

GLOSSARY

The number in parentheses is the number of the section in which the term first appears.

- A**
- absolute temperature scale.** A temperature scale on which absolute zero (0 K) is the lowest temperature (also called the Kelvin temperature scale). (5.3)
- absolute zero.** Theoretically the lowest attainable temperature. (5.3)
- accuracy.** The closeness of a measurement to the true value of the quantity that is being measured. (1.6)
- acid.** A substance that yields hydrogen ions (H^+) when dissolved in water. (2.7)
- acid ionization constant (K_a).** The equilibrium constant for acid ionization. (16.5)
- actinide series.** Elements that have incompletely filled 5f subshells or readily give rise to cations that have incompletely filled 5f subshells. (7.9)
- activated complex.** The species temporarily formed by reactant molecules as a result of a collision before they form the product. (14.4)
- activation energy.** The minimum amount of energy required to initiate a chemical reaction. (14.4)
- activity series.** A summary of the results of many possible displacement reactions. (4.4)
- actual yield.** The amount of product actually obtained in a reaction. (3.10)
- addition reaction.** A reaction in which one molecule is added to another. (11.2)
- adhesion.** Attraction between unlike molecules. (12.3)
- alcohol.** An organic compound containing the hydroxyl group ($-\text{OH}$). (11.4)
- aldehydes.** Compounds with a carbonyl functional group and the general formula RCHO , in which R is an H atom, an alkyl, or an aromatic hydrocarbon group. (11.4)
- aliphatic hydrocarbons.** Hydrocarbons that do not contain the benzene group or the benzene ring. (11.1)
- alkali metals.** The Group 1A elements (Li, Na, K, Rb, Cs, and Fr). (2.4)
- alkaline earth metals.** The Group 2A elements (Be, Mg, Ca, Sr, Ba, and Ra). (2.4)
- alkanes.** Hydrocarbons having the general formula $\text{C}_n\text{H}_{2n+2}$, in which $n = 1, 2, \dots$ (11.2)
- alkenes.** Hydrocarbons that contain one or more carbon-carbon double bonds. They have the general formula C_nH_{2n} in which $n = 2, 3, \dots$ (11.2)
- alkynes.** Hydrocarbons that contain one or more carbon-carbon triple bonds. They have the general formula $\text{C}_n\text{H}_{2n-2}$, in which $n = 2, 3, \dots$ (11.2)
- allotropes.** Two or more forms of the same element that differ significantly in chemical and physical properties. (2.6)
- alpha particles.** See alpha rays.
- alpha (α) rays.** Helium ions with a charge of $+2$. (2.2)
- amines.** Organic bases that have the functional group $-\text{NR}_2$, in which R may be H, an alkyl group, or an aromatic hydrocarbon group. (11.4)
- amino acid.** A special kind of carboxylic acid that contains at least one carboxyl group ($-\text{COOH}$) and at least one amino group ($-\text{NH}_2$). (22.3)
- amorphous solid.** A solid that lacks a regular three-dimensional arrangement of atoms or molecules. (12.5)
- amphoteric oxide.** An oxide that exhibits both acidic and basic properties. (8.6)
- amplitude.** The vertical distance from the middle of a wave to the peak or trough. (7.1)
- anion.** An ion with a net negative charge. (2.5)
- anode.** The electrode at which oxidation occurs. (19.2)
- antibonding molecular orbital.** A molecular orbital that is of higher energy and lower stability than the atomic orbitals from which it was formed. (10.6)
- aqueous solution.** A solution in which the solvent is water. (4.1)
- aromatic hydrocarbon.** A hydrocarbon that contains one or more benzene rings. (11.1)
- atmospheric pressure.** The pressure exerted by Earth's atmosphere. (5.2)
- atom.** The basic unit of an element that can enter into chemical combination. (2.2)
- atomic mass.** The mass of an atom in atomic mass units. (3.1)
- atomic mass unit (amu).** A mass exactly equal to one-twelfth the mass of one carbon-12 atom. (3.1)
- atomic number (Z).** The number of protons in the nucleus of an atom. (2.3)
- atomic orbital.** The wave function of an electron in an atom. (7.5)
- atomic radius.** One-half the distance between the nuclei in two adjacent atoms of the same element in a metal. For elements that exist as diatomic units, the atomic radius is one-half the distance between the nuclei of two atoms in a particular molecule. (8.3)
- Aufbau principle.** As protons are added one by one to the nucleus to build up the elements, electrons similarly are added to the atomic orbitals. (7.9)
- Avogadro's law.** At constant pressure and temperature, the volume of a gas is directly proportional to the number of moles of the gas present. (5.3)
- Avogadro's number (N_A).** 6.022×10^{23} ; the number of particles in a mole. (3.2)
- B**
- barometer.** An instrument that measures atmospheric pressure. (5.2)
- base.** A substance that yields hydroxide ions (OH^-) when dissolved in water. (2.7)
- base ionization constant (K_b).** The equilibrium constant for the ionization of a base. (16.6)
- battery.** A galvanic cell or a series of several connected galvanic cells that can be used as a source of direct electric current at a constant voltage. (19.6)
- beta particles.** See beta rays.
- beta (β) rays.** Streams of electrons emitted during the decay of certain radioactive substances. (2.2)
- bimolecular reaction.** An elementary step involving two molecules that is part of a reaction mechanism. (14.5)
- binary compounds.** Compounds containing just two elements. (2.7)
- boiling point.** The temperature at which the vapor pressure of a liquid is equal to the external atmospheric pressure. (12.6)
- bond enthalpy.** The enthalpy change required to break a bond in a mole of gaseous molecules. (9.10)
- bond length.** The distance between the centers of two bonded atoms in a molecule. (9.4)
- bond order.** The difference between the number of electrons in bonding molecular orbitals and antibonding molecular orbitals, divided by two. (10.6)
- bonding molecular orbital.** A molecular orbital that is of lower energy and greater stability than the atomic orbitals from which it was formed. (10.6)
- Born-Haber cycle.** The cycle that relates lattice energies of ionic compounds to ionization energies, electron affinities, and other atomic and molecular properties. (9.3)

boundary surface diagram. Diagram of the region containing about 90 percent of the electron density in an atomic orbital. (7.7)

Boyle's law. The volume of a fixed amount of gas is inversely proportional to the gas pressure at constant temperature. (5.3)

breeder reactor. A nuclear reactor that produces more fissionable material than it uses. (21.5)

Brønsted acid. A substance capable of donating a proton in a reaction. (4.3)

Brønsted base. A substance capable of accepting a proton in a reaction. (4.3)

buffer solution. A solution of (a) a weak acid or base and (b) its salt; both components must be present. A buffer solution has the ability to resist changes in pH when small amounts of either acid or base are added to it. (17.2)

C

calorimetry. The measurement of heat changes. (6.5)

carboxylic acids. Acids that contain the carboxyl group ($-\text{COOH}$). (11.4)

catalyst. A substance that increases the rate of a chemical reaction by providing an alternate reaction pathway without being consumed during the reaction. (14.6)

cathode. The electrode at which reduction occurs. (19.2)

cation. An ion with a net positive charge. (2.5)

cell voltage. Difference in electrical potential between the anode and the cathode of a galvanic cell. (19.2)

Charles's and Gay-Lussac's law. See Charles's law.

Charles's law. The volume of a fixed amount of gas is directly proportional to the absolute temperature of the gas when the pressure is held constant. (5.3)

chelating agent. A substance that forms complex ions with metal ions in solution. (20.2)

chemical energy. Energy stored within the structural units of chemical substances. (6.1)

chemical equation. An equation that uses chemical symbols to show what happens during a chemical reaction. (3.7)

chemical equilibrium. A state in which the rates of the forward and reverse reactions are equal and no net changes can be observed. (4.1, 15.1)

chemical formula. An expression showing the chemical composition of a compound in terms of the symbols for the atoms of the elements involved. (2.6)

chemical kinetics. The area of chemistry concerned with the speeds, or rates, at which chemical reactions occur. (14.1)

chemical property. Any property of a substance that cannot be studied without converting the substance into some other substance. (1.4)

chemical reaction. Chemical change. (3.7)

chemistry. The science that studies the properties of substances and how substances react with one another. (1.3)

chiral. Compounds or ions that are not superimposable with their mirror images. (11.5)

closed system. A system that allows the exchange of energy (usually in the form of heat) but not mass with its surroundings. (6.2)

closest packing. The most efficient arrangements for packing atoms, molecules, or ions in a crystal. (12.4)

cohesion. The intermolecular attraction between like molecules. (12.3)

colligative properties. Properties of solutions that depend on the number of solute particles in solution and not on the nature of the solute. (13.6)

combination reaction. A reaction in which two or more substances combine to form a single product. (4.4)

combustion reaction. A reaction in which a substance reacts with oxygen, usually with the release of heat and light to produce a flame. (4.4)

complex ion. An ion containing a central metal cation bonded to one or more molecules or ions. (17.7)

compound. A substance composed of two or more elements chemically united in fixed proportions. (1.3)

concentration of a solution. The amount of solute present in a given quantity of solution. (4.5)

condensation. The phenomenon of going from the gaseous state to the liquid state. (12.6)

condensation reaction. The joining of two molecules and the elimination of a small molecule, usually water. (11.4)

conformations. Different spatial arrangements of a molecule that are generated by rotation about single bonds. (11.2)

conjugate acid-base pair. An acid and its conjugate base or a base and its conjugate acid. (16.1)

coordinate covalent bond. A bond in which the pair of electrons is supplied by one of the two bonded atoms. (9.9)

coordination compound. A neutral species containing a complex ion. (20.2)

coordination number. In a crystal lattice it is defined as the number of atoms (or ions) surrounding an atom (or ion) (12.4). In coordination compounds it is defined as the number of donor atoms surrounding the central metal atom in a complex. (20.2)

copolymer. A polymer containing two or more different monomers. (22.2)

corrosion. The deterioration of metals by an electrochemical process. (19.7)

Coulomb's law. The potential energy between two ions is directly proportional

to the product of their charges and inversely proportional to the distance between them. (9.3)

covalent bond. A bond in which two electrons are shared by two atoms. (9.4)

covalent compounds. Compounds containing only covalent bonds. (9.4)

critical mass. The minimum mass of fissionable material required to generate a self-sustaining nuclear chain reaction. (21.5)

critical pressure (P_c). The minimum pressure necessary to bring about liquefaction at the critical temperature. (12.6)

critical temperature (T_c). The temperature above which a gas will not liquefy. (12.6)

crystal-field splitting. The energy difference between two sets of d orbitals of a metal atom in the presence of ligands. (20.4)

crystalline solid. A solid that possesses rigid and long-range structural order; its atoms, molecules, or ions occupy specific positions. (12.4)

crystallization. The process in which dissolved solute comes out of solution and forms crystals. (13.1)

cycloalkanes. Hydrocarbons having the general formula C_nH_{2n} in which $n = 3, 4, \dots$ (11.2)

D

Dalton's law of partial pressures. The total pressure of a mixture of gases is just the sum of the pressures that each gas would exert if it were present alone. (5.5)

decomposition reaction. The breakdown of a compound into two or more components. (4.4)

delocalized molecular orbital. A molecular orbital that is not confined between two adjacent bonding atoms but actually extends over three or more atoms. (11.3)

denatured protein. Protein that does not exhibit normal biological activities. (22.3)

density. The mass of a substance divided by its volume. (1.5)

deoxyribonucleic acid (DNA). A type of nucleic acid. (22.4)

deposition. The process in which vapor molecules are converted directly to the solid phase. (12.6)

diagonal relationship. Similarities between pairs of elements in different groups and periods of the periodic table. (8.6)

diamagnetic. Repelled by a magnet; a diamagnetic substance contains only paired electrons. (7.8)

diatomic molecule. A molecule that consists of two atoms. (2.5)

diffusion. The gradual mixing of molecules of one gas with the molecules of another by virtue of their kinetic properties. (5.6)

dilution. A procedure for preparing a less concentrated solution from a more concentrated solution. (4.5)

dipole moment. The product of charge and the distance between the charges in a molecule. (10.2)

dipole-dipole forces. Forces that act between polar molecules. (12.2)

diprotic acid. Each unit of the acid yields two hydrogen ions. (4.3)

dispersion forces. The attractive forces that arise as a result of temporary dipoles induced in the atoms or molecules. (12.2)

displacement reaction. A reaction in which an atom or an ion in a compound is replaced by an atom of another element. (4.4)

donor atom. The atom in a ligand that is bonded directly to the metal atom. (20.2)

double bond. A covalent bond in which two atoms share two pairs of electrons. (9.4)

dynamic equilibrium. The condition in which the rate of a forward process is exactly balanced by the rate of the reverse process. (12.6)

E

effective nuclear charge (Z_{eff}). The nuclear charge felt by an electron when both the actual charge (Z) and the repulsive effect (shielding) of the other electrons are taken into account. (8.3)

effusion. The process by which a gas under pressure escapes from one compartment of a container to another by passing through a small opening. (5.6)

electrochemistry. The branch of chemistry that deals with the interconversion of electrical energy and chemical energy. (19.1)

electrolysis. A process in which electrical energy is used to bring about a nonspontaneous chemical reaction. (19.8)

electrolyte. A substance that, when dissolved in water, results in a solution that can conduct electricity. (4.1)

electrolytic cell. An apparatus for carrying out electrolysis. (19.8)

electromagnetic radiation. The emission and transmission of energy in the form of electromagnetic waves. (7.1)

electromagnetic wave. A wave that has an electric field component and a mutually perpendicular magnetic field component. (7.1)

electromotive force (emf) (E). The voltage difference between electrodes. (19.2)

electron. A subatomic particle that has a very low mass and carries a single negative electric charge. (2.2)

electron affinity. The negative of the energy change that takes place when an electron is accepted by an atom (or an ion) in the gaseous state. (8.5)

electron configuration. The distribution of electrons among the various orbitals in an atom or molecule. (7.8)

electron density. The probability that an electron will be found at a particular region in an atomic orbital. (7.5)

electronegativity. The ability of an atom to attract electrons toward itself in a chemical bond. (9.5)

element. A substance that cannot be separated into simpler substances by chemical means. (1.3)

elementary steps. A series of simple reactions that represent the overall progress of a reaction at the molecular level. (14.5)

emission spectrum. The continuous or line spectrum of electromagnetic radiation emitted by a substance. (7.3)

empirical formula. An expression using chemical symbols to show the types of elements in a substance and the simplest ratios of the different kinds of atoms. (2.6)

enantiomers. Compounds and their nonsuperimposable mirror images. (11.5)

end point. Occurs in a titration when the indicator changes color. (17.4)

endothermic processes. Processes that absorb heat from the surroundings. (6.2)

energy. The capacity to do work or to produce change. (6.1)

enthalpy (H). A thermodynamic quantity used to describe heat changes taking place at constant pressure. (6.4)

enthalpy of reaction (ΔH). The difference between the enthalpies of the products and the enthalpies of the reactants. (6.4)

entropy (S). A measure of how spread out or dispersed the energy of a system is among the different possible ways that the system can contain energy. (18.3)

enzyme. A biological catalyst. (14.6)

equilibrium constant. A number equal to the ratio of the equilibrium concentrations of products to the equilibrium concentrations of reactants, each raised to the power of its stoichiometric coefficient. (15.1)

equilibrium vapor pressure. The vapor pressure measured for a dynamic equilibrium of condensation and evaporation. (12.6)

equivalence point. The point at which an acid is completely reacted with or neutralized by a base. (4.6)

esters. Compounds that have the general formula RCOOR' , in which R can be H or an alkyl group or an aromatic hydrocarbon group and R' is an alkyl group or an aromatic hydrocarbon group. (11.4)

ether. An organic compound containing the $\text{R}-\text{O}-\text{R}'$ linkage, in which R and R' are alkyl and/or aromatic hydrocarbon groups. (11.4)

evaporation. The escape of molecules from the surface of a liquid; also called vaporization. (12.6)

excess reagent. A reactant present in a quantity greater than necessary to react

with the amount of the limiting reagent present. (3.9)

excited level (or state). A state that has higher energy than the ground state of the system. (7.3)

exothermic processes. Processes that give off heat to the surroundings. (6.2)

extensive property. A property that depends on how much matter is being considered. (1.4)

F

family. The elements in a vertical column of the periodic table. (2.4)

Faraday constant (F). Charge contained in 1 mole of electrons, equivalent to 96,485 coulombs. (19.4)

first law of thermodynamics. Energy can be converted from one form to another, but cannot be created or destroyed. (6.3)

first-order reaction. A reaction whose rate depends on reactant concentration raised to the first power. (14.3)

formal charge. The electrical charge difference between the number of valence electrons in an isolated atom and the number of electrons assigned to that atom in a Lewis structure. (9.7)

formation constant (K_f). The equilibrium constant for the complex ion formation. (17.7)

free energy (G). The energy available to do useful work. (18.5)

frequency (ν). The number of waves that pass through a particular point per unit time. (7.1)

fuel cell. A galvanic cell that requires a continuous supply of reactants to keep functioning. (19.6)

functional group. That part of a molecule characterized by a special arrangement of atoms that is largely responsible for the chemical behavior of the parent molecule. (11.1)

G

galvanic cell. An electrochemical cell that generates electricity by means of a spontaneous redox reaction. (19.2)

gamma (γ) rays. High-energy radiation. (2.2)

gas constant (R). The constant that appears in the ideal gas equation ($PV = nRT$). It is expressed as $0.08206 \text{ L} \cdot \text{atm/K} \cdot \text{mol}$, or $8.314 \text{ J/K} \cdot \text{mol}$. (5.4)

geometric isomers. Compounds with the same type and number of atoms and the same chemical bonds but different spatial arrangements; such isomers cannot be interconverted without breaking a chemical bond. (11.2)

Gibbs free energy. See free energy.

Graham's law of diffusion. Under the same conditions of temperature and pressure, the rates of diffusion of gases are inversely proportional to the square roots of their molar masses. (5.6)

gravimetric analysis. An experimental procedure that involves the measurement of mass to identify an unknown component of a substance. (4.6)

ground level (or state). The lowest energy state of a system. (7.3)

group. The elements in a vertical column of the periodic table. (2.4)

H

half-cell reactions. Oxidation and reduction reactions that occur at the electrodes. (19.2)

half-life. The time required for the concentration of a reactant to decrease to half its initial concentration. (14.3)

half-reaction. A reaction that explicitly shows electrons involved in either oxidation or reduction. (4.4)

halogens. The nonmetallic elements in Group 7A (F, Cl, Br, I, and At). (2.4)

heat. Transfer of energy between two bodies that are at different temperatures. (6.2)

heat capacity (*C*). The amount of heat required to raise the temperature of a given quantity of a substance by one degree Celsius. (6.5)

Heisenberg uncertainty principle. It is impossible to know simultaneously both the momentum and the position of a particle with certainty. (7.5)

Henry's law. The solubility of a gas in a liquid is proportional to the pressure of the gas over the solution. (13.5)

Hess's law. When reactants are converted to products, the change in enthalpy is the same whether the reaction takes place in one step or in a series of steps. (6.6)

heterogeneous equilibrium. An equilibrium state in which the reacting species are not all in the same phase. (15.2)

heterogeneous mixture. The individual components of such a mixture remain physically separate and can be seen as separate components. (1.3)

homogeneous equilibrium. An equilibrium condition in which all reacting species are in the same phase. (15.2)

homogeneous mixture. The composition of the mixture is the same throughout the solution. (1.3)

homonuclear diatomic molecule. A diatomic molecule containing atoms of the same element. (10.6)

homopolymer. A polymer that is made from only one type of monomer. (22.2)

Hund's rule. The most stable arrangement of electrons in atomic subshells is the one with the greatest number of parallel spins. (7.8)

hybrid orbitals. Atomic orbitals obtained when two or more nonequivalent orbitals of the same atom combine before covalent bond formation. (10.4)

hybridization. The process of mixing the atomic orbitals in an atom (usually the

central atom) to generate a set of new atomic orbitals before covalent bond formation. (10.4)

hydrates. Compounds that have a specific number of water molecules attached to them. (2.7)

hydration. A process in which an ion or a molecule is surrounded by water molecules arranged in a specific manner. (4.1)

hydrocarbons. Compounds made up of only carbon and hydrogen. (11.1)

hydrogen bond. A special type of dipole-dipole interaction between the hydrogen atom bonded to an atom of a very electronegative element (F, N, O) and another atom of one of the three electronegative elements. (12.2)

hydrogenation. The addition of hydrogen, especially to compounds with double and triple carbon-carbon bonds. (11.2)

hydronium ion. H_3O^+ . (4.3)

hypothesis. A tentative explanation for a set of observations. (1.2)

I

ideal gas. A hypothetical gas whose pressure-volume-temperature behavior can be completely accounted for by the ideal gas equation. (5.4)

ideal gas equation. An equation expressing the relationships among pressure, volume, temperature, and amount of gas ($PV = nRT$, in which R is the gas constant). (5.4)

ideal solution. Any solution that obeys Raoult's law. (13.6)

indicators. Substances that have distinctly different colors in acidic and basic media. (4.6)

induced dipole. The separation of positive and negative charges in an atom (or a nonpolar molecule) caused by the proximity of an ion or a polar molecule. (12.2)

inert complex. A complex ion that undergoes very slow ligand exchange reactions. (20.5)

intensive property. A property that does not depend on how much matter is being considered. (1.4)

intermediate. A species that appears in the mechanism of the reaction (that is, in the elementary steps) but not in the overall balanced equation. (14.5)

intermolecular forces. Attractive forces that exist among molecules. (12.2)

International System of Units. A revised metric system (abbreviated SI) that is widely used in scientific research. (1.5)

intramolecular forces. Forces that hold atoms together in a molecule. (12.2)

ion. An atom or group of atoms that has a net positive or negative charge. (2.5)

ion pair. A species made up of at least one cation and at least one anion held together by electrostatic forces. (13.6)

ion-dipole forces. Forces that operate between an ion and a dipole. (12.2)

ionic bond. The electrostatic force that holds ions together in an ionic compound. (9.2)

ionic compound. Any neutral compound containing cations and anions. (2.5)

ionic equation. An equation that shows dissolved ionic compounds in terms of their free ions. (4.2)

ionic radius. The radius of a cation or an anion as measured in an ionic compound. (8.3)

ionization energy. The minimum energy required to remove an electron from an isolated atom (or an ion) in its ground state. (8.4)

ion-product constant. Product of hydrogen ion concentration and hydroxide ion concentration (both in molarity) at a particular temperature. (16.2)

isoelectronic. Ions, or atoms and ions, that possess the same number of electrons, and hence the same ground-state electron configuration, are said to be isoelectronic. (8.2)

isolated system. A system that does not allow the transfer of either mass or energy to or from its surroundings. (6.2)

isotopes. Atoms having the same atomic number but different mass numbers. (2.3)

J

Joule. Unit of energy given by newtons \times meters. (5.6)

K

Kelvin temperature scale. See absolute temperature scale.

ketones. Compounds with a carbonyl functional group and the general formula $\text{RR}'\text{CO}$, in which R and R' are alkyl and/or aromatic hydrocarbon groups. (11.4)

kinetic energy (KE). Energy available because of the motion of an object. (5.6)

kinetic molecular theory of gases. A theory that describes the physical behavior of gases at the molecular level. (5.6)

L

labile complex. Complexes that undergo rapid ligand exchange reactions. (20.5)

lanthanide series. Elements that have incompletely filled 4f subshells or readily give rise to cations that have incompletely filled 4f subshells. (7.9)

lattice energy. The energy required to completely separate one mole of a solid ionic compound into gaseous ions. (9.3)

lattice points. The positions occupied by atoms, molecules, or ions that define the geometry of a unit cell. (12.4)

law. A concise verbal or mathematical statement of a relationship between

phenomena that is always the same under the same conditions. (1.2)

law of conservation of energy. The total quantity of energy in the universe is constant. (6.1)

law of conservation of mass. Matter can be neither created nor destroyed. (2.1)

law of definite proportions. Different samples of the same compound always contain its constituent elements in the same proportions by mass. (2.1)

law of multiple proportions. If 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 ratios of small whole numbers. (2.1)

Le Châtelier's principle. If an external stress is applied to a system at equilibrium, the system will adjust itself in such a way as to partially offset the stress. (15.4)

Lewis acid. A substance that can accept a pair of electrons. (16.11)

Lewis base. A substance that can donate a pair of electrons. (16.11)

Lewis dot symbol. The symbol of an element with one or more dots that represent the number of valence electrons in an atom of the element. (9.1)

Lewis structure. A representation of covalent bonding using Lewis symbols. Shared electron pairs are shown either as lines or as pairs of dots between two atoms, and lone pairs are shown as pairs of dots on individual atoms. (9.4)

ligand. A molecule or an ion that is bonded to the metal ion in a complex ion. (20.2)

limiting reagent. The reactant used up first in a reaction. (3.9)

line spectrum. Spectrum produced when radiation is absorbed or emitted by a substance only at some wavelengths. (7.3)

liter. The volume occupied by 1 cubic decimeter. (1.5)

lone pairs. Valence electrons that are not involved in covalent bond formation. (9.4)

M

macroscopic properties. Properties that can be measured directly. (1.5)

manometer. A device used to measure the pressure of gases. (5.2)

many-electron atoms. Atoms that contain two or more electrons. (7.5)

mass. A measure of the quantity of matter contained in an object. (1.5)

mass defect. The difference between the mass of an atom and the sum of the masses of its protons, neutrons, and electrons. (21.2)

mass number. The total number of neutrons and protons present in the nucleus of an atom. (2.3)

matter. Anything that occupies space and possesses mass. (1.3)

melting point. The temperature at which solid and liquid phases coexist in equilibrium. (12.6)

metalloid. An element with properties intermediate between those of metals and nonmetals. (2.4)

metals. Elements that are good conductors of heat and electricity and have the tendency to form positive ions in ionic compounds. (2.4)

metathesis reaction. A reaction that involves the exchange of parts between two compounds. (4.2)

microscopic properties. Properties that must be measured indirectly with the aid of a microscope or other special instrument. (1.5)

miscible. Two liquids that are completely soluble in each other in all proportions are said to be miscible. (13.2)

mixture. A combination of two or more substances in which the substances retain their identities. (1.3)

moderator. A substance that can reduce the kinetic energy of neutrons. (21.5)

molality. The number of moles of solute dissolved in 1 kilogram of solvent. (13.3)

molar concentration. See molarity.

molar heat of fusion (ΔH_{fus}). The energy (in kilojoules) required to melt 1 mole of a solid. (12.6)

molar heat of sublimation (ΔH_{sub}). The energy (in kilojoules) required to sublime 1 mole of a solid. (12.6)

molar heat of vaporization (ΔH_{vap}). The energy (in kilojoules) required to vaporize 1 mole of a liquid. (12.6)

molar mass (M). The mass (in grams or kilograms) of 1 mole of atoms, molecules, or other particles. (3.2)

molar solubility. The number of moles of solute in 1 liter of a saturated solution (mol/L). (17.5)

molarity (M). The number of moles of solute in 1 liter of solution. (4.5)

mole (mol). The amount of substance that contains as many elementary entities (atoms, molecules, or other particles) as there are atoms in exactly 12 grams (or 0.012 kilograms) of the carbon-12 isotope. (3.2)

mole fraction. Ratio of the number of moles of one component of a mixture to the total number of moles of all components in the mixture. (5.5)

mole method. An approach for determining the amount of product formed in a reaction. (3.8)

molecular equations. Equations in which the formulas of the compounds are written as though all species existed as molecules or whole units. (4.2)

molecular formula. An expression showing the exact numbers of atoms of each element in a molecule. (2.6)

molecular mass. The sum of the atomic masses (in amu) present in a given molecule. (3.3)

molecular orbital. An orbital that results from the interaction of atomic orbitals of the bonding atoms. (10.6)

molecularity of a reaction. The number of molecules reacting in an elementary step. (14.5)

molecule. An aggregate of at least two atoms in a definite arrangement held together by special forces. (2.5)

monatomic ion. An ion that contains only one atom. (2.5)

monomer. Simple repeating units in a polymer. (22.2)

monoprotic acid. Each unit of the acid yields one hydrogen ion. (4.3)

multiple bonds. Bonds formed when two atoms share two or more pairs of electrons. (9.4)

N

Nernst equation. The relation between the emf of a galvanic cell and the standard emf and the concentrations of the oxidizing and reducing agents. (19.5)

net ionic equation. An equation that includes only the ionic species that actually take part in the reaction. (4.2)

neutralization reaction. A reaction between an acid and a base. (4.3)

neutron. A subatomic particle that bears no net electric charge. Its mass is slightly greater than a proton's. (2.2)

Newton (N). The SI unit for force. (5.2)

noble gas core. The noble gas that most nearly precedes the element being considered; used in writing electron configurations. (7.9)

noble gases. Nonmetallic elements in Group 8A (He, Ne, Ar, Kr, Xe, and Rn). (2.4)

node. A point at which the amplitude of a wave is zero. (7.4)

nonelectrolyte. A substance that, when dissolved in water, gives a solution that is not electrically conducting. (4.1)

nonmetals. Elements that are usually poor conductors of heat and electricity. (2.4)

nonpolar molecule. A molecule that does not possess a dipole moment. (10.2)

nonvolatile. Does not have a measurable vapor pressure. (13.6)

nuclear binding energy. The energy required to break up a nucleus into protons and neutrons. (21.2)

nuclear chain reaction. A self-sustaining sequence of nuclear fission reactions. (21.5)

nuclear fission. The process in which a heavy nucleus (mass number > 200) divides to form small nuclei of intermediate mass and one or more neutrons. (21.5)

nuclear fusion. The combining of small nuclei into larger ones. (21.6)

nuclear reaction. A reaction involving change in an atomic nucleus. (21.1)
nuclear transmutation. The change undergone by a nucleus as a result of bombardment by neutrons or other particles. (21.1)
nucleic acid. High molar mass polymers that play an essential role in protein synthesis. (22.4)
nucleotide. The repeating unit in each strand of a DNA molecule which consists of a base-deoxyribose-phosphate group. (22.4)
nucleus. The central core of an atom. (2.2)

O

octet rule. An atom other than hydrogen tends to form bonds until it is surrounded by eight valence electrons. (9.4)
open system. A system that can exchange mass and energy (usually in the form of heat) with its surroundings. (6.2)
organic chemistry. The branch of chemistry that deals with carbon compounds. (11.1)
osmosis. The net movement of solvent molecules through a semipermeable membrane from a pure solvent or from a dilute solution to a more concentrated solution. (13.6)
osmotic pressure (π). The pressure required to stop osmosis. (13.6)
overvoltage. The additional voltage required to cause electrolysis. (19.8)
oxidation number. The number of charges an atom would have in a molecule if electrons were transferred completely in the direction indicated by the difference in electronegativity. (4.4)
oxidation reaction. The half-reaction that represents the loss of electrons in a redox process. (4.4)
oxidation state. See oxidation number.
oxidation-reduction reaction. See redox reaction.
oxidizing agent. A substance that can accept electrons from another substance or increase the oxidation numbers of another substance. (4.4)
oxoacid. An acid containing hydrogen, oxygen, and another element (the central element). (2.7)
oxoanion. An anion derived from an oxoacid. (2.7)

P

paramagnetic. Attracted to a magnet. A paramagnetic substance contains one or more unpaired electrons. (7.8)
partial pressure. The pressure of one component in a mixture of gases. (5.5)
Pascal (Pa). A pressure of one newton per square meter (1 N/m²). (5.2)
Pauli exclusion principle. No two electrons in an atom can have the same four quantum numbers. (7.8)

percent composition. The percent by mass of each element in a compound. (3.5)
percent ionization. The ratio of ionized acid concentration at equilibrium to the initial concentration of acid. (16.5)
percent by mass. The ratio of the mass of a solute to the mass of the solution, multiplied by 100%. (13.3)
percent yield. The ratio of the actual yield of a reaction to the theoretical yield, multiplied by 100%. (3.10)
period. A horizontal row of the periodic table. (2.4)
periodic table. A tabular arrangement of the elements by similarities in properties and by increasing atomic number. (2.4)
pH. The negative logarithm of the hydrogen ion concentration in an aqueous solution. (16.3)
phase. A homogeneous part of a system that is in contact with other parts of the system but separated from them by a well-defined boundary. (12.6)
phase change. Transformation from one phase to another. (12.6)
phase diagram. A diagram showing the conditions at which a substance exists as a solid, liquid, and vapor. (12.7)
photoelectric effect. A phenomenon in which electrons are ejected from the surface of certain metals exposed to light of at least a certain minimum frequency. (7.2)
photon. A particle of light. (7.2)
physical equilibrium. An equilibrium in which only physical properties change. (15.1)
physical property. Any property of a substance that can be observed without transforming the substance into some other substance. (1.4)
pi bond (π bond). A covalent bond formed by sideways overlapping orbitals; its electron density is concentrated above and below the plane of the nuclei of the bonding atoms. (10.5)
pi molecular orbital. A molecular orbital in which the electron density is concentrated above and below the line joining the two nuclei of the bonding atoms. (10.6)
plasma. A gaseous state of matter consisting of positive ions and electrons. (21.6)
polar covalent bond. In such a bond, the electrons spend more time in the vicinity of one atom than the other. (9.5)
polar molecule. A molecule that possesses a dipole moment. (10.2)
polarimeter. The instrument for studying interaction between plane-polarized light and chiral molecules. (11.5)
polarizability. The ease with which the electron distribution in the atom (or molecule) can be distorted. (12.2)
polyatomic ion. An ion that contains more than one atom. (2.5)

polyatomic molecule. A molecule that consists of more than two atoms. (2.5)
polymer. A molecular compound distinguished by a high molar mass and made up of many repeating units. (22.1)
positron. A particle that has the same mass as the electron but bears a +1 charge. (21.1)
potential energy. Energy available by virtue of an object's position. (6.1)
precipitate. An insoluble solid that separates from a supersaturated solution. (4.2)
precipitation reaction. A reaction characterized by the formation of a precipitate. (4.2)
precision. The closeness of agreement of two or more measurements of the same quantity. (1.6)
pressure. Force applied per unit area. (5.2)
product. The substance formed as a result of a chemical reaction. (3.7)
protein. A polymer of amino acids. (22.3)
proton. A subatomic particle having a single positive electric charge. The mass of a proton is about 1840 times that of an electron. (2.2)

Q

qualitative. Consisting of general observations about the system. (1.2)
qualitative analysis. The determination of the types of ions present in a solution. (17.8)
quantitative. Comprising numbers obtained by various measurements of the system. (1.2)
quantum. The smallest quantity of energy that can be emitted or absorbed in the form of electromagnetic radiation. (7.1)
quantum numbers. Numbers that describe the distribution of electrons in atoms. (7.6)

R

racemic mixture. An equimolar mixture of two enantiomers. (11.5)
radiant energy. Energy transmitted in the form of waves. (6.1)
radiation. The emission and transmission of energy through space in the form of particles and/or waves. (2.2)
radical. A species that contains an unpaired electron. (11.2)
radioactive decay series. A sequence of nuclear reactions that ultimately result in the formation of a stable isotope. (21.3)
radioactivity. The spontaneous breakdown of a nucleus by the emission of particles and/or radiation. (2.2)
Raoult's law. The partial pressure of the solvent over a solution is given by the product of the vapor pressure of the pure solvent and the mole fraction of the solvent in the solution. (13.6)
rare earth series. See lanthanide series.
rare gases. See noble gases.

rate constant (k). Proportionality constant relating reaction rate to the concentrations of reactants. (14.2)

rate law. An expression relating the rate of a reaction to the rate constant and the concentrations of the reactants. (14.2)

rate-determining step. The slowest step in the sequence of steps leading to the formation of products. (14.5)

reactants. The starting substances in a chemical reaction. (3.7)

reaction mechanism. The sequence of elementary steps that leads to product formation. (14.5)

reaction order. The sum of the powers to which all reactant concentrations appearing in the rate law are raised. (14.2)

reaction quotient (Q). A number equal to the ratio of product concentrations to reactant concentrations, each raised to the power of its stoichiometric coefficient at some point other than equilibrium. (15.3)

reaction rate. The change in concentration of reactant or product with time. (14.1)

redox reaction. A reaction in which there is either a transfer of electrons or a change in the oxidation numbers of the substances taking part in the reaction. Also called oxidation-reduction reaction. (4.4)

reducing agent. A substance that can donate electrons to another substance or decrease the oxidation numbers in another substance. (4.4)

reduction reaction. The half-reaction that represents the gain of electrons in a redox process. (4.4)

representative elements. Elements in Groups 1A through 7A, all of which have at least an incompletely filled s or p subshell of the highest principal quantum number. (8.2)

resonance. The use of two or more Lewis structures to represent a particular molecule. (9.8)

resonance structure. One of two or more alternative Lewis structures for a single molecule that cannot be described fully with a single Lewis structure. (9.8)

reversible reaction. A reaction that can occur in both directions. (4.1)

ribonucleic acid (RNA). A type of nucleic acid. (22.4)

root-mean-square (rms) speed (u_{rms}). A measure of the average molecular speed at a given temperature. (5.6)

S

salt. An ionic compound made up of a cation other than H^+ and an anion other than OH^- or O^{2-} . (4.3)

salt hydrolysis. The reaction of the anion or cation, or both, of a salt with water. (16.9)

saponification. Soapmaking. (11.4)

saturated hydrocarbons. Hydrocarbons that contain only single covalent bonds. (11.2)

saturated solution. At a given temperature, the solution that results when the maximum amount of a substance has dissolved in a solvent. (13.1)

scientific method. A systematic approach to research. (1.2)

second law of thermodynamics. The entropy of the universe increases in a spontaneous process and remains unchanged in an equilibrium process. (18.4)

second-order reaction. A reaction whose rate depends on the concentration of one reactant raised to the second power or on the concentrations of two different reactants, each raised to the first power. (14.3)

semipermeable membrane. A membrane that allows solvent molecules to pass through, but blocks the movement of solute molecules. (13.6)

sigma bond (σ bond). A covalent bond formed by orbitals overlapping end-to-end; its electron density is concentrated between the nuclei of the bonding atoms. (10.5)

sigma molecular orbital. A molecular orbital in which the electron density is concentrated around a line between the two nuclei of the bonding atoms. (10.6)

significant figures. The number of meaningful digits in a measured or calculated quantity. (1.6)

single bond. Two atoms held together by one electron pair are joined by a single bond. (9.4)

solubility. The maximum amount of solute that can be dissolved in a given quantity of solvent at a specific temperature. (4.2, 17.5)

solubility product (K_{sp}). The product of the molar concentrations of constituent ions, each raised to the power of its stoichiometric coefficient in the equilibrium equation. (17.5)

solute. The substance present in the smaller amount in a solution. (4.1)

solution. A homogeneous mixture of two or more substances. (4.1)

solvation. The process in which an ion or molecule is surrounded by solvent molecules arranged in an ordered manner. (13.2)

solvent. The substance present in the larger amount in a solution. (4.1)

specific heat (s). The amount of heat energy required to raise the temperature of one gram of a substance by one degree Celsius. (6.5)

spectator ions. Ions that are not involved in the overall reaction. (4.2)

spectrochemical series. A list of ligands arranged in order of their abilities to split the d -orbital energies. (20.4)

standard atmospheric pressure (1 atm). The pressure that supports a column of

mercury exactly 76 cm high at 0°C at sea level. (5.2)

standard emf (E°). The difference of the standard reduction potentials of the substance that undergoes reduction and the substance that undergoes oxidation in a redox process. (19.3)

standard enthalpy of formation (ΔH_f°). The heat change that results when 1 mole of a compound is formed from its elements in their standard states. (6.6)

standard enthalpy of reaction ($\Delta H_{\text{rxn}}^\circ$). The enthalpy change that occurs when a reaction is carried out under standard-state conditions. (6.6)

standard entropy of reaction ($\Delta S_{\text{rxn}}^\circ$). The entropy change when the reaction is carried out under standard-state conditions. (18.4)

standard free energy of formation (ΔG_f°). The free-energy change when 1 mole of a compound is synthesized from its elements in their standard states. (18.5)

standard free energy of reaction ($\Delta G_{\text{rxn}}^\circ$). The free energy change when the reaction is carried out under standard-state conditions. (18.5)

standard reduction potential. The voltage measured as a reduction reaction occurs at the electrode when all solutes are 1 M and all gases are at 1 atm. (19.3)

standard solution. A solution of accurately known concentration. (4.6)

standard state. The condition of 1 atm of pressure. (6.6)

standard temperature and pressure (STP). 0°C and 1 atm. (5.4)

state function. A property that is determined by the state of the system. (6.3)

state of a system. The values of all pertinent macroscopic variables (for example, composition, volume, pressure, and temperature) of a system. (6.3)

stereoisomers. Compounds that are made up of the same types and numbers of atoms bonded together in the same sequence but with different spatial arrangements. (20.3)

stoichiometric amounts. The exact molar amounts of reactants and products that appear in a balanced chemical equation. (3.9)

stoichiometry. The mass relationships among reactants and products in chemical reactions. (3.8)

strong acid. An acid that is a strong electrolyte. (16.4)

strong base. A base that is a strong electrolyte. (16.4)

structural formula. A representation that shows how atoms are bonded to one another in a molecule. (2.6)

structural isomers. Molecules that have the same molecular formula but different structures. (11.2)

sublimation. The process in which molecules go directly from the solid phase into the vapor phase. (12.6)

substance. A form of matter that has a definite or constant composition (the number and type of basic units present) and distinct properties. (1.3)

supersaturated solution. A solution that contains more of the solute than is present in a saturated solution. (13.1)

surface tension. The amount of energy required to stretch or increase the surface of a liquid by a unit area. (12.3)

surroundings. The rest of the universe outside a system. (6.2)

system. Any specific part of the universe that is of interest to us. (6.2)

T

termolecular reaction. An elementary step involving three molecules. (14.5)

ternary compounds. Compounds consisting of three elements. (2.7)

theoretical yield. The amount of product predicted by the balanced equation when all of the limiting reagent has reacted. (3.10)

theory. A unifying principle that explains a body of facts and/or those laws that are based on them. (1.2)

thermal energy. Energy associated with the random motion of atoms and molecules. (6.1)

thermal pollution. The heating of the environment to temperatures that are harmful to its living inhabitants. (13.4)

thermochemical equation. An equation that shows both the mass and enthalpy relations. (6.4)

thermochemistry. The study of heat changes in chemical reactions. (6.2)

thermodynamics. The scientific study of the interconversion of heat and other forms of energy. (6.3)

thermonuclear reactions. Nuclear fusion reactions that occur at very high temperatures. (21.6)

third law of thermodynamics. The entropy of a perfect crystalline substance is zero at the absolute zero of temperature. (18.4)

titration. The gradual addition of a solution of accurately known concentration to another solution of unknown concentration until the chemical reaction between the two solutions is complete. (4.6)

tracers. Isotopes, especially radioactive isotopes, that are used to trace the path of the atoms of an element in a chemical or biological process. (21.7)

transition metals. Elements that have incompletely filled *d* subshells or readily give rise to cations that have incompletely filled *d* subshells. (7.9)

transition state. See activated complex.

transuranium elements. Elements with atomic numbers greater than 92. (21.4)

triple bond. A covalent bond in which two atoms share three pairs of electrons. (9.4)

triple point. The point at which the vapor, liquid, and solid states of a substance are in equilibrium. (12.7)

triprotic acid. Each unit of the acid yields three hydrogen ions. (4.3)

U

unimolecular reaction. An elementary step involving one molecule. (14.5)

unit cell. The basic repeating unit of the arrangement of atoms, molecules, or ions in a crystalline solid. (12.4)

unsaturated hydrocarbons. Hydrocarbons that contain carbon-carbon double bonds or carbon-carbon triple bonds. (11.2)

unsaturated solution. A solution that contains less solute than it has the capacity to dissolve. (13.1)

V

valence electrons. The outer electrons of an atom, which are the ones involved in chemical bonding. (8.2)

valence shell. The outermost electron-occupied shell of an atom, which holds the

electrons that are usually involved in bonding. (10.1)

valence-shell electron-pair repulsion (VSEPR) model. A model that accounts for the geometrical arrangements of shared and unshared electron pairs around a central atom in terms of the repulsive forces between electron pairs. (10.1)

van der Waals equation. An equation that describes the relationships among *P*, *V*, *n*, and *T* for a nonideal gas. (5.7)

van der Waals forces. The collective name for certain attractive forces between atoms and molecules, namely, dipole-dipole, dipole-induced dipole, and dispersion forces. (12.2)

vaporization. The escape of molecules from the surface of a liquid; also called evaporation. (12.6)

viscosity. A measure of a fluid's resistance to flow. (12.3)

volatile. Having a measurable vapor pressure. (13.6)

volume. Length cubed. (1.5)

W

wave. A vibrating disturbance by which energy is transmitted. (7.1)

wavelength (λ). The distance between identical points on successive waves. (7.1)

weak acid. An acid that is a weak electrolyte. (16.4)

weak base. A base that is a weak electrolyte. (16.4)

weight. The force that gravity exerts on an object. (1.5)

work. Directed energy change resulting from a process. (6.1)

X

X-ray diffraction. The scattering of X rays by the units of a regular crystalline solid. (12.4)