answer the following:

Chapter 1, Question numbers:

16, 18, 22, 34, 35, 71, 75, 76

Chapter 2, Question numbers:

12, 16, 21, 32, 33, 41, 46, 53, 56, 68, 72, 74, 82

3. Memorize the names and symbols of the sixty (60) elements listed below.
4. Memorize the formula, charge, and name of each of the polyatomic ions listed on the next page.
5. Know the cause of the solubility rules listed and explained on the next page.
6. Complete the chemical formula naming and compound writing worksheet.

Elements and symbols to be memorized:

H He Li Be B C N O F Ne Na Mg Al Si P S Cl Ar
K Ca Sc Ti V Cr Mn Fe Co Ni Cu Zn Ga Ge As Se Br
Kr Rb Sr Ag Cd Sn Sb Te I Xe Cs Ba Pt Au Hg Pb Bi
Po Rn Fr Ra Th Pa U Pu

Polyatomic ions (charge, name, and formula) to be memorized:

| | | | |
|---|----------------------|------------------------------|------------------|
| H_3O^+ | hydronium | CrO_4^{2-} | chromate |
| Hg_2^{2+} | dimercury (I) | $\text{Cr}_2\text{O}_7^{2-}$ | dichromate |
| NH_4^+ | ammonium | MnO_4^- | permanganate |
| $\text{C}_2\text{H}_3\text{O}_2^-$ CH_3COO^- | ethanoate or acetate | NO_2^- | nitrite |
| | | NO_3^- | nitrate |
| CN^- | cyanide | O_2^{2-} | peroxide |
| CO_3^{2-} | carbonate | OH^- | hydroxide |
| HCO_3^- | hydrogen carbonate | PO_4^{3-} | phosphate |
| $\text{C}_2\text{O}_4^{2-}$ | oxalate | SCN^- | thiocyanate |
| ClO^- | hypochlorite | SO_3^{2-} | sulfite |
| ClO_2^- | chlorite | SO_4^{2-} | sulfate |
| ClO_3^- | chlorate | HSO_4^- | hydrogen sulfate |
| ClO_4^- | perchlorate | $\text{S}_2\text{O}_3^{2-}$ | thiosulfate |

Spelling and font are significant!

Solubility Rules: Generally, ions that have HIGH charge density (high charge but small SIZE) form the STRONGEST ionic bonds, so they form INSOLUBLE salts; ions that have LOW charge density (low charge and large size) form weaker ionic bonds, so they form soluble salts in aqueous solution.

SOLUBILITY RULES

1. Salts of ammonium (NH_4^+) and Group IA are always soluble.
2. a. All chlorides (Cl^-) are soluble except AgCl , Hg_2Cl_2 , and PbCl_2 which are insoluble.
b. All bromides (Br^-) are soluble except AgBr , Hg_2Br_2 , HgBr_2 , and PbBr_2 which are insoluble.
c. All iodides (I^-) are soluble except AgI , Hg_2I_2 , HgI_2 , and PbI_2 which are insoluble.
3. ALL chlorates (ClO_3^-), nitrates (NO_3^-), and acetates (CH_3COO^-) are soluble.
4. Sulfates (SO_4^{-2}) are soluble except CaSO_4 , SrSO_4 , BaSO_4 , Hg_2SO_4 , HgSO_4 , PbSO_4 , and Ag_2SO_4 which are insoluble.
5. Phosphates (PO_4^{-3}), and carbonates (CO_3^{-2}) are insoluble except NH_4^+ and Group IA compounds.
6. All metallic hydroxides (OH^-) are insoluble except NH_4^+ and Group IA and $\text{Ca}(\text{OH})_2$, $\text{Sr}(\text{OH})_2$, and $\text{Ba}(\text{OH})_2$ of Group IIA.
7. All sulfides (S^{-2}) are insoluble except NH_4^+ and Groups IA and IIA.
For example, CuS is insoluble but CaS is soluble.

Summer Assignment Formula Writing and Compound Naming:

Complete the following chart of ALL possible salts formed from the following lattices of ions; write CLEARLY and NEATLY (you may use separate paper if you need more writing space):

| Cation with charge | Anion with charge | Salt Formula | Salt Name |
|--------------------|-------------------|-------------------|-----------------|
| Group 1 | | | |
| Li ⁺ | C ⁴⁻ | Li ₄ C | lithium carbide |
| | N ³⁻ | | |
| | P ³⁻ | | |
| | O ²⁻ | | |
| | S ²⁻ | | |
| | Se ²⁻ | | |
| | F ⁻ | | |
| | Cl ⁻ | | |
| | Br ⁻ | | |
| | I ⁻ | | |
| | H ⁻ | | |
| Na ⁺ | C ⁴⁻ | | |
| | N ³⁻ | | |
| | P ³⁻ | | |
| | O ²⁻ | | |
| | S ²⁻ | | |
| | Se ²⁻ | | |
| | F ⁻ | | |
| | Cl ⁻ | | |
| | Br ⁻ | | |
| | I ⁻ | | |
| | H ⁻ | | |
| K ⁺ | C ⁴⁻ | | |
| | N ³⁻ | | |
| | P ³⁻ | | |
| | O ²⁻ | | |
| | S ²⁻ | | |
| | Se ²⁻ | | |
| | F ⁻ | | |
| | Cl ⁻ | | |
| | Br ⁻ | | |
| | I ⁻ | | |
| | H ⁻ | | |

| Cation with charge | Anion with charge | Salt Formula | Salt Name |
|---|-------------------|--------------|-----------|
| Group 1 | | | |
| Rb ⁺ | C ⁴⁻ | | |
| | N ³⁻ | | |
| | P ³⁻ | | |
| | O ²⁻ | | |
| | S ²⁻ | | |
| | Se ²⁻ | | |
| | F ⁻ | | |
| | Cl ⁻ | | |
| | Br ⁻ | | |
| | I ⁻ | | |
| | H ⁻ | | |
| Cs ⁺ | C ⁴⁻ | | |
| | N ³⁻ | | |
| | P ³⁻ | | |
| | O ²⁻ | | |
| | S ²⁻ | | |
| | Se ²⁻ | | |
| | F ⁻ | | |
| | Cl ⁻ | | |
| | Br ⁻ | | |
| | I ⁻ | | |
| | H ⁻ | | |
| NH ₄ ⁺ (not in Group I) | C ⁴⁻ | | |
| | N ³⁻ | | |
| | P ³⁻ | | |
| | O ²⁻ | | |
| | S ²⁻ | | |
| | Se ²⁻ | | |
| | F ⁻ | | |
| | Cl ⁻ | | |
| | Br ⁻ | | |
| | I ⁻ | | |
| | H ⁻ | | |

| Cation with charge | Anion with charge | Salt Formula | Salt Name |
|--------------------|-------------------|-------------------|-------------------|
| Group 2 | | | |
| Mg ²⁺ | C ⁴⁻ | Mg ₂ C | magnesium carbide |
| | N ³⁻ | | |
| | P ³⁻ | | |
| | O ²⁻ | | |
| | S ²⁻ | | |
| | Se ²⁻ | | |
| | F ⁻ | | |
| | Cl ⁻ | | |
| | Br ⁻ | | |
| | I ⁻ | | |
| | H ⁻ | | |
| Ca ²⁺ | C ⁴⁻ | | |
| | N ³⁻ | | |
| | P ³⁻ | | |
| | O ²⁻ | | |
| | S ²⁻ | | |
| | Se ²⁻ | | |
| | F ⁻ | | |
| | Cl ⁻ | | |
| | Br ⁻ | | |
| | I ⁻ | | |
| | H ⁻ | | |
| Sr ²⁺ | C ⁴⁻ | | |
| | N ³⁻ | | |
| | P ³⁻ | | |
| | O ²⁻ | | |
| | S ²⁻ | | |
| | Se ²⁻ | | |
| | F ⁻ | | |
| | Cl ⁻ | | |
| | Br ⁻ | | |
| | I ⁻ | | |
| | H ⁻ | | |

Now for salts of POLYATOMIC ANIONS

| Cation with charge | Anion with charge | Salt Formula | Salt Name |
|--------------------|--|---------------------------------|-------------------|
| Group 1 | | | |
| Li ⁺ | PO ₄ ³⁻ | Li ₃ PO ₄ | lithium phosphate |
| | C ₂ O ₄ ²⁻ | | |
| | SO ₄ ²⁻ | | |
| | SO ₃ ²⁻ | | |
| | CO ₃ ²⁻ | | |
| | Cr ₂ O ₇ ²⁻ | | |
| | CrO ₄ ²⁻ | | |
| | HCO ₃ ⁻ | | |
| | CH ₃ COO ⁻ | LiCH ₃ COO | lithium ethanoate |
| | NO ₃ ⁻ | | |
| | NO ₂ ⁻ | | |
| Na ⁺ | PO ₄ ³⁻ | | |
| | C ₂ O ₄ ²⁻ | | |
| | SO ₄ ²⁻ | | |
| | SO ₃ ²⁻ | | |
| | CO ₃ ²⁻ | | |
| | Cr ₂ O ₇ ²⁻ | | |
| | CrO ₄ ²⁻ | | |
| | HCO ₃ ⁻ | | |
| | CH ₃ COO ⁻ | | |
| | NO ₃ ⁻ | | |
| | NO ₂ ⁻ | | |
| K ⁺ | PO ₄ ³⁻ | | |
| | C ₂ O ₄ ²⁻ | | |
| | SO ₄ ²⁻ | | |
| | SO ₃ ²⁻ | | |
| | CO ₃ ²⁻ | | |
| | Cr ₂ O ₇ ²⁻ | | |
| | CrO ₄ ²⁻ | | |
| | HCO ₃ ⁻ | | |
| | CH ₃ COO ⁻ | | |
| | NO ₃ ⁻ | | |
| | NO ₂ ⁻ | | |

| Cation with charge | Anion with charge | Salt Formula | Salt Name |
|---|--|--------------|-----------|
| Group 1 | | | |
| Rb ⁺ | PO ₄ ³⁻ | | |
| | C ₂ O ₄ ²⁻ | | |
| | SO ₄ ²⁻ | | |
| | SO ₃ ²⁻ | | |
| | CO ₃ ²⁻ | | |
| | Cr ₂ O ₇ ²⁻ | | |
| | CrO ₄ ²⁻ | | |
| | HCO ₃ ⁻ | | |
| | CH ₃ COO ⁻ | | |
| | NO ₃ ⁻ | | |
| | NO ₂ ⁻ | | |
| Cs ⁺ | PO ₄ ³⁻ | | |
| | C ₂ O ₄ ²⁻ | | |
| | SO ₄ ²⁻ | | |
| | SO ₃ ²⁻ | | |
| | CO ₃ ²⁻ | | |
| | Cr ₂ O ₇ ²⁻ | | |
| | CrO ₄ ²⁻ | | |
| | HCO ₃ ⁻ | | |
| | CH ₃ COO ⁻ | | |
| | NO ₃ ⁻ | | |
| | NO ₂ ⁻ | | |
| NH ₄ ⁺ (not in Group I) | PO ₄ ³⁻ | | |
| | C ₂ O ₄ ²⁻ | | |
| | SO ₄ ²⁻ | | |
| | SO ₃ ²⁻ | | |
| | CO ₃ ²⁻ | | |
| | Cr ₂ O ₇ ²⁻ | | |
| | CrO ₄ ²⁻ | | |
| | HCO ₃ ⁻ | | |
| | CH ₃ COO ⁻ | | |
| | NO ₃ ⁻ | | |
| | NO ₂ ⁻ | | |

| Cation with charge | Anion with charge | Salt Formula | Salt Name |
|--------------------|--|--------------|-----------|
| Group 2 | | | |
| Mg ²⁺ | PO ₄ ³⁻ | | |
| | C ₂ O ₄ ²⁻ | | |
| | SO ₄ ²⁻ | | |
| | SO ₃ ²⁻ | | |
| | CO ₃ ²⁻ | | |
| | Cr ₂ O ₇ ²⁻ | | |
| | CrO ₄ ²⁻ | | |
| | HCO ₃ ⁻ | | |
| | CH ₃ COO ⁻ | | |
| | NO ₃ ⁻ | | |
| | NO ₂ ⁻ | | |
| Ca ²⁺ | PO ₄ ³⁻ | | |
| | C ₂ O ₄ ²⁻ | | |
| | SO ₄ ²⁻ | | |
| | SO ₃ ²⁻ | | |
| | CO ₃ ²⁻ | | |
| | Cr ₂ O ₇ ²⁻ | | |
| | CrO ₄ ²⁻ | | |
| | HCO ₃ ⁻ | | |
| | CH ₃ COO ⁻ | | |
| | NO ₃ ⁻ | | |
| | NO ₂ ⁻ | | |
| Sr ²⁺ | PO ₄ ³⁻ | | |
| | C ₂ O ₄ ²⁻ | | |
| | SO ₄ ²⁻ | | |
| | SO ₃ ²⁻ | | |
| | CO ₃ ²⁻ | | |
| | Cr ₂ O ₇ ²⁻ | | |
| | CrO ₄ ²⁻ | | |
| | HCO ₃ ⁻ | | |
| | CH ₃ COO ⁻ | | |
| | NO ₃ ⁻ | | |
| | NO ₂ ⁻ | | |

Name the following acids:

HF

HCl

HBr

HI

H₂S

H₃P

CH₃COOH (do NOT write the non-IUPAC, common name "acetic acid")

HNO₂

H₂SO₄

Write the formula of the following acids:

phosphoric acid

nitric acid

dichromic acid

chromic acid

sulfurous acid

oxalic acid

permanganic acid

chloric acid

perchloric acid

chlorous acid

hypochlorous acid

thiosulfurous acid

Name the following MOLECULES:

(compounds of covalently bonded atoms that make up discrete molecules that have no charge; there are NO ions in ANY molecule!)

prefix name

Stock system name

PCl₃

SO₂

NO

N₂O₃

N₂O₅

CO

ClO₂

Write the following molecular formulas:

tetraphosphorus decoxide

dinitrogen tetroxide

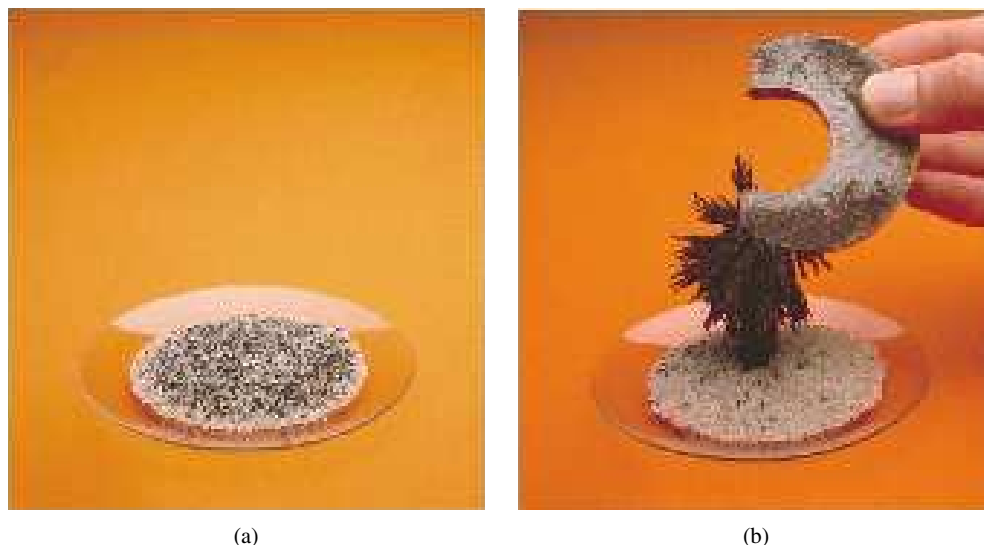
nitrogen triiodide

dichlorine heptafluoride

phosphorus pentachloride

phosphorus tribromide

FIGURE 1.4 Separating iron filings from a heterogeneous mixture. The same technique is used on a larger scale to separate iron and steel from nonmagnetic objects such as aluminum, glass, and plastics.



a dead end, and in which positive achievements came only after many wrong turns and at such a slow pace that they went unheralded. Yet even the dead ends contribute something to the continually growing body of knowledge about the physical universe. It is the love of the search that keeps many scientists in the laboratory.

1.4 CLASSIFICATIONS OF MATTER

We defined chemistry at the beginning of the chapter as the study of matter and the changes it undergoes. **Matter** is *anything that occupies space and has mass*. Matter includes things we can see and touch (such as water, earth, and trees), as well as things we cannot (such as air). Thus, everything in the universe has a “chemical” connection.

Chemists distinguish among several subcategories of matter based on composition and properties. The classifications of matter include substances, mixtures, elements, and compounds, as well as atoms and molecules, which we will consider in Chapter 2.

SUBSTANCES AND MIXTURES

A **substance** is a form of matter that has a definite (constant) composition and distinct properties. Examples are water, ammonia, table sugar (sucrose), gold, and oxygen. Substances differ from one another in composition and can be identified by their appearance, smell, taste, and other properties.

A **mixture** is a combination of two or more substances in which the substances retain their distinct identities. Some familiar examples are air, soft drinks, milk, and cement. Mixtures do not have constant composition. Therefore, samples of air collected in different cities would probably differ in composition because of differences in altitude, pollution, and so on.

Mixtures are either homogeneous or heterogeneous. When a spoonful of sugar dissolves in water we obtain a **homogeneous mixture** in which *the composition of the mixture is the same throughout*. If sand is mixed with iron filings, however, the sand grains and the iron filings remain separate (Figure 1.4). This type of mixture is called a **heterogeneous mixture** because *the composition is not uniform*.

Any mixture, whether homogeneous or heterogeneous, can be created and then separated by physical means into pure components without changing the identities of the components. Thus, sugar can be recovered from a water solution by heating the solution and evaporating it to dryness. Condensing the vapor will give us back the water component. To separate the iron-sand mixture, we can use a magnet to remove the iron filings from the sand, because sand is not attracted to the magnet [see Figure 1.4(b)]. After separation, the components of the mixture will have the same composition and properties as they did to start with.

ELEMENTS AND COMPOUNDS

Substances can be either elements or compounds. An *element* is a substance that cannot be separated into simpler substances by chemical means. To date, 112 elements have been positively identified. Eighty-three of them occur naturally on Earth. The others have been created by scientists via nuclear processes, which are the subject of Chapter 23 of this text.

For convenience, chemists use symbols of one, two, or three letters to represent the elements. The first letter of a symbol is *always* capitalized, but any following letters are not. For example, Co is the symbol for the element cobalt, whereas CO is the formula for the carbon monoxide molecule. Table 1.1 shows the names and symbols of some of the more common elements; a complete list of the elements and their symbols appears inside the front cover of this book. The symbols of some elements are derived from their Latin names—for example, Au from *aurum* (gold), Fe from *ferrum* (iron), and Na from *natrium* (sodium)—while most of them come from their English names. Appendix 1 gives the origin of the names and lists the discoverers of most of the elements.

Most elements can interact with one or more other elements to form compounds. Hydrogen gas, for example, burns in oxygen gas to form water, which has properties that are distinctly different from those of the starting materials. Water is made up of two parts hydrogen and one part oxygen. This composition does not change, regardless of whether the water comes from a faucet in the United States, a lake in Outer Mongolia, or the ice caps on Mars. Thus, water is a **compound**, a substance composed of atoms of two or more elements chemically united in fixed proportions. Unlike mixtures, compounds can be separated only by chemical means into their pure components.

TABLE 1.1 Some Common Elements and Their Symbols

| NAME | SYMBOL | NAME | SYMBOL | NAME | SYMBOL |
|----------|--------|-----------|--------|------------|--------|
| Aluminum | Al | Fluorine | F | Oxygen | O |
| Arsenic | As | Gold | Au | Phosphorus | P |
| Barium | Ba | Hydrogen | H | Platinum | Pt |
| Bismuth | Bi | Iodine | I | Potassium | K |
| Bromine | Br | Iron | Fe | Silicon | Si |
| Calcium | Ca | Lead | Pb | Silver | Ag |
| Carbon | C | Magnesium | Mg | Sodium | Na |
| Chlorine | Cl | Manganese | Mn | Sulfur | S |
| Chromium | Cr | Mercury | Hg | Tin | Sn |
| Cobalt | Co | Nickel | Ni | Tungsten | W |
| Copper | Cu | Nitrogen | N | Zinc | Zn |

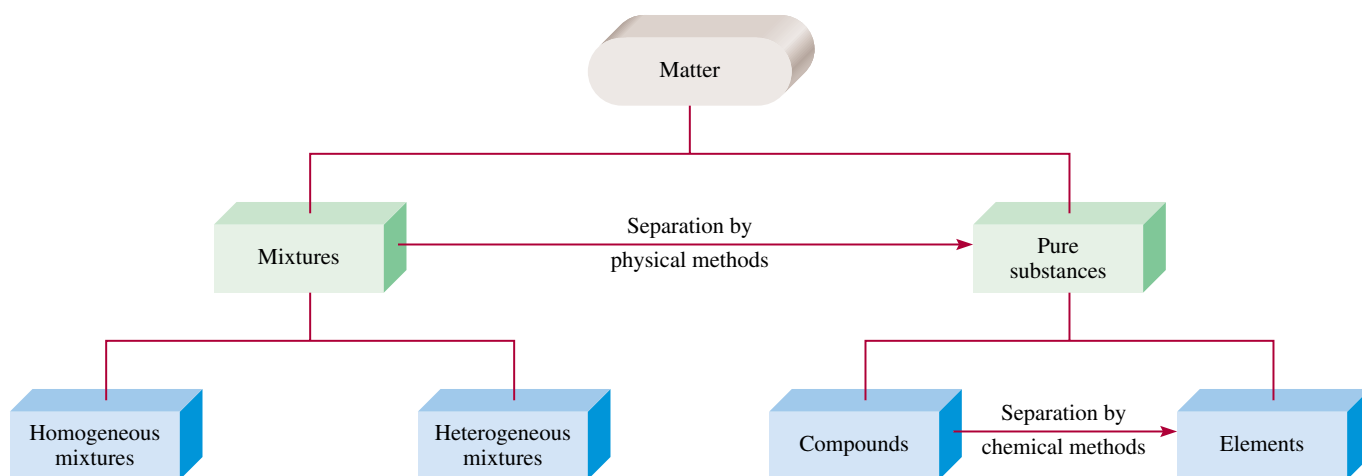


FIGURE 1.5 Classification of matter.

The relationships among elements, compounds, and other categories of matter are summarized in Figure 1.5.

1.5 THE THREE STATES OF MATTER

All substances, at least in principle, can exist in three states: solid, liquid, and gas. As Figure 1.6 shows, gases differ from liquids and solids in the distances between the molecules. In a solid, molecules are held close together in an orderly fashion with little freedom of motion. Molecules in a liquid are close together but are not held so rigidly in position and can move past one another. In a gas, the molecules are separated by distances that are large compared with the size of the molecules.

The three states of matter can be interconverted without changing the composition of the substance. Upon heating, a solid (for example, ice) will melt to form a liquid (water). (The temperature at which this transition occurs is called the *melting point*.) Further heating will convert the liquid into a gas. (This conversion takes place at the *boiling point* of the liquid.) On the other hand, cooling a gas will cause it to condense

FIGURE 1.6 Microscopic views of a solid, a liquid, and a gas.

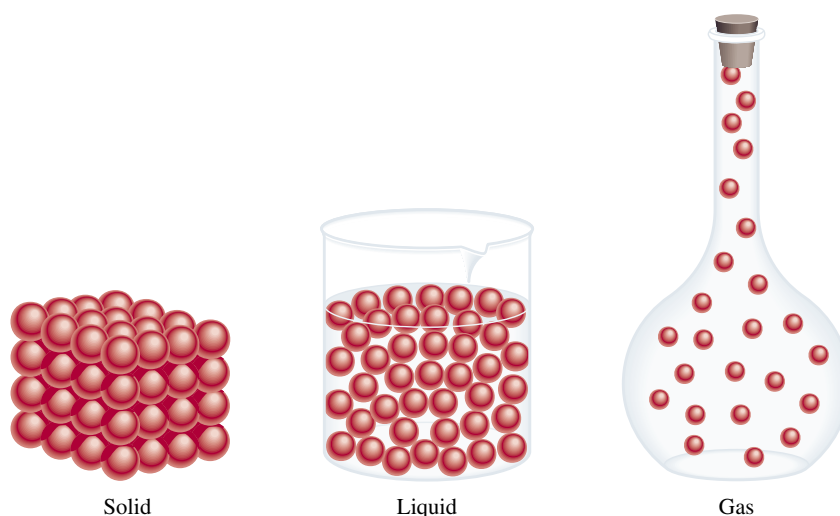




FIGURE 1.7 The three states of matter. A hot poker changes ice into water and steam.



into a liquid. When the liquid is cooled further, it will freeze into the solid form. Figure 1.7 shows the three states of water.

1.6 PHYSICAL AND CHEMICAL PROPERTIES OF MATTER

Substances are identified by their properties as well as by their composition. Color, melting point, and boiling point are physical properties. A **physical property** can be measured and observed without changing the composition or identity of a substance. For example, we can measure the melting point of ice by heating a block of ice and recording the temperature at which the ice is converted to water. Water differs from ice only in appearance, not in composition, so this is a physical change; we can freeze the water to recover the original ice. Therefore, the melting point of a substance is a physical property. Similarly, when we say that helium gas is lighter than air, we are referring to a physical property.

On the other hand, the statement “Hydrogen gas burns in oxygen gas to form water” describes a **chemical property** of hydrogen, because *in order to observe this property we must carry out a chemical change*, in this case burning. After the change, the original chemical substance, the hydrogen gas, will have vanished, and all that will be left is a different chemical substance—water. We *cannot* recover the hydrogen from the water by means of a physical change, such as boiling or freezing.

Every time we hard-boil an egg, we bring about a chemical change. When subjected to a temperature of about 100°C , the yolk and the egg white undergo changes that alter not only their physical appearance but their chemical makeup as well. When eaten, the egg is changed again, by substances in our bodies called *enzymes*. This digestive action is another example of a chemical change. What happens during digestion depends on the chemical properties of both the enzymes and the food.



Hydrogen burning in air to form water.

All measurable properties of matter fall into one of two additional categories: extensive properties and intensive properties. The measured value of an **extensive property** depends on how much matter is being considered. **Mass**, which is the quantity of matter in a given sample of a substance, is an extensive property. More matter means more mass. Values of the same extensive property can be added together. For example, two copper pennies will have a combined mass that is the sum of the masses of each penny, and the length of two tennis courts is the sum of the lengths of each tennis court. **Volume**, defined as *length cubed*, is another extensive property. The value of an extensive quantity depends on the amount of matter.

The measured value of an **intensive property** does not depend on how much matter is being considered. **Density**, defined as the mass of an object divided by its volume, is an intensive property. So is temperature. Suppose that we have two beakers of water at the same temperature. If we combine them to make a single quantity of water in a larger beaker, the temperature of the larger quantity of water will be the same as it was in two separate beakers. Unlike mass, length, and volume, temperature and other intensive properties are not additive.

1.7 MEASUREMENT

The measurements chemists make are often used in calculations to obtain other related quantities. Different instruments enable us to measure a substance's properties: The meter stick measures length or scale; the buret, the pipet, the graduated cylinder, and the volumetric flask measure volume (Figure 1.8); the balance measures mass; the thermometer measures temperature. These instruments provide measurements of **macroscopic properties**, which can be determined directly. **Microscopic properties**, on the atomic or molecular scale, must be determined by an indirect method, as we will see in the next chapter.

FIGURE 1.8 Some common measuring devices found in a chemistry laboratory. These devices are not drawn to scale relative to one another. We will discuss the uses of these measuring devices in Chapter 4.

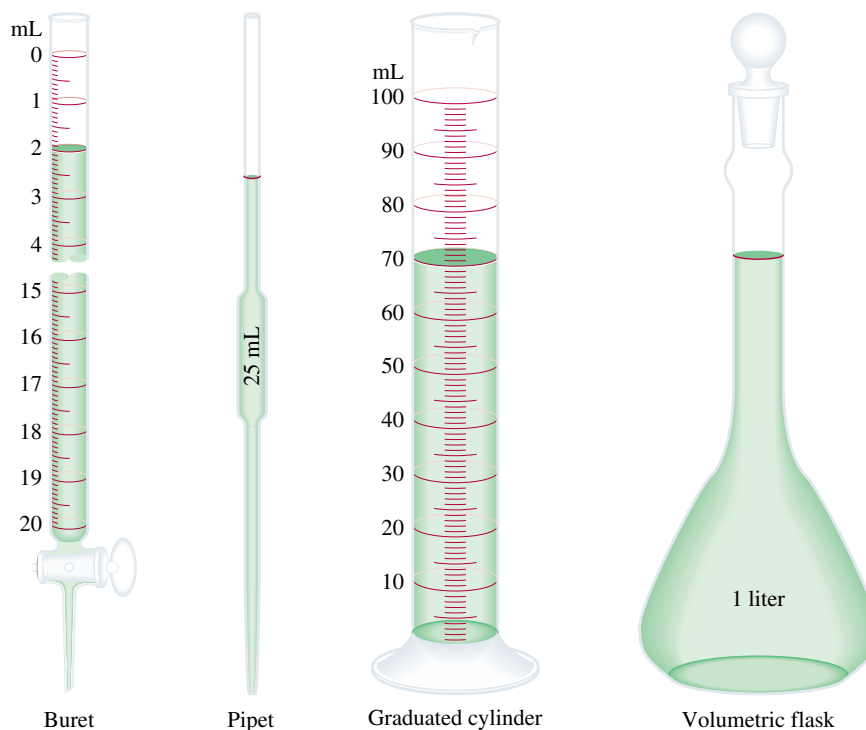


TABLE 1.2 SI Base Units

| BASE QUANTITY | NAME OF UNIT | SYMBOL |
|---------------------|--------------|--------|
| Length | Meter | m |
| Mass | Kilogram | kg |
| Time | Second | s |
| Electrical current | Ampere | A |
| Temperature | Kelvin | K |
| Amount of substance | Mole | mol |
| Luminous intensity | Candela | cd |

TABLE 1.3 Prefixes Used with SI Units

| PREFIX | SYMBOL | MEANING | EXAMPLE |
|--------|--------|------------------------------------|--|
| Tera- | T | 1,000,000,000,000, or 10^{12} | 1 terameter (Tm) = 1×10^{12} m |
| Giga- | G | 1,000,000,000, or 10^9 | 1 gigameter (Gm) = 1×10^9 m |
| Mega- | M | 1,000,000, or 10^6 | 1 megameter (Mm) = 1×10^6 m |
| Kilo- | k | 1,000, or 10^3 | 1 kilometer (km) = 1×10^3 m |
| Deci- | d | 1/10, or 10^{-1} | 1 decimeter (dm) = 0.1 m |
| Centi- | c | 1/100, or 10^{-2} | 1 centimeter (cm) = 0.01 m |
| Milli- | m | 1/1,000, or 10^{-3} | 1 millimeter (mm) = 0.001 m |
| Micro- | μ | 1/1,000,000, or 10^{-6} | 1 micrometer (μ m) = 1×10^{-6} m |
| Nano- | n | 1/1,000,000,000, or 10^{-9} | 1 nanometer (nm) = 1×10^{-9} m |
| Pico- | p | 1/1,000,000,000,000, or 10^{-12} | 1 picometer (pm) = 1×10^{-12} m |

A measured quantity is usually written as a number with an appropriate unit. To say that the distance between New York and San Francisco by car along a certain route is 5166 is meaningless. We must specify that the distance is 5166 kilometers. The same is true in chemistry; units are essential to stating measurements correctly.

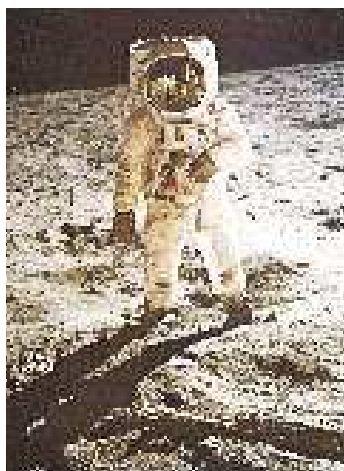
SI UNITS

For many years scientists recorded measurements in *metric units*, which are related decimally, that is, by powers of 10. In 1960, however, the General Conference of Weights and Measures, the international authority on units, proposed a revised metric system called the *International System of Units* (abbreviated **SI**, from the French *Système International d'Unités*). Table 1.2 shows the seven SI base units. All other units of measurement can be derived from these base units. Like metric units, SI units are modified in decimal fashion by a series of prefixes, as shown in Table 1.3. We will use both metric and SI units in this book.

Measurements that we will utilize frequently in our study of chemistry include time, mass, volume, density, and temperature.

MASS AND WEIGHT

The terms “mass” and “weight” are often used interchangeably, although, strictly speaking, they are different quantities. Whereas mass is a measure of the amount of matter in an object, *weight*, technically speaking, is *the force that gravity exerts on an object*. An apple that falls from a tree is pulled downward by Earth’s gravity. The mass of the



An astronaut on the surface of the moon.

apple is constant and does not depend on its location, but its weight does. For example, on the surface of the moon the apple would weigh only one-sixth what it does on Earth, because the moon's gravity is only one-sixth that of Earth. The moon's smaller gravity enables astronauts to jump about rather freely on its surface despite their bulky suits and equipment. Chemists are interested primarily in mass, which can be determined readily with a balance; the process of measuring mass, oddly, is called *weighing*.

The SI base unit of mass is the *kilogram* (kg), but in chemistry the smaller *gram* (g) is more convenient:

$$1 \text{ kg} = 1000 \text{ g} = 1 \times 10^3 \text{ g}$$

VOLUME

The SI unit of length is the *meter* (m), and the SI-derived unit for volume is the *cubic meter* (m^3). Generally, however, chemists work with much smaller volumes, such as the cubic centimeter (cm^3) and the cubic decimeter (dm^3):

$$1 \text{ cm}^3 = (1 \times 10^{-2} \text{ m})^3 = 1 \times 10^{-6} \text{ m}^3$$

$$1 \text{ dm}^3 = (1 \times 10^{-1} \text{ m})^3 = 1 \times 10^{-3} \text{ m}^3$$

Another common unit of volume is the liter (L). A *liter* is the volume occupied by one cubic decimeter. One liter of volume is equal to 1000 milliliters (mL) or 1000 cm^3 :

$$\begin{aligned} 1 \text{ L} &= 1000 \text{ mL} \\ &= 1000 \text{ cm}^3 \\ &= 1 \text{ dm}^3 \end{aligned}$$

and one milliliter is equal to one cubic centimeter:

$$1 \text{ mL} = 1 \text{ cm}^3$$

Figure 1.9 compares the relative sizes of two volumes. Even though the liter is not an SI unit, volumes are usually expressed in liters and milliliters.

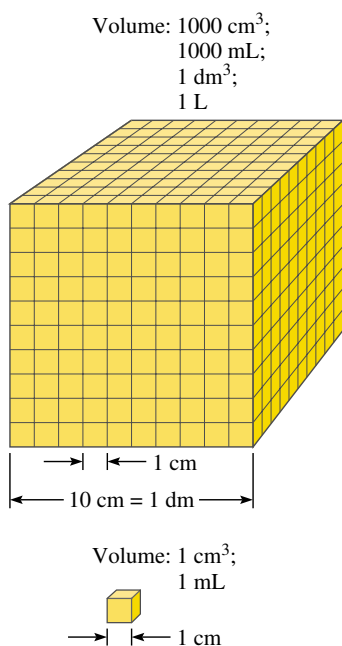


FIGURE 1.9 Comparison of two volumes, 1 mL and 1000 mL.

DENSITY

The equation for density is

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

or

$$d = \frac{m}{V} \quad (1.1)$$

where d , m , and V denote density, mass, and volume, respectively. Because density is an intensive property and does not depend on the quantity of mass present, for a given material the ratio of mass to volume always remains the same; in other words, V increases as m does.

The SI-derived unit for density is the kilogram per cubic meter (kg/m^3). This unit is awkwardly large for most chemical applications. Therefore, grams per cubic centimeter (g/cm^3) and its equivalent, grams per milliliter (g/mL), are more commonly used for solid and liquid densities. Because gas densities are often very low, we express them in units of grams per liter (g/L):

$$1 \text{ g/cm}^3 = 1 \text{ g/mL} = 1000 \text{ kg/m}^3$$

$$1 \text{ g/L} = 0.001 \text{ g/mL}$$

The following examples show density calculations.



Gold bars.

Similar problems: 1.21, 1.22.

EXAMPLE 1.1

Gold is a precious metal that is chemically unreactive. It is used mainly in jewelry, dentistry, and electronic devices. A piece of gold ingot with a mass of 301 g has a volume of 15.6 cm³. Calculate the density of gold.

Answer The density of the gold metal is given by

$$\begin{aligned} d &= \frac{m}{V} \\ &= \frac{301 \text{ g}}{15.6 \text{ cm}^3} \\ &= 19.3 \text{ g/cm}^3 \end{aligned}$$

PRACTICE EXERCISE

A piece of platinum metal with a density of 21.5 g/cm³ has a volume of 4.49 cm³. What is its mass?



Ethanol is produced during the fermentation of bread. It evaporates during baking and produces the fragrant aroma.

Similar problems: 1.21, 1.22.

EXAMPLE 1.2

The density of ethanol, a colorless liquid that is commonly known as grain alcohol, is 0.798 g/mL. Calculate the mass of 17.4 mL of the liquid.

Answer The mass of ethanol is found by rearranging the density equation $d = m/V$ as follows:

$$\begin{aligned} m &= d \times V \\ &= 0.798 \frac{\text{g}}{\text{mL}} \times 17.4 \text{ mL} \\ &= 13.9 \text{ g} \end{aligned}$$

PRACTICE EXERCISE

The density of sulfuric acid in a certain car battery is 1.41 g/mL. Calculate the mass of 242 mL of the liquid.

TEMPERATURE SCALES

Three temperature scales are currently in use. Their units are °F (degrees Fahrenheit), °C (degrees Celsius), and K (kelvin). The Fahrenheit scale, which is the most commonly used scale in the United States outside the laboratory, defines the normal freezing and boiling points of water to be exactly 32°F and 212°F, respectively. The Celsius scale divides the range between the freezing point (0°C) and boiling point (100°C) of water into 100 degrees. As Table 1.2 shows, the *kelvin* is the *SI base unit of temperature*; it is the *absolute* temperature scale. By absolute we mean that the zero on the kelvin scale, denoted by 0 K, is the lowest temperature that can be attained theoreti-

Note that the kelvin scale does not have the degree sign. Also, temperatures expressed in kelvin can never be negative.

cally. On the other hand, 0°F and 0°C are based on the behavior of an arbitrarily chosen substance, water. Figure 1.10 compares the three temperature scales.

The size of a degree on the Fahrenheit scale is only 100/180, or 5/9, of a degree on the Celsius scale. To convert degrees Fahrenheit to degrees Celsius, we write

$$?^{\circ}\text{C} = (^{\circ}\text{F} - 32^{\circ}\text{F}) \times \frac{5^{\circ}\text{C}}{9^{\circ}\text{F}} \quad (1.2)$$

The following equation is used to convert degrees Celsius to degrees Fahrenheit

$$?^{\circ}\text{F} = \frac{9^{\circ}\text{F}}{5^{\circ}\text{C}} \times (^{\circ}\text{C}) + 32^{\circ}\text{F} \quad (1.3)$$

Both the Celsius and the Kelvin scales have units of equal magnitude; that is, one degree Celsius is equivalent to one kelvin. Experimental studies have shown that absolute zero on the kelvin scale is equivalent to -273.15°C on the Celsius scale. Thus we can use the following equation to convert degrees Celsius to kelvin:

$$? \text{ K} = (^{\circ}\text{C} + 273.15^{\circ}\text{C}) \frac{1 \text{ K}}{1^{\circ}\text{C}} \quad (1.4)$$

We will frequently find it necessary to convert between degrees Celsius and degrees Fahrenheit and between degrees Celsius and kelvin. The following example illustrates these conversions.



Solder is used extensively in the construction of electronic circuits.

Similar problems: 1.24, 1.25, 1.26.

EXAMPLE 1.3

(a) Solder is an alloy made of tin and lead that is used in electronic circuits. A certain solder has a melting point of 224°C . What is its melting point in degrees Fahrenheit? (b) Helium has the lowest boiling point of all the elements at -452°F . Convert this temperature to degrees Celsius. (c) Mercury, the only metal that exists as a liquid at room temperature, melts at -38.9°C . Convert its melting point to kelvins.

Answer (a) This conversion is carried out by writing

$$\frac{9^{\circ}\text{F}}{5^{\circ}\text{C}} \times (224^{\circ}\text{C}) + 32^{\circ}\text{F} = 435^{\circ}\text{F}$$

(b) Here we have

$$(-452^{\circ}\text{F} - 32^{\circ}\text{F}) \times \frac{5^{\circ}\text{C}}{9^{\circ}\text{F}} = -269^{\circ}\text{C}$$

(c) The melting point of mercury in kelvins is given by

$$(-38.9^{\circ}\text{C} + 273.15^{\circ}\text{C}) \times \frac{1 \text{ K}}{1^{\circ}\text{C}} = 234.3 \text{ K}$$

PRACTICE EXERCISE

Convert (a) 327.5°C (the melting point of lead) to degrees Fahrenheit; (b) 172.9°F (the boiling point of ethanol) to degrees Celsius; and (c) 77 K, the boiling point of liquid nitrogen, to degrees Celsius.

that the decimal point must be moved to give the number N (which is between 1 and 10). If the decimal point has to be moved to the left, then n is a positive integer; if it has to be moved to the right, n is a negative integer. The following examples illustrate the use of scientific notation:

(a) Express 568.762 in scientific notation:

$$568.762 = 5.68762 \times 10^2$$

Note that the decimal point is moved to the left by two places and $n = 2$.

(b) Express 0.00000772 in scientific notation:

$$0.00000772 = 7.72 \times 10^{-6}$$

Here the decimal point is moved to the right by six places and $n = -6$.

Any number raised to the power zero is equal to one.

Keep in mind the following two points. First, $n = 0$ is used for numbers that are not expressed in scientific notation. For example, 74.6×10^0 ($n = 0$) is equivalent to 74.6. Second, the usual practice is to omit the superscript when $n = 1$. Thus the scientific notation for 74.6 is 7.46×10 and not 7.46×10^1 .

Next, we consider how scientific notation is handled in arithmetic operations.

Addition and Subtraction

To add or subtract using scientific notation, we first write each quantity—say N_1 and N_2 —with the same exponent n . Then we combine N_1 and N_2 ; the exponents remain the same. Consider the following examples:

$$\begin{aligned}(7.4 \times 10^3) + (2.1 \times 10^3) &= 9.5 \times 10^3 \\(4.31 \times 10^4) + (3.9 \times 10^3) &= (4.31 \times 10^4) + (0.39 \times 10^4) \\ &= 4.70 \times 10^4 \\(2.22 \times 10^{-2}) - (4.10 \times 10^{-3}) &= (2.22 \times 10^{-2}) - (0.41 \times 10^{-2}) \\ &= 1.81 \times 10^{-2}\end{aligned}$$

Multiplication and Division

To multiply numbers expressed in scientific notation, we multiply N_1 and N_2 in the usual way, but *add* the exponents together. To divide using scientific notation, we divide N_1 and N_2 as usual and subtract the exponents. The following examples show how these operations are performed:

$$\begin{aligned}(8.0 \times 10^4) \times (5.0 \times 10^2) &= (8.0 \times 5.0)(10^{4+2}) \\ &= 40 \times 10^6 \\ &= 4.0 \times 10^7 \\(4.0 \times 10^{-5}) \times (7.0 \times 10^3) &= (4.0 \times 7.0)(10^{-5+3}) \\ &= 28 \times 10^{-2} \\ &= 2.8 \times 10^{-1} \\ \frac{6.9 \times 10^7}{3.0 \times 10^{-5}} &= \frac{6.9}{3.0} \times 10^{7-(-5)} \\ &= 2.3 \times 10^{12} \\ \frac{8.5 \times 10^4}{5.0 \times 10^9} &= \frac{8.5}{5.0} \times 10^{4-9} \\ &= 1.7 \times 10^{-5}\end{aligned}$$

FIGURE 1.11 A single-pan balance.



SIGNIFICANT FIGURES

Except when all the numbers involved are integers (for example, in counting the number of students in a class), it is often impossible to obtain the exact value of the quantity under investigation. For this reason, it is important to indicate the margin of error in a measurement by clearly indicating the number of *significant figures*, which are *the meaningful digits in a measured or calculated quantity*. When significant figures are used, the last digit is understood to be uncertain. For example, we might measure the volume of a given amount of liquid using a graduated cylinder with a scale that gives an uncertainty of 1 mL in the measurement. If the volume is found to be 6 mL, then the actual volume is in the range of 5 mL to 7 mL. We represent the volume of the liquid as (6 ± 1) mL. In this case, there is only one significant figure (the digit 6) that is uncertain by either plus or minus 1 mL. For greater accuracy, we might use a graduated cylinder that has finer divisions, so that the volume we measure is now uncertain by only 0.1 mL. If the volume of the liquid is now found to be 6.0 mL, we may express the quantity as (6.0 ± 0.1) mL, and the actual value is somewhere between 5.9 mL and 6.1 mL. We can further improve the measuring device and obtain more significant figures, but in every case, the last digit is always uncertain; the amount of this uncertainty depends on the particular measuring device we use.

Figure 1.11 shows a modern balance. Balances such as this one are available in many general chemistry laboratories; they readily measure the mass of objects to four decimal places. Therefore the measured mass typically will have four significant figures (for example, 0.8642 g) or more (for example, 3.9745 g). Keeping track of the number of significant figures in a measurement such as mass ensures that calculations involving the data will reflect the precision of the measurement.

Guidelines for Using Significant Figures

We must always be careful in scientific work to write the proper number of significant figures. In general, it is fairly easy to determine how many significant figures a number has by following these rules:



- Any digit that is not zero is significant. Thus 845 cm has three significant figures, 1.234 kg has four significant figures, and so on.
- Zeros between nonzero digits are significant. Thus 606 m contains three significant figures, 40,501 kg contains five significant figures, and so on.
- Zeros to the left of the first nonzero digit are not significant. Their purpose is to indicate the placement of the decimal point. For example, 0.08 L contains one significant figure, 0.0000349 g contains three significant figures, and so on.
- If a number is greater than 1, then all the zeros written to the right of the decimal point count as significant figures. Thus 2.0 mg has two significant figures, 40.062 mL has five significant figures, and 3.040 dm has four significant figures. If a number is less than 1, then only the zeros that are at the end of the number and the zeros that are between nonzero digits are significant. This means that 0.090 kg has two significant figures, 0.3005 L has four significant figures, 0.00420 min has three significant figures, and so on.
- For numbers that do not contain decimal points, the trailing zeros (that is, zeros after the last nonzero digit) may or may not be significant. Thus 400 cm may have one significant figure (the digit 4), two significant figures (40), or three significant figures (400). We cannot know which is correct without more information. By using scientific notation, however, we avoid this ambiguity. In this particular case, we can express the number 400 as 4×10^2 for one significant figure, 4.0×10^2 for two significant figures, or 4.00×10^2 for three significant figures.

The following example shows the determination of significant figures.

EXAMPLE 1.4

Determine the number of significant figures in the following measurements: (a) 478 cm, (b) 6.01 g, (c) 0.825 m, (d) 0.043 kg, (e) 1.310×10^{22} atoms, (f) 7000 mL.

Answer (a) Three, (b) Three, (c) Three, (d) Two, (e) Four, (f) This is an ambiguous case. The number of significant figures may be four (7.000×10^3), three (7.00×10^3), two (7.0×10^3), or one (7×10^3).

Similar problems: 1.33, 1.34.

PRACTICE EXERCISE

Determine the number of significant figures in each of the following measurements: (a) 24 mL, (b) 3001 g, (c) 0.0320 m³, (d) 6.4×10^4 molecules, (e) 560 kg.

A second set of rules specifies how to handle significant figures in calculations.

- In addition and subtraction, the number of significant figures to the right of the decimal point in the final sum or difference is determined by the *smallest* number of significant figures to the right of the decimal point in any of the original numbers. Consider these examples:

$$\begin{array}{r}
 89.332 \\
 + 1.1 \quad \leftarrow \text{one significant figure after the decimal point} \\
 \hline
 90.432 \quad \leftarrow \text{round off to 90.4}
 \end{array}$$

$$\begin{array}{r}
 2.097 \\
 - 0.12 \quad \leftarrow \text{two significant figures after the decimal point} \\
 \hline
 1.977 \quad \leftarrow \text{round off to 1.98}
 \end{array}$$

The rounding-off procedure is as follows. To round off a number at a certain point we simply drop the digits that follow if the first of them is less than 5. Thus 8.724 rounds off to 8.72 if we want only two figures after the decimal point. If the first digit following the point of rounding off is equal to or greater than 5, we add 1 to the preceding digit. Thus 8.727 rounds off to 8.73, and 0.425 rounds off to 0.43.

- In multiplication and division, the number of significant figures in the final product or quotient is determined by the original number that has the *smallest* number of significant figures. The following examples illustrate this rule:

$$2.8 \times 4.5039 = 12.61092 \leftarrow \text{round off to } 13$$

$$\frac{6.85}{112.04} = 0.0611388789 \leftarrow \text{round off to } 0.0611$$

- Keep in mind that *exact numbers* obtained from definitions or by counting numbers of objects can be considered to have an infinite number of significant figures. If an object has a mass of 0.2786 g, then the mass of eight such objects is

$$0.2786 \text{ g} \times 8 = 2.229 \text{ g}$$

We do *not* round off this product to one significant figure, because the number 8 is 8.00000 . . . , by definition. Similarly, to take the average of the two measured lengths 6.64 cm and 6.68 cm, we write

$$\frac{6.64 \text{ cm} + 6.68 \text{ cm}}{2} = 6.66 \text{ cm}$$

because the number 2 is 2.00000 . . . , by definition.

The following example shows how significant figures are handled in arithmetic operations.

EXAMPLE 1.5

Carry out the following arithmetic operations: (a) $11,254.1 \text{ g} + 0.1983 \text{ g}$, (b) $66.59 \text{ L} - 3.113 \text{ L}$, (c) $8.16 \text{ m} \times 5.1355$, (d) $0.0154 \text{ kg} \div 88.3 \text{ mL}$, (e) $2.64 \times 10^3 \text{ cm} + 3.27 \times 10^2 \text{ cm}$.

Answer

$$\begin{array}{r} \text{(a)} \quad 11,254.1 \text{ g} \\ + \quad 0.1983 \text{ g} \\ \hline 11,254.2983 \text{ g} \leftarrow \text{round off to } 11,254.3 \text{ g} \end{array}$$

$$\begin{array}{r} \text{(b)} \quad 66.59 \text{ L} \\ - \quad 3.113 \text{ L} \\ \hline 63.477 \text{ L} \leftarrow \text{round off to } 63.48 \text{ L} \end{array}$$

$$\text{(c)} \quad 8.16 \text{ m} \times 5.1355 = 41.90568 \text{ m} \leftarrow \text{round off to } 41.9 \text{ m}$$

$$\text{(d)} \quad \frac{0.0154 \text{ kg}}{88.3 \text{ mL}} = 0.000174405436 \text{ kg/mL} \leftarrow \text{round off to } 0.000174 \text{ kg/mL} \\ \text{or } 1.74 \times 10^{-4} \text{ kg/mL}$$

(e) First we change $3.27 \times 10^2 \text{ cm}$ to $0.327 \times 10^3 \text{ cm}$ and then carry out the addition $(2.64 \text{ cm} + 0.327 \text{ cm}) \times 10^3$. Following the procedure in (a), we find the answer is $2.97 \times 10^3 \text{ cm}$.

Similar problems: 1.35, 1.36.

PRACTICE EXERCISE

Carry out the following arithmetic operations and round off the answers to the appropriate number of significant figures: (a) $26.5862 \text{ L} + 0.17 \text{ L}$, (b) $9.1 \text{ g} - 4.682 \text{ g}$, (c) $7.1 \times 10^4 \text{ dm} \times 2.2654 \times 10^2$, (d) $6.54 \text{ g} \div 86.5542 \text{ mL}$, (e) $(7.55 \times 10^4 \text{ m}) - (8.62 \times 10^3 \text{ m})$.

The above rounding-off procedure applies to one-step calculations. In *chain calculations*, that is, calculations involving more than one step, we use a modified procedure. Consider the following two-step calculation:

$$\text{First step: } A \times B = C$$

$$\text{Second step: } C \times D = E$$

Let us suppose that $A = 3.66$, $B = 8.45$, and $D = 2.11$. Depending on whether we round off C to three or four significant figures, we obtain a different number for E :

Method 1

$$3.66 \times 8.45 = 30.9$$

$$30.9 \times 2.11 = 65.2$$

Method 2

$$3.66 \times 8.45 = 30.93$$

$$30.93 \times 2.11 = 65.3$$

However, if we had carried out the calculation as $3.66 \times 8.45 \times 2.11$ on a calculator without rounding off the intermediate result, we would have obtained 65.3 as the answer for E . In general, we will show the correct number of significant figures in each step of the calculation. However, in some worked examples, only the final answer is rounded off to the correct number of significant figures. The answers for all intermediate calculations will be carried to one extra figure.

Accuracy and Precision

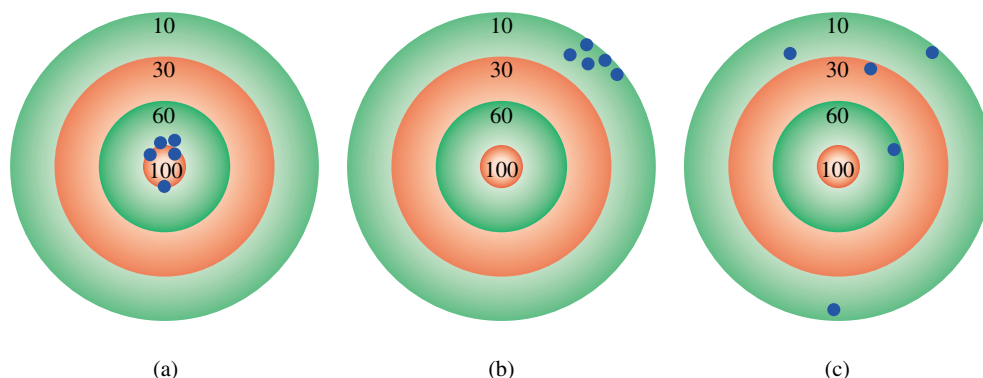
In discussing measurements and significant figures it is useful to distinguish between *accuracy* and *precision*. **Accuracy** tells us *how close a measurement is to the true value of the quantity that was measured*. To a scientist there is a distinction between accuracy and precision. **Precision** refers to *how closely two or more measurements of the same quantity agree with one another* (Figure 1.12).

The difference between accuracy and precision is a subtle but important one. Suppose, for example, that three students are asked to determine the mass of a piece of copper wire. The results of two successive weighings by each student are

| | STUDENT A | STUDENT B | STUDENT C |
|---------------|----------------|----------------|----------------|
| | 1.964 g | 1.972 g | 2.000 g |
| | <u>1.978 g</u> | <u>1.968 g</u> | <u>2.002 g</u> |
| Average value | 1.971 g | 1.970 g | 2.001 g |

The true mass of the wire is 2.000 g. Therefore, Student B's results are more *precise* than those of Student A (1.972 g and 1.968 g deviate less from 1.970 g than 1.964 g and 1.978 g from 1.971 g), but neither set of results is very *accurate*. Student C's results are not only the most *precise*, but also the most *accurate*, since the average value is closest to the true value. Highly accurate measurements are usually precise too. On the other hand, highly precise measurements do not necessarily guarantee accurate results. For example, an improperly calibrated meter stick or a faulty balance may give precise readings that are in error.

FIGURE 1.12 The distribution of darts on a dart board shows the difference between precise and accurate. (a) Good accuracy and good precision. (b) Poor accuracy and good precision. (c) Poor accuracy and poor precision.



1.9 THE FACTOR-LABEL METHOD OF SOLVING PROBLEMS

Careful measurements and the proper use of significant figures, along with correct calculations, will yield accurate numerical results. But to be meaningful, the answers also must be expressed in the desired units. The procedure we will use to convert between units in solving chemistry problems is called the *factor-label method*, or *dimensional analysis*. A simple technique requiring little memorization, the factor-label method is based on the relationship between different units that express the same physical quantity.

We know, for example, that the unit “dollar” for money is different from the unit “penny.” However, we say that 1 dollar is *equivalent* to 100 pennies because they both represent the same amount of money. This equivalence allows us to write

$$1 \text{ dollar} = 100 \text{ pennies}$$

Because 1 dollar is equal to 100 pennies, it follows that their ratio has a value of 1; that is,

$$\frac{1 \text{ dollar}}{100 \text{ pennies}} = 1$$

This ratio can be read as 1 dollar per 100 pennies. The fraction is called a *unit factor* (equal to 1) because the numerator and denominator describe the same amount of money.

We can also write the ratio as 100 pennies per dollar:

$$\frac{100 \text{ pennies}}{1 \text{ dollar}} = 1$$

This fraction is also a unit factor. We see that the reciprocal of any unit factor is also a unit factor. The usefulness of unit factors is that they allow us to carry out conversions between different units that measure the same quantity. Suppose that we want to convert 2.46 dollars into pennies. This problem may be expressed as

$$? \text{ pennies} = 2.46 \text{ dollars}$$

Since this is a dollar-to-penny conversion, we choose the unit factor that has the unit “dollar” in the denominator (to cancel the “dollars” in 2.46 dollars) and write

$$2.46 \cancel{\text{ dollars}} \times \frac{100 \text{ pennies}}{1 \cancel{\text{ dollar}}} = 246 \text{ pennies}$$

Note that the unit factor 100 pennies/1 dollar contains exact numbers, so it does not affect the number of significant figures in the final answer.

Next let us consider the conversion of 57.8 meters to centimeters. This problem may be expressed as

$$? \text{ cm} = 57.8 \text{ m}$$

By definition,

$$1 \text{ cm} = 1 \times 10^{-2} \text{ m}$$

Since we are converting “m” to “cm,” we choose the unit factor that has meters in the denominator,

$$\frac{1 \text{ cm}}{1 \times 10^{-2} \text{ m}} = 1$$

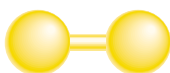
and write the conversion as

$$\begin{aligned} ? \text{ cm} &= 57.8 \cancel{\text{ m}} \times \frac{1 \text{ cm}}{1 \times 10^{-2} \cancel{\text{ m}}} \\ &= 5780 \text{ cm} \\ &= 5.78 \times 10^3 \text{ cm} \end{aligned}$$

Note that scientific notation is used to indicate that the answer has three significant figures. The unit factor $1 \text{ cm}/1 \times 10^{-2} \text{ m}$ contains exact numbers; therefore, it does not affect the number of significant figures.

In the factor-label method the units are carried through the entire sequence of calculations. Therefore, if the equation is set up correctly, then all the units will cancel except the desired one. If this is not the case, then an error must have been made somewhere, and it can usually be spotted by reviewing the solution.

The following examples illustrate the use of the factor-label method.



A hydrogen molecule.

EXAMPLE 1.6

The distance between two hydrogen atoms in a hydrogen molecule is 74 pm. Convert this distance to meters.

Answer The problem is

$$? \text{ m} = 74 \text{ pm}$$

By definition,

$$1 \text{ pm} = 1 \times 10^{-12} \text{ m}$$

The unit factor is

$$\frac{1 \times 10^{-12} \text{ m}}{1 \text{ pm}} = 1$$

Therefore we write

$$? \text{ m} = 74 \cancel{\text{ pm}} \times \frac{1 \times 10^{-12} \text{ m}}{1 \cancel{\text{ pm}}} = 7.4 \times 10^{-11} \text{ m}$$

Similar problem: 1.37(a).

PRACTICE EXERCISE

Convert 197 pm, the radius of a calcium (Ca) atom, to centimeters.

Conversion factors for some of the English system units commonly used in the United States for nonscientific measurements (for example, pounds and inches) are provided inside the back cover of this book.

Similar problem: 1.43.

EXAMPLE 1.7

A person's average daily intake of glucose (a form of sugar) is 0.0833 pound (lb). What is this mass in milligrams (mg)? (1 lb = 453.6 g)

Answer The problem can be expressed as

$$? \text{ mg} = 0.0833 \text{ lb}$$

so the unit factor is

$$\frac{453.6 \text{ g}}{1 \text{ lb}} = 1$$

and

$$1 \text{ mg} = 1 \times 10^{-3} \text{ g}$$

so we must also include the unit factor

$$\frac{1 \text{ mg}}{1 \times 10^{-3} \text{ g}} = 1$$

Thus

$$? \text{ mg} = 0.0833 \text{ lb} \times \frac{453.6 \cancel{\text{g}}}{1 \cancel{\text{lb}}} \times \frac{1 \text{ mg}}{1 \times 10^{-3} \cancel{\text{g}}} = 3.78 \times 10^4 \text{ mg}$$

PRACTICE EXERCISE

A roll of aluminum foil has a mass of 1.07 kg. What is its mass in pounds?

Note that unit factors may be squared or cubed, because $1^2 = 1^3 = 1$. The use of such factors is illustrated in Examples 1.8 and 1.9.

EXAMPLE 1.8

An average adult has 5.2 liters of blood. What is the volume of blood in m^3 ?

Answer Since $1 \text{ L} = 1000 \text{ cm}^3$, 5.2 L is equivalent to $5.2 \times 10^3 \text{ cm}^3$. The problem can be stated as

$$? \text{ m}^3 = 5.2 \times 10^3 \text{ cm}^3$$

By definition

$$1 \text{ m} = 100 \text{ cm}$$

The unit factor is

$$\frac{1 \text{ m}}{100 \text{ cm}} = 1$$

It follows that

$$\left(\frac{1 \text{ m}}{100 \text{ cm}} \right)^3 = 1^3 = 1$$

Therefore we write

$$? \text{ m}^3 = 5.2 \times 10^3 \text{ cm}^3 \times \left(\frac{1 \text{ m}}{100 \text{ cm}} \right)^3 = 5.2 \times 10^{-3} \text{ m}^3$$

Primordial Helium and The Big Bang Theory

Where did we come from? How did the universe begin? Humans have asked these questions for as long as we have been able to think. The search for answers provides an example of the scientific method.

In the 1940s the Russian-American physicist George Gamow hypothesized that our universe burst into being billions of years ago in a gigantic explosion, or *Big Bang*. In its earliest moments, the universe occupied a tiny volume and was unimaginably hot. This blistering fireball of radiation mixed with microscopic particles of matter gradually cooled enough for atoms to form. Under the influence of gravity, these atoms clumped together to make billions of galaxies including our own Milky Way Galaxy.

Gamow's idea is interesting and highly provocative. It has been tested experimentally in a number of ways. First, measurements showed that the universe is expanding; that is, galaxies are all moving away from one another at high speeds. This fact is consistent with the universe's explosive birth. By imagining the expansion running backwards, like a movie in reverse, astronomers have deduced that the universe was born about 15 billion years ago. The second observation that supports Gamow's hypothesis is the detection of *cosmic background radiation*. Over billions of years, the searingly hot universe has cooled down to a mere 3 K (or -270°C)! At this temperature, most energy is in the microwave region. Because the Big Bang would have occurred simultaneously throughout the tiny volume of the forming universe, the radiation it generated should have filled the entire universe. Thus the radiation should be the same in any direction that we observe. Indeed, the microwave signals recorded by astronomers are *independent* of direction.

The third piece of evidence supporting Gamow's hypothesis is the discovery of primordial helium. Scientists believe that helium and hydrogen (the lightest elements) were the first elements formed in the early stages of cosmic evolution. (The heavier elements, like carbon, nitrogen, and oxygen, are thought to have originated later via nuclear reactions involving hydrogen and helium in the center of stars.) If so, a dif-

fuse gas of hydrogen and helium would have spread through the early universe before much of the galaxies formed. In 1995 astronomers analyzed ultraviolet light from a distant *quasar* (a strong source of light and radio signals that is thought to be an exploding galaxy at the edge of the universe) and found that some of the light was absorbed by helium atoms on the way to Earth. Since this particular quasar is more than 10 billion light years away (a light year is the distance traveled by light in a year), the light reaching Earth reveals events that took place 10 billion years ago. Why wasn't the more abundant hydrogen detected? A hydrogen atom has only one electron, which is stripped by the light from a quasar in a process known as *ionization*. Ionized hydrogen atoms cannot absorb any of the quasar's light. A helium atom, on the other hand, has two electrons. Radiation may strip a helium atom of one electron, but not always both. Singly ionized helium atoms can still absorb light and are therefore detectable.

Proponents of Gamow's explanation rejoiced at the detection of helium in the far reaches of the universe. In recognition of all the supporting evidence, scientists now refer to Gamow's hypothesis as the Big Bang theory.

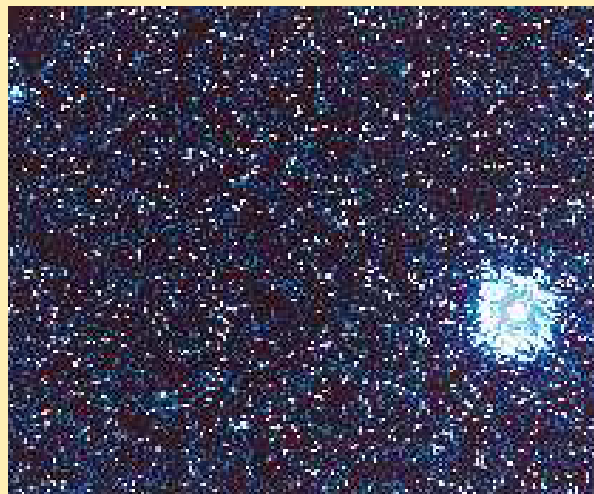


Photo showing some distant galaxy, including the position of a quasar.

Similar problem: 1.48(d).



A silver coin.

Comment Notice that in cubing the quantity $[1 \text{ m}/(100 \text{ cm})]$ we cube both the numbers and the units.

PRACTICE EXERCISE

The volume of a room is $1.08 \times 10^8 \text{ dm}^3$. What is the volume in m^3 ?

EXAMPLE 1.9

The density of silver is 10.5 g/cm^3 . Convert the density to units of kg/m^3 .

Answer The problem can be stated as

$$? \text{ kg/m}^3 = 10.5 \text{ g/cm}^3$$

We need two unit factors—one to convert g to kg and the other to convert cm^3 to m^3 . We know that

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

so

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

Second, since $1 \text{ cm} = 1 \times 10^{-2} \text{ m}$, the following unit factors can be generated:

$$\frac{1 \text{ cm}}{1 \times 10^{-2} \text{ m}} = 1 \quad \text{and} \quad \left(\frac{1 \text{ cm}}{1 \times 10^{-2} \text{ m}}\right)^3 = 1$$

Finally we can calculate the density in the desired units as follows:

$$\begin{aligned} ? \text{ kg/m}^3 &= \frac{10.5 \text{ g}}{1 \text{ cm}^3} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \left(\frac{1 \text{ cm}}{1 \times 10^{-2} \text{ m}}\right)^3 = 10,500 \text{ kg/m}^3 \\ &= 1.05 \times 10^4 \text{ kg/m}^3 \end{aligned}$$

Similar problem: 1.49.

Comment The units kg/m^3 give inconveniently large values for density.

PRACTICE EXERCISE

The density of the lightest metal, lithium (Li), is $5.34 \times 10^2 \text{ kg/m}^3$. Convert the density to g/cm^3 .

SUMMARY OF KEY EQUATIONS

- $d = \frac{m}{V}$ (1.1) Equation for density
- ${}^\circ\text{C} = ({}^\circ\text{F} - 32{}^\circ\text{F}) \times \frac{5{}^\circ\text{C}}{9{}^\circ\text{F}}$ (1.2) Converting ${}^\circ\text{F}$ to ${}^\circ\text{C}$
- ${}^\circ\text{F} = \frac{9{}^\circ\text{F}}{5{}^\circ\text{C}} \times ({}^\circ\text{C}) + 32{}^\circ\text{F}$ (1.3) Converting ${}^\circ\text{C}$ to ${}^\circ\text{F}$
- ${}^\circ\text{K} = ({}^\circ\text{C} + 273.15{}^\circ\text{C}) \frac{1 \text{ K}}{1{}^\circ\text{C}}$ (1.4) Converting ${}^\circ\text{C}$ to K

SUMMARY OF FACTS AND CONCEPTS

1. The study of chemistry involves three basic steps: observation, representation, and interpretation. Observation refers to measurements in the macroscopic world; representation involves the use of shorthand notation symbols and equations for communication; interpretations are based on atoms and molecules, which belong to the microscopic world.
2. The scientific method is a systematic approach to research that begins with the gathering of information through observation and measurements. In the process, hypotheses, laws, and theories are devised and tested.
3. Chemists study matter and the changes it undergoes. The substances that make up matter have unique physical properties that can be observed without changing their identity and unique chemical properties that, when they are demonstrated, do change the identity of the substances. Mixtures, whether homogeneous or heterogeneous, can be separated into pure components by physical means.
4. The simplest substances in chemistry are elements. Compounds are formed by the chemical combination of atoms of different elements in fixed proportions.
5. All substances, in principle, can exist in three states: solid, liquid, and gas. The interconversion between these states can be effected by changing the temperature.
6. SI units are used to express physical quantities in all sciences, including chemistry.
7. Numbers expressed in scientific notation have the form $N \times 10^n$, where N is between 1 and 10, and n is a positive or negative integer. Scientific notation helps us handle very large and very small quantities.

KEY WORDS

| | | | |
|------------------------------|---|-----------------------------|----------------------------|
| Accuracy, p. 24 | Homogeneous mixture, p. 10 | Macroscopic property, p. 14 | Quantitative, p. 8 |
| Chemical property, p. 13 | Hypothesis, p. 9 | Mass, p. 14 | Scientific method, p. 8 |
| Chemistry, p. 4 | Intensive property, p. 14 | Matter, p. 10 | Significant figures, p. 21 |
| Compound, p. 11 | International System of Units (SI), p. 15 | Microscopic property, p. 14 | Substance, p. 10 |
| Density, p. 14 | Kelvin, p. 17 | Mixture, p. 10 | Theory, p. 9 |
| Element, p. 11 | Law, p. 9 | Physical property, p. 13 | Volume, p. 14 |
| Extensive property, p. 14 | Liter, p. 16 | Precision, p. 24 | Weight, p. 15 |
| Heterogeneous mixture, p. 10 | | Qualitative, p. 8 | |

QUESTIONS AND PROBLEMS

THE SCIENTIFIC METHOD

Review Questions

- 1.1 Explain what is meant by the scientific method.
- 1.2 What is the difference between qualitative data and quantitative data?

Problems

- 1.3 Classify the following as qualitative or quantitative statements, giving your reasons. (a) The sun is approximately 93 million miles from Earth. (b) Leonardo da Vinci was a better painter than Michelangelo. (c) Ice is less dense than water. (d) Butter tastes better than margarine. (e) A stitch in time saves nine.
- 1.4 Classify each of the following statements as a hypothesis, a law, or a theory. (a) Beethoven's contribu-

tion to music would have been much greater if he had married. (b) An autumn leaf gravitates toward the ground because there is an attractive force between the leaf and Earth. (c) All matter is composed of very small particles called atoms.

CLASSIFICATION AND PROPERTIES OF MATTER

Review Questions

- 1.5 Give an example for each of the following terms: (a) matter, (b) substance, (c) mixture.
- 1.6 Give an example of a homogeneous mixture and an example of a heterogeneous mixture.
- 1.7 Using examples, explain the difference between a physical property and a chemical property?
- 1.8 How does an intensive property differ from an extensive property? Which of the following properties are

intensive and which are extensive? (a) length, (b) volume, (c) temperature, (d) mass.

- 1.9 Give an example of an element and a compound. How do elements and compounds differ?
- 1.10 What is the number of known elements?

Problems

- 1.11 Do the following statements describe chemical or physical properties? (a) Oxygen gas supports combustion. (b) Fertilizers help to increase agricultural production. (c) Water boils below 100°C on top of a mountain. (d) Lead is denser than aluminum. (e) Sugar tastes sweet.
- 1.12 Does each of the following describe a physical change or a chemical change? (a) The helium gas inside a balloon tends to leak out after a few hours. (b) A flashlight beam slowly gets dimmer and finally goes out. (c) Frozen orange juice is reconstituted by adding water to it. (d) The growth of plants depends on the sun's energy in a process called photosynthesis. (e) A spoonful of table salt dissolves in a bowl of soup.
- 1.13 Give the names of the elements represented by the chemical symbols Li, F, P, Cu, As, Zn, Cl, Pt, Mg, U, Al, Si, Ne. (See Table 1.1 and the inside front cover.)
- 1.14 Give the chemical symbols for the following elements: (a) potassium, (b) tin, (c) chromium, (d) boron, (e) barium, (f) plutonium, (g) sulfur, (h) argon, (i) mercury. (See Table 1.1 and the inside front cover.)
- 1.15 Classify each of the following substances as an element or a compound: (a) hydrogen, (b) water, (c) gold, (d) sugar.
- 1.16 Classify each of the following as an element, a compound, a homogeneous mixture, or a heterogeneous mixture: (a) seawater, (b) helium gas, (c) sodium chloride (table salt), (d) a bottle of soft drink, (e) a milkshake, (f) air, (g) concrete.

MEASUREMENT

Review Questions

- 1.17 Name the SI base units that are important in chemistry. Give the SI units for expressing the following: (a) length, (b) volume, (c) mass, (d) time, (e) energy, (f) temperature.
- 1.18 Write the numbers represented by the following prefixes: (a) mega-, (b) kilo-, (c) deci-, (d) centi-, (e) milli-, (f) micro-, (g) nano-, (h) pico-.
- 1.19 What units do chemists normally use for density of liquids and solids? For gas density? Explain the difference?
- 1.20 Describe the three temperature scales used in the laboratory and in every day life: the Fahrenheit scale, the Celsius scale, and the Kelvin scale.

Problems

- 1.21 Bromine is a reddish-brown liquid. Calculate its density (in g/mL) if 586 g of the substance occupies 188 mL.
- 1.22 Mercury is the only metal that is a liquid at room temperature. Its density is 13.6 g/mL . How many grams of mercury will occupy a volume of 95.8 mL?
- 1.23 Convert the following temperatures to degrees Celsius: (a) 95°F , the temperature on a hot summer day; (b) 12°F , the temperature on a cold winter day; (c) a 102°F fever; (d) a furnace operating at 1852°F .
- 1.24 (a) Normally the human body can endure a temperature of 105°F for only short periods of time without permanent damage to the brain and other vital organs. What is this temperature in degrees Celsius? (b) Ethylene glycol is a liquid organic compound that is used as an antifreeze in car radiators. It freezes at -11.5°C . Calculate its freezing temperature in degrees Fahrenheit. (c) The temperature on the surface of the sun is about 6300°C . What is this temperature in degrees Fahrenheit?
- 1.25 Convert the following temperatures to Kelvin: (a) 113°C , the melting point of sulfur, (b) 37°C , the normal body temperature, (c) 357°C , the boiling point of mercury.
- 1.26 Convert the following temperatures to degrees Celsius: (a) 77 K, the boiling point of liquid nitrogen, (b) 4.2 K, the boiling point of liquid helium, (c) 601 K, the melting point of lead.

HANDLING NUMBERS

Review Questions

- 1.27 What is the advantage of using scientific notation over decimal notation?
- 1.28 Define significant figure. Discuss the importance of using the proper number of significant figures in measurements and calculations.

Problems

- 1.29 Express the following numbers in scientific notation: (a) 0.000000027, (b) 356, (c) 47,764, (d) 0.096.
- 1.30 Express the following numbers as decimals: (a) 1.52×10^{-2} , (b) 7.78×10^{-8} .
- 1.31 Express the answers to the following calculations in scientific notation:
(a) $145.75 + (2.3 \times 10^{-1})$
(b) $79,500 \div (2.5 \times 10^2)$
(c) $(7.0 \times 10^{-3}) - (8.0 \times 10^{-4})$
(d) $(1.0 \times 10^4) \times (9.9 \times 10^6)$
- 1.32 Express the answers to the following calculations in scientific notation:
(a) $0.0095 + (8.5 \times 10^{-3})$

- (b) $653 \div (5.75 \times 10^{-8})$
 (c) $850,000 - (9.0 \times 10^5)$
 (d) $(3.6 \times 10^{-4}) \times (3.6 \times 10^6)$
- 1.33** What is the number of significant figures in each of the following measurements?
 (a) 4867 mi
 (b) 56 mL
 (c) 60,104 ton
 (d) 2900 g
 (e) 40.2 g/cm^3
 (f) 0.0000003 cm
 (g) 0.7 min
 (h) 4.6×10^{19} atoms
- 1.34** How many significant figures are there in each of the following? (a) 0.006 L, (b) 0.0605 dm, (c) 60.5 mg, (d) 605.5 cm^2 , (e) 960×10^{-3} g, (f) 6 kg, (g) 60 m.
- 1.35** Carry out the following operations as if they were calculations of experimental results, and express each answer in the correct units with the correct number of significant figures:
 (a) $5.6792 \text{ m} + 0.6 \text{ m} + 4.33 \text{ m}$
 (b) $3.70 \text{ g} - 2.9133 \text{ g}$
 (c) $4.51 \text{ cm} \times 3.6666 \text{ cm}$
- 1.36** Carry out the following operations as if they were calculations of experimental results, and express each answer in the correct units with the correct number of significant figures:
 (a) $7.310 \text{ km} \div 5.70 \text{ km}$
 (b) $(3.26 \times 10^{-3} \text{ mg}) - (7.88 \times 10^{-5} \text{ mg})$
 (c) $(4.02 \times 10^6 \text{ dm}) + (7.74 \times 10^7 \text{ dm})$

THE FACTOR-LABEL METHOD

Problems

- 1.37** Carry out the following conversions: (a) 22.6 m to decimeters, (b) 25.4 mg to kilograms.
- 1.38** Carry out the following conversions: (a) 242 lb to milligrams, (b) 68.3 cm^3 to cubic meters.
- 1.39** The price of gold on November 3, 1995, was \$384 per ounce. How much did 1.00 g of gold cost that day? (1 ounce = 28.4 g.)
- 1.40** How many seconds are there in a solar year (365.24 days)?
- 1.41** How many minutes does it take light from the sun to reach Earth? (The distance from the sun to Earth is 93 million mi; the speed of light = 3.00×10^8 m/s.)
- 1.42** A slow jogger runs a mile in 13 min. Calculate the speed in (a) in/s, (b) m/min, (c) km/h. (1 mi = 1609 m; 1 in = 2.54 cm.)
- 1.43** A 6.0-ft person weighs 168 lb. Express this person's height in meters and weight in kilograms. (1 lb = 453.6 g; 1 m = 3.28 ft.)
- 1.44** The current speed limit in some states in the U.S. is 55 miles per hour. What is the speed limit in kilometers per hour? (1 mi = 1609 m.)
- 1.45** For a fighter jet to take off from the deck of an aircraft carrier, it must reach a speed of 62 m/s. Calculate the speed in mph.
- 1.46** The "normal" lead content in human blood is about 0.40 part per million (that is, 0.40 g of lead per million grams of blood). A value of 0.80 part per million (ppm) is considered to be dangerous. How many grams of lead are contained in 6.0×10^3 g of blood (the amount in an average adult) if the lead content is 0.62 ppm?
- 1.47** Carry out the following conversions: (a) 1.42 light-years to miles (a light-year is an astronomical measure of distance—the distance traveled by light in a year, or 365 days; the speed of light is 3.00×10^8 m/s), (b) 32.4 yd to centimeters, (c) 3.0×10^{10} cm/s to ft/s.
- 1.48** Carry out the following conversions: (a) 47.4°F to degrees Celsius, (b) -273.15°C (the lowest attainable temperature) to degrees Fahrenheit, (c) 71.2 cm^3 to m^3 , (d) 7.2 m^3 to liters.
- 1.49** Aluminum is a lightweight metal (density = 2.70 g/cm^3) used in aircraft construction, high-voltage transmission lines, beverage cans, and foils. What is its density in kg/m^3 ?
- 1.50** The density of ammonia gas under certain conditions is 0.625 g/L. Calculate its density in g/cm^3 .

ADDITIONAL PROBLEMS

- 1.51** Give one qualitative and one quantitative statement about each of the following: (a) water, (b) carbon, (c) iron, (d) hydrogen gas, (e) sucrose (cane sugar), (f) table salt (sodium chloride), (g) mercury, (h) gold, (i) air.
- 1.52** Which of the following statements describe physical properties and which describe chemical properties? (a) Iron has a tendency to rust. (b) Rainwater in industrialized regions tends to be acidic. (c) Hemoglobin molecules have a red color. (d) When a glass of water is left out in the sun, the water gradually disappears. (e) Carbon dioxide in air is converted to more complex molecules by plants during photosynthesis.
- 1.53** In 1995, 95.4 billion pounds of sulfuric acid were produced in the United States. Convert this quantity to tons.
- 1.54** In determining the density of a rectangular metal bar, a student made the following measurements: length, 8.53 cm; width, 2.4 cm; height, 1.0 cm; mass, 52.7064 g. Calculate the density of the metal to the correct number of significant figures.
- 1.55** Calculate the mass of each of the following: (a) a sphere of gold with a radius of 10.0 cm [the volume

of a sphere with a radius r is $V = (4/3)\pi r^3$; the density of gold = 19.3 g/cm^3], (b) a cube of platinum of edge length 0.040 mm (the density of platinum = 21.4 g/cm^3), (c) 50.0 mL of ethanol (the density of ethanol = 0.798 g/mL).

- 1.56** A cylindrical glass tube 12.7 cm in length is filled with mercury. The mass of mercury needed to fill the tube is 105.5 g . Calculate the inner diameter of the tube. (The density of mercury = 13.6 g/mL .)
- 1.57** The following procedure was used to determine the volume of a flask. The flask was weighed dry and then filled with water. If the masses of the empty flask and filled flask were 56.12 g and 87.39 g , respectively, and the density of water is 0.9976 g/cm^3 , calculate the volume of the flask in cm^3 .
- 1.58** The speed of sound in air at room temperature is about 343 m/s . Calculate this speed in miles per hour (mph). ($1 \text{ mi} = 1609 \text{ m}$.)
- 1.59** A piece of silver (Ag) metal weighing 194.3 g is placed in a graduated cylinder containing 242.0 mL of water. The volume of water now reads 260.5 mL . From these data calculate the density of silver.
- 1.60** The experiment described in Problem 1.59 is a crude but convenient way to determine the density of some solids. Describe a similar experiment that would allow you to measure the density of ice. Specifically, what would be the requirements for the liquid used in your experiment?
- 1.61** A lead sphere has a mass of $1.20 \times 10^4 \text{ g}$, and its volume is $1.05 \times 10^3 \text{ cm}^3$. Calculate the density of lead.
- 1.62** Lithium is the least dense metal known (density: 0.53 g/cm^3). What is the volume occupied by $1.20 \times 10^3 \text{ g}$ of lithium?
- 1.63** The medicinal thermometer commonly used in homes can be read $\pm 0.1^\circ\text{F}$, while those in the doctor's office may be accurate to $\pm 0.1^\circ\text{C}$. In degrees Celsius, express the percent error expected from each of these thermometers in measuring a person's body temperature of 38.9°C .
- 1.64** Vanillin (used to flavor vanilla ice cream and other foods) is the substance whose aroma the human nose detects in the smallest amount. The threshold limit is $2.0 \times 10^{-11} \text{ g}$ per liter of air. If the current price of 50 g of vanillin is $\$112$, determine the cost to supply enough vanillin so that the aroma could be detected in a large aircraft hangar with a volume of $5.0 \times 10^7 \text{ ft}^3$.
- 1.65** At what temperature does the numerical reading on a Celsius thermometer equal that on a Fahrenheit thermometer?
- 1.66** Suppose that a new temperature scale has been devised on which the melting point of ethanol (-117.3°C) and the boiling point of ethanol (78.3°C) are taken as 0°S

and 100°S , respectively, where S is the symbol for the new temperature scale. Derive an equation relating a reading on this scale to a reading on the Celsius scale. What would this thermometer read at 25°C ?

- 1.67** A resting adult requires about 240 mL of pure oxygen/min and breathes about 12 times every minute. If inhaled air contains 20 percent oxygen by volume and exhaled air 16 percent, what is the volume of air per breath? (Assume that the volume of inhaled air is equal to that of exhaled air.)
- 1.68** (a) Referring to Problem 1.67, calculate the total volume (in liters) of air an adult breathes in a day. (b) In a city with heavy traffic, the air contains $2.1 \times 10^{-6} \text{ L}$ of carbon monoxide (a poisonous gas) per liter. Calculate the average daily intake of carbon monoxide in liters by a person.
- 1.69** The total volume of seawater is $1.5 \times 10^{21} \text{ L}$. Assume that seawater contains 3.1 percent sodium chloride by mass and that its density is 1.03 g/mL . Calculate the total mass of sodium chloride in kilograms and in tons. ($1 \text{ ton} = 2000 \text{ lb}$; $1 \text{ lb} = 453.6 \text{ g}$)
- 1.70** Magnesium (Mg) is a valuable metal used in alloys, in batteries, and in the manufacture of chemicals. It is obtained mostly from seawater, which contains about 1.3 g of Mg for every kilogram of seawater. Referring to Problem 1.69, calculate the volume of seawater (in liters) needed to extract $8.0 \times 10^4 \text{ tons}$ of Mg, which is roughly the annual production in the United States.
- 1.71** A student is given a crucible and asked to prove whether it is made of pure platinum. She first weighs the crucible in air and then weighs it suspended in water (density = 0.9986 g/mL). The readings are 860.2 g and 820.2 g , respectively. Based on these measurements and given that the density of platinum is 21.45 g/cm^3 , what should her conclusion be? (*Hint:* An object suspended in a fluid is buoyed up by the mass of the fluid displaced by the object. Neglect the buoyance of air.)
- 1.72** The surface area and average depth of the Pacific Ocean are $1.8 \times 10^8 \text{ km}^2$ and $3.9 \times 10^3 \text{ m}$, respectively. Calculate the volume of water in the ocean in liters.
- 1.73** The unit "troy ounce" is often used for precious metals such as gold (Au) and platinum (Pt). ($1 \text{ troy ounce} = 31.103 \text{ g}$) (a) A gold coin weighs 2.41 troy ounces . Calculate its mass in grams. (b) Is a troy ounce heavier or lighter than an ounce. ($1 \text{ lb} = 16 \text{ oz}$; $1 \text{ lb} = 453.6 \text{ g}$)
- 1.74** Osmium (Os) is the densest element known (density = 22.57 g/cm^3). Calculate the mass in pounds and in kilograms of an Os sphere 15 cm in diameter (about the size of a grapefruit). See Problem 1.55 for volume of a sphere.

- 1.75** Percent error is often expressed as the absolute value of the difference between the true value and the experimental value, divided by the true value:

$$\text{percent error} = \frac{|\text{true value} - \text{experimental value}|}{|\text{true value}|} \times 100\%$$

The vertical lines indicate absolute value. Calculate the percent error for the following measurements: (a) The density of alcohol (ethanol) is found to be 0.802 g/mL. (True value: 0.798 g/mL.) (b) The mass of gold in an earring is analyzed to be 0.837 g. (True value: 0.864 g.)

- 1.76** The natural abundances of elements in the human body, expressed as percent by mass, are: oxygen (O), 65%; carbon (C), 18%; hydrogen (H), 10%; nitrogen (N), 3%; calcium (Ca), 1.6%; phosphorus (P), 1.2%; all other elements, 1.2%. Calculate the mass in grams of each element in the body of a 62-kg person.
- 1.77** The men's world record for running a mile outdoors (as of 1997) is 3 minutes 44.39 seconds. At this rate, how long would it take to run a 1500-m race? (1 mi = 1609 m.)
- 1.78** Venus, the second closest planet to the sun, has a surface temperature of 7.3×10^2 K. Convert this temperature to °C and °F.
- 1.79** Chalcopyrite, the principal ore of copper (Cu), contains 34.63% Cu by mass. How many grams of Cu can be obtained from 5.11×10^3 kg of the ore?
- 1.80** It has been estimated that 8.0×10^4 tons of gold (Au) have been mined. Assume gold costs \$350 per ounce. What is the total worth of this quantity of gold?
- 1.81** A 1.0-mL volume of seawater contains about 4.0×10^{-12} g of gold. The total volume of ocean water is 1.5×10^{21} L. Calculate the total amount of gold (in grams) that is present in seawater, and the worth of the gold in dollars (see Problem 1.80). With so much gold out there, why hasn't someone become rich by mining gold from the ocean?
- 1.82** Measurements show that 1.0 g of iron (Fe) contains 1.1×10^{22} Fe atoms. How many Fe atoms are in 4.9 g of Fe, which is the total amount of iron in the body of an average adult?
- 1.83** The thin outer layer of Earth, called the crust, contains only 0.50% of Earth's total mass and yet is the source of almost all the elements (the atmosphere provides elements such as oxygen, nitrogen, and a few other gases). Silicon (Si) is the second most abundant element in Earth's crust (27.2% by mass). Calculate the mass of silicon in kilograms in Earth's crust. (The mass of Earth is 5.9×10^{21} tons. 1 ton = 2000 lb; 1 lb = 453.6 g.)
- 1.84** The diameter of a copper (Cu) atom is roughly 1.3×10^{-12} m. How many times can you divide evenly a piece of 10-cm copper wire until it is reduced to two separate copper atoms? (Assume there are appropriate tools for this procedure and that copper atoms are lined up in a straight line, in contact with each other.) Round off your answer to an integer.)
- 1.85** One gallon of gasoline in an automobile's engine produces on the average 9.5 kg of carbon dioxide, which is a greenhouse gas, that is, it promotes the warming of Earth's atmosphere. Calculate the annual production of carbon dioxide in kilograms if there are 40 million cars in the United States and each car covers a distance of 5000 miles at a consumption rate of 20 miles per gallon.
- 1.86** A sheet of aluminum (Al) foil has a total area of 1.000 ft² and a mass of 3.636 g. What is the thickness of the foil in millimeters? (Density of Al = 2.699 g/cm³.)
- 1.87** Comment on whether each of the following is a homogeneous mixture or a heterogeneous mixture: (a) air in a closed bottle and (b) air over New York City.
- 1.88** It has been proposed that dinosaurs and many other organisms became extinct 65 million years ago because Earth was struck by a large asteroid. The idea is that dust from the impact was lofted into the upper atmosphere all around the globe, where it lingered for at least several months and blocked the sunlight from reaching Earth's surface. In the dark and cold conditions that temporarily resulted, many forms of life became extinct. Available evidence suggests that about 20% of the asteroid's mass turned to dust and spread uniformly over Earth after eventually settling out of the upper atmosphere. This dust amounted to about 0.02 g/cm² of Earth's surface. The asteroid very likely had a density of about 2 g/cm³. Calculate the mass (in kilograms and tons) of the asteroid and its radius in meters, assuming that it was a sphere. (The area of Earth is 5.1×10^{14} m²; 1 lb = 453.6 g.) (Source: *Consider a Spherical Cow—A Course in Environmental Problem Solving* by J. Harte, University Science Books, Mill Valley, CA, 1988. Used with permission.)
- 1.89** The world's total petroleum reserve is estimated at 2.0×10^{22} J (Joule is the unit of energy where 1 J = 1 kg m²/s²). At the present rate of consumption, 1.8×10^{20} J/yr, how long would it take to exhaust the supply?
- 1.90** Chlorine is used to disinfect swimming pools. The accepted concentration for this purpose is 1 ppm chlorine, or one gram of chlorine per million grams of water. Calculate the volume of a chlorine solution (in milliliters) a homeowner should add to her swimming pool if the solution contains 6.0% chlorine by mass and there are 2.0×10^4 gallons of water in the pool. (1 gallon = 3.79 L; density of liquids = 1.0 g/mL.)

1.91 Fluoridation is the process of adding fluorine compounds to drinking water to help fight tooth decay. A concentration of 1 ppm of fluorine is sufficient for the purpose. (1 ppm means one part per million, or 1 g of fluorine per one million grams of water.) The compound normally chosen for fluoridation is sodium fluoride, which is also added to some toothpastes. Calculate the quantity of sodium fluoride in kilograms needed per year for a city of 50,000 people if the daily consumption of water per person is 150 gallons. What percent of the sodium fluoride is “wasted” if each person uses only 6.0 L of water a day for drinking and cooking? (Sodium fluoride is 45.0% fluorine by mass. 1 gallon = 3.79 L; 1 year = 365 days; 1 ton = 2000 lb; 1 lb = 453.6 g; density of water = 1.0 g/mL.)

1.92 In water conservation, chemists spread a thin film of certain inert material over the surface of water to cut down the rate of evaporation of water in reservoirs. This technique was pioneered by Benjamin Franklin two centuries ago. Franklin found that 0.10 mL of oil could spread over the surface of about 40 m² of water. Assuming that the oil forms a *monolayer*, that is,

a layer that is only one molecule thick, estimate the length of each oil molecule in nanometers. (1 nm = 1×10^{-9} m.)

1.93 Pheromones are compounds secreted by females of many insect species to attract mates. Typically, 1.0×10^{-8} g of a pheromone is sufficient to reach all targeted males within a radius of 0.50 mi. Calculate the density of the pheromone (in grams per liter) in a cylindrical air space having a radius of 0.50 mi and a height of 40 ft.

1.94 A gas company in Massachusetts charges \$1.30 for 15.0 ft³ of natural gas. (a) Convert this rate to dollars per liter of gas. (b) It takes 0.304 ft³ of gas to boil a liter of water, starting at room temperature (25°C), how much would it cost to boil a 2.1-liter kettle of water?

Answers to Practice Exercises: **1.1** 96.5 g. **1.2** 341 g. **1.3** (a) 621.5°F, (b) 78.3°C, (c) -196°C. **1.4** (a) Two, (b) four, (c) three, (d) two, (e) three or two. **1.5** (a) 26.76 L, (b) 4.4 g, (c) 1.6×10^7 dm, (d) 0.0756 g/mL, (e) 6.69×10^4 m. **1.6** 1.97×10^{-8} cm. **1.7** 2.36 lb. **1.8** 1.08×10^5 m³. **1.9** 0.534 g/cm³.

CHAPTER
2

Atoms, Molecules, and Ions

INTRODUCTION

SINCE ANCIENT TIMES HUMANS HAVE PONDERED THE NATURE OF MATTER. OUR MODERN IDEAS OF THE STRUCTURE OF MATTER BEGAN TO TAKE SHAPE IN THE EARLY NINETEENTH CENTURY WITH DALTON'S ATOMIC THEORY. WE NOW KNOW THAT ALL MATTER IS MADE OF ATOMS, MOLECULES, AND IONS. ALL OF CHEMISTRY IS CONCERNED IN ONE WAY OR ANOTHER WITH THESE SPECIES.

- 2.1** THE ATOMIC THEORY
- 2.2** THE STRUCTURE OF THE ATOM
- 2.3** ATOMIC NUMBER, MASS NUMBER, AND ISOTOPES
- 2.4** THE PERIODIC TABLE
- 2.5** MOLECULES AND IONS
- 2.6** CHEMICAL FORMULAS
- 2.7** NAMING COMPOUNDS



2.1 THE ATOMIC THEORY

In the fifth century B.C. the Greek philosopher Democritus expressed the belief that all matter consists of very small, indivisible particles, which he named *atomos* (meaning uncuttable or indivisible). Although Democritus' idea was not accepted by many of his contemporaries (notably Plato and Aristotle), somehow it endured. Experimental evidence from early scientific investigations provided support for the notion of “atomism” and gradually gave rise to the modern definitions of elements and compounds. It was in 1808 that an English scientist and school teacher, John Dalton,[†] formulated a precise definition of the indivisible building blocks of matter that we call atoms.

Dalton's work marked the beginning of the modern era of chemistry. The hypotheses about the nature of matter on which Dalton's atomic theory is based can be summarized as follows:

1. Elements are composed of extremely small particles called atoms. All atoms of a given element are identical, having the same size, mass, and chemical properties. The atoms of one element are different from the atoms of all other elements.
2. Compounds are composed of atoms of more than one element. In any compound, the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.
3. A chemical reaction involves only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction.

Figure 2.1 is a schematic representation of the first two hypotheses.

Dalton's concept of an atom was far more detailed and specific than Democritus'. The first hypothesis states that atoms of one element are different from atoms of all other elements. Dalton made no attempt to describe the structure or composition of atoms—he had no idea what an atom is really like. But he did realize that the different properties shown by elements such as hydrogen and oxygen can be explained by assuming that hydrogen atoms are not the same as oxygen atoms.

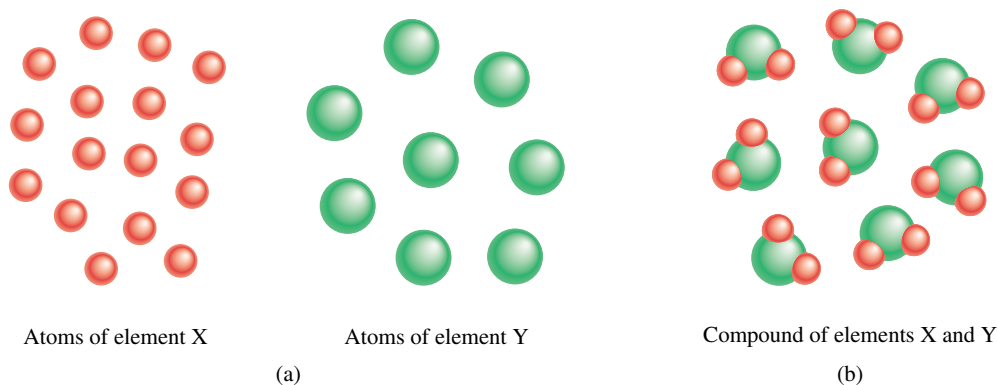
The second hypothesis suggests that, in order to form a certain compound, we need not only atoms of the right kinds of elements, but specific numbers of these atoms as well. This idea is an extension of a law published in 1799 by Joseph Proust,[‡] a French chemist. Proust's *law of definite proportions* states that *different samples of the same compound always contain its constituent elements in the same proportion by mass*. Thus, if we were to analyze samples of carbon dioxide gas obtained from different sources, we would find in each sample the same ratio by mass of carbon to oxygen. It stands to reason, then, that if the ratio of the masses of different elements in a given compound is fixed, the ratio of the atoms of these elements in the compound also must be constant.

Dalton's second hypothesis supports another important law, the *law of multiple proportions*. According to the law, *if two elements can combine to form more than one*

[†]John Dalton (1766–1844). English chemist, mathematician, and philosopher. In addition to the atomic theory, he also formulated several gas laws and gave the first detailed description of color blindness, from which he suffered. Dalton was described as an indifferent experimenter, and singularly wanting in the language and power of illustration. His only recreation was lawn bowling on Thursday afternoons. Perhaps it was the sight of those wooden balls that provided him with the idea of the atomic theory.

[‡]Joseph Louis Proust (1754–1826). French chemist. Proust was the first person to isolate sugar from grapes.

FIGURE 2.1 (a) According to Dalton's atomic theory, atoms of the same element are identical, but atoms of one element are different from atoms of other elements. (b) Compound formed from atoms of elements X and Y. In this case, the ratio of the atoms of element X to the atoms of element Y is 2:1.



compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers. Dalton's theory explains the law of multiple proportions quite simply: Different compounds made up of the same elements differ in the number of atoms of each kind that combine. For example, carbon forms two stable compounds with oxygen, namely, carbon monoxide and carbon dioxide. Modern measurement techniques indicate that one atom of carbon combines with one atom of oxygen in carbon monoxide and with two atoms of oxygen in carbon dioxide. Thus, the ratio of oxygen in carbon monoxide to oxygen in carbon dioxide is 1:2. This result is consistent with the law of multiple proportions.

Dalton's third hypothesis is another way of stating the *law of conservation of mass*, which is that *matter can be neither created nor destroyed*. Since matter is made of atoms that are unchanged in a chemical reaction, it follows that mass must be conserved as well. Dalton's brilliant insight into the nature of matter was the main stimulus for the rapid progress of chemistry during the nineteenth century.

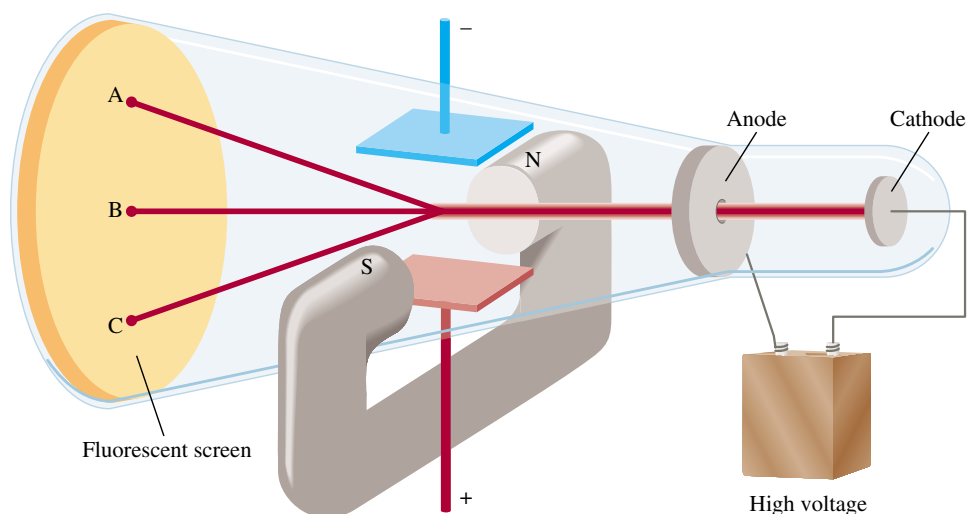
2.2 THE STRUCTURE OF THE ATOM

On the basis of Dalton's atomic theory, we can define an *atom* as *the basic unit of an element that can enter into chemical combination*. Dalton imagined an atom that was both extremely small and indivisible. However, a series of investigations that began in the 1850s and extended into the twentieth century clearly demonstrated that atoms actually possess internal structure; that is, they are made up of even smaller particles, which are called *subatomic particles*. This research led to the discovery of three such particles—electrons, protons, and neutrons.

THE ELECTRON

In the 1890s many scientists became caught up in the study of *radiation*, *the emission and transmission of energy through space in the form of waves*. Information gained from this research contributed greatly to our understanding of atomic structure. One device used to investigate this phenomenon was a cathode ray tube, the forerunner of the television tube (Figure 2.2). It is a glass tube from which most of the air has been evacuated. When the two metal plates are connected to a high-voltage source, the negatively charged plate, called the *cathode*, emits an invisible ray. The cathode ray is

FIGURE 2.2 A cathode ray tube with an electric field perpendicular to the direction of the cathode rays and an external magnetic field. The symbols *N* and *S* denote the north and south poles of the magnet. The cathode rays will strike the end of the tube at *A* in the presence of a magnetic field, at *C* in the presence of an electric field, and at *B* when there are no external fields present or when the effects of the electric field and magnetic field cancel each other.



drawn to the positively charged plate, called the *anode*, where it passes through a hole and continues traveling to the other end of the tube. When the ray strikes the specially coated surface, it produces a strong fluorescence, or bright light.

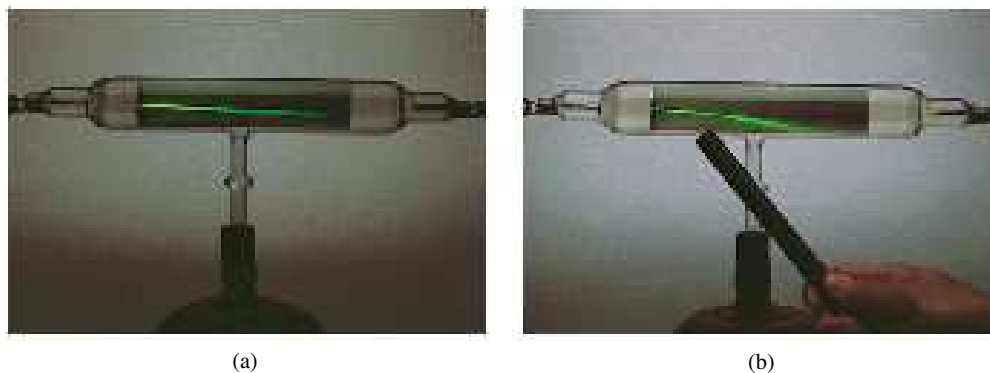
In some experiments two electrically charged plates and a magnet were added to the *outside* of the cathode ray tube (see Figure 2.2). When the magnetic field is on and the electric field is off, the cathode ray strikes point *A*. When only the electric field is on, the ray strikes point *C*. When both the magnetic and the electric fields are off or when they are both on but balanced so that they cancel each other's influence, the ray strikes point *B*. According to electromagnetic theory, a moving charged body behaves like a magnet and can interact with electric and magnetic fields through which it passes. Since the cathode ray is attracted by the plate bearing positive charges and repelled by the plate bearing negative charges, it must consist of negatively charged particles. We know these *negatively charged particles* as **electrons**. Figure 2.3 shows the effect of a bar magnet on the cathode ray.

An English physicist, J. J. Thomson,[†] used a cathode ray tube and his knowledge of electromagnetic theory to determine the ratio of electric charge to the mass of an

[†]Joseph John Thomson (1856–1940). British physicist who received the Nobel Prize in Physics in 1906 for discovering the electron.

Electrons are normally associated with atoms. However, they can also be studied individually.

FIGURE 2.3 (a) A cathode ray produced in a discharge tube. The ray itself is invisible, but the fluorescence of a zinc sulfide coating on the glass causes it to appear green. (b) The cathode ray is bent in the presence of a magnet.



individual electron. The number he came up with is -1.76×10^8 C/g, where C stands for *coulomb*, which is the unit of electric charge. Thereafter, in a series of experiments carried out between 1908 and 1917, R. A. Millikan[†] found the charge of an electron to be -1.60×10^{-19} C. From these data he calculated the mass of an electron:

$$\begin{aligned} \text{mass of an electron} &= \frac{\text{charge}}{\text{charge/mass}} \\ &= \frac{-1.60 \times 10^{-19} \text{ C}}{-1.76 \times 10^8 \text{ C/g}} \\ &= 9.09 \times 10^{-28} \text{ g} \end{aligned}$$

This is an exceedingly small mass.

RADIOACTIVITY

In 1895, the German physicist Wilhelm Röntgen[‡] noticed that cathode rays caused glass and metals to emit very unusual rays. This highly energetic radiation penetrated matter, darkened covered photographic plates, and caused a variety of substances to fluoresce. Since these rays could not be deflected by a magnet, they could not contain charged particles as cathode rays do. Röntgen called them X rays because their nature was not known.

Not long after Röntgen's discovery, Antoine Becquerel,[§] a professor of physics in Paris, began to study the fluorescent properties of substances. Purely by accident, he found that exposing thickly wrapped photographic plates to a certain uranium compound caused them to darken, even without the stimulation of cathode rays. Like X rays, the rays from the uranium compound were highly energetic and could not be deflected by a magnet, but they differed from X rays because they arose spontaneously. One of Becquerel's students, Marie Curie,[¶] suggested the name *radioactivity* to describe this *spontaneous emission of particles and/or radiation*. Since then, any element that spontaneously emits radiation is said to be *radioactive*.

Three types of rays are produced by the *decay*, or breakdown, of radioactive substances such as uranium. Two of the three are deflected by oppositely charged metal plates (Figure 2.4). **Alpha (α) rays** consist of *positively charged particles*, called **α particles**, and therefore are deflected by the positively charged plate. **Beta (β) rays**, or **β particles**, are *electrons* and are deflected by the negatively charged plate. The third type of radioactive radiation consists of high-energy rays called **γ rays**. Like X rays, γ rays have no charge and are not affected by an external field.

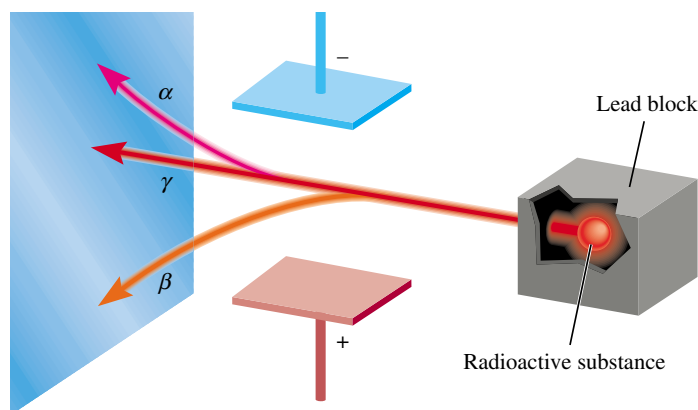
[†]Robert Andrews Millikan (1868–1953). American physicist who was awarded the Nobel Prize in Physics in 1923 for determining the charge of the electron.

[‡]Wilhelm Konrad Röntgen (1845–1923). German physicist who received the Nobel Prize in Physics in 1901 for the discovery of X rays.

[§]Antoine Henri Becquerel (1852–1908). French physicist who was awarded the Nobel Prize in Physics in 1903 for discovering radioactivity in uranium.

[¶]Marie (Marya Skłodowska) Curie (1867–1934). Polish-born chemist and physicist. In 1903 she and her French husband, Pierre Curie, were awarded the Nobel Prize in Physics for their work on radioactivity. In 1911, she again received the Nobel prize, this time in chemistry, for her work on the radioactive elements radium and polonium. She is one of only three people to have received two Nobel prizes in science. Despite her great contribution to science, her nomination to the French Academy of Sciences in 1911 was rejected by one vote because she was a woman! Her daughter Irene, and son-in-law Frederic Joliot-Curie, shared the Nobel Prize in Chemistry in 1935.

FIGURE 2.4 Three types of rays emitted by radioactive elements. β Rays consist of negatively charged particles (electrons) and are therefore attracted by the positively charged plate. The opposite holds true for α rays—they are positively charged and are drawn to the negatively charged plate. Because γ rays have no charges, their path is unaffected by an external electric field.



THE PROTON AND THE NUCLEUS

By the early 1900s, two features of atoms had become clear: they contain electrons, and they are electrically neutral. To maintain electric neutrality, an atom must contain an equal number of positive and negative charges. Therefore, Thomson proposed that an atom could be thought of as a uniform, positive sphere of matter in which electrons are embedded like raisins in a cake (Figure 2.5). This so-called “plum-pudding model” was the accepted theory for a number of years.

In 1910 the New Zealand physicist Ernest Rutherford,[†] who had studied with Thomson at Cambridge University, decided to use α particles to probe the structure of atoms. Together with his associate Hans Geiger[‡] and an undergraduate named Ernest Marsden,[§] Rutherford carried out a series of experiments using very thin foils of gold and other metals as targets for α particles from a radioactive source (Figure 2.6). They observed that the majority of particles penetrated the foil either undeflected or with only a slight deflection. But every now and then an α particle was scattered (or deflected) at a large angle. In some instances, an α particle actually bounced back in the direction from which it had come! This was a most surprising finding, for in Thomson’s model the positive charge of the atom was so diffuse that the positive α particles should have passed through the foil with very little deflection. To quote Rutherford’s initial reaction when told of this discovery: “It was as incredible as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you.”

Rutherford was later able to explain the results of the α -scattering experiment in terms of a new model for the atom. According to Rutherford, most of the atom must

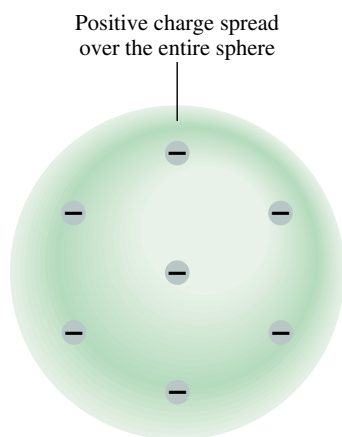


FIGURE 2.5 Thomson’s model of the atom, sometimes described as the “plum pudding” model, after a traditional English dessert containing raisins. The electrons are embedded in a uniform, positively charged sphere.

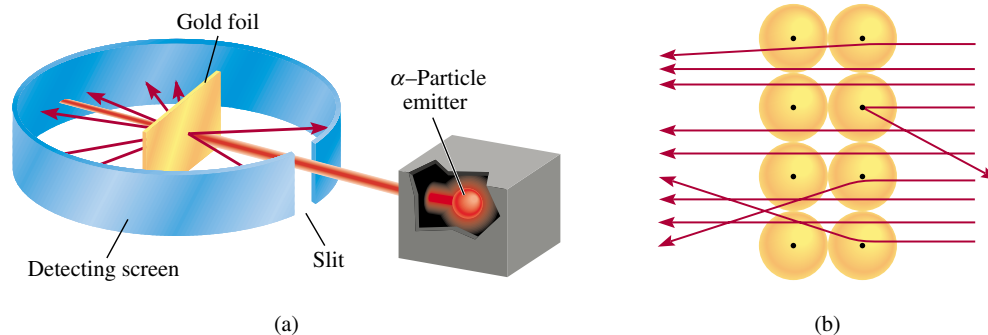
[†]Ernest Rutherford (1871–1937). New Zealand physicist. Rutherford did most of his work in England (Manchester and Cambridge universities). He received the Nobel Prize in Chemistry in 1908 for his investigations into the structure of the atomic nucleus. His often-quoted comment to his students was that “all science is either physics or stamp-collecting.”

[‡]Johannes Hans Wilhelm Geiger (1882–1945). German physicist. Geiger’s work focused on the structure of the atomic nucleus and on radioactivity. He invented a device for measuring radiation that is now commonly called the Geiger counter.

[§]Ernest Marsden (1889–1970). English physicist. It is gratifying to know that at times an undergraduate can assist in winning a Nobel Prize. Marsden went on to contribute significantly to the development of science in New Zealand.



FIGURE 2.6 (a) Rutherford's experimental design for measuring the scattering of α particles by a piece of gold foil. Most of the α particles passed through the gold foil with little or no deflection. A few were deflected at wide angles. Occasionally an α particle was turned back. (b) Magnified view of α particles passing through and being deflected by nuclei.



be empty space. This explains why the majority of α particles passed through the gold foil with little or no deflection. The atom's positive charges, Rutherford proposed, are all concentrated in the **nucleus**, which is a *dense central core within the atom*. Whenever an α particle came close to a nucleus in the scattering experiment, it experienced a large repulsive force and therefore a large deflection. Moreover, an α particle traveling directly toward a nucleus would be completely repelled and its direction would be reversed.

The positively charged particles in the nucleus are called **protons**. In separate experiments, it was found that each proton carries the same *quantity* of charge as an electron and has a mass of 1.67252×10^{-24} g—about 1840 times the mass of the oppositely charged electron.

At this stage of investigation, scientists perceived the atom as follows: The mass of a nucleus constitutes most of the mass of the entire atom, but the nucleus occupies only about $1/10^{13}$ of the volume of the atom. We express atomic (and molecular) dimensions in terms of the SI unit called the *picometer (pm)*, where

$$1 \text{ pm} = 1 \times 10^{-12} \text{ m}$$

A common non-SI unit for atomic length is the angstrom (\AA ; $1 \text{ \AA} = 100 \text{ pm}$).

A typical atomic radius is about 100 pm, whereas the radius of an atomic nucleus is only about 5×10^{-3} pm. You can appreciate the relative sizes of an atom and its nucleus by imagining that if an atom were the size of the Houston Astrodome, the volume of its nucleus would be comparable to that of a small marble. While the protons are confined to the nucleus of the atom, the electrons are conceived of as being spread out about the nucleus at some distance from it.

The concept of atomic radius is useful experimentally, but it should not be inferred that atoms have well-defined boundaries or surfaces. We will learn later that the outer regions of atoms are relatively “fuzzy.”

THE NEUTRON

Rutherford's model of atomic structure left one major problem unsolved. It was known that hydrogen, the simplest atom, contains only one proton and that the helium atom contains two protons. Therefore, the ratio of the mass of a helium atom to that of a hydrogen atom should be 2:1. (Because electrons are much lighter than protons, their contribution to atomic mass can be ignored.) In reality, however, the ratio is 4:1. Rutherford and others postulated that there must be another type of subatomic particle in the atomic nucleus; the proof was provided by another English physicist, James

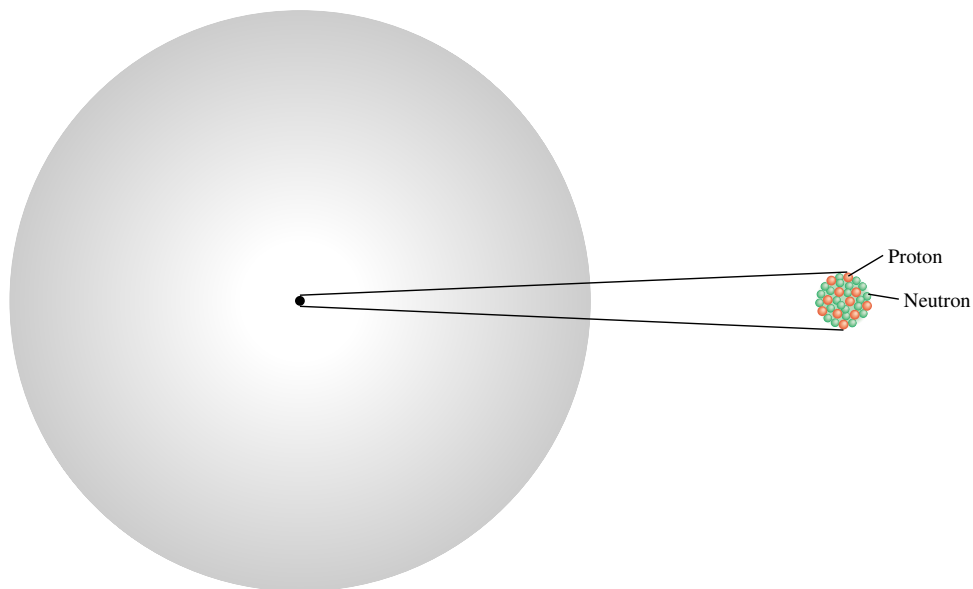


Figure 2.7 The protons and neutrons of an atom are packed in an extremely small nucleus. Electrons are shown as “clouds” around the nucleus.

Chadwick,[†] in 1932. When Chadwick bombarded a thin sheet of beryllium with α particles, a very high-energy radiation similar to γ rays was emitted by the metal. Later experiments showed that the rays actually consisted of a third type of subatomic particles, which Chadwick named **neutrons**, because they proved to be *electrically neutral particles having a mass slightly greater than that of protons*. The mystery of the mass ratio could now be explained. In the helium nucleus there are two protons and two neutrons, but in the hydrogen nucleus there is only one proton and no neutrons; therefore, the ratio is 4:1.

Figure 2.7 shows the location of the elementary particles (protons, neutrons, and electrons) in an atom. There are other subatomic particles, but the electron, the proton, and the neutron are the three fundamental components of the atom that are important in chemistry. Table 2.1 shows the masses and charges of these three elementary particles.

[†]James Chadwick (1891–1972). British physicist. In 1935 he received the Nobel Prize in Physics for proving the existence of neutrons.

TABLE 2.1 Mass and Charge of Subatomic Particles

| PARTICLE | MASS (g) | CHARGE | |
|-----------|---------------------------|---------------------------|-------------|
| | | COULOMB | CHARGE UNIT |
| Electron* | 9.1095×10^{-28} | -1.6022×10^{-19} | -1 |
| Proton | 1.67252×10^{-24} | $+1.6022 \times 10^{-19}$ | +1 |
| Neutron | 1.67495×10^{-24} | 0 | 0 |

*More refined measurements have given us a more accurate value of an electron’s mass than Millikan’s.

2.3 ATOMIC NUMBER, MASS NUMBER, AND ISOTOPES

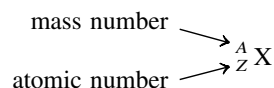
All atoms can be identified by the number of protons and neutrons they contain. The **atomic number** (Z) is the number of protons in the nucleus of each atom of an element. In a neutral atom the number of protons is equal to the number of electrons, so the atomic number also indicates the number of electrons present in the atom. The chemical identity of an atom can be determined solely from its atomic number. For example, the atomic number of nitrogen is 7. This means that each neutral nitrogen atom has 7 protons and 7 electrons. Or, viewed another way, every atom in the universe that contains 7 protons is correctly named “nitrogen.”

The **mass number** (A) is the total number of neutrons and protons present in the nucleus of an atom of an element. Except for the most common form of hydrogen, which has one proton and no neutrons, all atomic nuclei contain both protons and neutrons. In general the mass number is given by

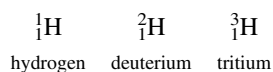
$$\begin{aligned}\text{mass number} &= \text{number of protons} + \text{number of neutrons} \\ &= \text{atomic number} + \text{number of neutrons}\end{aligned}$$

The number of neutrons in an atom is equal to the difference between the mass number and the atomic number, or ($A - Z$). For example, the mass number of fluorine is 19 and the atomic number is 9 (indicating 9 protons in the nucleus). Thus the number of neutrons in an atom of fluorine is $19 - 9 = 10$. Note that the atomic number, number of neutrons, and mass number all must be positive integers (whole numbers).

Atoms of a given element do not all have the same mass. Most elements have two or more **isotopes**, atoms that have the same atomic number but different mass numbers. For example, there are three isotopes of hydrogen. One, simply known as hydrogen, has one proton and no neutrons. The *deuterium* isotope contains one proton and one neutron, and *tritium* has one proton and two neutrons. The accepted way to denote the atomic number and mass number of an atom of an element (X) is as follows:



Thus, for the isotopes of hydrogen, we write



As another example, consider two common isotopes of uranium with mass numbers of 235 and 238, respectively:



The first isotope is used in nuclear reactors and atomic bombs, whereas the second isotope lacks the properties necessary for these applications. With the exception of hydrogen, which has different names for each of its isotopes, isotopes of elements are identified by their mass numbers. Thus the above two isotopes are called uranium-235 (pronounced “uranium two thirty-five”) and uranium-238 (pronounced “uranium two thirty-eight”).

The chemical properties of an element are determined primarily by the protons and electrons in its atoms; neutrons do not take part in chemical changes under nor-

mal conditions. Therefore, isotopes of the same element have similar chemistries, forming the same types of compounds and displaying similar reactivities.

The following example shows how to calculate the number of protons, neutrons, and electrons using atomic numbers and mass numbers.

EXAMPLE 2.1

Give the number of protons, neutrons, and electrons in each of the following species:

(a) $^{17}_8\text{O}$, (b) $^{199}_{80}\text{Hg}$, (c) $^{200}_{80}\text{Hg}$.

Answer (a) The atomic number is 8, so there are 8 protons. The mass number is 17, so the number of neutrons is $17 - 8 = 9$. The number of electrons is the same as the number of protons, that is, 8.

(b) The atomic number is 80, so there are 80 protons. The mass number is 199, so the number of neutrons is $199 - 80 = 119$. The number of electrons is 80.

(c) Here the number of protons is the same as in (b), or 80. The number of neutrons is $200 - 80 = 120$. The number of electrons is also the same as in (b), 80. The species in (b) and (c) are chemically similar isotopes of mercury.

Similar problems: 2.15, 2.16.

PRACTICE EXERCISE

How many protons, neutrons, and electrons are in the following isotope of copper: $^{63}_{29}\text{Cu}$?

2.4 THE PERIODIC TABLE

More than half of the elements known today were discovered between 1800 and 1900. During this period, chemists noted that many elements show very strong similarities to one another. Recognition of periodic regularities in physical and chemical behavior and the need to organize the large volume of available information about the structure and properties of elemental substances led to the development of the *periodic table*, a chart in which elements having similar chemical and physical properties are grouped together. Figure 2.8 shows the modern periodic table in which the elements are arranged by atomic number (shown above the element symbol) in *horizontal rows* called *periods* and in *vertical columns* known as *groups* or *families*, according to similarities in their chemical properties. Note that Elements 110, 111, and 112 have recently been synthesized, although they have not yet been named.

The elements can be divided into three categories—metals, nonmetals, and metalloids. A *metal* is a good conductor of heat and electricity while a *nonmetal* is usually a poor conductor of heat and electricity. A *metalloid* has properties that are intermediate between those of metals and nonmetals. Figure 2.8 shows that the majority of known elements are metals; only seventeen elements are nonmetals, and eight elements are metalloids. From left to right across any period, the physical and chemical properties of the elements change gradually from metallic to nonmetallic.

Elements are often referred to collectively by their periodic table group number (Group 1A, Group 2A, and so on). However, for convenience, some element groups have been given special names. The Group 1A elements (Li, Na, K, Rb, Cs, and Fr) are called *alkali metals*, and the Group 2A elements (Be, Mg, Ca, Sr, Ba, and Ra) are called *alkaline earth metals*. Elements in Group 7A (F, Cl, Br, I, and At) are known as

We will discuss the nature of chemical bonds in Chapters 9 and 10.

MOLECULES

A **molecule** is an aggregate of at least two atoms in a definite arrangement held together by chemical forces (also called *chemical bonds*). A molecule may contain atoms of the same element or atoms of two or more elements joined in a fixed ratio, in accordance with the law of definite proportions stated in Section 2.1. Thus, a molecule is not necessarily a compound, which, by definition, is made up of two or more elements (see Section 1.2). Hydrogen gas, for example, is a pure element, but it consists of molecules made up of two H atoms each. Water, on the other hand, is a molecular compound that contains hydrogen and oxygen in a ratio of two H atoms and one O atom. Like atoms, molecules are electrically neutral.

The hydrogen molecule, symbolized as H_2 , is called a **diatomic molecule** because it contains only two atoms. Other elements that normally exist as diatomic molecules are nitrogen (N_2) and oxygen (O_2), as well as the Group 7A elements—fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2). Of course, a diatomic molecule can contain atoms of different elements. Examples are hydrogen chloride (HCl) and carbon monoxide (CO).

The vast majority of molecules contain more than two atoms. They can be atoms of the same element, as in ozone (O_3), which is made up of three atoms of oxygen, or they can be combinations of two or more different elements. Molecules containing more than two atoms are called **polyatomic molecules**. Like ozone, water (H_2O) and ammonia (NH_3) are polyatomic molecules.

Molecular Models

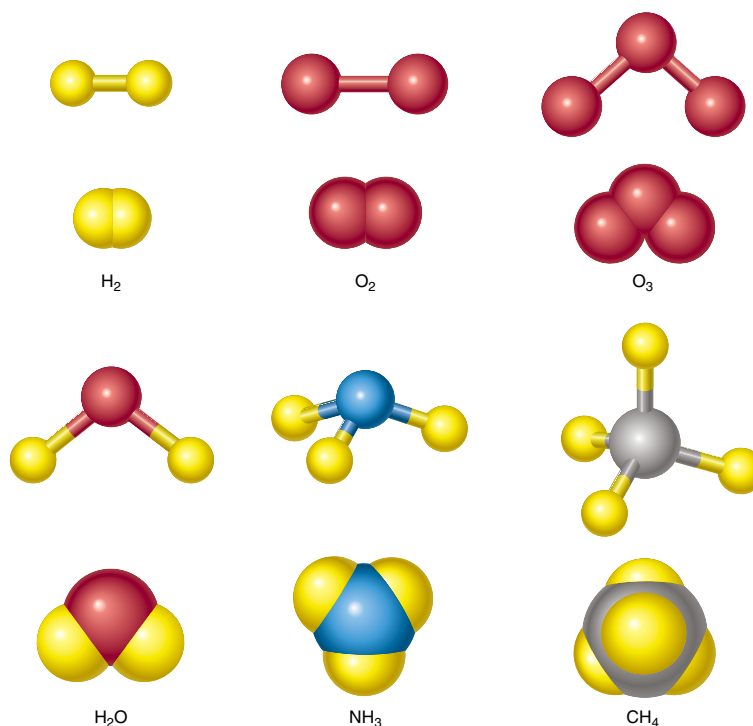
Molecules are too small for us to observe directly. An effective means of visualizing them is by the use of molecular models. Two standard types of molecular models are currently in use: *ball-and-stick* models and *space-filling* models (Figure 2.9). In ball-and-stick model kits the atoms are wooden or plastic balls with holes in them. Sticks or springs are used to represent chemical bonds. The angles they form between atoms approximate the bond angles in actual molecules. The balls are all the same size and each type of atom is represented by a specific color. In space-filling models atoms are represented by truncated balls held together by snap fasteners, so that the bonds are not visible. The balls are proportional in size to atoms.

Ball-and-stick models show the three-dimensional arrangement of atoms clearly, and they are fairly easy to construct. However, the balls are not proportional to the size of atoms. Furthermore, the sticks greatly exaggerate the space between atoms in a molecule. Space-filling models are more accurate because they show the variation in atomic size. Their drawbacks are that they are time-consuming to put together, and they do not show the three-dimensional positions of atoms very well. We will use mostly the ball-and-stick model in this text.

IONS

An **ion** is a charged species formed from a neutral atom or molecule when electrons are gained or lost as the result of a chemical change. The number of positively charged protons in the nucleus of an atom remains the same during ordinary chemical changes (called chemical reactions), but negatively charged electrons may be lost or gained. The loss of one or more electrons from a neutral atom results in a **cation**, an ion with

FIGURE 2.9 Ball-and-stick and space-filling models of some simple molecules.



In Chapter 8 we will see why atoms of different elements gain (or lose) a specific number of electrons.

a net positive charge. For example, a sodium atom (Na) can readily lose an electron to become sodium cation, which is represented by Na^+ :

| Na ATOM | Na^+ ION |
|--------------|-------------------|
| 11 protons | 11 protons |
| 11 electrons | 10 electrons |

On the other hand, an **anion** is an ion whose net charge is negative due to an increase in the number of electrons. A chlorine atom (Cl), for instance, can gain an electron to become the chloride ion Cl^- :

| Cl ATOM | Cl^- ION |
|--------------|-------------------|
| 17 protons | 17 protons |
| 17 electrons | 18 electrons |

Sodium chloride (NaCl), ordinary table salt, is called an **ionic compound** because it is formed from cations and anions.

An atom can lose or gain more than one electron. Examples of ions formed by the loss or gain of more than one electron are Mg^{2+} , Fe^{3+} , S^{2-} , and N^{3-} . These ions, as well as Na^+ and Cl^- , are called **monatomic ions** because they contain only one atom. Figure 2.10 shows the charges of a number of monatomic ions. With very few exceptions, metals tend to form cations and nonmetals form anions.

In addition, two or more atoms can combine to form an ion that has a net positive or net negative charge. **Polyatomic ions** such as OH^- (hydroxide ion), CN^- (cyanide ion), and NH_4^+ (ammonium ion) are ions containing more than one atom.

| 1 1A | 2 2A | 3 3B | 4 4B | 5 5B | 6 6B | 7 7B | 8 8B | 9 8B | 10 8B | 11 1B | 12 2B | 13 3A | 14 4A | 15 5A | 16 6A | 17 7A | 18 8A |
|-----------------|------------------|---------|---------|---------|------------------|------------------|--------------------------------------|------------------|------------------|-------------------------------------|---|------------------|------------------|-----------------|------------------|-----------------|----------|
| Li ⁺ | | | | | | | | | | | | Al ³⁺ | C ⁴⁻ | N ³⁻ | O ²⁻ | F ⁻ | |
| Na ⁺ | Mg ²⁺ | | | | Cr ³⁺ | Mn ²⁺ | Fe ²⁺ Fe ³⁺ | Co ²⁺ | Ni ²⁺ | Cu ⁺ Cu ²⁺ | Zn ²⁺ | | | P ³⁻ | S ²⁻ | Cl ⁻ | |
| K ⁺ | Ca ²⁺ | | | | | | | | | | | | | | Se ²⁻ | Br ⁻ | |
| Rb ⁺ | Sr ²⁺ | | | | | | | | | Ag ⁺ | Cd ²⁺ | | Sn ²⁺ | | Te ²⁻ | I ⁻ | |
| Cs ⁺ | Ba ²⁺ | | | | | | | | | | Hg ₂ ²⁺ Hg ²⁺ | | Pb ²⁺ | | | | |

FIGURE 2.10 Common monatomic ions arranged according to their positions in the periodic table. Note that the Hg_2^{2+} ion contains two atoms.

2.6 CHEMICAL FORMULAS

Chemists use **chemical formulas** to express the composition of molecules and ionic compounds in terms of chemical symbols. By composition we mean not only the elements present but also the ratios in which the atoms are combined. Here we are concerned with two types of formulas: molecular formulas and empirical formulas.

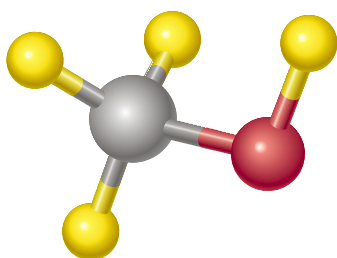
MOLECULAR FORMULAS

A **molecular formula** shows the exact number of atoms of each element in the smallest unit of a substance. In our discussion of molecules, each example was given with its molecular formula in parentheses. Thus H_2 is the molecular formula for hydrogen, O_2 is oxygen, O_3 is ozone, and H_2O is water. The subscript numeral indicates the number of atoms of an element present. There is no subscript for O in H_2O because there is only one atom of oxygen in a molecule of water, and so the number “one” is omitted from the formula. Note that oxygen (O_2) and ozone (O_3) are allotropes of oxygen. An **allotrope** is one of two or more distinct forms of an element. Two allotropic forms of the element carbon—diamond and graphite—are dramatically different not only in properties but also in their relative cost.

EMPIRICAL FORMULAS

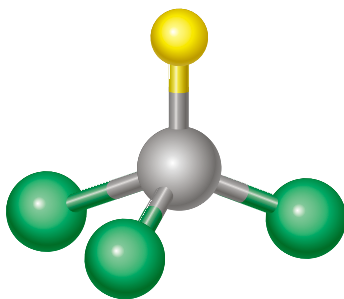
The molecular formula of hydrogen peroxide, a substance used as an antiseptic and as a bleaching agent for textiles and hair, is H_2O_2 . This formula indicates that each hydrogen peroxide molecule consists of two hydrogen atoms and two oxygen atoms. The ratio of hydrogen to oxygen atoms in this molecule is 2:2 or 1:1. The empirical formula of hydrogen peroxide is HO. Thus the **empirical formula** tells us which elements are present and the simplest whole-number ratio of their atoms, but not necessarily the

The word “empirical” means “derived from experiment.” As we will see in Chapter 3, empirical formulas are determined experimentally.



Methanol. (Gray=carbon, yellow=hydrogen, red=oxygen.)

Similar problems: 2.41, 2.42.



Chloroform. (Gray=carbon, yellow=hydrogen, green=chlorine.)

Similar problems: 2.43, 2.44.

actual number of atoms in a given molecule. As another example, consider the compound hydrazine (N_2H_4), which is used as a rocket fuel. The empirical formula of hydrazine is NH_2 . Although the ratio of nitrogen to hydrogen is 1:2 in both the molecular formula (N_2H_4) and the empirical formula (NH_2), only the molecular formula tells us the actual number of N atoms (two) and H atoms (four) present in a hydrazine molecule.

Empirical formulas are the *simplest* chemical formulas; they are written by reducing the subscripts in molecular formulas to the smallest possible whole numbers. Molecular formulas are the *true* formulas of molecules. As we will see in Chapter 3, when chemists analyze an unknown compound, the first step is usually the determination of the compound’s empirical formula.

For many molecules, the molecular formula and the empirical formula are one and the same. Some examples are water (H_2O), ammonia (NH_3), carbon dioxide (CO_2), and methane (CH_4).

The following two examples deal with writing molecular formulas from molecular models and writing empirical formulas from molecular formulas.

EXAMPLE 2.2

Write the molecular formula of methanol, an organic solvent and antifreeze, from its ball-and-stick model, shown in the margin.

Answer There is one C atom, four H atoms, and one O atom. Therefore, the molecular formula is CH_4O . However, the standard way of writing the molecular formula for methanol is CH_3OH because it shows how the atoms are joined in the molecule.

PRACTICE EXERCISE

Write the molecular formula of chloroform, which is used as a solvent and a cleansing agent. The ball-and-stick model of chloroform is shown in the margin.

EXAMPLE 2.3

Write the empirical formulas for the following molecules: (a) acetylene (C_2H_2), which is used in welding torches; (b) glucose ($\text{C}_6\text{H}_{12}\text{O}_6$), a substance known as blood sugar; and (c) nitrous oxide (N_2O), a gas that is used as an anesthetic gas (“laughing gas”) and as an aerosol propellant for whipped creams.

Answer (a) There are two carbon atoms and two hydrogen atoms in acetylene. Dividing the subscripts by 2, we obtain the empirical formula CH.

(b) In glucose there are six carbon atoms, twelve hydrogen atoms, and six oxygen atoms. Dividing the subscripts by 6, we obtain the empirical formula CH_2O . Note that if we had divided the subscripts by 3, we would have obtained the formula $\text{C}_2\text{H}_4\text{O}_2$. Although the ratio of carbon to hydrogen to oxygen atoms in $\text{C}_2\text{H}_4\text{O}_2$ is the same as that in $\text{C}_6\text{H}_{12}\text{O}_6$ (1:2:1), $\text{C}_2\text{H}_4\text{O}_2$ is not the simplest formula because its subscripts are not in the smallest whole-number ratio.

(c) Since the subscripts in N_2O are already the smallest possible whole numbers, the empirical formula for nitrous oxide is the same as its molecular formula.

PRACTICE EXERCISE

Write the empirical formula for caffeine ($C_8H_{10}N_4O_2$), a stimulant found in tea and coffee.

The formulas of ionic compounds are always the same as their empirical formulas because ionic compounds do not consist of discrete molecular units. For example, a solid sample of sodium chloride ($NaCl$) consists of equal numbers of Na^+ and Cl^- ions arranged in a three-dimensional network (Figure 2.11). In such a compound there is a 1:1 ratio of cations to anions so that the compound is electrically neutral. As you can see in Figure 2.11, no Na^+ ion in $NaCl$ is associated with just one particular Cl^- ion. In fact, each Na^+ ion is equally held by six surrounding Cl^- ions and vice versa. Thus $NaCl$ is the empirical formula for sodium chloride. In other ionic compounds the actual structure may be different, but the arrangement of cations and anions is such that the compounds are all electrically neutral. Note that the charges on the cation and anion are not shown in the formula for an ionic compound.

In order for ionic compounds to be electrically neutral, the sum of the charges on the cation and anion in each formula unit must be zero. If the charges on the cation and anion are numerically different, we apply the following rule to make the formula electrically neutral: *The subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation.* If the charges are numerically equal, then no subscripts are necessary. This rule follows from the fact that because the formulas of ionic compounds are empirical formulas, the subscripts must always be reduced to the smallest ratios. Let us consider some examples.

- *Potassium bromide.* The potassium cation K^+ and the bromine anion Br^- combine to form the ionic compound potassium bromide. The sum of the charges is $+1 + (-1) = 0$, so no subscripts are necessary. The formula is KBr .
- *Zinc iodide.* The zinc cation Zn^{2+} and the iodine anion I^- combine to form zinc iodide. The sum of the charges of one Zn^{2+} ion and one I^- ion is $+2 + (-1) = +1$. To make the charges add up to zero we multiply the -1 charge of the anion by 2 and add the subscript "2" to the symbol for iodine. Therefore the formula for zinc iodide is ZnI_2 .
- *Aluminum oxide.* The cation is Al^{3+} and the oxygen anion is O^{2-} . The following

FIGURE 2.11 (a) Structure of solid $NaCl$. (b) In reality, the cations are in contact with the anions. In both (a) and (b), the smaller spheres represent Na^+ ions and the larger spheres, Cl^- ions.

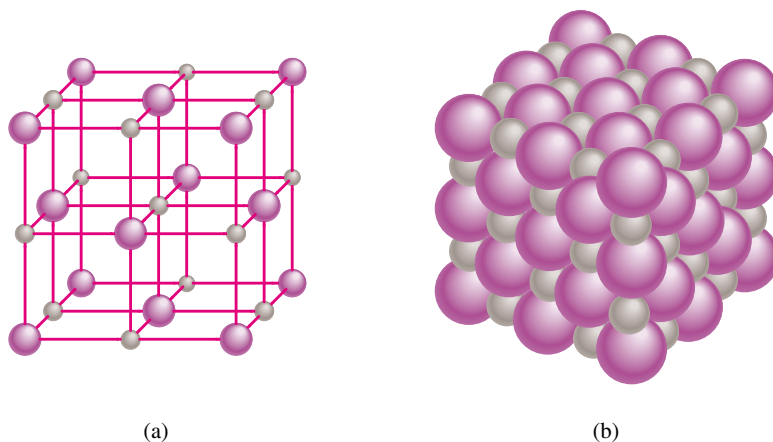
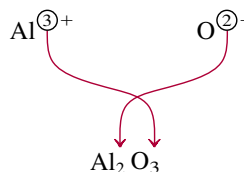


diagram helps us determine the subscripts for the compound formed by the cation and the anion:



The sum of the charges is $2(+3) + 3(-2) = 0$. Thus the formula for aluminum oxide is Al_2O_3 .

2.7 NAMING COMPOUNDS

When chemistry was a young science and the number of known compounds was small, it was possible to memorize their names. Many of the names were derived from their physical appearance, properties, origin, or application—for example, milk of magnesia, laughing gas, limestone, caustic soda, lye, washing soda, and baking soda.

Today the number of known compounds is well over 13 million. Fortunately, it is not necessary to memorize their names. Over the years chemists have devised a clear system for naming chemical substances. The rules are accepted worldwide, facilitating communication among chemists and providing a useful way of labeling an overwhelming variety of substances. Mastering these rules now will prove beneficial almost immediately as we proceed with our study of chemistry.

To begin our discussion of chemical *nomenclature*, the naming of chemical compounds, we must first distinguish between inorganic and organic compounds. *Organic compounds* contain carbon, usually in combination with elements such as hydrogen, oxygen, nitrogen, and sulfur. All other compounds are classified as *inorganic compounds*. For convenience, some carbon-containing compounds, such as carbon monoxide (CO), carbon dioxide (CO₂), carbon disulfide (CS₂), compounds containing the cyanide group (CN⁻), and carbonate (CO₃²⁻) and bicarbonate (HCO₃⁻) groups are considered to be inorganic compounds. Although the nomenclature of organic compounds will not be discussed until Chapter 24, we will use some organic compounds to illustrate chemical principles throughout this book.

To organize and simplify our venture into naming compounds, we can divide inorganic compounds into four categories: ionic compounds, molecular compounds, acids and bases, and hydrates.

IONIC COMPOUNDS

In Section 2.5 we learned that ionic compounds are made up of cations (positive ions) and anions (negative ions). With the important exception of the ammonium ion, NH₄⁺, all cations of interest to us are derived from metal atoms. Metal cations take their names from the elements. For example:

| ELEMENT | NAME OF CATION |
|--------------|--|
| Na sodium | Na ⁺ sodium ion (or sodium cation) |
| K potassium | K ⁺ potassium ion (or potassium cation) |
| Mg magnesium | Mg ²⁺ magnesium ion (or magnesium cation) |
| Al aluminum | Al ³⁺ aluminum ion (or aluminum cation) |

Many ionic compounds are **binary compounds**, or *compounds formed from just two elements*. For binary compounds the first element named is the metal cation, followed by the nonmetallic anion. Thus NaCl is sodium chloride. The anion is named by taking the first part of the element name (chlorine) and adding “-ide.” Potassium bromide (KBr), zinc iodide (ZnI₂), and aluminum oxide (Al₂O₃) are also binary compounds. Table 2.2 shows the “-ide” nomenclature of some common monatomic anions according to their positions in the periodic table.

The “-ide” ending is also used for certain anion groups containing different elements, such as hydroxide (OH⁻) and cyanide (CN⁻). Thus the compounds LiOH and KCN are named lithium hydroxide and potassium cyanide, respectively. These and a number of other such ionic substances are called **ternary compounds**, meaning *compounds consisting of three elements*. Table 2.3 lists alphabetically the names of a number of common cations and anions.

Certain metals, especially the *transition metals*, can form more than one type of cation. Take iron as an example. Iron can form two cations: Fe²⁺ and Fe³⁺. An older nomenclature system that is still in limited use assigns the ending “-ous” to the cation with fewer positive charges and the ending “-ic” to the cation with more positive charges:

| | |
|------------------|-------------|
| Fe ²⁺ | ferrous ion |
| Fe ³⁺ | ferric ion |

The names of the compounds that these iron ions form with chlorine would thus be

| | |
|-------------------|------------------|
| FeCl ₂ | ferrous chloride |
| FeCl ₃ | ferric chloride |

This method of naming ions has some distinct limitations. First, the “-ous” and “-ic” suffixes do not provide information regarding the actual charges of the two cations involved. Thus the ferric ion is Fe³⁺, but the cation of copper named cupric has the formula Cu²⁺. In addition, the “-ous” and “-ic” designations provide names for only two different elemental cations. Some metallic elements can assume three or more different positive charges in compounds. Therefore, it has become increasingly common to designate different cations with Roman numerals. This is called the Stock[†] system. In this system, the Roman numeral I indicates one positive charge, II means two positive charges, and so on. For example, manganese (Mn) atoms can assume several different positive charges:

| | |
|---|----------------------|
| Mn ²⁺ : MnO | manganese(II) oxide |
| Mn ³⁺ : Mn ₂ O ₃ | manganese(III) oxide |
| Mn ⁴⁺ : MnO ₂ | manganese(IV) oxide |

These names are pronounced “manganese-two oxide,” “manganese-three oxide,” and “manganese-four oxide.” Using the Stock system, we denote the ferrous ion and the ferric ion as iron(II) and iron(III), respectively; ferrous chloride becomes iron(II) chloride; and ferric chloride is called iron(III) chloride. In keeping with modern practice, we will favor the Stock system of naming compounds in this textbook.

The following examples illustrate how to name ionic compounds and write formulas for ionic compounds based on the information given in Figure 2.10 and Tables 2.2 and 2.3.

[†]Alfred E. Stock (1876–1946). German chemist. Stock did most of his research in the synthesis and characterization of boron, beryllium, and silicon compounds. He was the first scientist to explore the dangers of mercury poisoning.

The transition metals are the elements in Groups 1B and 3B-8B (see Figure 2.8).



FeCl₂ (left) and FeCl₃ (right).

TABLE 2.2 The “-ide” Nomenclature of Some Common Monatomic Anions According to Their Positions in the Periodic Table

| GROUP 4A | GROUP 5A | GROUP 6A | GROUP 7A |
|---------------------------|--------------------------|----------------------------|--------------------------|
| C Carbide (C^{4-})* | N Nitride (N^{3-}) | O Oxide (O^{2-}) | F Fluoride (F^{-}) |
| Si Silicide (Si^{4-}) | P Phosphide (P^{3-}) | S Sulfide (S^{2-}) | Cl Chloride (Cl^{-}) |
| | | Se Selenide (Se^{2-}) | Br Bromide (Br^{-}) |
| | | Te Telluride (Te^{2-}) | I Iodide (I^{-}) |

*The word “carbide” is also used for the anion C_2^{2-} .

TABLE 2.3 Names and Formulas of Some Common Inorganic Cations and Anions

| CATION | ANION |
|--|---|
| Aluminum (Al^{3+}) | Bromide (Br^{-}) |
| Ammonium (NH_4^{+}) | Carbonate (CO_3^{2-}) |
| Barium (Ba^{2+}) | Chlorate (ClO_3^{-}) |
| Cadmium (Cd^{2+}) | Chloride (Cl^{-}) |
| Calcium (Ca^{2+}) | Chromate (CrO_4^{2-}) |
| Cesium (Cs^{+}) | Cyanide (CN^{-}) |
| Chromium(III) or chromic (Cr^{3+}) | Dichromate ($Cr_2O_7^{2-}$) |
| Cobalt(II) or cobaltous (Co^{2+}) | Dihydrogen phosphate ($H_2PO_4^{-}$) |
| Copper(I) or cuprous (Cu^{+}) | Fluoride (F^{-}) |
| Copper(II) or cupric (Cu^{2+}) | Hydride (H^{-}) |
| Hydrogen (H^{+}) | Hydrogen carbonate or bicarbonate (HCO_3^{-}) |
| Iron(II) or ferrous (Fe^{2+}) | Hydrogen phosphate (HPO_4^{2-}) |
| Iron(III) or ferric (Fe^{3+}) | Hydrogen sulfate or bisulfate (HSO_4^{-}) |
| Lead(II) or plumbous (Pb^{2+}) | Hydroxide (OH^{-}) |
| Lithium (Li^{+}) | Iodide (I^{-}) |
| Magnesium (Mg^{2+}) | Nitrate (NO_3^{-}) |
| Manganese(II) or manganous (Mn^{2+}) | Nitride (N^{3-}) |
| Mercury(I) or mercurous (Hg_2^{2+})* | Nitrite (NO_2^{-}) |
| Mercury(II) or mercuric (Hg^{2+}) | Oxide (O^{2-}) |
| Potassium (K^{+}) | Permanganate (MnO_4^{-}) |
| Silver (Ag^{+}) | Peroxide (O_2^{2-}) |
| Sodium (Na^{+}) | Phosphate (PO_4^{3-}) |
| Strontium (Sr^{2+}) | Sulfate (SO_4^{2-}) |
| Tin(II) or stannous (Sn^{2+}) | Sulfide (S^{2-}) |
| Zinc (Zn^{2+}) | Sulfite (SO_3^{2-}) |
| | Thiocyanate (SCN^{-}) |

*Mercury(I) exists as a pair as shown.

EXAMPLE 2.4

Name the following ionic compounds: (a) $Cu(NO_3)_2$, (b) KH_2PO_4 , and (c) NH_4ClO_3 .

Answer (a) Since the nitrate ion (NO_3^{-}) bears one negative charge (see Table 2.3), the copper ion must have two positive charges. Therefore, the compound is copper(II) nitrate.

Similar problems: 2.53(a), (b), (e).

(b) The cation is K^+ and the anion is $H_2PO_4^-$ (dihydrogen phosphate). Since potassium only forms one type of ion (K^+), there is no need to use potassium(I) in the name. The compound is potassium dihydrogen phosphate.

(c) The cation is NH_4^+ (ammonium ion) and the anion is ClO_3^- . The compound is ammonium chlorate.

PRACTICE EXERCISE

Name the following compounds: (a) PbO , (b) Li_2SO_3 .

EXAMPLE 2.5

Write chemical formulas for the following compounds: (a) mercury(I) nitrite, (b) cesium sulfide, and (c) calcium phosphate.

Answer: (a) The mercury(I) ion is diatomic, namely, Hg_2^{2+} (see Table 2.3), and the nitrite ion is NO_2^- . Therefore, the formula is $Hg_2(NO_2)_2$.

(b) Each sulfide ion bears two negative charges, and each cesium ion bears one positive charge (cesium is in Group 1A, as is sodium). Therefore, the formula is Cs_2S .

(c) Each calcium ion (Ca^{2+}) bears two positive charges, and each phosphate ion (PO_4^{3-}) bears three negative charges. To make the sum of the charges equal zero, we must adjust the numbers of cations and anions:

$$3(+2) + 2(-3) = 0$$

Thus the formula is $Ca_3(PO_4)_2$.

PRACTICE EXERCISE

Write formulas for the following ionic compounds: (a) rubidium sulfate, (b) barium hydride.

Similar problems: 2.55(a), (b), (h).

TABLE 2.4 Greek Prefixes Used in Naming Molecular Compounds

| PREFIX | MEANING |
|--------|---------|
| Mono- | 1 |
| Di- | 2 |
| Tri- | 3 |
| Tetra- | 4 |
| Penta- | 5 |
| Hexa- | 6 |
| Hepta- | 7 |
| Octa- | 8 |
| Nona- | 9 |
| Deca- | 10 |

MOLECULAR COMPOUNDS

Unlike ionic compounds, molecular compounds contain discrete molecular units. They are usually composed of nonmetallic elements (see Figure 2.8). Many molecular compounds are binary compounds. Naming binary molecular compounds is similar to naming binary ionic compounds. We place the name of the first element in the formula first, and the second element is named by adding -ide to the root of the element name. Some examples are

| | |
|-----|-------------------|
| HCl | hydrogen chloride |
| HBr | hydrogen bromide |
| SiC | silicon carbide |

It is quite common for one pair of elements to form several different compounds. In these cases, confusion in naming the compounds is avoided by the use of Greek prefixes to denote the number of atoms of each element present (see Table 2.4). Consider the following examples:

| | |
|-----------------|-----------------|
| CO | carbon monoxide |
| CO ₂ | carbon dioxide |

| | |
|-------------------------------|----------------------|
| SO ₂ | sulfur dioxide |
| SO ₃ | sulfur trioxide |
| NO ₂ | nitrogen dioxide |
| N ₂ O ₄ | dinitrogen tetroxide |

The following guidelines are helpful in naming compounds with prefixes:

- The prefix “mono-” may be omitted for the first element. For example, PCl₃ is named phosphorus trichloride, not monophosphorus trichloride. Thus the absence of a prefix for the first element usually means there is only one atom of that element present in the molecule.
- For oxides, the ending “a” in the prefix is sometimes omitted. For example, N₂O₄ may be called dinitrogen tetroxide rather than dinitrogen tetraoxide.

Exceptions to the use of Greek prefixes are molecular compounds containing hydrogen. Traditionally, many of these compounds are called either by their common, nonsystematic names or by names that do not specifically indicate the number of H atoms present:

| | |
|-------------------------------|------------------|
| B ₂ H ₆ | diborane |
| CH ₄ | methane |
| SiH ₄ | silane |
| NH ₃ | ammonia |
| PH ₃ | phosphine |
| H ₂ O | water |
| H ₂ S | hydrogen sulfide |

Note that even the order of writing the elements in the formulas for hydrogen compounds is irregular. In water and hydrogen sulfide, H is written first, whereas it appears last in the other compounds.

Writing formulas for molecular compounds is usually straightforward. Thus the name arsenic trifluoride means that there are one As atom and three F atoms in each molecule, and the molecular formula is AsF₃. Note that the order of elements in the formula is the same as in its name.

EXAMPLE 2.6

Name the following molecular compounds: (a) SiCl₄ and (b) P₄O₁₀.

Answer (a) Since there are four chlorine atoms present, the compound is silicon tetrachloride.

(b) There are four phosphorus atoms and ten oxygen atoms present, so the compound is tetraphosphorus decoxide. Note that the “a” is omitted in “deca.”

PRACTICE EXERCISE

Name the following molecular compounds: (a) NF₃ and (b) Cl₂O₇.

EXAMPLE 2.7

Write chemical formulas for the following molecular compounds: (a) carbon disulfide and (b) disilicon hexabromide.

Similar problems: 2.53(c), (h), (j).

Similar problems: 2.55(g), (j).

Answer (a) Since there are one carbon atom and two sulfur atoms present, the formula is CS_2 .
 (b) There are two silicon atoms and six bromine atoms present, so the formula is Si_2Br_6 .

PRACTICE EXERCISE

Write chemical formulas for the following molecular compounds: (a) sulfur tetrafluoride, (b) dinitrogen pentoxide.

ACIDS AND BASES

Naming Acids

H^+ is equivalent to one *proton*, and is sometimes referred to that way.

An **acid** can be described as a substance that yields hydrogen ions (H^+) when dissolved in water. Formulas for acids contain one or more hydrogen atoms as well as an anionic group. Anions whose names end in “-ide” form acids with a “hydro-” prefix and an “-ic” ending, as shown in Table 2.5. In some cases two different names seem to be assigned to the same chemical formula.

HCl hydrogen chloride
 HCl hydrochloric acid

The name assigned to the compound depends on its physical state. In the gaseous or pure liquid state, HCl is a molecular compound called hydrogen chloride. When it is dissolved in water, the molecules break up into H^+ and Cl^- ions; in this state, the substance is called hydrochloric acid.

Oxoacids are acids that contain hydrogen, oxygen, and another element (the central element). The formulas of oxoacids are usually written with the H first, followed by the central element and then O, as illustrated by the following examples:

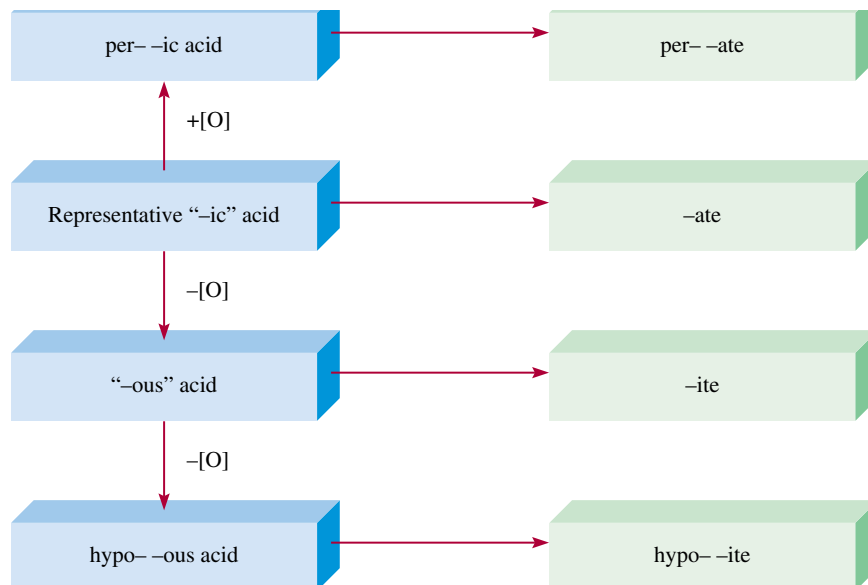
H_2CO_3 carbonic acid
 HNO_3 nitric acid
 H_2SO_4 sulfuric acid
 HClO_3 chloric acid

Often two or more oxoacids have the same central atom but a different number of O atoms. Starting with the oxoacids whose names end with “-ic,” we use the following rules to name these compounds.

TABLE 2.5 Some Simple Acids

| ANION | CORRESPONDING ACID |
|---------------------------|---|
| F^- (fluoride) | HF (hydrofluoric acid) |
| Cl^- (chloride) | HCl (hydrochloric acid) |
| Br^- (bromide) | HBr (hydrobromic acid) |
| I^- (iodide) | HI (hydroiodic acid) |
| CN^- (cyanide) | HCN (hydrocyanic acid) |
| S^{2-} (sulfide) | H_2S (hydrosulfuric acid) |

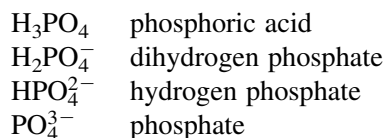
FIGURE 2.12 Naming oxoacids and oxoanions.



- Addition of one O atom to the “-ic” acid: The acid is called “per...-ic” acid. Thus adding an O atom to HClO_3 changes chloric acid to perchloric acid, HClO_4 .
- Removal of one O atom from the “-ic” acid: The acid is called “-ous” acid. Thus nitric acid, HNO_3 , becomes nitrous acid, HNO_2 .
- Removal of two O atoms from the “-ic” acid: The acid is called “hypo...-ous” acid. Thus when HBrO_3 is converted to HBrO , the acid is called hypobromous acid.

The rules for naming *oxoanions*, anions of oxoacids, are as follows:

- When all the H ions are removed from the “-ic” acid, the anion’s name ends with “-ate.” For example, the anion CO_3^{2-} derived from H_2CO_3 is called carbonate.
- When all the H ions are removed from the “-ous” acid, the anion’s name ends with “-ite.” Thus the anion ClO_2^- derived from HClO_2 is called chlorite.
- The names of anions in which one or more but not all the hydrogen ions have been removed must indicate the number of H ions present. For example, consider the anions derived from phosphoric acid:



Note that we usually omit the prefix “mono-” when there is only one H in the anion. Figure 2.12 summarizes the nomenclature for the oxoacids and oxoanions, and Table 2.6 gives the names of the oxoacids and oxoanions that contain chlorine.

TABLE 2.6 Names of Oxoacids and Oxoanions That Contain Chlorine

| ACID | ANION |
|-------------------------------------|---|
| HClO ₄ (perchloric acid) | ClO ₄ ⁻ (perchlorate) |
| HClO ₃ (chloric acid) | ClO ₃ ⁻ (chlorate) |
| HClO ₂ (chlorous acid) | ClO ₂ ⁻ (chlorite) |
| HClO (hypochlorous acid) | ClO ⁻ (hypochlorite) |

The following example deals with the nomenclature for an oxoacid and an oxoanion.

EXAMPLE 2.8

Name the following oxoacid and oxoanion: (a) H₃PO₃, (b) IO₄⁻.

Answer (a) We start with our reference acid, phosphoric acid (H₃PO₄). Since H₃PO₃ has one fewer O atom, it is called phosphorous acid.

(b) The parent acid is HIO₄. Since the acid has one more O atom than our reference iodic acid (HIO₃), it is called periodic acid. Therefore, the anion derived from HIO₄ is called periodate.

Similar problems: 2.53(h), 2.54(c).

PRACTICE EXERCISE

Name the following oxoacid and oxoanion: (a) HBrO, (b) HSO₄⁻.

Naming Bases

A *base* can be described as a substance that yields hydroxide ions (OH⁻) when dissolved in water. Some examples are

| | |
|---------------------|---------------------|
| NaOH | sodium hydroxide |
| KOH | potassium hydroxide |
| Ba(OH) ₂ | barium hydroxide |

Ammonia (NH₃), a molecular compound in the gaseous or pure liquid state, is also classified as a common base. At first glance this may seem to be an exception to the definition of a base. But note that as long as a substance *yields* hydroxide ions when dissolved in water, it need not contain hydroxide ions in its structure to be considered a base. In fact, when ammonia dissolves in water, NH₃ reacts partially with water to yield NH₄⁺ and OH⁻ ions. Thus it is properly classified as a base.

HYDRATES

Hydrates are compounds that have a specific number of water molecules attached to them. For example, in its normal state, each unit of copper(II) sulfate has five water molecules associated with it. The systematic name for this compound is copper(II) sulfate pentahydrate, and its formula is written as CuSO₄ · 5H₂O. The water molecules can be driven off by heating. When this occurs, the resulting compound is CuSO₄, which is sometimes called *anhydrous* copper(II) sulfate; “anhydrous” means that the compound no longer has water molecules associated with it (Figure 2.13). Some other

FIGURE 2.13 $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ (left) is blue; CuSO_4 (right) is white.



TABLE 2.7 Common and Systematic Names of Some Compounds

| FORMULA | COMMON NAME | SYSTEMATIC NAME |
|---|--------------------------|----------------------------------|
| H_2O | Water | Dihydrogen oxide |
| NH_3 | Ammonia | Trihydrogen nitride |
| CO_2 | Dry ice | Solid carbon dioxide |
| NaCl | Table salt | Sodium chloride |
| N_2O | Laughing gas | Dinitrogen oxide (nitrous oxide) |
| CaCO_3 | Marble, chalk, limestone | Calcium carbonate |
| CaO | Quicklime | Calcium oxide |
| $\text{Ca}(\text{OH})_2$ | Slaked lime | Calcium hydroxide |
| NaHCO_3 | Baking soda | Sodium hydrogen carbonate |
| $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ | Washing soda | Sodium carbonate decahydrate |
| $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ | Epsom salt | Magnesium sulfate heptahydrate |
| $\text{Mg}(\text{OH})_2$ | Milk of magnesia | Magnesium hydroxide |
| $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ | Gypsum | Calcium sulfate dihydrate |

hydrates are

| | |
|--|--------------------------------|
| $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$ | barium chloride dihydrate |
| $\text{LiCl} \cdot \text{H}_2\text{O}$ | lithium chloride monohydrate |
| $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$ | magnesium sulfate heptahydrate |
| $\text{Sr}(\text{NO}_3)_2 \cdot 4\text{H}_2\text{O}$ | strontium nitrate tetrahydrate |

FAMILIAR INORGANIC COMPOUNDS

Some compounds are better known by their common names than by their systematic

SUMMARY OF FACTS AND CONCEPTS

1. Modern chemistry began with Dalton's atomic theory, which states that all matter is composed of tiny, indivisible particles called atoms; that all atoms of the same element are identical; that compounds contain atoms of different elements combined in whole-number ratios; and that atoms are neither created nor destroyed in chemical reactions (the law of conservation of mass).
2. Atoms of constituent elements in a particular compound are always combined in the same proportions by mass (law of definite proportions). When two elements can combine to form more than one type of compound, the masses of one element that combine with a fixed mass of the other element are in a ratio of small whole numbers (law of multiple proportions).

Distribution of Elements on Earth and in Living Systems

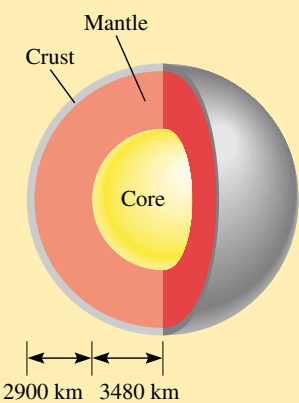
The majority of elements are naturally occurring. How are these elements distributed on Earth, and which are essential to living systems?

Earth's crust extends from the surface to a depth of about 40 km (about 25 miles). Because of technical difficulties, scientists have not been able to study the inner portions of Earth as easily as the crust. Nevertheless, it is believed that there is a solid core consisting mostly of iron at the center of Earth. Surrounding the core is a layer called the *mantle*, which consists of hot fluid containing iron, carbon, silicon, and sulfur.

Of the 83 elements that are found in nature, 12 of them make up 99.7 percent of Earth's crust by mass. They are, in decreasing order of natural abundance, oxygen (O), silicon (Si), aluminum (Al), iron (Fe), calcium (Ca), magnesium (Mg), sodium (Na), potassium (K), titanium (Ti), hydrogen (H), phosphorus (P), and

manganese (Mn). In discussing the natural abundance of the elements, we should keep in mind that (1) the elements are not evenly distributed throughout Earth's crust, and (2) most elements occur in combined forms. These facts provide the basis for most methods of obtaining pure elements from their compounds, as we will see in later chapters.

The accompanying table lists the essential elements in the human body. Of special interest are the *trace elements*, such as iron (Fe), copper (Cu), zinc (Zn), iodine (I), and cobalt (Co), which together make up about one percent of the body's mass. These elements are necessary for biological functions such as growth, transport of oxygen for metabolism, and defense against disease. There is a delicate balance in the amounts of these elements in our bodies. Too much or too little over an extended period of time can lead to serious illness, retardation, or even death.



Structure of Earth's interior.

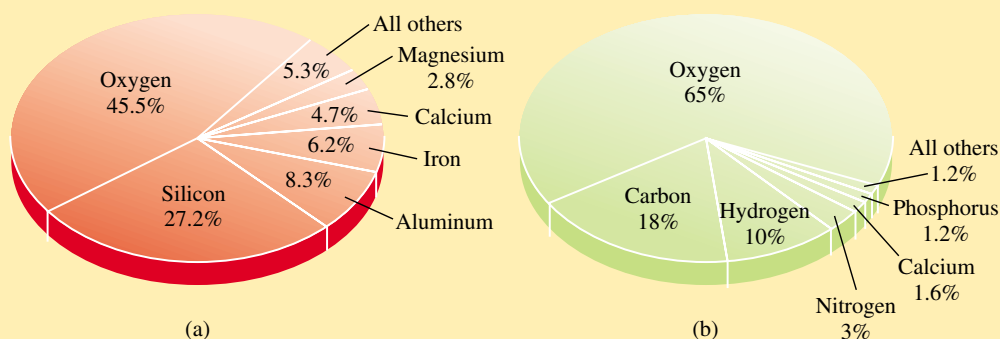
Essential Elements in the Human Body

| ELEMENT | PERCENT BY MASS* | ELEMENT | PERCENT BY MASS* |
|------------|------------------|-----------|------------------|
| Oxygen | 65 | Sodium | 0.1 |
| Carbon | 18 | Magnesium | 0.05 |
| Hydrogen | 10 | Iron | <0.05 |
| Nitrogen | 3 | Cobalt | <0.05 |
| Calcium | 1.5 | Copper | <0.05 |
| Phosphorus | 1.2 | Zinc | <0.05 |
| Potassium | 0.2 | Iodine | <0.05 |
| Sulfur | 0.2 | Selenium | <0.01 |
| Chlorine | 0.2 | Fluorine | <0.01 |

*Percent by mass gives the mass of the element in grams present in a 100-g sample.

chemical names. Familiar examples are listed in Table 2.7.

(a) Natural abundance of the elements in percent by mass. For example, oxygen's abundance is 45.5 percent. This means that in a 100-g sample of Earth's crust there are, on the



3. An atom consists of a very dense central nucleus containing protons and neutrons, with electrons moving about the nucleus at a relatively large distance from it.
4. Protons are positively charged, neutrons have no charge, and electrons are negatively charged. Protons and neutrons have roughly the same mass, which is about 1840 times greater than the mass of an electron.
5. The atomic number of an element is the number of protons in the nucleus of an atom of the element; it determines the identity of an element. The mass number is the sum of the number of protons and the number of neutrons in the nucleus.
6. Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons.
7. Chemical formulas combine the symbols for the constituent elements with whole-number subscripts to show the type and number of atoms contained in the smallest unit of a compound.
8. The molecular formula conveys the specific number and type of atoms combined in each molecule of a compound. The empirical formula shows the simplest ratios of the atoms combined in a molecule.
9. Chemical compounds are either molecular compounds (in which the smallest units are discrete, individual molecules) or ionic compounds (in which positive and negative ions are held together by mutual attraction). Ionic compounds are made up of cations and anions, formed when atoms lose and gain electrons, respectively.
10. The names of many inorganic compounds can be deduced from a set of simple rules. The formulas can be written from the names of the compounds.

KEY WORDS

| | | | |
|-------------------------------------|--------------------------------|------------------------------------|----------------------------|
| Acid, p. 58 | Cation, p. 48 | Law of conservation of mass, p. 39 | Noble gases, p. 47 |
| Alkali metals, p. 46 | Chemical formula, p. 50 | Law of definite proportions, p. 38 | Nonmetal, p. 46 |
| Alkaline earth metals, p. 46 | Diatomic molecule, p. 48 | Law of multiple proportions, p. 38 | Nucleus, p. 43 |
| Allotrope, p. 50 | Electron, p. 40 | Mass number (<i>A</i>), p. 45 | Oxoacid, p. 58 |
| Alpha (α) particles, p. 41 | Empirical formula, p. 50 | Metal, p. 46 | Oxoanion, p. 59 |
| Alpha (α) rays, p. 41 | Families, p. 46 | Metalloid, p. 46 | Period, p. 46 |
| Anion, p. 48 | Gamma (γ) rays, p. 41 | Molecular formula, p. 50 | Periodic table, p. 46 |
| Atom, p. 39 | Groups, p. 46 | Molecule, p. 48 | Polyatomic ion, p. 49 |
| Atomic number (<i>Z</i>), p. 45 | Halogens, p. 47 | Monatomic ion, p. 49 | Polyatomic molecule, p. 48 |
| Base, p. 60 | Hydrate, p. 60 | Neutron, p. 44 | Proton, p. 43 |
| Beta (β) particles, p. 41 | Ion, p. 48 | | Radiation, p. 39 |
| Beta (β) rays, p. 41 | Ionic compound, p. 49 | | Radioactivity, p. 41 |
| Binary compound, p. 54 | Isotope, p. 45 | | Ternary compound, p. 54 |

QUESTIONS AND PROBLEMS

STRUCTURE OF THE ATOM

Review Questions

- 2.1 Define the following terms: (a) α particle, (b) β particle, (c) γ ray, (d) X ray.
- 2.2 Name the types of radiation known to be emitted by radioactive elements.
- 2.3 Compare the properties of the following: α particles, cathode rays, protons, neutrons, electrons.
- 2.4 What is meant by the term “fundamental particle”?
- 2.5 Describe the contributions of the following scientists to our knowledge of atomic structure: J. J. Thomson,

R. A. Millikan, Ernest Rutherford, James Chadwick.

- 2.6 Describe the experimental basis for believing that the nucleus occupies a very small fraction of the volume of the atom.

Problems

- 2.7 The diameter of a neutral helium atom is about 1×10^2 pm. Suppose that we could line up helium atoms side by side in contact with one another. Approximately how many atoms would it take to make the distance from end to end 1 cm?

- 2.8** Roughly speaking, the radius of an atom is about 10,000 times greater than that of its nucleus. If an atom were magnified so that the radius of its nucleus became 2.0 cm, about the size of a marble, what would be the radius of the atom in miles? (1 mi = 1609 m.)

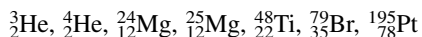
ATOMIC NUMBER, MASS NUMBER, AND ISOTOPES

Review Questions

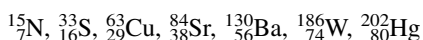
- 2.9** Use the helium-4 isotope to define atomic number and mass number. Why does a knowledge of atomic number enable us to deduce the number of electrons present in an atom?
- 2.10** Why do all atoms of an element have the same atomic number, although they may have different mass numbers?
- 2.11** What do we call atoms of the same elements with different mass numbers?
- 2.12** Explain the meaning of each term in the symbol A_ZX .

Problems

- 2.13** What is the mass number of an iron atom that has 28 neutrons?
- 2.14** Calculate the number of neutrons of Pu-239.
- 2.15** For each of the following species, determine the number of protons and the number of neutrons in the nucleus:



- 2.16** Indicate the number of protons, neutrons, and electrons in each of the following species:



- 2.17** Write the appropriate symbol for each of the following isotopes: (a) $Z = 11$, $A = 23$; (b) $Z = 28$, $A = 64$.
- 2.18** Write the appropriate symbol for each of the following isotopes: (a) $Z = 74$, $A = 186$; (b) $Z = 80$, $A = 201$.

THE PERIODIC TABLE

Review Questions

- 2.19** What is the periodic table, and what is its significance in the study of chemistry?
- 2.20** State two differences between a metal and a nonmetal.
- 2.21** Write the names and symbols for four elements in each of the following categories: (a) nonmetal, (b) metal, (c) metalloid.
- 2.22** Define, with two examples, the following terms: (a) alkali metals, (b) alkaline earth metals, (c) halogens, (d) noble gases.

Problems

- 2.23** Elements whose names end with *ium* are usually metals; sodium is one example. Identify a nonmetal whose name also ends with *ium*.
- 2.24** Describe the changes in properties (from metals to nonmetals or from nonmetals to metals) as we move (a) down a periodic group and (b) across the periodic table.
- 2.25** Consult a handbook of chemical and physical data (ask your instructor where you can locate a copy of the handbook) to find (a) two metals less dense than water, (b) two metals more dense than mercury, (c) the densest known solid metallic element, (d) the densest known solid nonmetallic element.
- 2.26** Group the following elements in pairs that you would expect to show similar chemical properties: K, F, P, Na, Cl, and N.

MOLECULES AND IONS

Review Questions

- 2.27** What is the difference between an atom and a molecule?
- 2.28** What are allotropes? Give an example. How are allotropes different from isotopes?
- 2.29** Describe the two commonly used molecular models.
- 2.30** Give an example of each of the following: (a) a monatomic cation, (b) a monatomic anion, (c) a polyatomic cation, (d) a polyatomic anion.

Problems

- 2.31** Identify the following as elements or compounds: NH_3 , N_2 , S_8 , NO , CO , CO_2 , H_2 , SO_2 .
- 2.32** Give two examples of each of the following: (a) a diatomic molecule containing atoms of the same element, (b) a diatomic molecule containing atoms of different elements, (c) a polyatomic molecule containing atoms of the same element, (d) a polyatomic molecule containing atoms of different elements.
- 2.33** Give the number of protons and electrons in each of the following common ions: Na^+ , Ca^{2+} , Al^{3+} , Fe^{2+} , I^- , F^- , S^{2-} , O^{2-} , and N^{3-} .
- 2.34** Give the number of protons and electrons in each of the following common ions: K^+ , Mg^{2+} , Fe^{3+} , Br^- , Mn^{2+} , C^{4-} , Cu^{2+} .

CHEMICAL FORMULAS

Review Questions

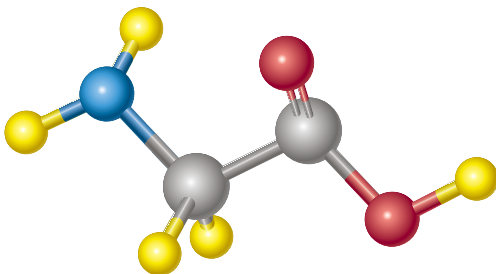
- 2.35** What does a chemical formula represent? What is the ratio of the atoms in the following molecular formula?

las? (a) NO, (b) NCl₃, (c) N₂O₄, (d) P₄O₆

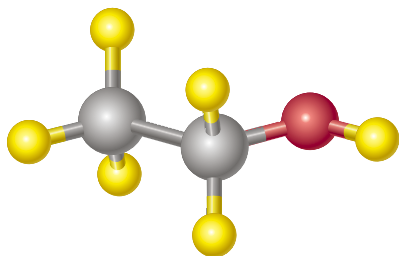
- 2.36** Define molecular formula and empirical formula. What are the similarities and differences between the empirical formula and molecular formula of a compound?
- 2.37** Give an example of a case in which two molecules have different molecular formulas but the same empirical formula.
- 2.38** What does P₄ signify? How does it differ from 4P?
- 2.39** What is an ionic compound? How is electrical neutrality maintained in an ionic compound?
- 2.40** Explain why the chemical formulas of ionic compounds are always the same as their empirical formulas.

Problems

- 2.41** What are the empirical formulas of the following compounds? (a) C₂N₂, (b) C₆H₆, (c) C₉H₂₀, (d) P₄O₁₀, (e) B₂H₆
- 2.42** What are the empirical formulas of the following compounds? (a) Al₂Br₆, (b) Na₂S₂O₄, (c) N₂O₅, (d) K₂Cr₂O₇
- 2.43** Write the molecular formula of glycine, an amino acid present in proteins. The color codes are: gray (carbon), blue (nitrogen), red (oxygen), and yellow (hydrogen).



- 2.44** Write the molecular formula of ethanol. The color codes are: gray (carbon), red (oxygen), and yellow (hydrogen).



- 2.45** Which of the following compounds are likely to be ionic? Which are likely to be molecular? SiCl₄, LiF, BaCl₂, B₂H₆, KCl, C₂H₄

- 2.46** Which of the following compounds are likely to be ionic? Which are likely to be molecular? CH₄, NaBr, BaF₂, CCl₄, ICl, CsCl, NF₃

NAMING INORGANIC COMPOUNDS

Review Questions

- 2.47** What is the difference between inorganic compounds and organic compounds?
- 2.48** What are the four major categories of inorganic compounds?
- 2.49** Give an example each for a binary compound and a ternary compound.
- 2.50** What is the Stock system? What are its advantages over the older system of naming cations?
- 2.51** Explain why the formula HCl can represent two different chemical systems.
- 2.52** Define acids, bases, oxoacids, oxoanions, and hydrates.

Problems

- 2.53** Name the following compounds: (a) KH₂PO₄, (b) K₂HPO₄, (c) HBr (gas), (d) HBr (in water), (e) Li₂CO₃, (f) K₂Cr₂O₇, (g) NH₄NO₂, (h) HIO₃, (i) PF₅, (j) P₄O₆, (k) CdI₂, (l) SrSO₄, (m) Al(OH)₃.
- 2.54** Name the following compounds: (a) KClO, (b) Ag₂CO₃, (c) HNO₂, (d) KMnO₄, (e) CsClO₃, (f) KNH₄SO₄, (g) FeO, (h) Fe₂O₃, (i) TiCl₄, (j) NaH, (k) Li₃N, (l) Na₂O, (m) Na₂O₂.
- 2.55** Write the formulas for the following compounds: (a) rubidium nitrite, (b) potassium sulfide, (c) sodium hydrogen sulfide, (d) magnesium phosphate, (e) calcium hydrogen phosphate, (f) potassium dihydrogen phosphate, (g) iodine heptafluoride, (h) ammonium sulfate, (i) silver perchlorate, (j) boron trichloride.
- 2.56** Write the formulas for the following compounds: (a) copper(I) cyanide, (b) strontium chlorite, (c) perbromic acid, (d) hydroiodic acid, (e) disodium ammonium phosphate, (f) lead(II) carbonate, (g) tin(II) fluoride, (h) tetraphosphorus decasulfide, (i) mercury(II) oxide, (j) mercury(I) iodide, (k) selenium hexafluoride.

ADDITIONAL PROBLEMS

- 2.57** A sample of an uranium compound is found to be losing mass gradually. Explain what is happening to the sample.
- 2.58** In which one of the following pairs do the two species resemble each other most closely in chemical properties? Explain. (a) ¹H and ¹H⁺, (b) ¹⁴N and ¹⁴N³⁻, (c) ¹²C and ¹³C

- 2.59** One isotope of a metallic element has mass number 65 and 35 neutrons in the nucleus. The cation derived from the isotope has 28 electrons. Write the symbol for this cation.
- 2.60** One isotope of a nonmetallic element has mass number 127 and 74 neutrons in the nucleus. The anion derived from the isotope has 54 electrons. Write the symbol for this anion.
- 2.61** The table below gives numbers of electrons, protons, and neutrons in atoms or ions of a number of elements. Answer the following: (a) Which of the species are neutral? (b) Which are negatively charged? (c) Which are positively charged? (d) What are the conventional symbols for all the species?

| ATOM OR ION OF ELEMENT | A | B | C | D | E | F | G |
|------------------------|---|----|----|----|----|---|----|
| Number of electrons | 5 | 10 | 18 | 28 | 36 | 5 | 9 |
| Number of protons | 5 | 7 | 19 | 30 | 35 | 5 | 9 |
| Number of neutrons | 5 | 7 | 20 | 36 | 46 | 6 | 10 |

- 2.62** What is wrong with or ambiguous about the phrase “four molecules of NaCl”?
- 2.63** The following phosphorus sulfides are known: P_4S_3 , P_4S_7 , and P_4S_{10} . Do these compounds obey the law of multiple proportions?
- 2.64** Which of the following are elements, which are molecules but not compounds, which are compounds but not molecules, and which are both compounds and molecules? (a) SO_2 , (b) S_8 , (c) Cs, (d) N_2O_5 , (e) O, (f) O_2 , (g) O_3 , (h) CH_4 , (i) KBr, (j) S, (k) P_4 , (l) LiF
- 2.65** Why is magnesium chloride ($MgCl_2$) not called magnesium(II) chloride?
- 2.66** Some compounds are better known by their common names than by their systematic chemical names. Give the chemical formulas of the following substances: (a) dry ice, (b) table salt, (c) laughing gas, (d) marble (chalk, limestone), (e) quicklime, (f) slaked lime, (g) baking soda, (h) washing soda, (i) gypsum, (j) milk of magnesia.
- 2.67** Fill in the blanks in the following table:

| | | | | | |
|------------|---|-----------------------|----|-----|-----|
| SYMBOL | | ${}^{54}_{26}Fe^{2+}$ | | | |
| PROTONS | 5 | | | 79 | 86 |
| NEUTRONS | 6 | | 16 | 117 | 136 |
| ELECTRONS | 5 | | 18 | 79 | |
| NET CHARGE | | | -3 | | 0 |

- 2.68** (a) Which elements are most likely to form ionic compounds? (b) Which metallic elements are most likely to form cations with different charges?
- 2.69** What ion is each of the following most likely to form in ionic compounds: (a) Li, (b) S, (c) I, (d) N, (e) Al, (f) Cs, (g) Mg?
- 2.70** Which of the following symbols provides more information about the atom: ${}^{23}Na$ or ${}_{11}Na$? Explain.
- 2.71** Write the chemical formulas and names of acids that contain Group 7A elements. Do the same for elements in Groups 3A, 4A, 5A, and 6A.
- 2.72** Of the 112 elements known, only two are liquids at room temperature ($25^\circ C$). What are they? (*Hint*: One element is a familiar metal and the other element is in Group 7A.)
- 2.73** For the noble gases (the Group 8A elements), 4He , ${}^{20}_{10}Ne$, ${}^{40}_{18}Ar$, ${}^{84}_{36}Kr$, and ${}^{132}_{54}Xe$, (a) determine the number of protons and neutrons in the nucleus of each atom, and (b) determine the ratio of neutrons to protons in the nucleus of each atom. Describe any general trend you discover in the way this ratio changes with increasing atomic number.
- 2.74** List the elements that exist as gases at room temperature. (*Hint*: These elements can be found in Groups 5A, 6A, 7A, and 8A.)
- 2.75** The Group 1B metals, Cu, Ag, and Au, are called coinage metals. What chemical properties make them specially suitable for making coins and jewels?
- 2.76** The elements in Group 8A of the periodic table are called noble gases. Can you suggest what “noble” means in this context?
- 2.77** The formula for calcium oxide is CaO. What are the formulas for magnesium oxide and strontium oxide?
- 2.78** A common mineral of barium is barytes, or barium sulfate ($BaSO_4$). Because elements in the same periodic group have similar chemical properties, we might expect to find some radium sulfate ($RaSO_4$) mixed with barytes since radium is the last member of Group 2A. However, the only source of radium compounds in nature is in uranium minerals. Why?
- 2.79** List five elements each that are (a) named after places, (b) named after people, (c) named after a color. (*Hint*: See Appendix 1.)
- 2.80** Name the only country that is named after an element. (*Hint*: This country is in South America.)
- 2.81** Fluorine reacts with hydrogen (H) and deuterium (D) to form hydrogen fluoride (HF) and deuterium fluoride (DF), where deuterium (2H) is an isotope of hydrogen. Would a given amount of fluorine react with different masses of the two hydrogen isotopes? Does this violate the law of definite proportion? Explain.

- 2.82** Predict the formula and name of a binary compound formed from the following elements: (a) Na and H, (b) B and O, (c) Na and S, (d) Al and F, (e) F and O, (f) Sr and Cl.
- 2.83** Identify each of the following elements: (a) a halogen whose anion contains 36 electrons, (b) a radioactive noble gas with 86 protons, (c) a Group 6A element whose anion contains 36 electrons, (d) An alkali metal cation that contains 36 electrons, (e) a Group 4A cation that contains 80 electrons.

Answers to Practice Exercises: **2.1** 29 protons, 34 neutrons, and 29 electrons. **2.2** CHCl_3 . **2.3** $\text{C}_4\text{H}_5\text{N}_2\text{O}$. **2.4** (a) Lead(II) oxide, (b) lithium sulfite. **2.5** (a) Rb_2SO_4 , (b) BaH_2 . **2.6** (a) Nitrogen trifluoride, (b) dichlorine heptoxide. **2.7** (a) SF_4 , (b) N_2O_5 . **2.8** (a) Hypobromous acid, (b) hydrogen sulfate ion.

AP Chemistry Summer Assignment Checklist:

Carefully read and highlighted the two assigned text chapters that were given out as part of this package.

Answered the assigned questions at the end of each of the two text chapters; CLEARLY and LEGIBLY showed a detailed, step-by-step answer to each question, including units, cancellations, and sketches where appropriate.

Memorized the element names, polyatomic ions, and solubility rules.

Know the fundamental cause of the solubility rules.

Completed the compound naming and formula writing worksheet (either on the worksheet or on separate paper).