BIOCHEMISTRY: LS2101

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Water

- Weak Interactions in Aqueous Systems
- Ionization of Water, Weak Acids, and Weak Bases
- Buffering against pH Changes in Biological Systems
- Water as a Reactant
- The Fitness of the Aqueous Environment for Living Organisms

Water

Weak bonds:

Hydrogen bonds and ionic, hydrophobic (Greek, “water-fearing”), and van der Waals

Water

- Weak Interactions in Aqueous Systems

TABLE 2–1 Melting Point, Boiling Point, and Heat of Vaporization of Some Common Solvents

	Melting point (°C)	*Boiling point (°C)*	*Heat of vaporization (J/g)**
Water	0	100	2,260
Methanol (CH_3OH)	−98	65	1,100
Ethanol (CH_3CH_2OH)	−117	78	854
Propanol ($CH_3CH_2CH_2OH$)	−127	97	687
Butanol ($CH_3(CH_2)_2CH_2OH$)	−90	117	590
Acetone (CH_3COCH_3)	−95	56	523
Hexane ($CH_3(CH_2)_4CH_3$)	−98	69	423
Benzene (C_6H_6)	6	80	394
Butane ($CH_3(CH_2)_2CH_3$)	−135	−0.5	381
Chloroform ($CHCl_3$)	−63	61	247

*The heat energy required to convert 1.0 g of a liquid at its boiling point, at atmospheric pressure, into its gaseous state at the same temperature. It is a direct measure of the energy required to overcome attractive forces between molecules in the liquid phase.

Water

Hydrogen Bonding Gives Water Its Unusual Properties

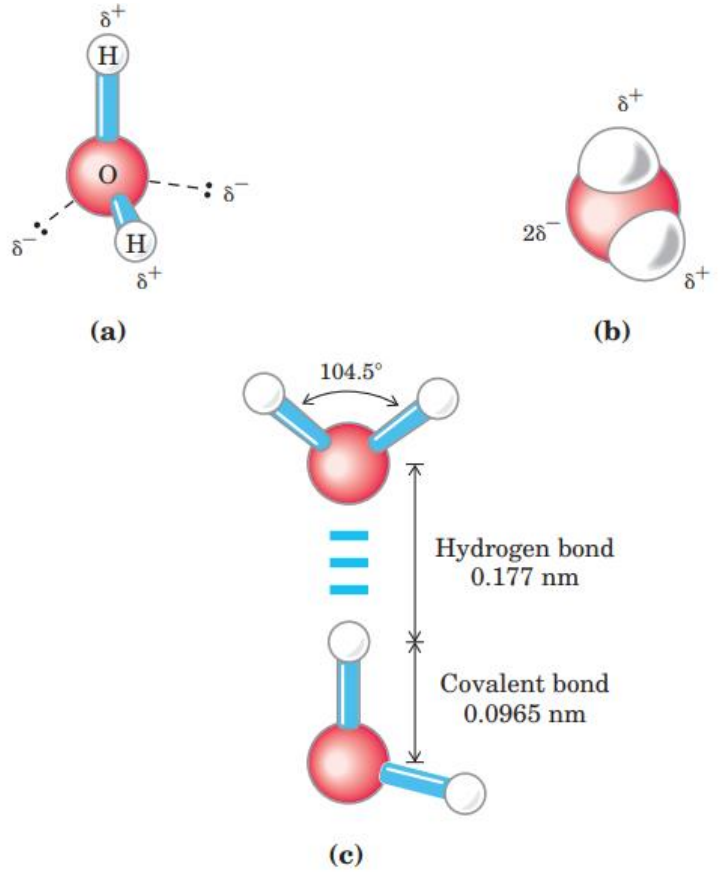


FIGURE 2–1 Structure of the water molecule. The dipolar nature of the H_2O molecule is shown by **(a)** ball-and-stick and **(b)** space-filling models. The dashed lines in **(a)** represent the nonbonding orbitals. There is a nearly tetrahedral arrangement of the outer-shell electron pairs around the oxygen atom; the two hydrogen atoms have localized partial positive charges (δ^+) and the oxygen atom has a partial negative charge ($2\delta^-$). **(c)** Two H_2O molecules joined by a hydrogen bond (designated here, and throughout this book, by three blue lines) between the oxygen atom of the upper molecule and a hydrogen atom of the lower one. Hydrogen bonds are longer and weaker than covalent O—H bonds.

Water

$H_2O(solid) \longrightarrow H_2O(liquid) \qquad \Delta H = +5.9\ kJ/mol$

$H_2O(liquid) \longrightarrow H_2O(gas) \qquad \Delta H = +44.0\ kJ/mol$

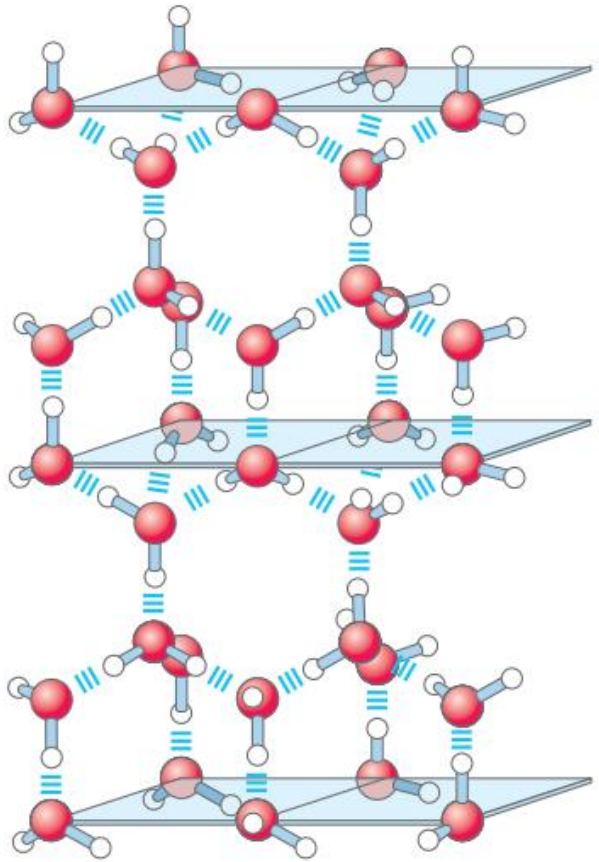


FIGURE 2–2 Hydrogen bonding in ice. In ice, each water molecule forms the maximum of four hydrogen bonds, creating a regular crystal lattice. By contrast, in liquid water at room temperature and atmospheric pressure, each water molecule hydrogen-bonds with an average of 3.4 other water molecules. This crystal lattice of ice makes it less dense than liquid water, and thus ice floats on liquid water.

Water

Water Forms Hydrogen Bonds with Polar Solutes

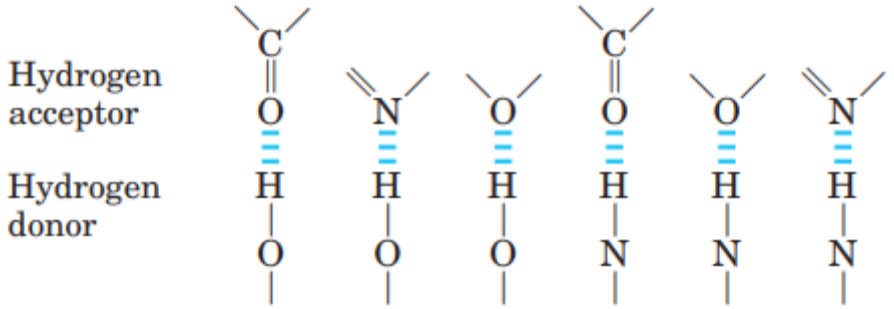


FIGURE 2–3 Common hydrogen bonds in biological systems. The hydrogen acceptor is usually oxygen or nitrogen; the hydrogen donor is another electronegative atom.

Water

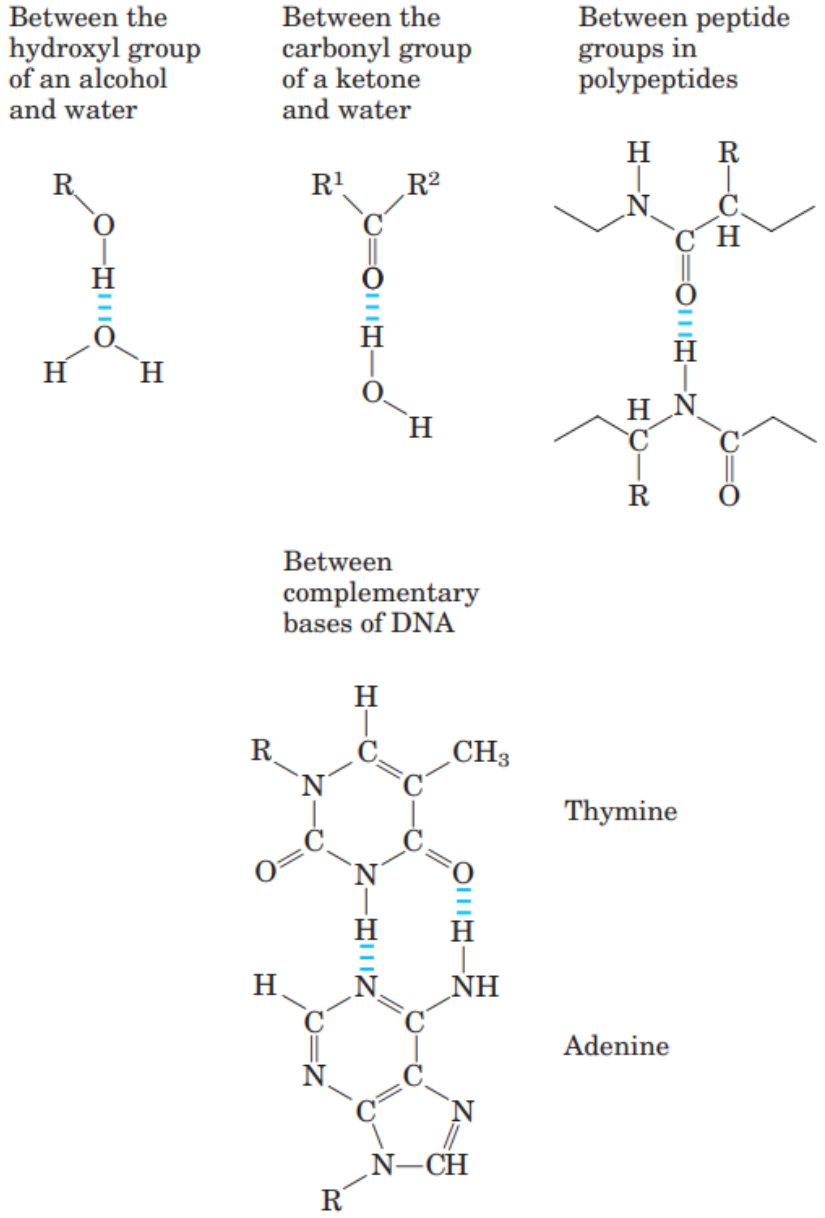


FIGURE 2–4 Some biologically important hydrogen bonds.

Water

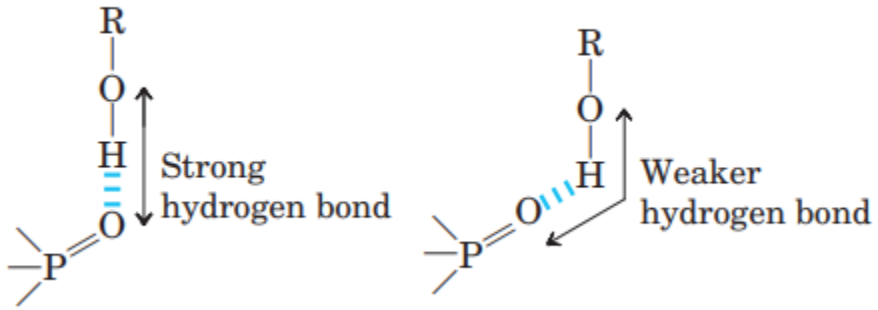


FIGURE 2–5 Directionality of the hydrogen bond. The attraction between the partial electric charges (see Fig. 2–1) is greatest when the three atoms involved (in this case O, H, and O) lie in a straight line. When the hydrogen-bonded moieties are structurally constrained (as when they are parts of a single protein molecule, for example), this ideal geometry may not be possible and the resulting hydrogen bond is weaker.

Water

Water Interacts Electrostatically with Charged Solutes

Hydrophilic: water-loving molecules, polar molecules, NaCl

Hydrophobic: avoids water: nonpolar molecules such as lipids and waxes.

Amphipathic: These compounds contain regions that are polar (or charged) and regions that are nonpolar

Electrostatic interactions

The strength, or force (F), of ionic interactions in a solution depends upon the magnitude of

the charges (Q),
the distance between the charged groups (r),
and the dielectric constant (ε) of the solvent in which the interactions occur

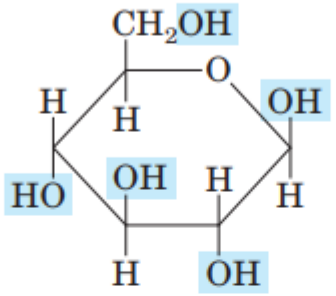
$$F = \frac{Q_1 Q_2}{\varepsilon r^2}$$

Water

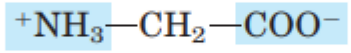
TABLE 2–2 Some Examples of Polar, Nonpolar, and Amphipathic Biomolecules (Shown as Ionic Forms at pH 7)

Polar

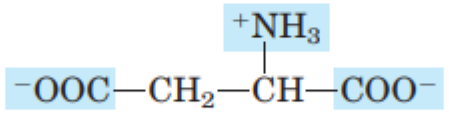
Glucose



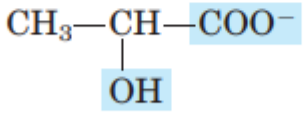
Glycine



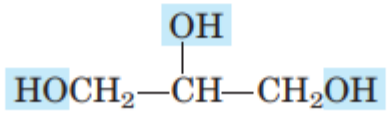
Aspartate



Lactate

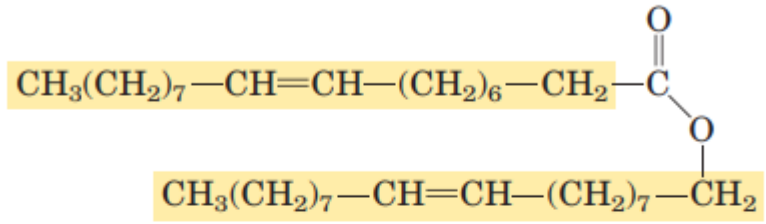


Glycerol



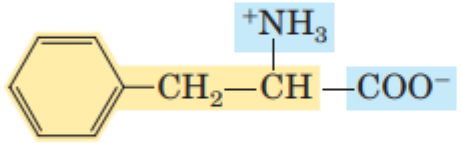
Nonpolar

Typical wax

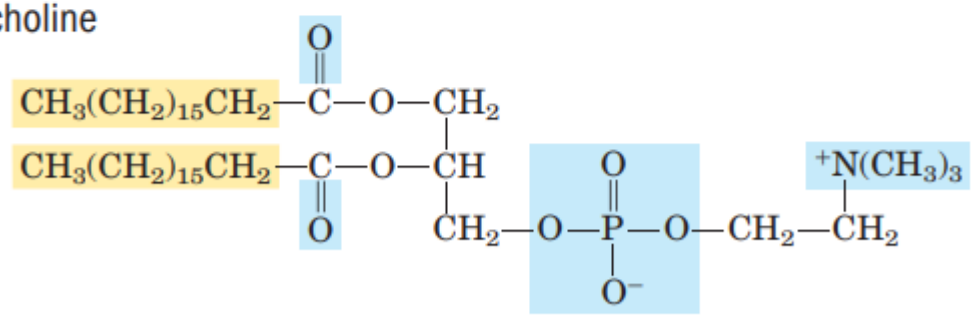


Amphipathic

Phenylalanine



Phosphatidylcholine



Polar groups Nonpolar groups

Water

Entropy Increases as Crystalline Substances Dissolve

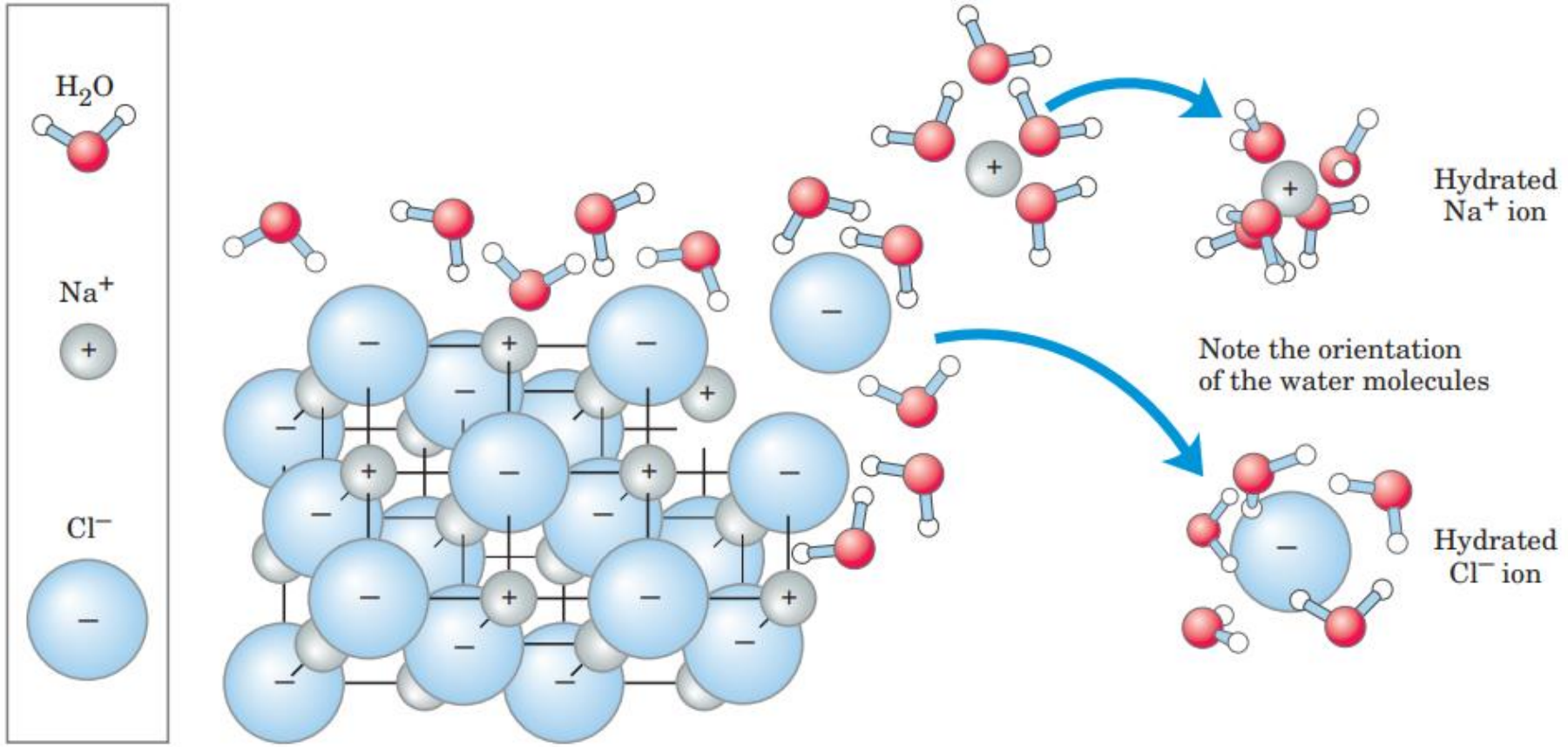


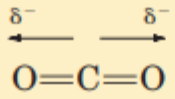
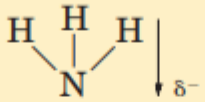
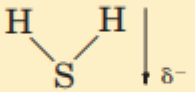
FIGURE 2–6 Water as solvent. Water dissolves many crystalline salts by hydrating their component ions. The NaCl crystal lattice is disrupted as water molecules cluster about the Cl^- and Na^+ ions. The ionic charges are partially neutralized, and the electrostatic attractions necessary for lattice formation are weakened.

Water

Nonpolar Gases Are Poorly Soluble in Water

Nonpolar Compounds Force Energetically Unfavorable Changes in the Structure of Water

TABLE 2–3 Solubilities of Some Gases in Water

Gas	*Structure**	*Polarity*	*Solubility in water (g/L)†*
Nitrogen	N≡N	Nonpolar	0.018 (40 °C)
Oxygen	O=O	Nonpolar	0.035 (50 °C)
Carbon dioxide		Nonpolar	0.97 (45 °C)
Ammonia		Polar	900 (10 °C)
Hydrogen sulfide		Polar	1,860 (40 °C)

*The arrows represent electric dipoles; there is a partial negative charge (δ^-) at the head of the arrow, a partial positive charge (δ^+; not shown here) at the tail.

†Note that polar molecules dissolve far better even at low temperatures than do nonpolar molecules at relatively high temperatures.

Water

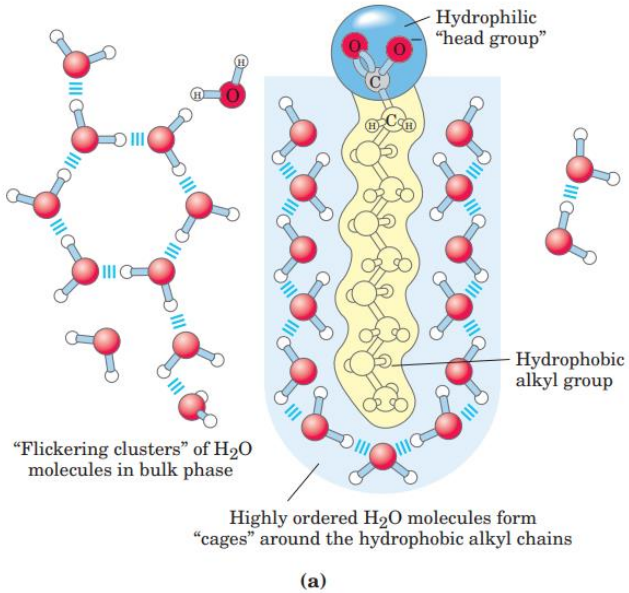
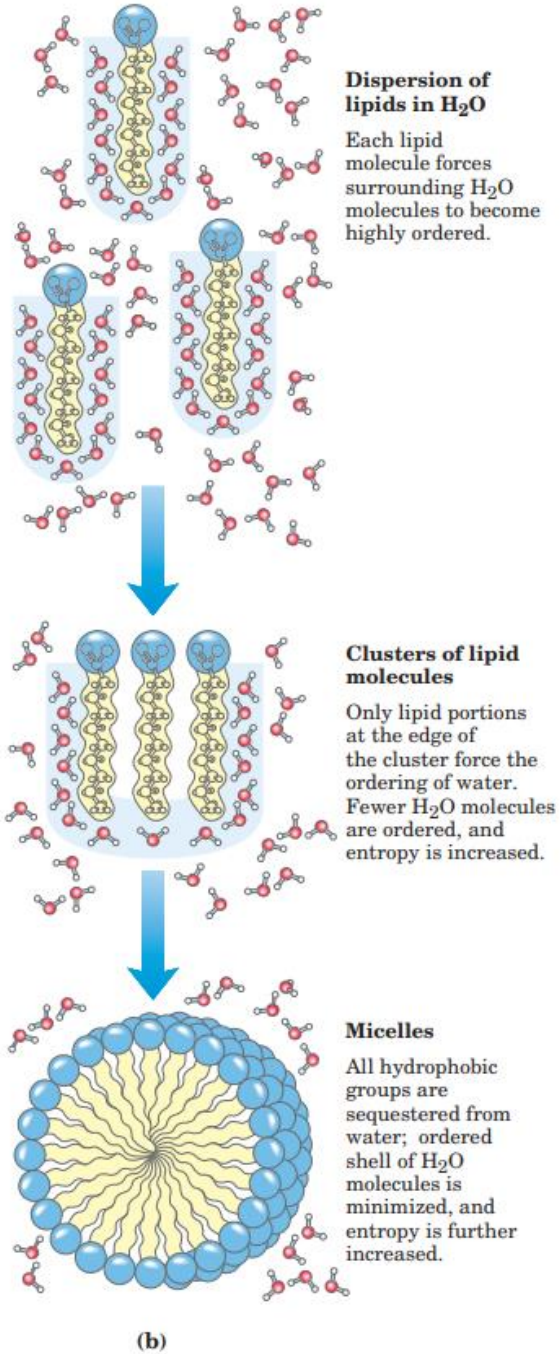


FIGURE 2–7 **Amphipathic compounds in aqueous solution.** **(a)** Long-chain fatty acids have very hydrophobic alkyl chains, each of which is surrounded by a layer of highly ordered water molecules. **(b)** By clustering together in micelles, the fatty acid molecules expose the smallest possible hydrophobic surface area to the water, and fewer water molecules are required in the shell of ordered water. The energy gained by freeing immobilized water molecules stabilizes the micelle.



Water

Formation of an
enzyme-substrate complex in water

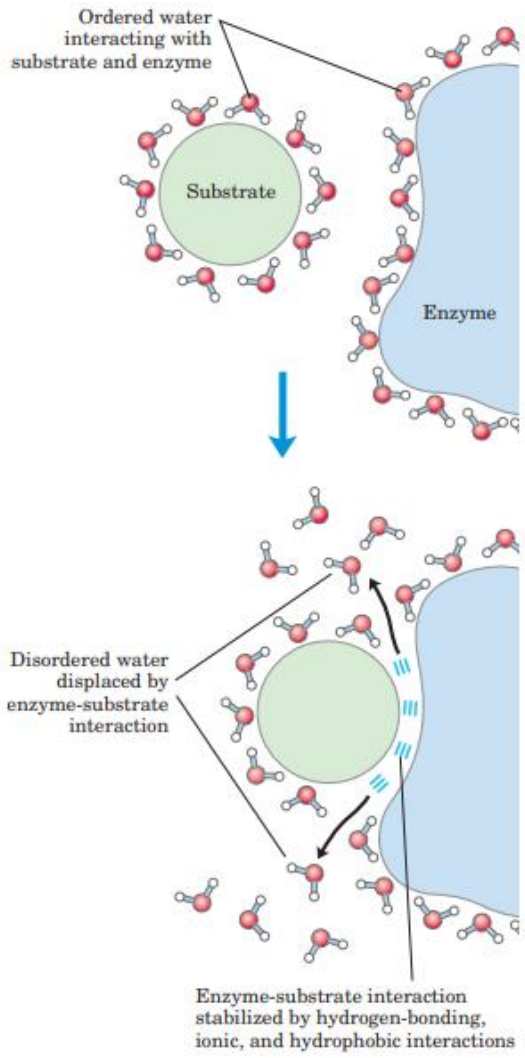


FIGURE 2-8 Release of ordered water favors formation of an enzyme-substrate complex. While separate, both enzyme and substrate force neighboring water molecules into an ordered shell. Binding of substrate to enzyme releases some of the ordered water, and the resulting increase in entropy provides a thermodynamic push toward formation of the enzyme-substrate complex.

Water

van der Waals Interactions Are Weak Interatomic Attractions

van der Waals radius

TABLE 2–4 van der Waals Radii and Covalent (Single-Bond) Radii of Some Elements

Element	*van der Waals radius (nm)*	*Covalent radius for single bond (nm)*
H	0.11	0.030
O	0.15	0.066
N	0.15	0.070
C	0.17	0.077
S	0.18	0.104
P	0.19	0.110
I	0.21	0.133

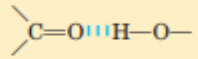
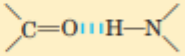
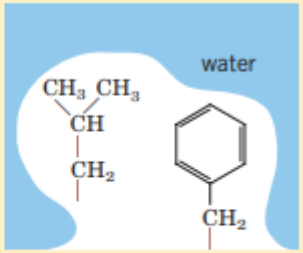
Sources: For van der Waals radii, Chauvin, R. (1992) Explicit periodic trend of van der Waals radii. *J. Phys. Chem.* **96,** 9194–9197. For covalent radii, Pauling, L. (1960) *Nature of the Chemical Bond,* 3rd edn, Cornell University Press, Ithaca, NY.

Note: van der Waals radii describe the space-filling dimensions of atoms. When two atoms are joined covalently, the atomic radii at the point of bonding are less than the van der Waals radii, because the joined atoms are pulled together by the shared electron pair. The distance between nuclei in a van der Waals interaction or a covalent bond is about equal to the sum of the van der Waals or covalent radii, respectively, for the two atoms. Thus the length of a carbon-carbon single bond is about 0.077 nm + 0.077 nm = 0.154 nm.

Water

Weak Interactions Are Crucial to Macromolecular Structure and Function

TABLE 2–5 Four Types of Noncovalent ("Weak") Interactions among Biomolecules in Aqueous Solvent

Hydrogen bonds	
Between neutral groups	
Between peptide bonds	
Ionic interactions	
Attraction	
Repulsion	
Hydrophobic interactions	
van der Waals interactions	Any two atoms in close proximity

Water

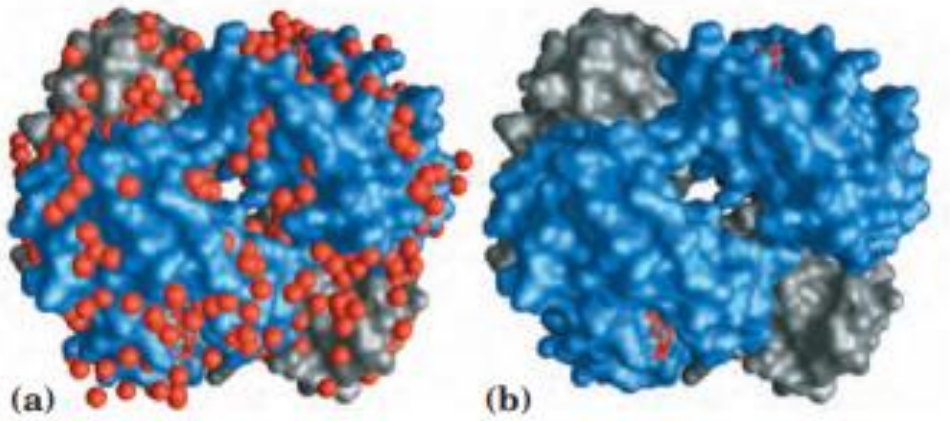


FIGURE 2-9 Water binding in hemoglobin. The crystal structure of hemoglobin, shown **(a)** with bound water molecules (red spheres) and **(b)** without the water molecules. These water molecules are so firmly bound to the protein that they affect the x-ray diffraction pattern as though they were fixed parts of the crystal. The gray structures with red and orange atoms are the four hemes of hemoglobin, discussed in detail in Chapter 5.

Water

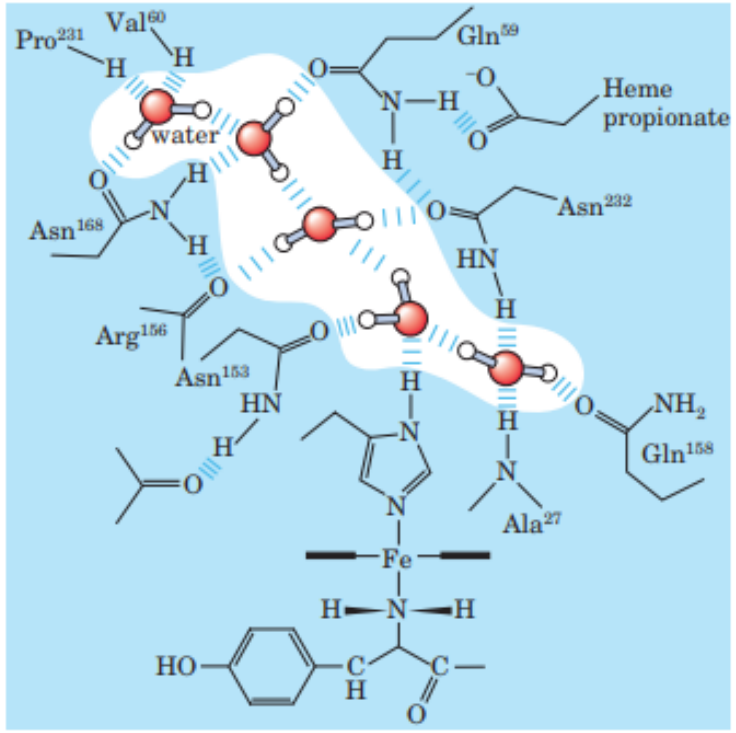


FIGURE 2–10 Water chain in cytochrome *f*. Water is bound in a proton channel of the membrane protein cytochrome *f*, which is part of the energy-trapping machinery of photosynthesis in chloroplasts (see Fig. 19–XX). Five water molecules are hydrogen-bonded to each other and to functional groups of the protein, which include the side chains of valine, proline, arginine, alanine, two asparagine, and two glutamine residues. The protein has a bound heme (see Fig. 5–1), its iron ion facilitating electron flow during photosynthesis. Electron flow is coupled to the movement of protons across the membrane, which probably involves "electron hopping" (see Fig. 2–14) through this chain of bound water molecules.

Water

Solutes Affect the Colligative Properties of Aqueous Solutions

Colligative properties of water: (“Tied together”)
vapor pressure,
boiling point,
melting point (freezing point),
and osmotic pressure,

Solutes alter the colligative properties of aqueous solutions by lowering the effective concentration of water.

Water

Solutes Affect the Colligative Properties of Aqueous Solutions

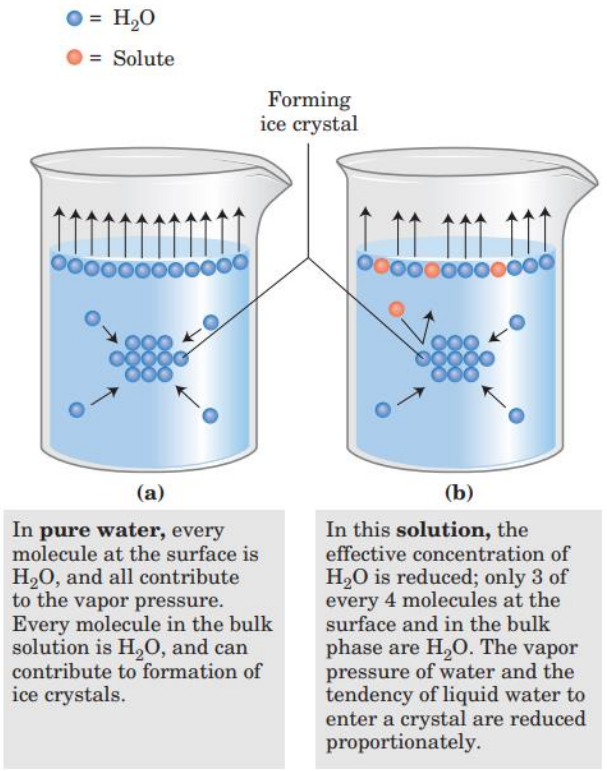


FIGURE 2–11 **Solutes alter the colligative properties of aqueous solutions.** **(a)** At 101 kPa (1 atm) pressure, pure water boils at 100 °C and freezes at 0 °C. **(b)** The presence of solute molecules reduces the probability of a water molecule leaving the solution and entering the gas phase, thereby reducing the vapor pressure of the solution and increasing the boiling point. Similarly, the probability of a water molecule colliding with and joining a forming ice crystal is reduced when some of the molecules colliding with the crystal are solute, not water, molecules. The effect is depression of the freezing point.

Water

Osmosis and

Osmotic pressure: Π (PI)

The osmotic pressure measured as the force necessary to resist water movement is approximated by the van't Hoff equation

$$\Pi = icRT$$

R is the gas constant and T is the absolute temperature.
Osmolarity of the solution: *ic*
It is the the product of the solute's molar concentration c and the van't Hoff factor *i*, which is a measure of the extent to which the solute dissociates into two or more ionic species.

For several solutes

$$\Pi = RT(i_1c_1 + i_2c_2 + \cdots + i_nc_n)$$

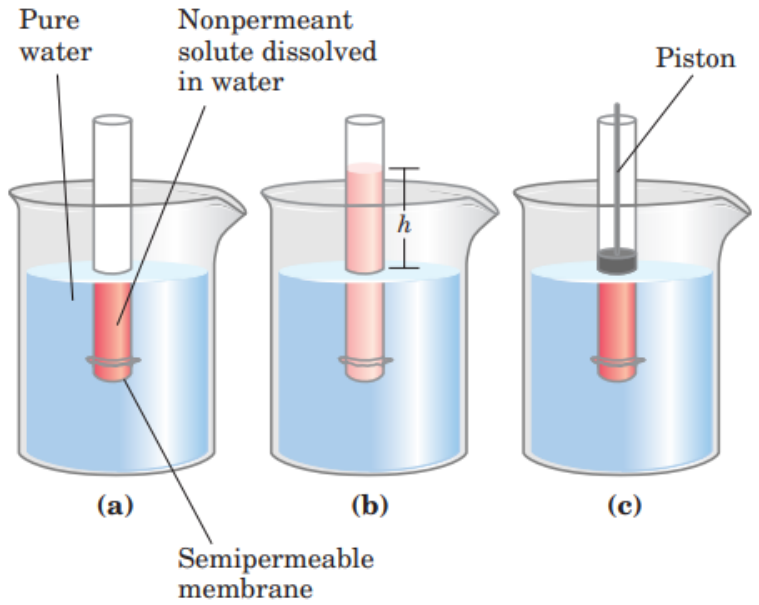


FIGURE 2–12 Osmosis and the measurement of osmotic pressure. **(a)** The initial state. The tube contains an aqueous solution, the beaker contains pure water, and the semipermeable membrane allows the passage of water but not solute. Water flows from the beaker into the tube to equalize its concentration across the membrane. **(b)** The final state. Water has moved into the solution of the nonpermeant compound, diluting it and raising the column of water within the tube. At equilibrium, the force of gravity operating on the solution in the tube exactly balances the tendency of water to move into the tube, where its concentration is lower. **(c)** Osmotic pressure (Π) is measured as the force that must be applied to return the solution in the tube to the level of that in the beaker. This force is proportional to the height, *h*, of the column in **(b)**.

Water

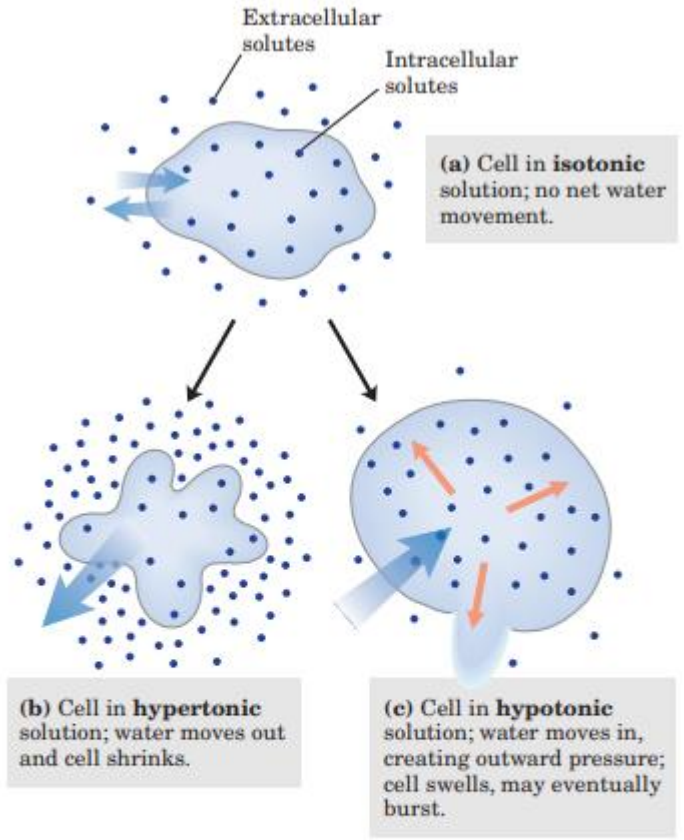
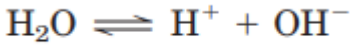


FIGURE 2–13 Effect of extracellular osmolarity on water movement across a plasma membrane. When a cell in osmotic balance with its surrounding medium (that is, in an isotonic medium) **(a)** is transferred into a hypertonic solution **(b)** or hypotonic solution **(c),** water moves across the plasma membrane in the direction that tends to equalize osmolarity outside and inside the cell.

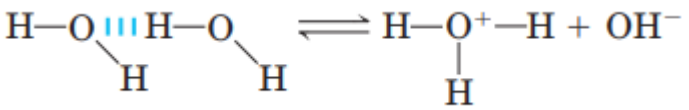
Water

Ionization of Water, Weak Acids, and Weak Bases

Pure Water Is Slightly Ionized



Hydronium ions (H3O)



Water

Proton hopping in water is
Very fast

Faster than Na^+ or K^+.

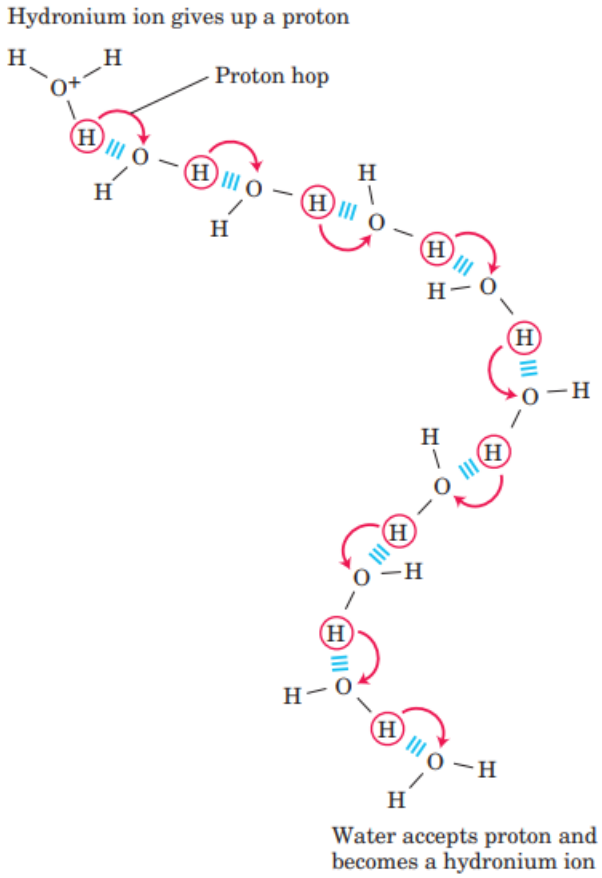


FIGURE 2–14 **Proton hopping.** Short "hops" of protons between a series of hydrogen-bonded water molecules effect an extremely rapid net movement of a proton over a long distance. As a hydronium ion (upper left) gives up a proton, a water molecule some distance away (lower right) acquires one, becoming a hydronium ion. Proton hopping is much faster than true diffusion and explains the remarkably high ionic mobility of H^+ ions compared with other monovalent cations such as Na^+ or K^+.

Water

$$A + B \rightleftharpoons C + D$$

Equilibrium constant, Keq $K_{eq} = \frac{[C][D]}{[A][B]}$

The Ionization of Water Is Expressed by an **Equilibrium Constant**

$$K_{eq} = \frac{[H^+][OH^-]}{[H_2O]}$$

In pure water at 25 C, the concentration of water is 55.5 M (grams of H2O in 1 L divided by its gram molecular weight: (1,000 g/L)/(18.015 g/mol))

Substituting 55.5 M in the equilibrium constant formula $K_{eq} = \frac{[H^+][OH^-]}{55.5\ \text{M}}$ or $(55.5\ \text{M})(K_{eq}) = [H^+][OH^-] = K_w$

Kw designates the product (55.5 M)(Keq), **the ion product of water** at 25 °C.

Water

The value for Keq, determined by electrical-conductivity measurements of pure water, is 1.8×10^{-16} M at 25 C.

Since $$(55.5\ \text{M})(K_{eq}) = [\text{H}^+][\text{OH}^-] = K_w$$

So the ion product of water at 25 °C

$$K_w = [\text{H}^+][\text{OH}^-] = (55.5\ \text{M})(1.8 \times 10^{-16}\ \text{M})$$
$$= 1.0 \times 10^{-14}\ \text{M}^2$$

At neutral pH there are exactly equal concentrations of H and OH, in pure water

$$K_w = [\text{H}^+][\text{OH}^-] = [\text{H}^+]^2$$

$$[\text{H}^+] = \sqrt{K_w} = \sqrt{1 \times 10^{-14}\ \text{M}^2}$$

$$[\text{H}^+] = [\text{OH}^-] = 10^{-7}\ \text{M}$$

Water

The pH scale

$$\mathrm{pH} = \log \frac{1}{[\mathrm{H}^+]} = -\log [\mathrm{H}^+]$$

The symbol p denotes “negative logarithm of.”

pH of neutral solution at 25 C

$$\mathrm{pH} = \log \frac{1}{1.0 \times 10^{-7}} = \log (1.0 \times 10^{7})$$

$$= \log 1.0 + \log 10^{7} = 0 + 7 = 7$$

Strong acid – Stronger dissociation of proton- Lower the pH value

Water

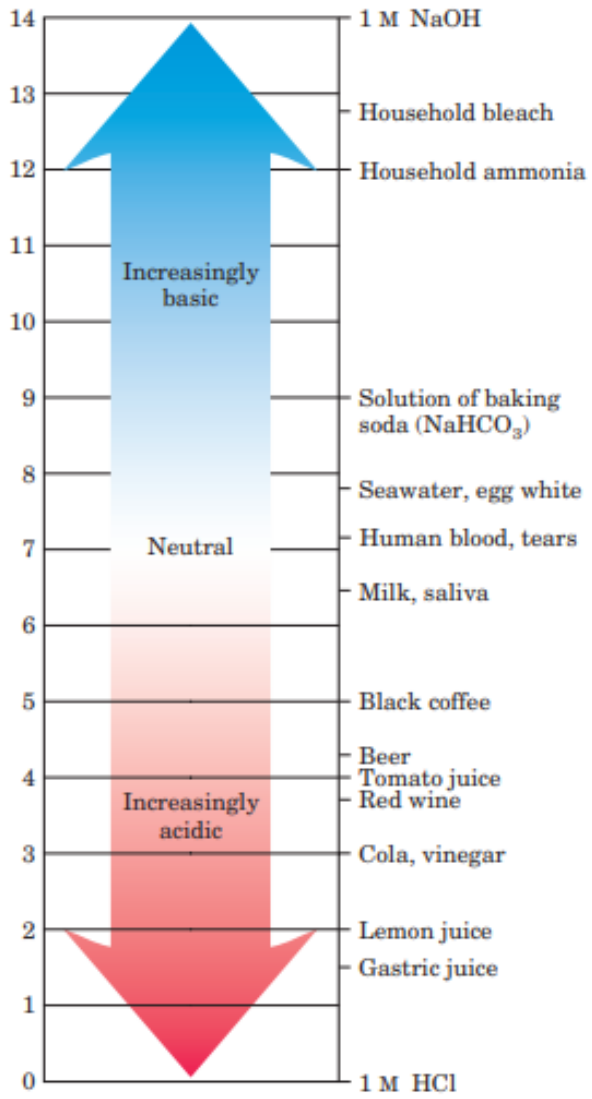


FIGURE 2–15 The pH of some aqueous fluids.

Water

Weak Acids and Bases Have Characteristic Dissociation Constants

Acids: proton donors

Bases: proton acceptors

Strong acids or bases ionize completely in dilute aqueous solution: HCl

Weak acids: Not completely

A proton donor and its corresponding proton acceptor make up a **conjugate acid-base pair**.

Acetic acid (CH3COOH), a proton donor, and the acetate anion (CH3COO), the corresponding proton acceptor, constitute a conjugate acid base pair:

$$CH_3COOH \rightleftharpoons H^+ + CH_3COO^- \qquad HA \rightleftharpoons H^+ + A^-,$$

$$K_{eq} = \frac{[H^+][A^-]}{[HA]} = K_a \qquad K_a \text{ : Dissociation constants}$$

Water

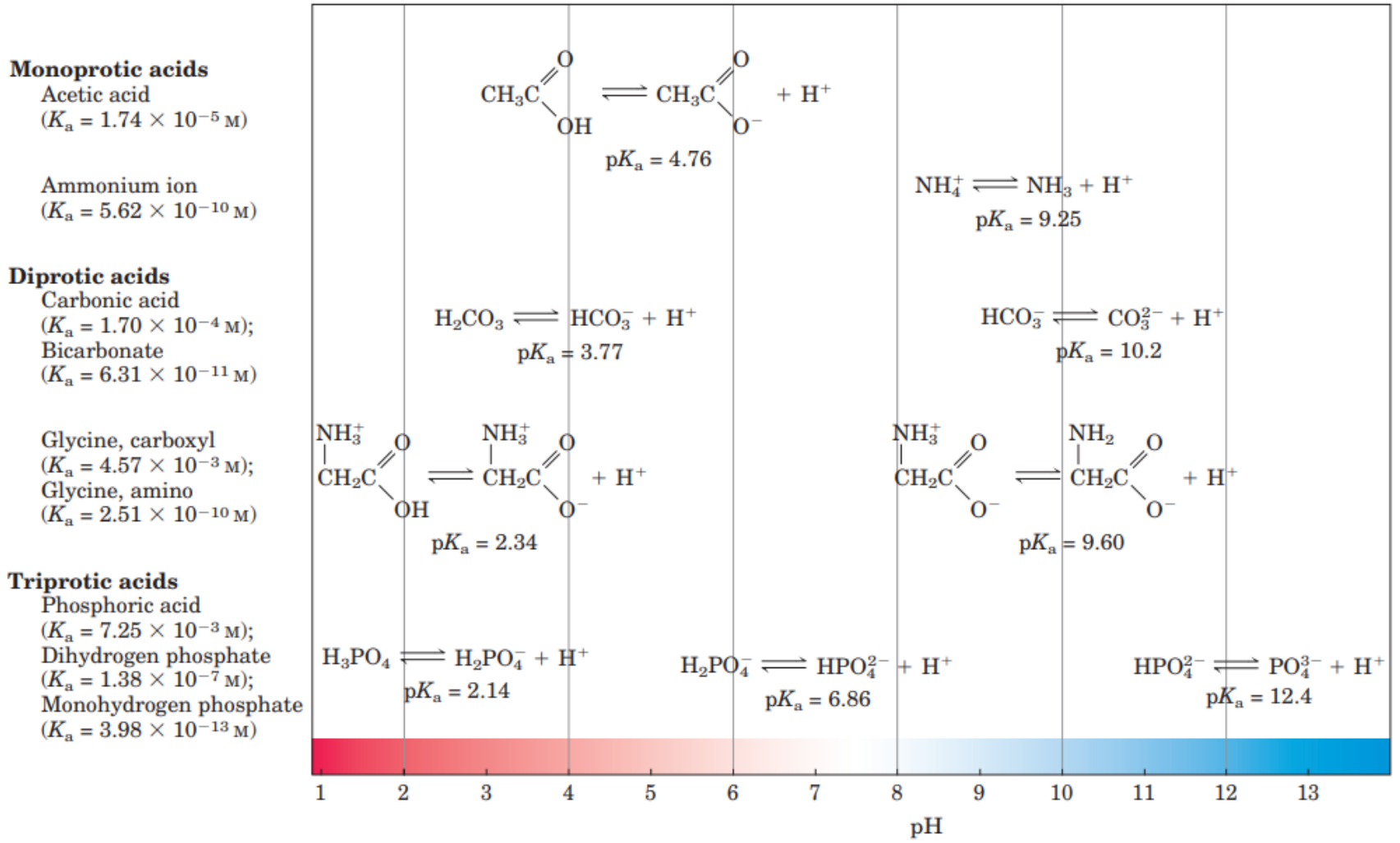


FIGURE 2–16 Conjugate acid-base pairs consist of a proton donor and a proton acceptor. Some compounds, such as acetic acid and ammonium ion, are monoprotic; they can give up only one proton. Others are diprotic (H_2CO_3 (carbonic acid) and glycine) or triprotic (H_3PO_4 (phosphoric acid)). The dissociation reactions for each pair are shown where they occur along a pH gradient. The equilibrium or dissociation constant (K_a) and its negative logarithm, the pK_a, are shown for each reaction.

Water

$$pK_a = \log \frac{1}{K_a} = -\log K_a$$

Strong acid – Stronger dissociation of proton- Lower the pH vale- Lower the pKa value

Water

Titration Curves Reveal the pKa of Weak Acids

Buffer:

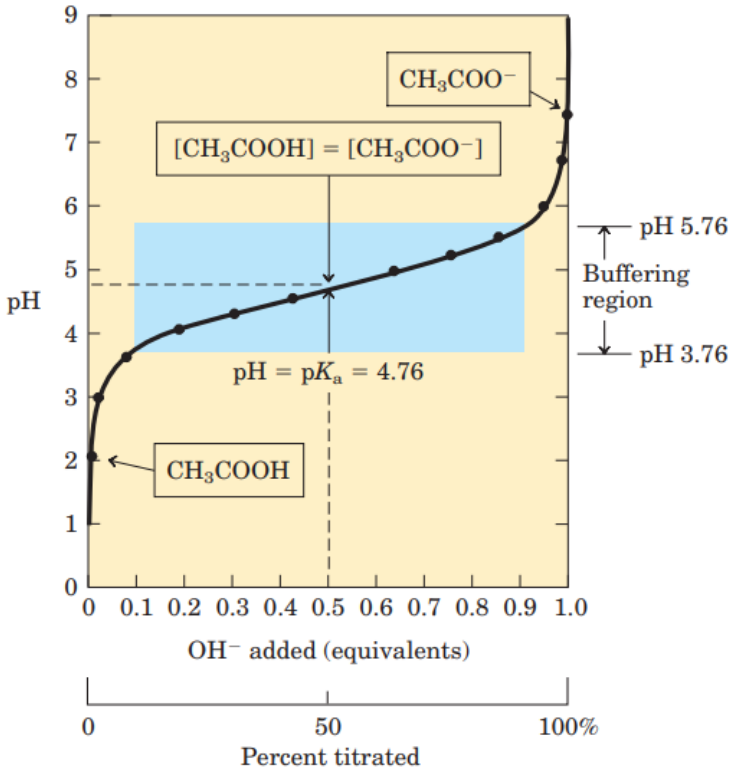


FIGURE 2–17 The titration curve of acetic acid. After addition of each increment of NaOH to the acetic acid solution, the pH of the mixture is measured. This value is plotted against the amount of NaOH expressed as a fraction of the total NaOH required to convert all the acetic acid to its deprotonated form, acetate. The points so obtained yield the titration curve. Shown in the boxes are the predominant ionic forms at the points designated. At the midpoint of the titration, the concentrations of the proton donor and proton acceptor are equal, and the pH is numerically equal to the pK_a. The shaded zone is the useful region of buffering power, generally between 10% and 90% titration of the weak acid.

Water

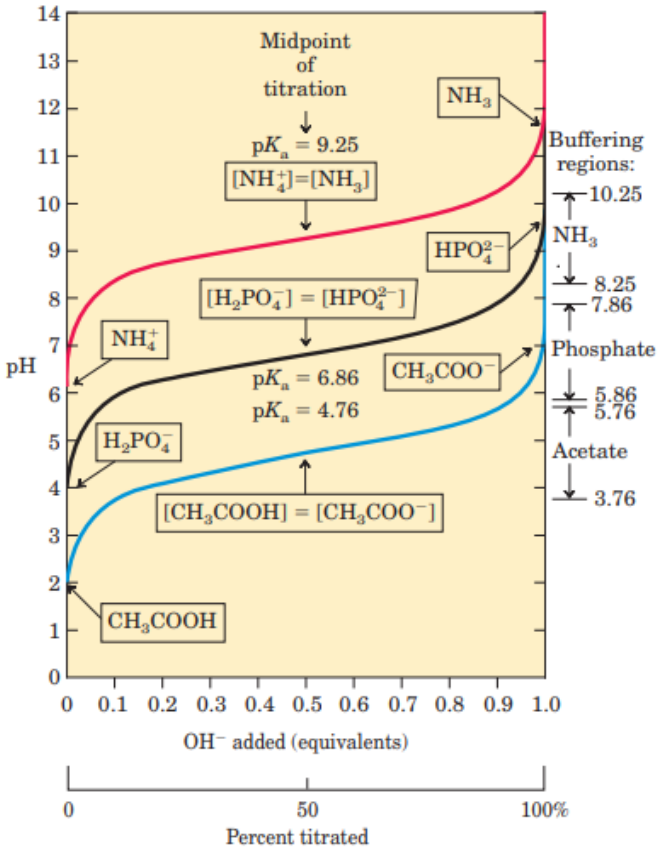


FIGURE 2–18 Comparison of the titration curves of three weak acids. Shown here are the titration curves for CH_3COOH, $H_2PO_4^-$, and NH_4^+. The predominant ionic forms at designated points in the titration are given in boxes. The regions of buffering capacity are indicated at the right. Conjugate acid-base pairs are effective buffers between approximately 10% and 90% neutralization of the proton-donor species.

Water

Buffering against pH Changes in Biological Systems

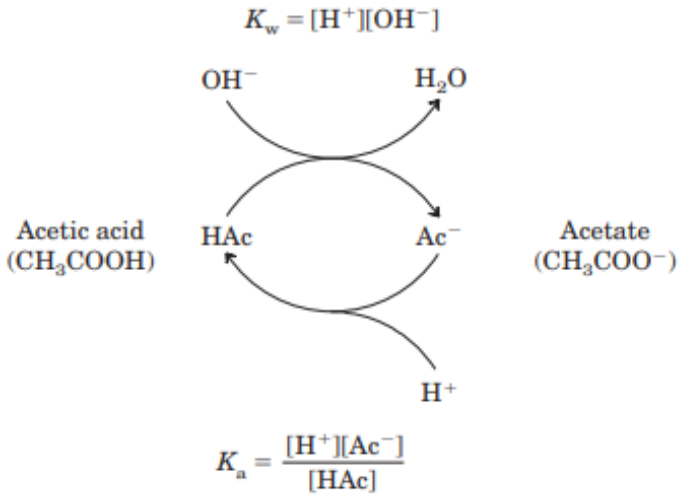


FIGURE 2–19 The acetic acid–acetate pair as a buffer system. The system is capable of absorbing either H^+ or OH^- through the reversibility of the dissociation of acetic acid. The proton donor, acetic acid (HAc), contains a reserve of bound H^+, which can be released to neutralize an addition of OH^- to the system, forming H_2O. This happens because the product $[H^+][OH^-]$ transiently exceeds K_w (1×10^{-14} M^2). The equilibrium quickly adjusts so that this product equals 1×10^{-14} M^2 (at 25 °C), thus transiently reducing the concentration of H^+. But now the quotient $[H^+][Ac^-] / [HAc]$ is less than K_a, so HAc dissociates further to restore equilibrium. Similarly, the conjugate base, Ac^-, can react with H^+ ions added to the system; again, the two ionization reactions simultaneously come to equilibrium. Thus a conjugate acid-base pair, such as acetic acid and acetate ion, tends to resist a change in pH when small amounts of acid or base are added. Buffering action is simply the consequence of two reversible reactions taking place simultaneously and reaching their points of equilibrium as governed by their equilibrium constants, K_W and K_a.

Water

A Simple Expression Relates pH, pKa, and Buffer Concentration

Henderson Hasselbalch equation

$$K_a = \frac{[H^+][A^-]}{[HA]} \qquad \text{so} \qquad [H^+] = K_a \frac{[HA]}{[A^-]}$$

Put Negative log on both side

$$-\log [H^+] = -\log K_a - \log \frac{[HA]}{[A^-]} \qquad \text{Then} \quad pH = pK_a - \log \frac{[HA]}{[A^-]} \qquad \text{Invert} \quad pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$pH = pK_a + \log \frac{[\text{proton acceptor}]}{[\text{proton donor}]}$$

At that point, [HA] equals [A] $\quad pH = pK_a + \log 1 = pK_a + 0 = pK_a$

Water

Weak Acids or Bases Buffer Cells and Tissues against pH Changes

Two especially important biological buffers are

The phosphate and

The bicarbonate systems

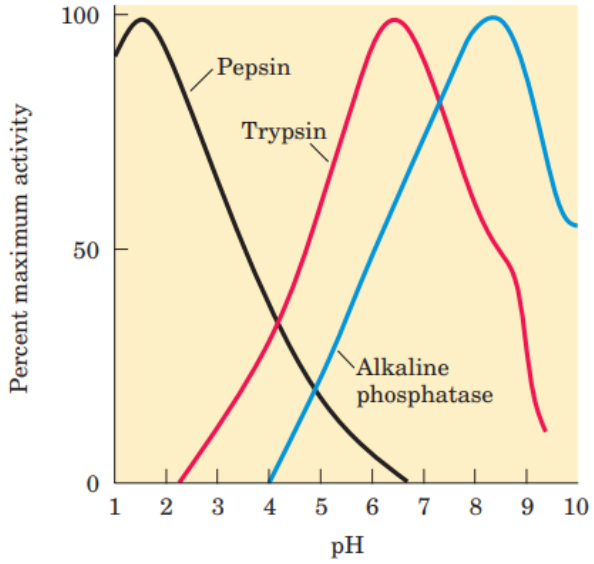


FIGURE 2–21 The pH optima of some enzymes. Pepsin is a digestive enzyme secreted into gastric juice; trypsin, a digestive enzyme that acts in the small intestine; alkaline phosphatase of bone tissue, a hydrolytic enzyme thought to aid in bone mineralization.

Water

Water can act as a Reactant

Condensation reaction

Hydrolysis reaction

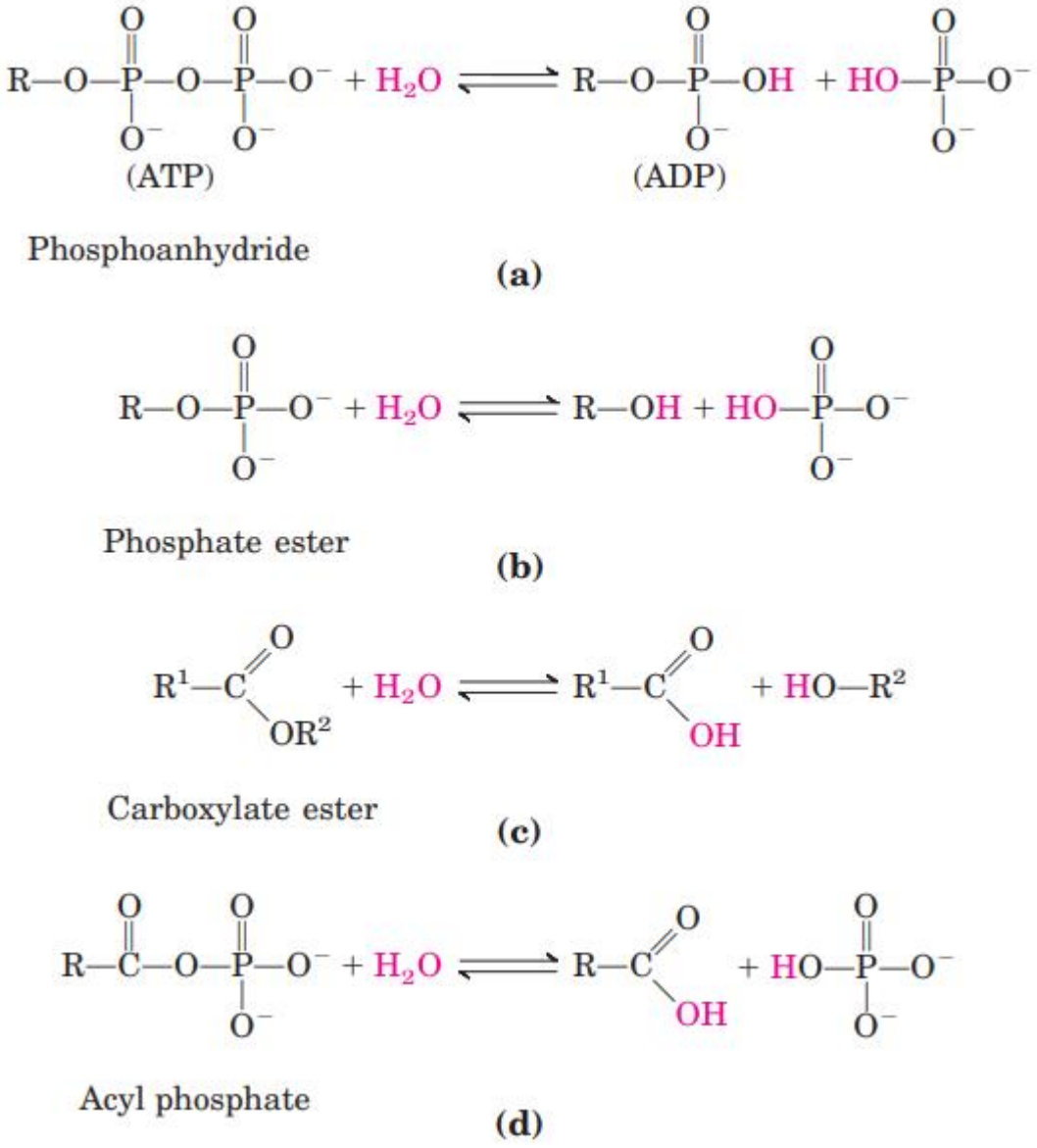


FIGURE 2–22 **Participation of water in biological reactions.** **(a)** ATP is a phosphoanhydride formed by a condensation reaction (loss of the elements of water) between ADP and phosphate. R represents adenosine monophosphate (AMP). This condensation reaction requires energy. The hydrolysis of (addition of the elements of water to) ATP to form ADP and phosphate releases an equivalent amount of energy. Also shown are some other condensation and hydrolysis reactions common in biological systems **(b), (c), (d).**