APPENDIX B

Basic Structure of Matter

For the convenience of students who have not had a course in basic chemistry, or for those who wish to review the material, we are providing a brief introduction here.

Elements and Atoms

All matter is composed of **elements,** which are substances that cannot be subdivided further by ordinary chemical reactions. Only 92 elements occur naturally, but elements may be combined by chemical bonds into a vast number of different compounds. Elements are designated by one or two letters derived from their Latin or English names (Table B-1). Elements are composed of discrete units called **atoms,** which are the smallest units of elements. Combination of atoms of an element with each other or with those of other elements by chemical bonds creates **molecules.** When molecules are composed of atoms of two or more different kinds of elements, they are a **compound.**

In a chemical formula, the symbol for an element stands for one atom of the element, with additional atoms indicated by appropriately placed numbers. Thus atmospheric nitrogen is N_2 (each molecule is composed of two atoms of nitrogen), and water is H_2O (two atoms of hydrogen and one of oxygen in each molecule), and so on.

TABLE B-1

Some of the Most Important Elements In Living Organisms

Element	Symbol	Atomic Number	Approximate Atomic Weight
Carbon	C	6	12
Oxygen	O	8	16
Hydrogen	H	1	1
Nitrogen	N	7	14
Phosphorus	P	15	31
Sodium	Na	11	23
Sulfur	S	16	32
Chlorine	Cl	17	35
Potassium	K	19	39
Calcium	Ca	20	40
Iron	Fe	26	56
Iodine	I	53	127

Subatomic Particles

Each atom is composed of subatomic particles, of these there are three with which we need concern ourselves: protons, neutrons, and electrons. Every atom consists of a positively charged nucleus surrounded by a negatively charged system of electrons (Figure B-1). The nucleus, containing most of the atom's mass, is made up of protons and neutrons clustered together in a very small volume. These two particles have about the same mass, each being about 2000 times heavier than an electron. Protons bear positive charges, and neutrons are uncharged (neutral). Although the number of protons in a nucleus is the same as the number of electrons around the nucleus, the number of neutrons may vary. For every positively charged proton in the nucleus, there is a negatively charged electron; the total charge of the atom is thus neutral.

The **atomic number** of an element is equal to the number of protons in the nucleus, whereas the **atomic mass** is nearly equal to the number of protons plus the number of neutrons (explanation of why atomic mass is not exactly equal to protons plus neutrons can be found in any introductory chemistry text). The mass of the electrons may be neglected because it is only 1/1836 that of a proton or neutron.

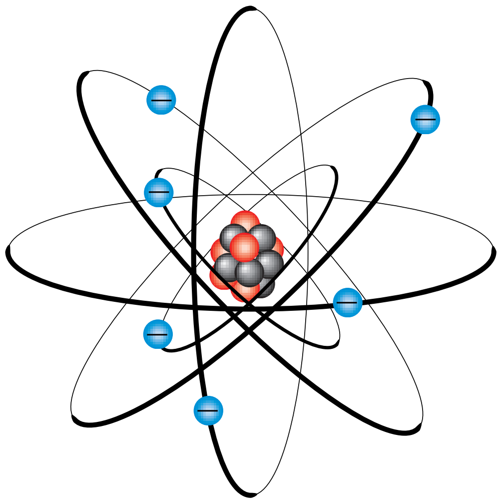


Figure B-1
Structure of an atom. Planetary system of negatively charged electrons around a dense nucleus of positively charged protons and uncharged neutrons.

Isotopes

It is possible for two atoms of the same element to have the same number of protons in their nuclei but have a different number of neutrons. Such different forms, having the same number of protons but different atomic masses, are called **isotopes.** For example, the predominant form of hydrogen in nature has 1 proton and no neutron (^{1}H) (Figure B-2). Another form (deuterium [^{2}H]) has 1 proton and 1 neutron. Tritium (^{3}H) has 1 proton and 2 neutrons. Some isotopes are unstable, undergoing a spontaneous disintegration with the emission of one or more of three types of particles, or rays: gamma rays (a form of electromagnetic radiation), beta rays (electrons), and alpha rays (positively charged helium nuclei stripped of their electrons). These unstable isotopes are said to be **radioactive.** Using radioisotopes, biologists are able to trace movements of elements and tagged compounds through organisms. Our present understanding of metabolic pathways in animals and plants is in large part a result of this powerful analytical tool. Among commonly used radioisotopes are carbon 14 (^{14}C), tritium, and phosphorus 32 (^{32}P).

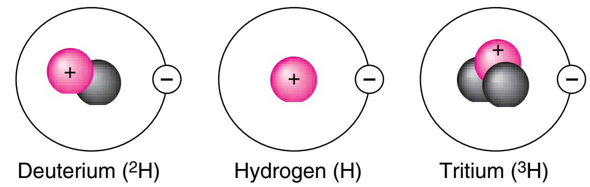


Figure B-2
Three isotopes of hydrogen. Of the three isotopes, hydrogen 1 makes up about 99.98% of all hydrogen, and deuterium (heavy hydrogen) makes up about 0.02%. Tritium is radioactive and is found only in traces in water. Numbers indicate approximate atomic weights. Most elements are mixtures of isotopes. Some elements (for example, tin) have as many as 10 isotopes.

Electron "Shells" of Atoms

According to Niels Bohr's planetary model of atoms, electrons revolve around the nucleus of an atom in circular orbits of precise energy and size. All orbits of any one energy and size compose an electron shell. This simplified picture of atoms has been greatly modified by more recent experimental evidence; definite electron pathways are no longer hypothesized, and an electron shell is more vaguely understood as a thick region of space around the nucleus rather than a narrow shell of a particular radius.

However, the old planetary model with the idea of electronic shells is still useful in interpreting chemical phenomena. The number of concentric shells required to contain an element's electrons varies with the element. Each shell can hold a maximum number of electrons. The shell closest to the atomic nucleus can hold a maximum of 2 electrons, and the second shell can hold 8; other shells also have a maximum number, but no atom can have more than 8 electrons in its outermost shell. Inner shells are filled first, and if there are not enough electrons to fill all the shells, the outer shell is left incomplete. Hydrogen has 1 proton in its nucleus and 1 electron in its single orbit but no neutron. Since its shell can hold 2 electrons, it has an incomplete shell. Helium has 2 electrons in its single shell, and its nucleus is made up of 2 protons and 2 neutrons. Since the 2-electron arrangement in helium's shell is the maximum number for this shell, the shell is closed and precludes all chemical activity. There is no known compound of helium. Neon is another inert (chemically inactive) gas because its outer shell contains 8 electrons, the maximum number (Figure B-3). However, stable compounds of xenon (another inert gas) with fluorine and oxygen are formed under special conditions. Oxygen has an atomic number of 8. Its 8 electrons are arranged with 2 in the first shell and 6 in the second shell (Figure B-3). It is active chemically, forming compounds with almost all elements except inert gases.

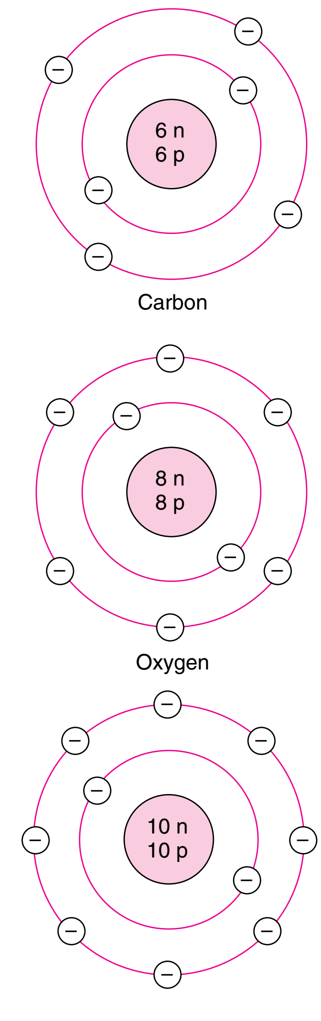


Figure B-3
Electron shells of three common atoms. Since no atom can have more than 8 electrons in its outermost shell and two electrons in its innermost shell, neon is chemically inactive. However, the second shells of carbon and oxygen, with 4 and 6 electrons, respectively, are open so that these elements are electronically unstable and react chemically whenever appropriate atoms come into contact. Chemical properties of atoms are determined by their outermost electron shells.

Chemical Bonds

As we noted above, atoms joined to each other by chemical bonds form molecules, and atoms of each element form molecules with each other or with atoms of other elements in particular ways, depending on the number of electrons in their outer orbits.

Ionic Bonds

Elements react in such a way as to gain a stable configuration of electrons in their outer shells. The number of electrons in the outer shell varies from 0 to 8. With either 0 or 8 in this shell, the element is chemically inactive. When there are fewer than 8 electrons in the outer shell, the atom will tend to lose or gain electrons to result in an outer shell of 8. This will give the atom a net electrical charge because the number of protons does not equal the number of electrons, and the atom is now an **ion.** Atoms with 1 to 3 electrons in the outer shell tend to lose them to other atoms and to become positively charged ions because of the excess protons in the nucleus. Atoms with 5 to 7 electrons in the outer orbit tend to gain electrons from other atoms and to become negatively charged ions because of the greater number of electrons than protons. Positive and negative ions tend to unite.

Every atom has a tendency to complete its outer shell to increase its stability in the presence of other atoms. Let us examine how two atoms with incomplete outer shells, sodium and chlorine, can interact to fill their outer shells. Sodium, with 11 electrons, has 2 electrons in its first shell, 8 in its second shell, and only 1 in the third shell. The third shell is highly incomplete; if this third-shell electron were lost, the second shell would be the outermost shell and would produce a stable atom. Chlorine, with 17 electrons, has 2 in the first shell, 8 in the second, and 7 in the incomplete third shell. Chlorine must gain an electron to fill the outer shell and become a stable atom. Clearly, the transfer of the third-shell sodium electron to the incomplete chlorine third shell would yield simultaneous stability to both atoms.

Sodium, now with 11 protons but only 10 electrons, becomes electropositive (Na^+). In gaining an electron from sodium, chlorine contains 18 electrons but only 17 protons and thus becomes an electronegative chloride ion (Cl^-). Because unlike charges attract, a strong electrostatic force, called an **ionic bond** (Figure B-4), is formed. The ionic compound formed, sodium chloride, can be represented in electron dot notation ("fly-speck formulas") as:

$$Na\bullet + \bullet\overset{\bullet\bullet}{\underset{\bullet\bullet}{Cl}}\!: \longrightarrow Na^+ + (:\overset{\bullet\bullet}{\underset{\bullet\bullet}{Cl}}:)^-$$

The number of dots shows the number of electrons present in the outer shell of an atom: 7 in the case of neutral chlorine atom and 8 for chloride ion; 1 in the case of neutral sodium atom and none for sodium ion.

If an element with 2 electrons in its outer shell, such as calcium, reacts with chlorine, it must give them both up, one to each of two chlorine atoms, and calcium becomes doubly positive:

$$Ca\!: + 2\bullet\overset{\bullet\bullet}{\underset{\bullet\bullet}{Cl}}\!: \longrightarrow Ca^{2+} + 2(:\overset{\bullet\bullet}{\underset{\bullet\bullet}{Cl}}:)^-$$

Processes that involve a **loss of electrons** are called **oxidation** reactions; those that involve a **gain of electrons** are **reduction** reactions. Since oxidation and reduction always occur simultaneously, each of these processes is really a "half-reaction." The entire reaction is called an **oxidation-reduction** reaction, or simply a **redox** reaction. The terminology is confusing because oxidation-reduction reactions involve electron transfers, rather than (necessarily) any reaction with oxy-gen. However, it is easier to learn the system than to try to change accepted usage.

Covalent Bonds

Stability can also be achieved when two atoms share electrons. Let us again consider a chlorine atom, which, as we have seen, has an incomplete 7-electron

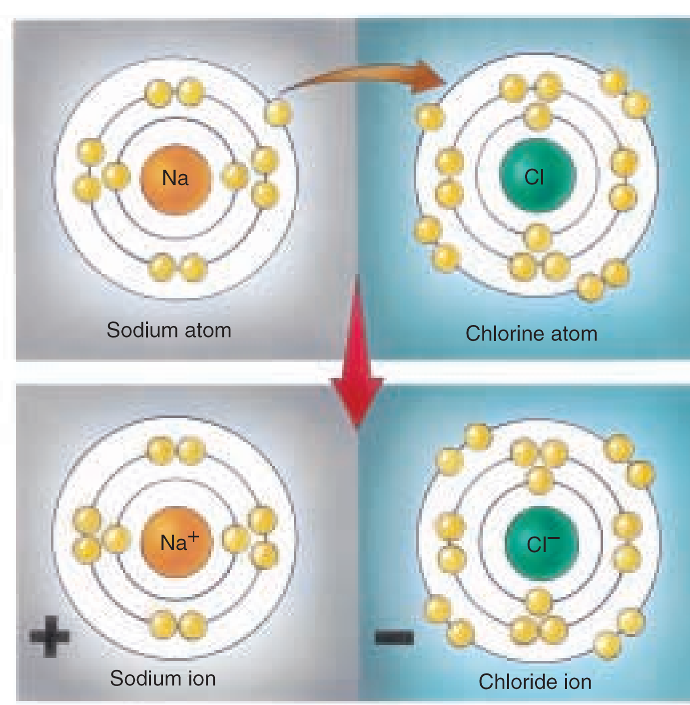
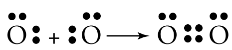


Figure B-4
Ionic bond. When one atom of sodium and one of chlorine react to form a molecule, a single electron in the outer shell of sodium is transferred to the outer shell of chlorine. This causes the outer or second shell (third shell is now empty) of sodium to have 8 electrons and also chlorine to have 8 electrons in its outer or third shell. The compound thus formed is sodium chloride (NaCl). By losing 1 electron, sodium becomes a positive ion, and by gaining 1 electron, chlorine becomes a negative ion (chloride). This ionic bond is the strong electrostatic force acting between positively and negatively charged ions.

outer shell. Stability is attained by gaining an electron. One way this can be done is for two chlorine atoms to *share* one pair of electrons (Figure B-5). To do this, the two chlorine atoms must *overlap* their third shells so that the electrons in these shells can now spread themselves over both atoms, thereby completing the filling of both shells. Many other elements can form covalent (or electron-pair) bonds. Examples are hydrogen (H_2):

$$H\bullet + H\bullet \longrightarrow H{:}H$$

and oxygen (O_2):



In this case oxygen must share two pairs of electrons to achieve stability. Each atom now has 8 electrons available to its outer shell, the stable number.

Covalent bonds are of great significance in living systems, since the major elements of living matter (carbon, oxygen, nitrogen, hydrogen) almost always share electrons in strong covalent bonds. Stability of these bonds is essential to integrity of DNA and other macromolecules, which, if easily dissociated, would result in biological disorder.

The outer shell of carbon contains 4 electrons. This element is endowed with great potential for forming a variety of atomic configurations with itself

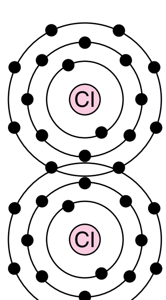


Figure B-5
Covalent bond. Each chlorine atom has 7 electrons in its outer shell, and by sharing one pair of electrons, each atom acquires a complete outer shell of 8 electrons, thus forming a molecule of chlorine (Cl_2).

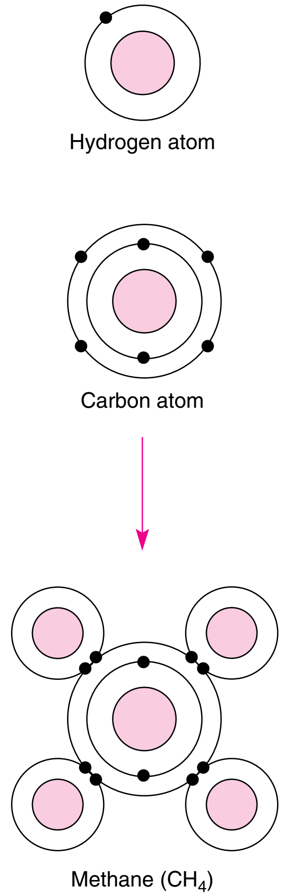
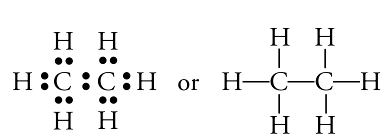
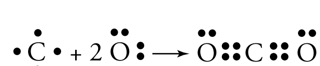


Figure B-6
In methane four hydrogen atoms each share an electron with a carbon atom. They are arranged symmetrically around the carbon atom and form a pyramid-shaped tetrahedron in which each of the hydrogen atoms is equally distant from others.

and other molecules. It can, for example, share its electrons with hydrogen to form methane (Figure B-6). Carbon now achieves stability with 8 electrons, and each hydrogen atom becomes stable with two electrons. Carbon can also bond with itself (and hydrogen) to form, for example, ethane:



Carbon also forms covalent bonds with oxygen:



This is a "double-bond" configuration usually written as O = C = O. Carbon can even form triple bonds as, for example, in acetylene:

$$H{:}C{:::}C{:}H \text{ or } H{\equiv}C{-}C{-}H$$

The significant aspect of each of these molecules is that each carbon gains a share in 4 electrons from atoms nearby, thus attaining the stability of 8 electrons. The sharing may occur between carbon and other elements or other carbon atoms, and in many instances 8-electron stability is achieved by means of multiple bonds.

These examples only begin to illustrate the amazing versatility of carbon. It is a part of virtually all compounds comprising living substance, and without carbon, life as we know it would not exist.

Hydrogen Bonds

Hydrogen bonds are described as "weak" bonds because they require little energy to break. They do not form by transfer or sharing of electrons, but result from unequal charge distribution on a molecule, so that the molecule is polar. For example, the two hydrogen atoms that share electrons with an oxygen atom to form water (H_2O) are not 180 degrees away from each other around the oxygen, but form an angle of about 105 degrees (Figure B-7). Thus the

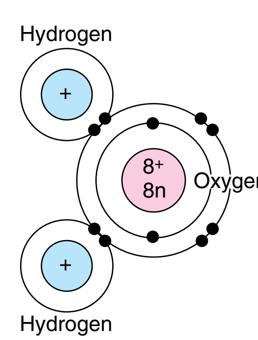


Figure B-7
Molecular structure of water. Two hydrogen atoms bonded covalently to an oxygen atom are arranged at an angle of about 105 degrees to each other. Since the electrical charge is not symmetrical, the molecule is polar with positively and negatively charged ends.

side of the molecule away from the hydrogen atoms is more negative, and the hydrogen side is more positive (contrast the methane molecule [Figure B-6], in which equidistant placement of hydrogen atoms cancels out charge displacements). Electrostatic attraction between the electropositive part of one molecule forms a hydrogen bond with the electronegative part of an adjacent molecule. The ability of water molecules to form hydrogen bonds with each other (Figure B-8) accounts for many unusual properties of this unique substance (p. 28, text). Hydrogen bonds are important in formation and function of other biologically active substances, such as proteins and nucleic acids (pp. 25 through 27, text).

Acids, Bases, and Salts

The hydrogen ion (H^+) is one of the most important ions in living organisms. Hydrogen atoms contain a single electron. When this electron is completely transferred to another atom (not just shared with another atom as in the covalent bonds with carbon), only the hydrogen nucleus with its positive proton remains. Any molecule that dissociates in solution and gives rise to a hydrogen ion is an **acid.** An acid is classified as strong or weak, depending on the extent to which the acid molecule dissociates in solution. Examples of strong acids that dissociate completely in water are hydrochloric acid ($HCl \rightarrow H^+ + Cl^-$) and nitric acid ($HNO_3 \rightarrow H^+ + NO_3^-$). Weak acids, such as carbonic acid ($H_2CO_3 \rightarrow H^+ + HCO_3^-$), dissociate only slightly. A solution of carbonic acid is mostly undissociated carbonic acid molecules with only a small number of bicarbonate (HCO_3^-) and hydrogen ions (H^+) present.

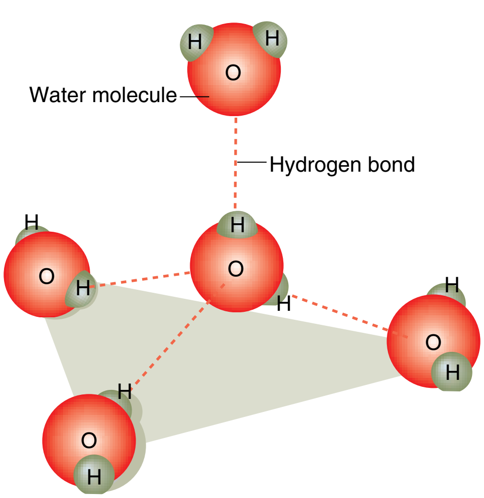


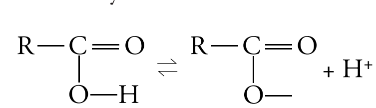
Figure B-8
Geometry of water molecules. Each water molecule is linked by hydrogen bonds (*dashed lines*) to four other water molecules. If imaginary lines are used to connect the divergent oxygen atoms, a tetrahedron is obtained. In ice, the individual tetrahedrons associate to form an open lattice structure.

A **base** contains negative ions called hydroxide ions and may be defined as a molecule or ion that will accept a proton (hydrogen ion). Bases are produced when compounds containing them are dissolved in water. Sodium hydroxide (NaOH) is a strong base because it will dissociate completely in water into sodium (Na^+) and hydroxide (OH^-) ions. Among the characteristics of bases is their ability to combine with hydrogen ions, thus decreasing the concentration of the hydrogen ions. Like acids, bases vary in the extent to which they dissociate in aqueous solutions into hydroxide ions.

A **salt** is a compound resulting from a chemical interaction of an acid and a base. Common salt, sodium chloride (NaCl), is formed by interaction of hydrochloric acid (HCl) and sodium hydroxide (NaOH). In water HCl is dissociated into H^+ and Cl^- ions. Hydrogen and hydroxide ions combine to form water (H_2O), and sodium and chloride ions remain as a dissolved form of salt (Na^+Cl^-):

$$\underset{\text{Acid}}{H^+Cl^-} + \underset{\text{Base}}{Na^+OH^-} \rightarrow \underset{\text{Salt}}{Na^+Cl^-} + H_2O$$

Organic acids are usually characterizedby having in their molecule a carboxyl group (—COOH). They are weak acids because a relatively small proportion of H^+ reversibly dissociates from carboxyl:



R refers to an atomic grouping unique to the molecule. Some common organic acids are acetic, citric, formic, lactic, and oxalic.

Hydrogen Ion Concentration (pH)

Solutions are classified as acid, basic, or neutral according to the proportion of hydrogen (H^+) and hydroxide (OH^-) ions they possess. In acid solutions there is an excess of hydrogen ions; in alkaline, or basic, solutions hydroxide ion is more common; and in neutral solutions both hydrogen and hydroxide ions are present in equal numbers.

To express acidity or alkalinity of a substance, a logarithmic scale, a type of mathematical shorthand, is employed that uses the numbers 1 to 14. This is pH, defined as:

$$pH = \log_{10}\frac{1}{[H]}$$

or

$$pH = -\log_{10}[H^+]$$

Thus pH is the negative logarithm of the hydrogen ion concentration in moles per liter. In other words, when the hydrogen ion concentration is expressed exponentially, pH is the exponent, but with the *opposite* sign; if $[H^+] = 10^{-2}$, then $pH = -(-2) = +2$. Unfortunately, pH can be a confusing concept because, as the $[H^+]$ decreases, the pH increases. Numbers below 7 indicate an acid range, and numbers above 7 indicate alkalinity (Figure B-9). A pH of 7 indicates neutrality, that is, the presence of equal numbers of H^+ and OH^- ions. According to this logarithmic scale, a pH of 3 is 10 times more acid than one of 4; a pH of 9 is 10 times more alkaline than one of 8.

Buffer Action

The hydrogen ion concentration in the extracellular fluids must be regulated so that metabolic reactions within the cell will not be adversely affected by a constantly changing hydrogen ion concentration, to which they are extremely sensitive. A change in pH of only 0.2 from the normal mammalian blood pH of about 7.35 can cause serious metabolic disturbances. To maintain pH within physiological limits, there are certain substances in cells and organisms that tend to compensate for any change in pH when acids or alkalies are produced in metabolic reactions or are added to body fluids. These substances are called **buffers.** A buffer is a mixture of slightly ionized weak acid and its completely ionized salt. In such a system, added H^+ combines with the anion of the salt to form undissociated acid, and added OH^- combines with H^+ from the weak acid molecule to form water. The most important buffers in blood and other extracellular fluids are the bicarbonates and phosphates, and organic molecules such as amino acids and proteins are important buffers within cells. The bicarbonate buffer system consists of carbonic acid (H_2CO_3, a weak acid) and its salt, sodium bicarbonate ($NaHCO_3$). Sodium bicarbonate is strongly ionized into sodium ions (Na^+) and bicarbonate ions (HCO_3^-). When a strong acid (for example, HCl) is added to the fluid, hydrogen ions (H^+) of the dissociated acid will react with bicarbonate ion (HCO_3^-) to form a very weak acid, carbonic acid, which dissociates only slightly. Thus H^+ ions from the HCl are removed and pH is little altered. When a strong base (for example, NaOH) is added to the fluid, OH^- ions of the strong base will react with carbonic acid by removing H^+ ions from H_2CO_3, to form water and bicarbonate ions. Again H^+ ion concentration in solution is little altered and pH remains nearly unchanged.

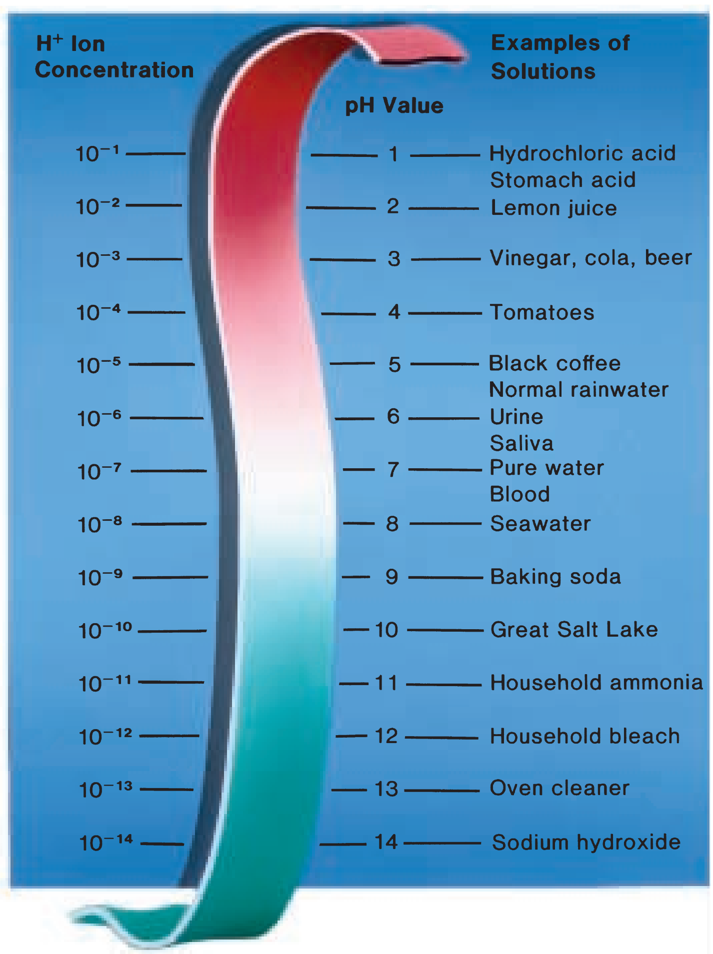


Figure B-9

The pH scale. A pH of 7 is neutral. Values below 7 are acidic, and the lower the value, the more acidic the solution. Values above 7 are basic or alkaline, and the higher the value, the more basic the solution. Representative fluids with approximate pH values are listed.