WATER

- Water is the most abundant substance in living systems, making up 70% or more of the weight of most organisms.

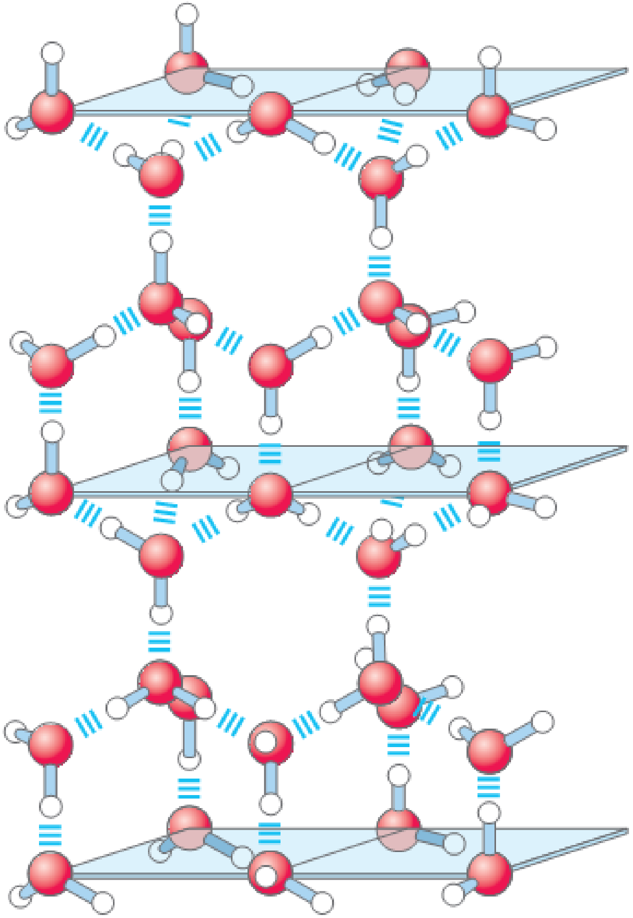
- The first living organisms doubtless arose in an aqueous environment, and the course of evolution has been shaped by the properties of the aqueous medium in which life began.

- Hydrogen bonds between water molecules provide the cohesive forces that make water a liquid at room temperature and that favor the extreme ordering of molecules that is typical of crystalline water (ice).

- Polar biomolecules dissolve readily in water because they can replace water-water interactions with more energetically favorable water-solute interactions.

- In contrast, nonpolar biomolecules interfere with water-water interactions but are unable to form water-solute interactions—consequently, nonpolar molecules are poorly soluble in water.

- In aqueous solutions, nonpolar molecules tend to cluster together.



- Hydrogen bonds and ionic, hydrophobic, and van der Waals interactions are individually weak, but collectively they have a very significant influence on the three-dimensional structures of proteins, nucleic acids, polysaccharides, and membrane lipids.

- In a crystal of salt, the positive and negative ions are held together by **ionic bonds**.

- Ionic bonds and covalent bonds are the strongest bonds, being many times stronger than the next weakest one.

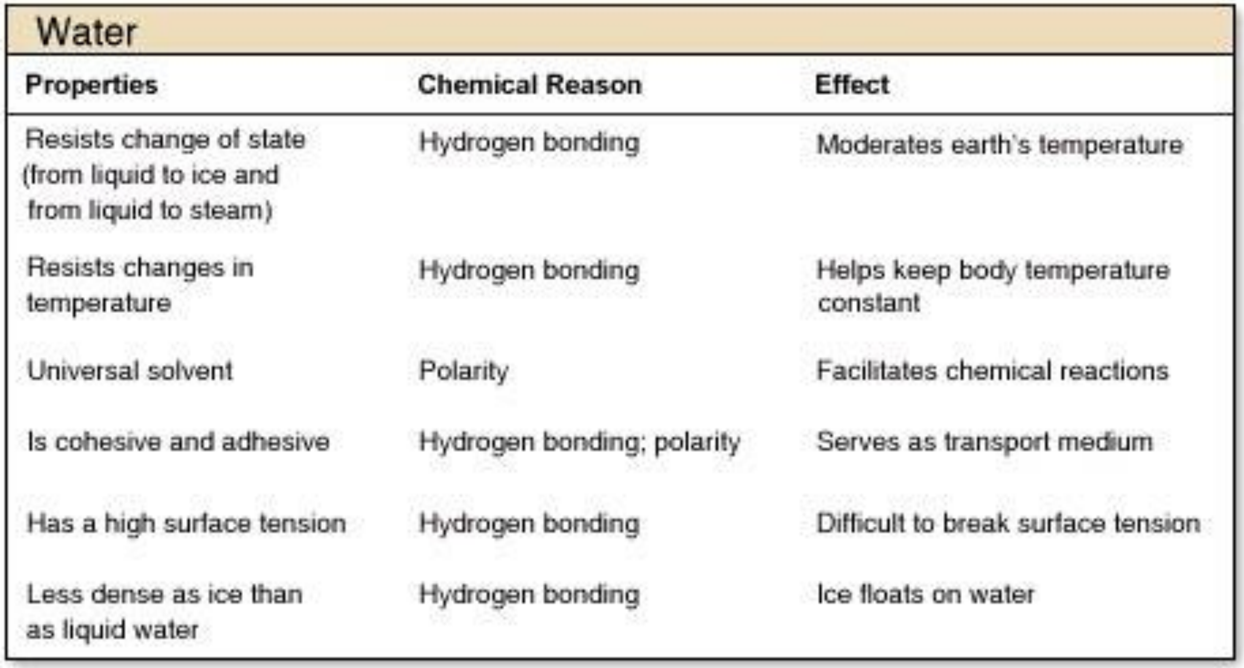
- Similar to ionic bonds found in salt crystals, biomolecules often have ionizable groups on them.

- When these oppositely charged molecules are in close proximity to one another, we call the resulting bond a **salt bridge**.

- Water has a higher melting point, boiling point, and heat of vaporization than most other common solvents.

- These unusual properties are a consequence of attractions between adjacent water molecules that give liquid water great internal cohesion.

- A look at the electron structure of the H_2O molecule reveals the cause of these intermolecular attractions.



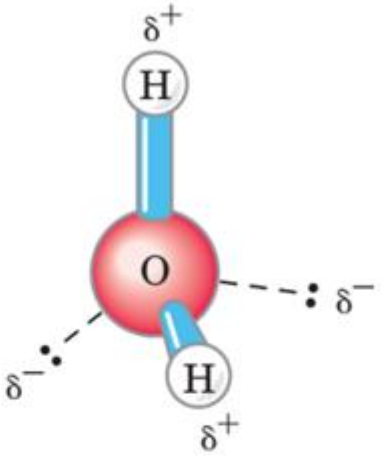
Water

Properties	Chemical Reason	Effect
Resists change of state (from liquid to ice and from liquid to steam)	Hydrogen bonding	Moderates earth's temperature
Resists changes in temperature	Hydrogen bonding	Helps keep body temperature constant
Universal solvent	Polarity	Facilitates chemical reactions
Is cohesive and adhesive	Hydrogen bonding; polarity	Serves as transport medium
Has a high surface tension	Hydrogen bonding	Difficult to break surface tension
Less dense as ice than as liquid water	Hydrogen bonding	Ice floats on water

- Each hydrogen atom of a water molecule shares an electron pair with the central oxygen atom.

- The geometry of the molecule is dictated by the shapes of the outer electron orbitals of the oxygen atom, which are similar to the sp3 bonding orbitals of carbon.

- The oxygen nucleus attracts electrons more strongly than does the hydrogen nucleus; that is, oxygen is more electronegative.



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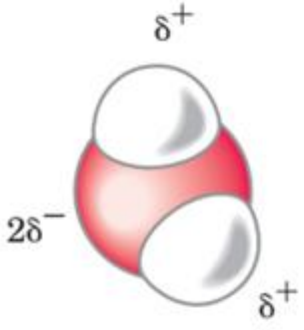

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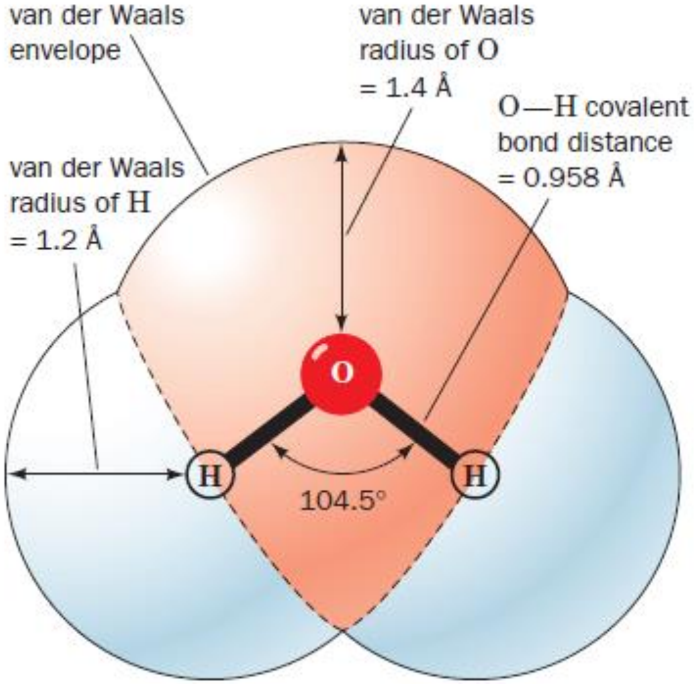
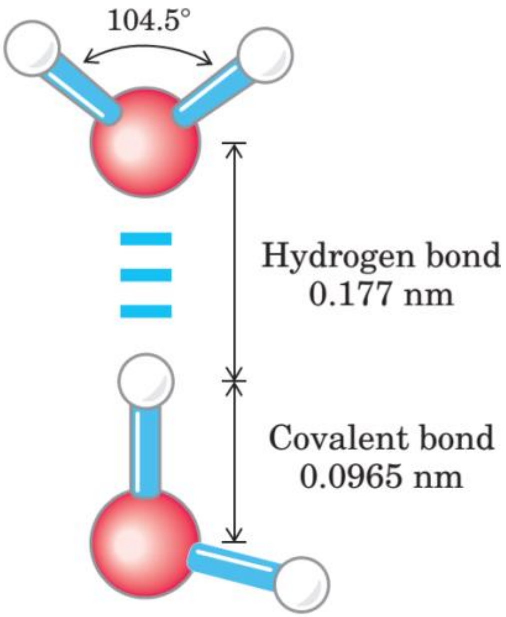


- The result of this unequal electron sharing is two electric dipoles in the water molecule, one along each of the H-O bonds; each hydrogen bears a partial positive charge (δ^+) and the oxygen atom bears a partial negative charge equal to the sum of the two partial positives ($2\delta^-$).

- As a result, there is an electrostatic attraction between the oxygen atom of one water molecule and the hydrogen of another, called a **hydrogen bond**.



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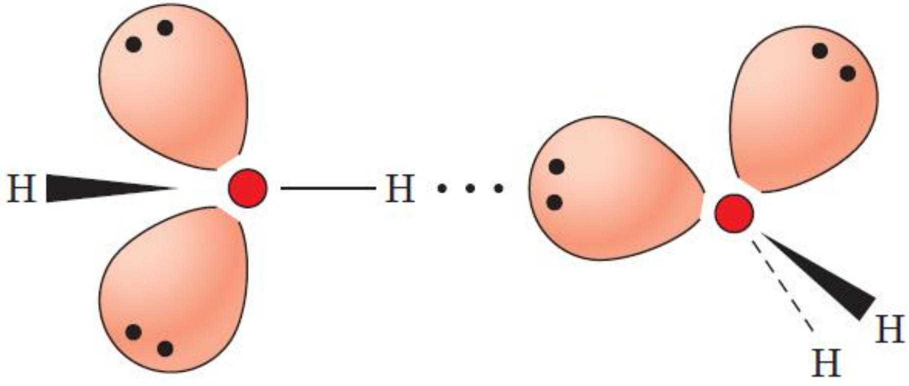


- Hydrogen bonds are relatively weak. Those in liquid water have a bond dissociation energy (the energy required to break a bond) of about 23 kJ/mol, compared with 470 kJ/mol for the covalent O-H bond in water.

- The hydrogen bond is about 10% covalent, due to overlaps in the bonding orbitals, and about 90% electrostatic.

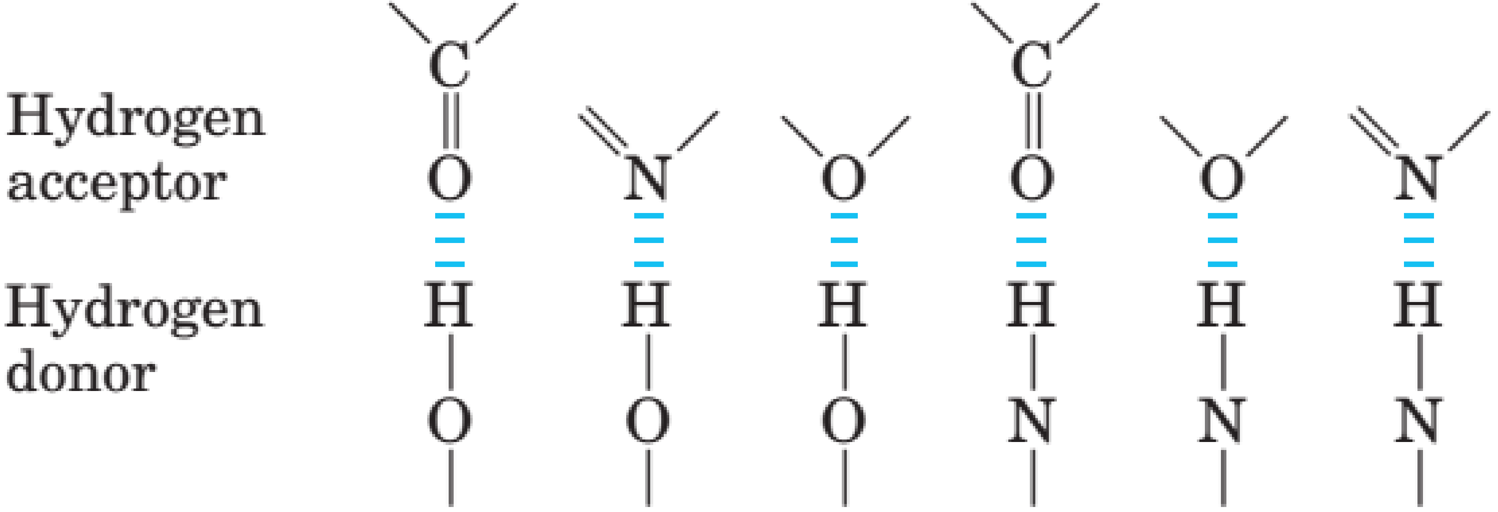
- At room temperature, hydrogen bonds between water molecules constantly break and reform.

- In ice, on the other hand, each water molecule is fixed in space and forms hydrogen bonds with a full complement of four other water molecules to yield a regular lattice structure

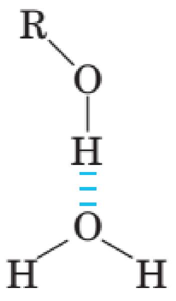
- Breaking a sufficient proportion of hydrogen bonds to destabilize the crystal lattice of ice requires much thermal energy, which accounts for the relatively high melting point of water.

- Hydrogen bonds are not unique to water.

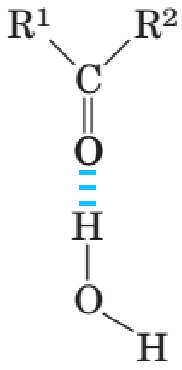
- They readily form between an electronegative atom (the hydrogen acceptor, usually oxygen or nitrogen with a lone pair of electrons) and a hydrogen atom covalently bonded to another electronegative atom (the hydrogen donor) in the same or another molecule.



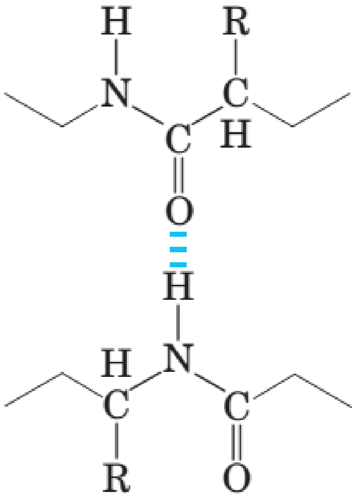
Between the hydroxyl group of an alcohol and water



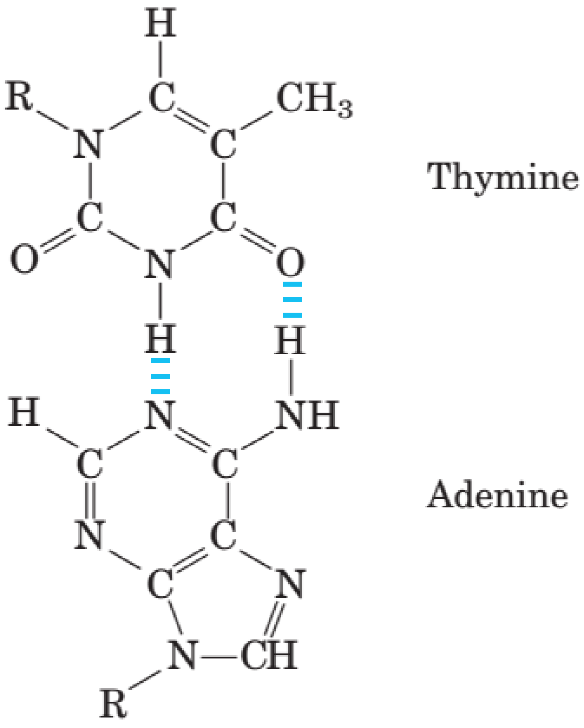
Between the carbonyl group of a ketone and water

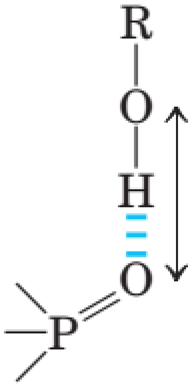


Between peptide groups in polypeptides

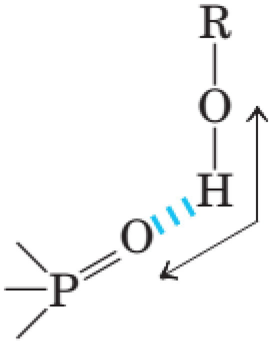


Between complementary bases of DNA





Strong hydrogen bond



Weaker hydrogen bond

- Hydrogen atoms covalently bonded to carbon atoms do not participate in hydrogen bonding, because carbon is only slightly more electronegative than hydrogen and thus the C-H bond is only very weakly polar.

- The distinction explains why butanol has a relatively high boiling point of 117 °C, whereas butane has a boiling point of only -0.5 ° C.

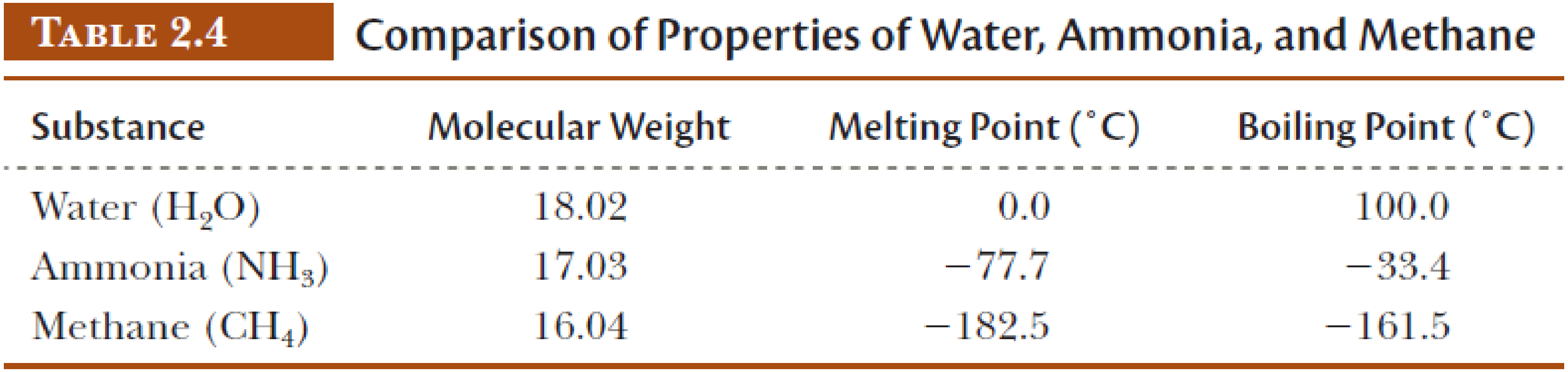


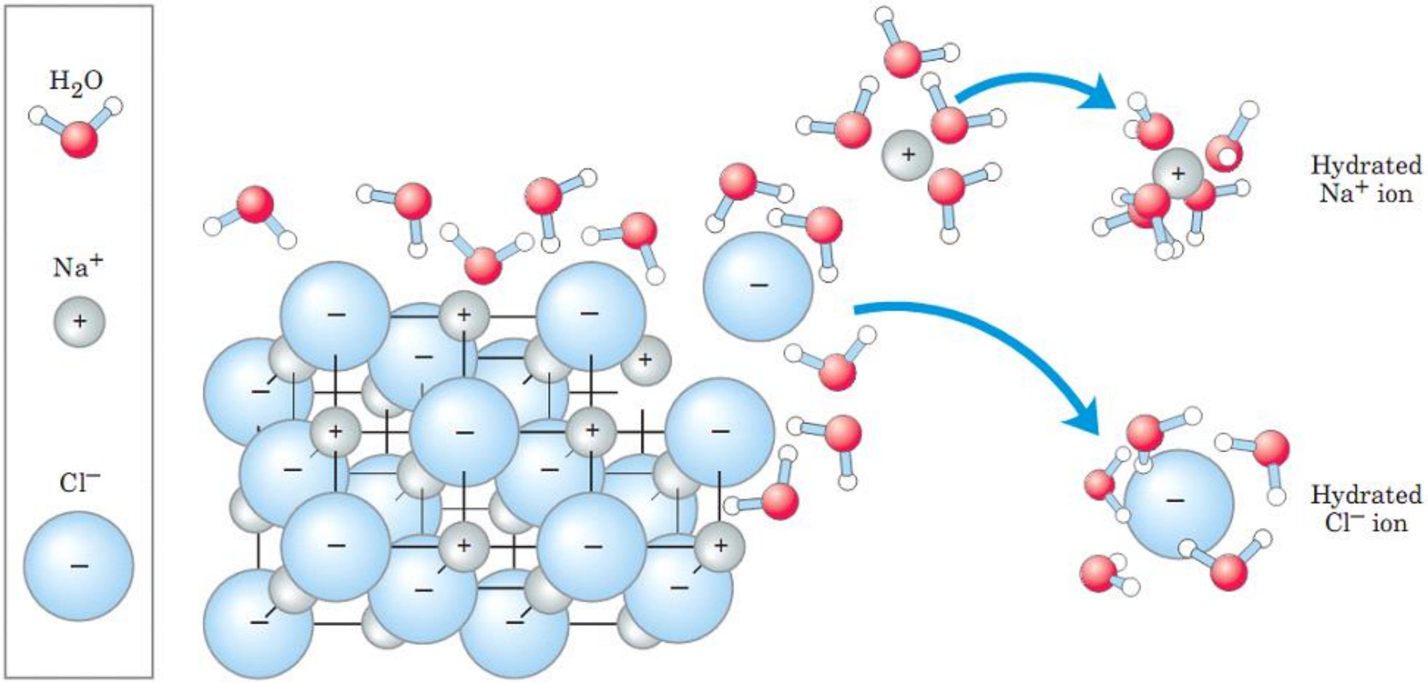
TABLE 2.4 Comparison of Properties of Water, Ammonia, and Methane

Substance	Molecular Weight	Melting Point (°C)	Boiling Point (°C)
Water (H_2O)	18.02	0.0	100.0
Ammonia (NH_3)	17.03	−77.7	−33.4
Methane (CH_4)	16.04	−182.5	−161.5

- Uncharged but polar biomolecules such as sugars dissolve readily in water because of the stabilizing effect of hydrogen bonds between the hydroxyl groups or carbonyl oxygen of the sugar and the polar water molecules. Alcohols, aldehydes, ketones, and compounds containing N-H bonds all form hydrogen bonds with water molecules and tend to be soluble in water.

- ❖ Water is a polar solvent. It readily dissolves most biomolecules, which are generally charged or polar compounds. Compounds that dissolve easily in water are hydrophilic.

- ❖ Water dissolves salts such as NaCl by hydrating and stabilizing the Na^+ and Cl^- ions, weakening the electrostatic interactions between them and thus counteracting their tendency to associate in a crystalline lattice.



- In contrast, nonpolar solvents such as chloroform and benzene are poor solvents for polar biomolecules but easily dissolve those that are hydrophobic—nonpolar molecules such as lipids and waxes.

- The same factors apply to charged biomolecules, compounds with functional groups such as ionized carboxylic acids, protonated amines, and phosphate esters or anhydrides.

- Water readily dissolves such compounds by replacing solute-solute hydrogen bonds with solute-water hydrogen bonds, thus screening the electrostatic interactions between solute molecules.

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

(F) force of ionic interactions

(Q) the magnitude of the charges

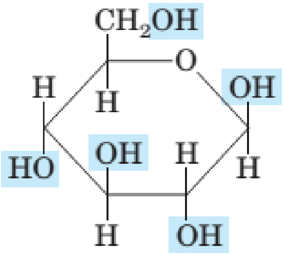
(r) the distance between the charged groups

(ε) dielectric constant

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

Polar

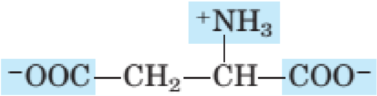
Glucose



Glycine

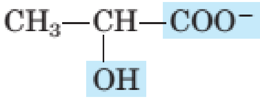
$^{+}NH_3-CH_2-COO^{-}$

Aspartate



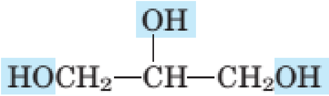
$$^{-}OOC-CH_2-\overset{^{+}NH_3}{\overset{|}{CH}}-COO^{-}$$

Lactate



$$CH_3-\underset{OH}{\underset{|}{CH}}-COO^{-}$$

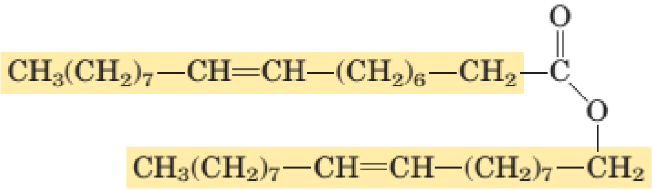
Glycerol



$$HOCH_2-\overset{OH}{\overset{|}{CH}}-CH_2OH$$

Nonpolar

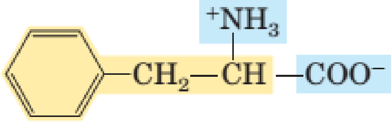
Typical wax



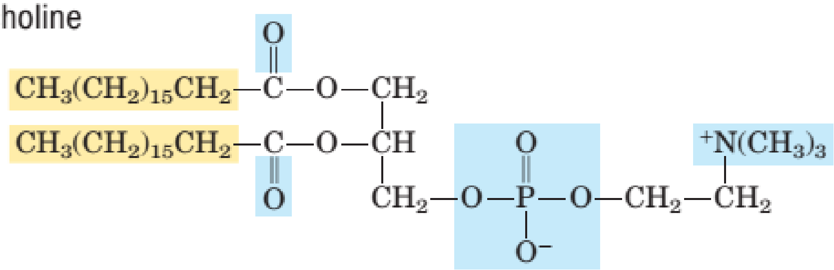
$$\begin{array}{r} CH_3(CH_2)_7-CH=CH-(CH_2)_6-CH_2-\overset{O}{\overset{\|}{C}} \\ \diagdown\, O \\ | \\ CH_3(CH_2)_7-CH=CH-(CH_2)_7-CH_2 \end{array}$$

Amphipathic

Phenylalanine



Phosphatidylcholine



🟦 Polar groups 🟨 Nonpolar groups

- The molecules of the biologically important gases CO_2, O_2, and N_2 are nonpolar.

- In O_2 and N_2, electrons are shared equally by both atoms. In CO_2, each C-O bond is polar, but the two dipoles are oppositely directed and cancel each other.

- The movement of molecules from the disordered gas phase into aqueous solution constrains their motion and the motion of water molecules and therefore represents a decrease in entropy.

- The nonpolar nature of these gases and the decrease in entropy when they enter solution combine to make them very poorly soluble in water.

- Some organisms have water-soluble carrier proteins (hemoglobin and myoglobin, for example) that facilitate the transport of O_2.

- Carbon dioxide forms carbonic acid (H_2CO_3) in aqueous solution and is transported as the HCO_3^- (bicarbonate) ion, either free—bicarbonate is very soluble in water (~100 g/L at 25 °C)—or bound to hemoglobin.

- Two other gases, NH_3 and H_2S, also have biological roles in some organisms; these gases are polar and dissolve readily in water.

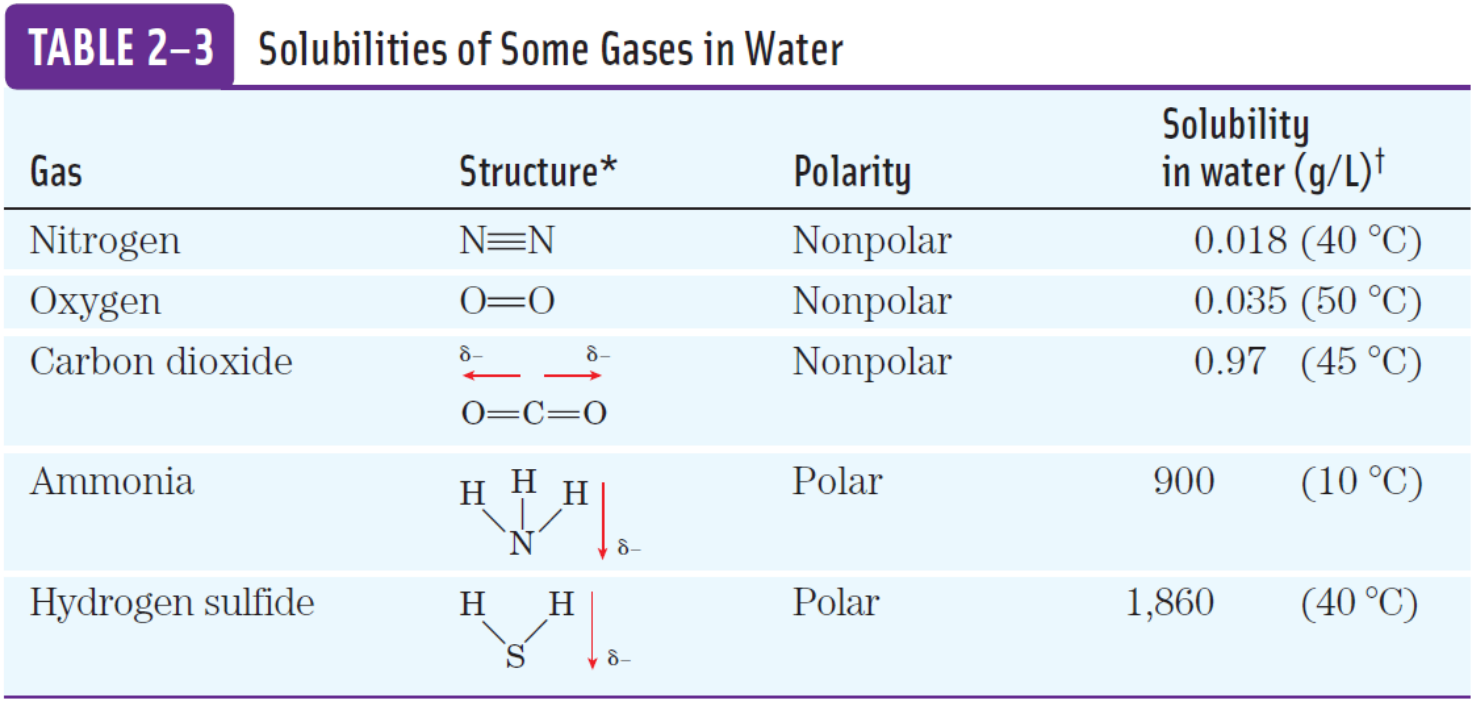


TABLE 2–3 Solubilities of Some Gases in Water

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	δ– δ– O═C═O	Nonpolar	0.97 (45 °C)
Ammonia	H H H N δ–	Polar	900 (10 °C)
Hydrogen sulfide	H H S δ–	Polar	1,860 (40 °C)

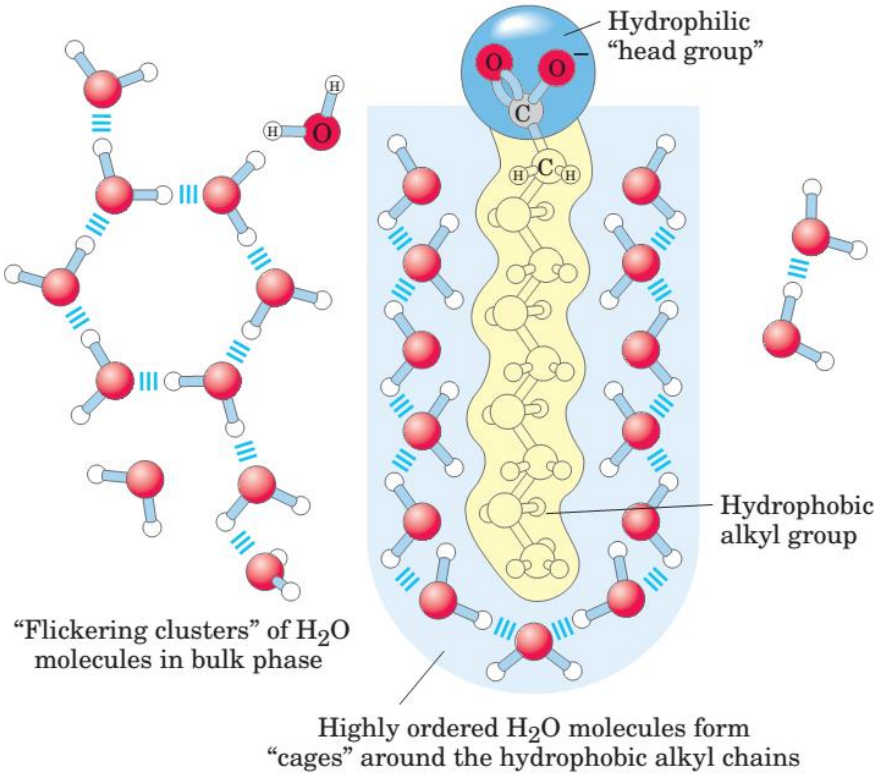
- When water is mixed with benzene or hexane, two phases form; neither liquid is soluble in the other.

- Nonpolar compounds such as benzene and hexane are hydrophobic—they are unable to undergo energetically favorable interactions with water molecules, and they interfere with the hydrogen bonding among water molecules.

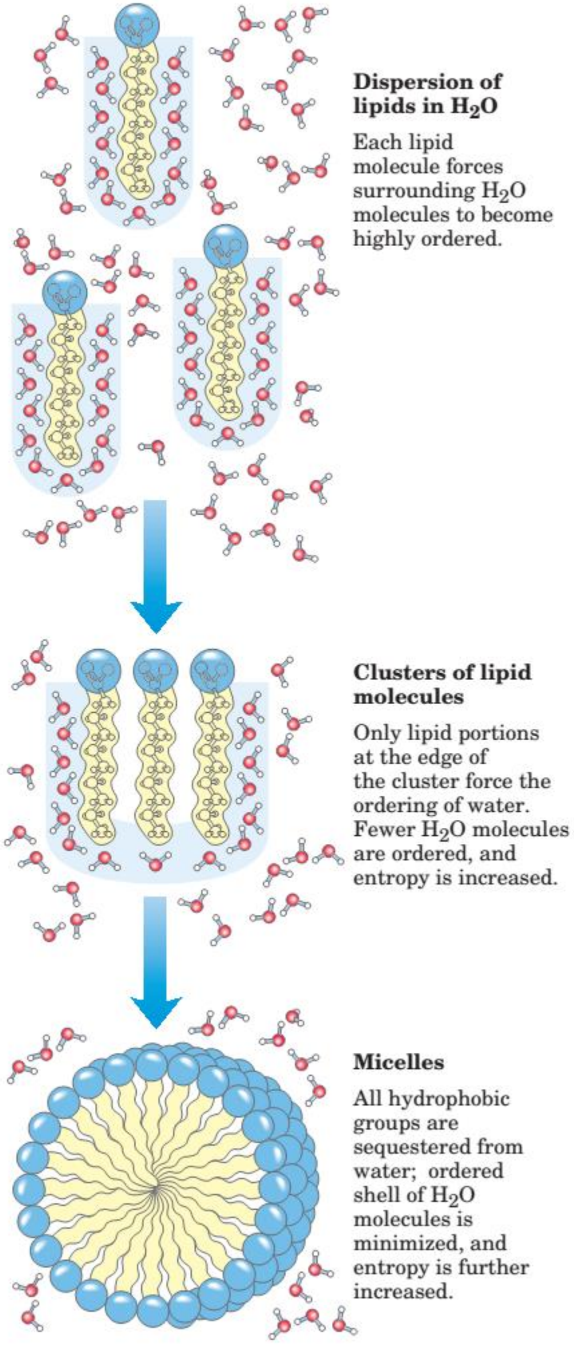
- All molecules or ions in aqueous solution interfere with the hydrogen bonding of some water molecules in their immediate vicinity, but polar or charged solutes (such as NaCl) compensate for lost water-water hydrogen bonds by forming new solute-water interactions.

- The net change in enthalpy (H) for dissolving these solutes is generally small.

- Hydrophobic solutes, however, offer no such compensation, and their addition to water may therefore result in a small gain of enthalpy.



(a)



Dispersion of lipids in H_2O

Each lipid molecule forces surrounding H_2O molecules to become highly ordered.

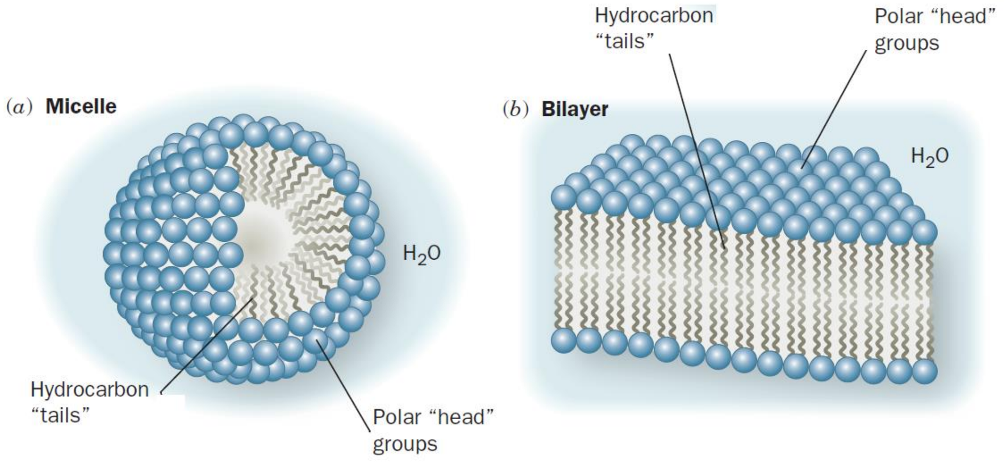
Clusters of lipid molecules

Only lipid portions at the edge of the cluster force the ordering of water. Fewer H_2O molecules are ordered, and entropy is increased.

Micelles

All hydrophobic groups are sequestered from water; ordered shell of H_2O molecules is minimized, and entropy is further increased.

(b)



(*a*) **Micelle**

(*b*) **Bilayer**

- **Amphipathic** compounds contain regions that are polar (or charged) and regions that are nonpolar .

- When an amphipathic compound is mixed with water, the polar, hydrophilic region interacts favorably with the solvent and tends to dissolve, but the nonpolar, hydrophobic region tends to avoid contact with the water.

- The nonpolar regions of the molecules cluster together to present the smallest hydrophobic area to the aqueous solvent, and the polar regions are arranged to maximize their interaction with the solvent.

- These stable structures of amphipathic compounds in water, called **micelles**, may contain hundreds or thousands of molecules.

- The forces that hold the nonpolar regions of the molecules together are called **hydrophobic interactions**.

- The strength of hydrophobic interactions is not due to any intrinsic attraction between nonpolar moieties.

- Rather, it results from the system’s achieving the greatest thermodynamic stability by minimizing the number of ordered water molecules required to surround hydrophobic portions of the solute molecules.

- Many biomolecules are amphipathic; they have both polar and nonpolar surface regions.

- Hydrophobic interactions among lipids, and between lipids and proteins, are the most important determinants of structure in biological membranes.

- Hydrophobic interactions between nonpolar amino acids also stabilize the three-dimensional structures of proteins.

- When two uncharged atoms are brought very close together, their surrounding electron clouds influence each other.

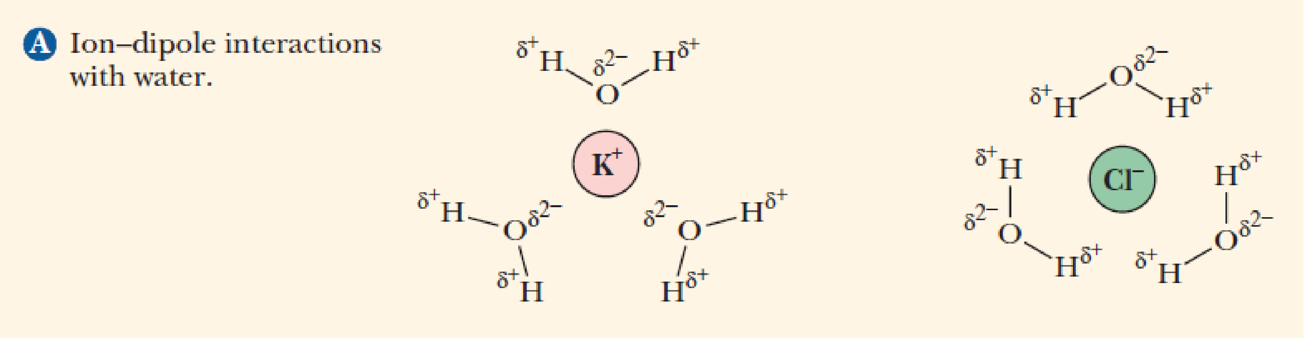
- Random variations in the positions of the electrons around one nucleus may create a transient electric dipole, which induces a transient, opposite electric dipole in the nearby atom.

- The two dipoles weakly attract each other, bringing the two nuclei closer.

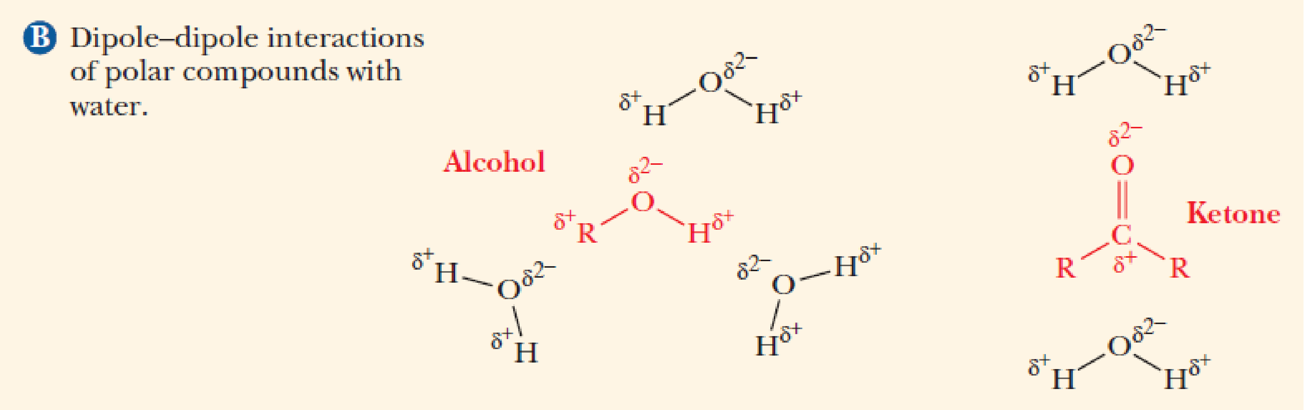
- These weak attractions are called **van der Waals** interactions (also known as London forces).

- As the two nuclei draw closer together, their electron clouds begin to repel each other. At the point where the van der Waals attraction exactly balances this repulsive force, the nuclei are said to be in van der Waals contact.

- Each atom has a characteristic van der Waals radius, a measure of how close that atom will allow another to approach.

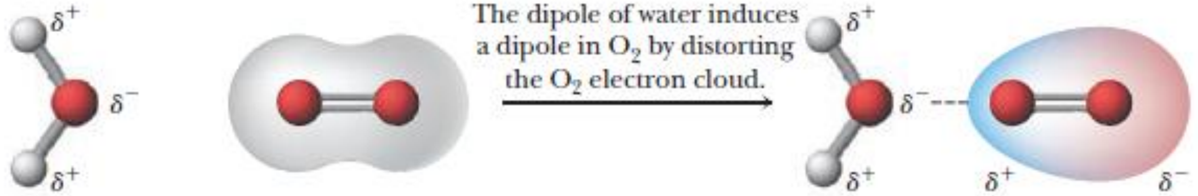


A Ion–dipole interactions with water.



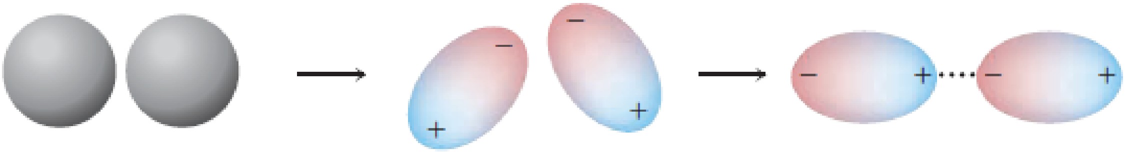
B Dipole–dipole interactions of polar compounds with water.

Dipole–induced dipole interaction.



The dipole of water induces a dipole in O_2 by distorting the O_2 electron cloud.

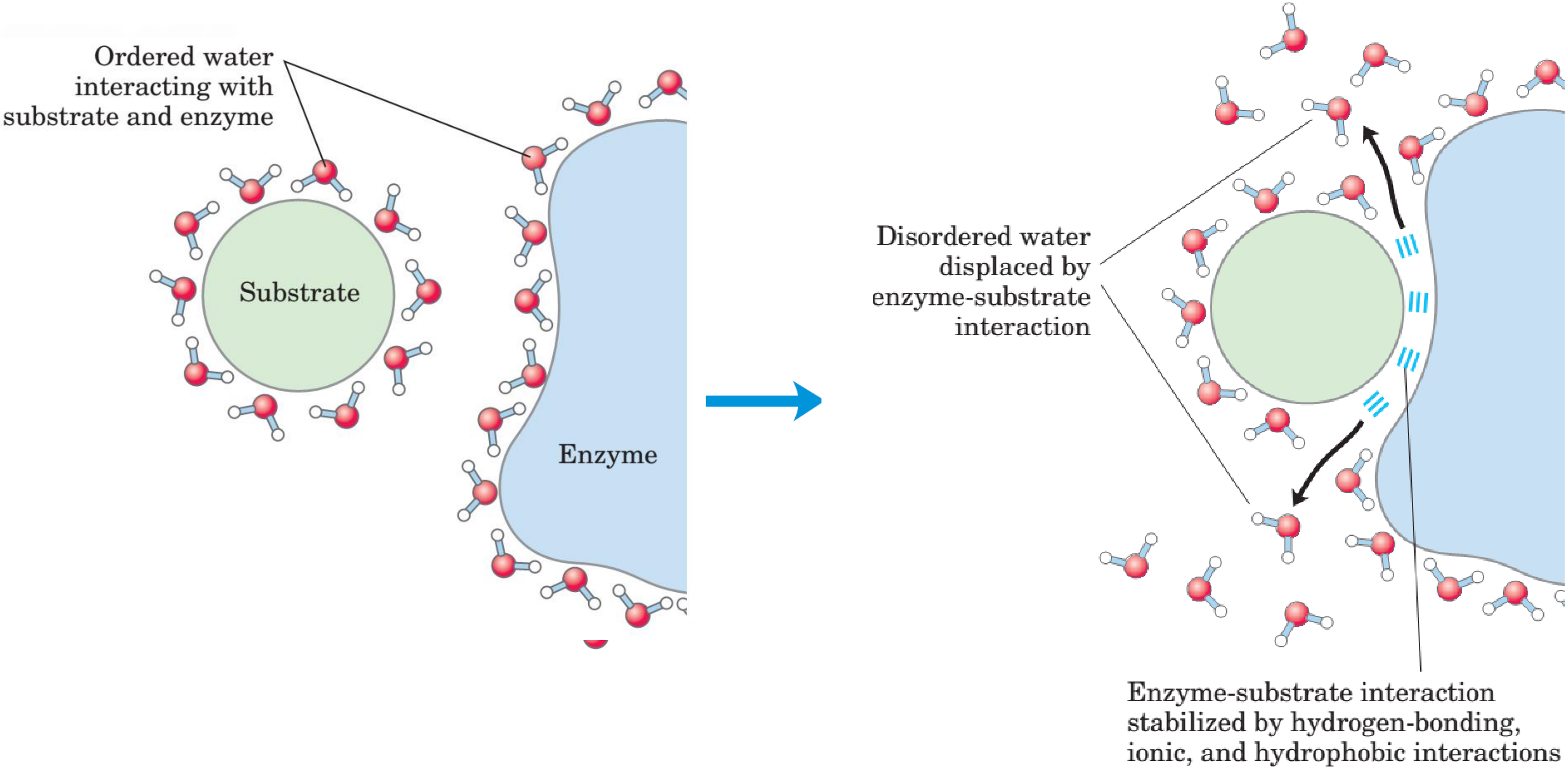
Induced dipole–induced dipole interactions or London dispersion forces



Two nonpolar atoms or molecules (depicted as having an electron cloud that has a time-averaged spherical shape).

Momentary attractions and repulsions between nuclei and electrons in neighboring molecules lead to induced dipoles.

Correlation of the electron motions between the two atoms or molecules (which are now polar) leads to a lower energy and stabilizes the system.



Ordered water interacting with substrate and enzyme

Disordered water displaced by enzyme-substrate interaction

Enzyme-substrate interaction stabilized by hydrogen-bonding, ionic, and hydrophobic interactions

- Hydrogen bonds and ionic, hydrophobic, and van der Waals interactions are much weaker than covalent bonds.

- An input of about 350 kJ of energy is required to break a mole of C-C single bonds, and about 410 kJ to break a mole of C-H bonds, but as little as 4 kJ is sufficient to disrupt a mole of typical van der Waals interactions.

- Hydrophobic interactions are also much weaker than covalent bonds, although they are substantially strengthened by a highly polar solvent (a concentrated salt solution, for example).

- Ionic interactions and hydrogen bonds are variable in strength, depending on the polarity of the solvent and the alignment of the hydrogen-bonded atoms, but they are always significantly weaker than covalent bonds.

- Although these four types of interactions are individually weak relative to covalent bonds, the cumulative effect of many such interactions can be very significant.

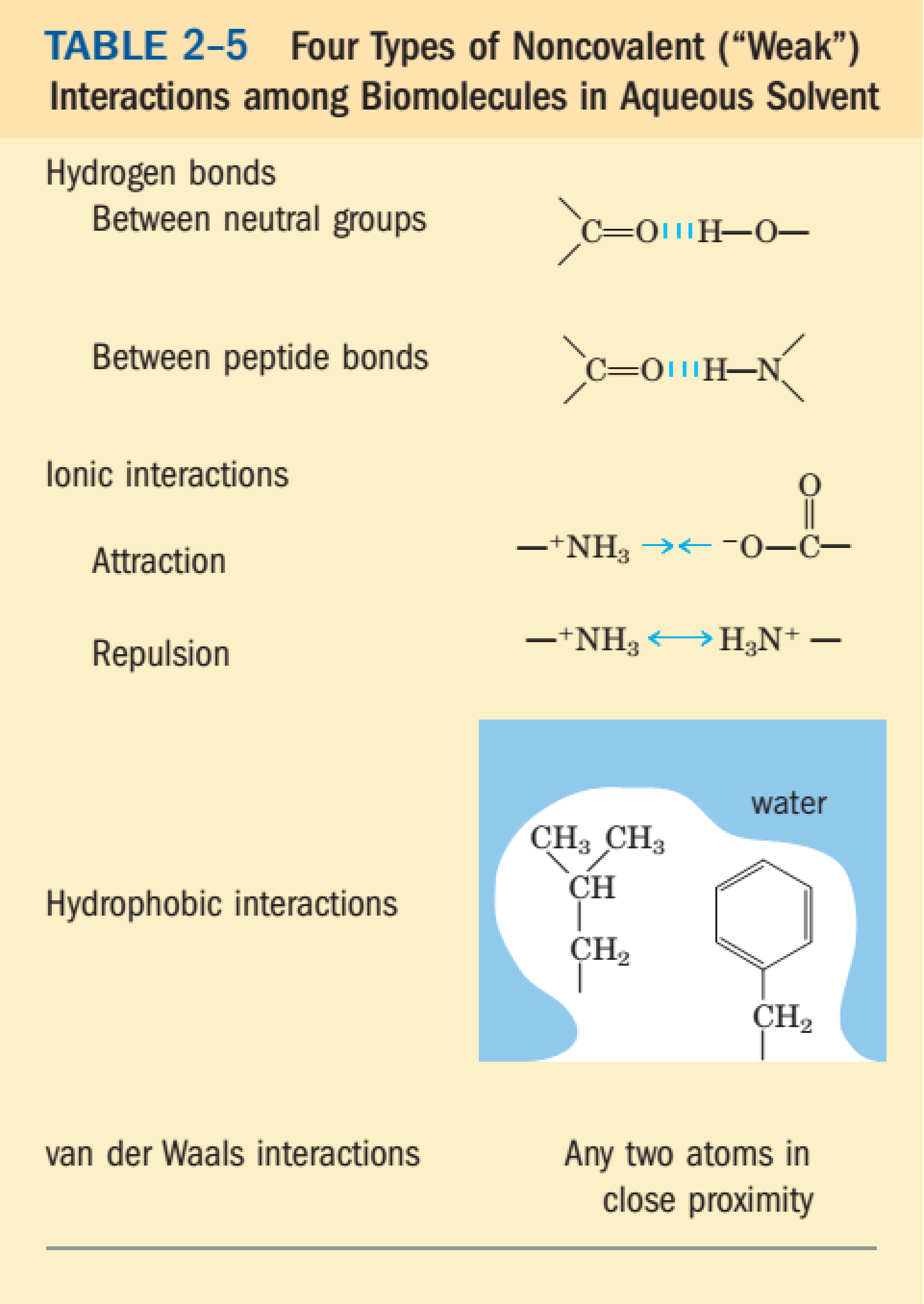
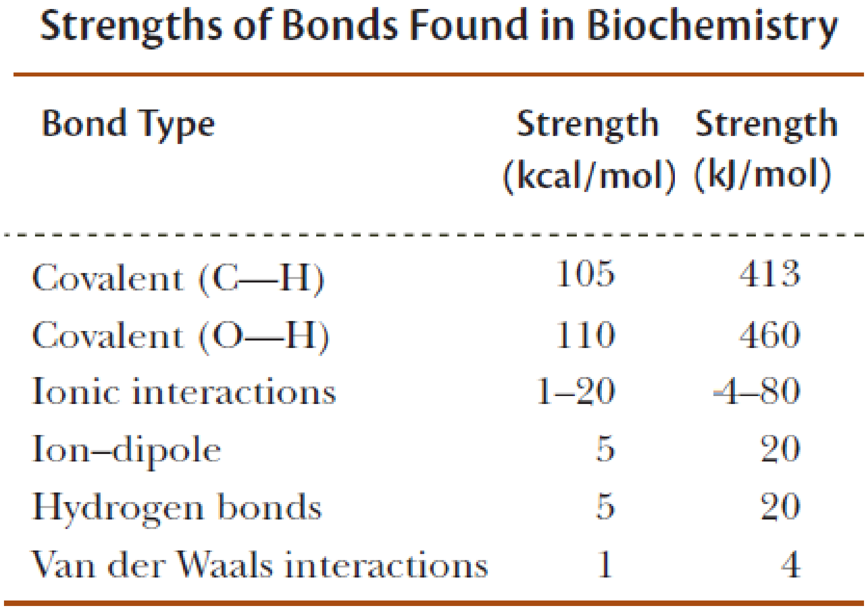


TABLE 2–5 Four Types of Noncovalent ("Weak") Interactions among Biomolecules in Aqueous Solvent

Hydrogen bonds	
Between neutral groups	$>C{=}O \cdots H{-}O{-}$
Between peptide bonds	$>C{=}O \cdots H{-}N<$
Ionic interactions	
Attraction	$-{}^{+}NH_3 \rightarrow\leftarrow {}^{-}O{-}\overset{\overset{O}{\|}}{C}{-}$
Repulsion	$-{}^{+}NH_3 \longleftrightarrow H_3N^{+}-$
Hydrophobic interactions	
van der Waals interactions	Any two atoms in close proximity

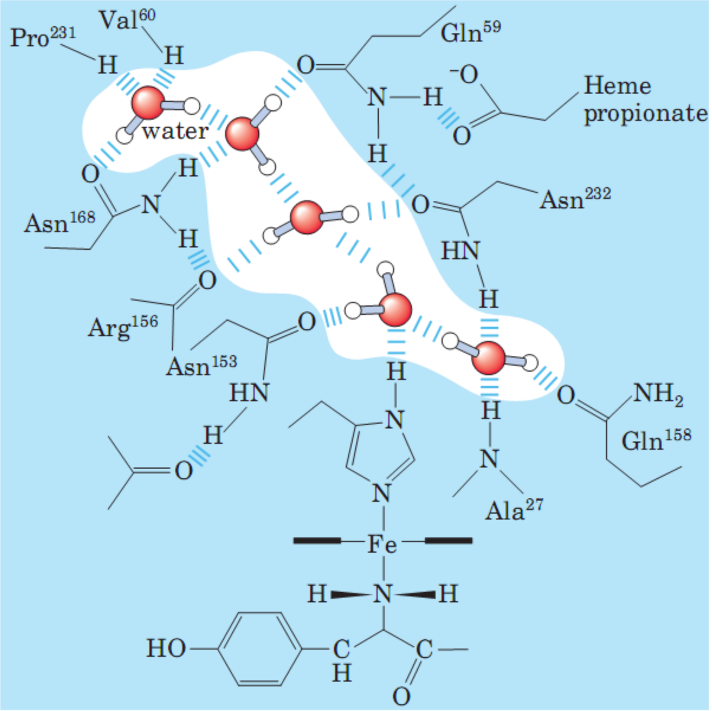


Strengths of Bonds Found in Biochemistry

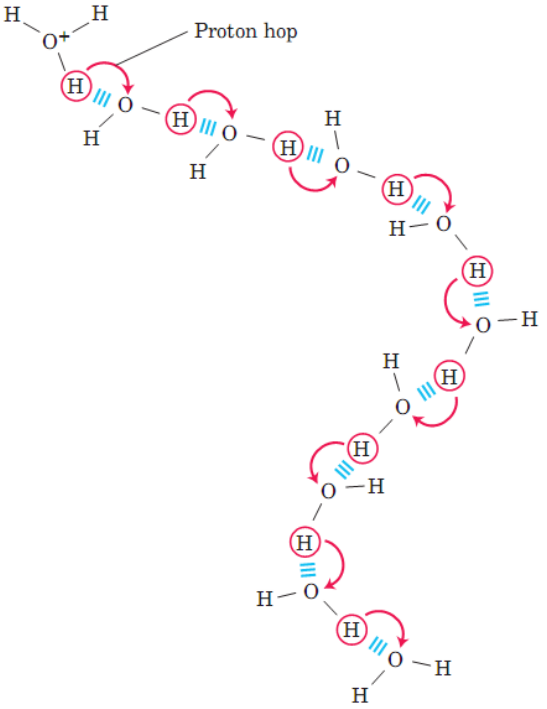
Bond Type	Strength (kcal/mol)	Strength (kJ/mol)
Covalent (C—H)	105	413
Covalent (O—H)	110	460
Ionic interactions	1–20	4–80
Ion–dipole	5	20
Hydrogen bonds	5	20
Van der Waals interactions	1	4

- When the structure of a protein is determined, water molecules are often found to be bound so tightly as to be part of the crystal structure. For many proteins, tightly bound water molecules are essential to their function.

- In a reaction central to the process of photosynthesis, for example, cytochrome f, has a chain of five bound water molecules that may provide a path for protons to move through the membrane by a process known as “proton hopping”.



Hydronium ion gives up a proton



Water accepts proton and becomes a hydronium ion

- Solutes of all kinds alter certain physical properties of the solvent, water: its vapor pressure, boiling point, melting point (freezing point), and osmotic pressure.

- These are called **colligative** (“tied together”) properties, because the effect of solutes on all four properties has the same basis: the concentration of water is lower in solutions than in pure water.

- The effect of solute concentration on the colligative properties of water is independent of the chemical properties of the solute; it depends only on the number of solute particles (molecules, ions) in a given amount of water.

- When two different aqueous solutions are separated by a semipermeable membrane, water molecules diffusing from the region of higher water concentration to that of lower water concentration produce osmotic pressure.

- This pressure, measured as the force necessary to resist water movement is approximated by the van’t Hoff equation: $\Pi = icRT$

$$\Pi = icRT$$

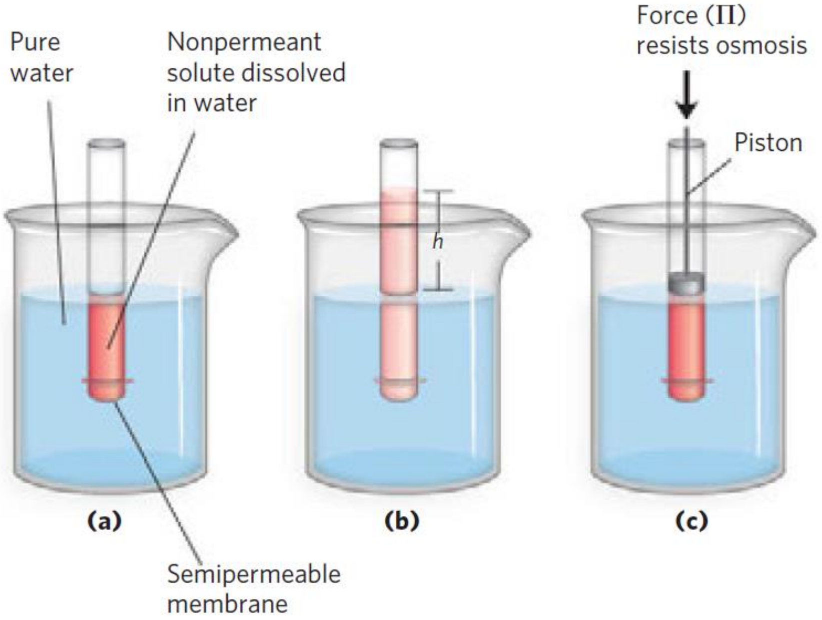
- in which R is the gas constant and T is the absolute temperature.

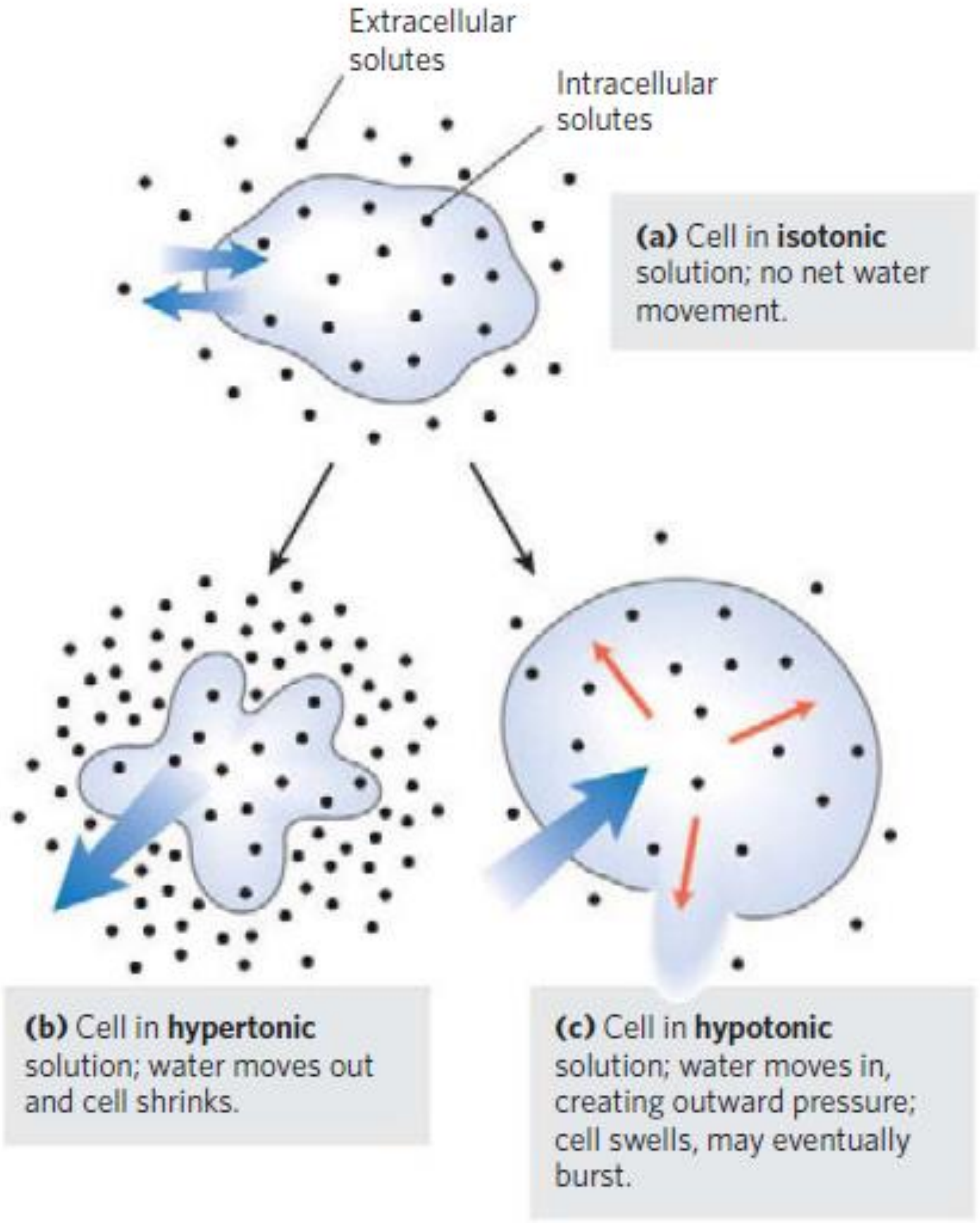
- The term ic is the osmolarity of the solution, the product of the solute’s molar concentration c and the van’t Hoff factor i, which is a measure of the extent to which the solute dissociates into two or more ionic species.

- In dilute NaCl solutions, the solute completely dissociates into Na^+ and Cl^-, doubling the number of solute particles, and thus i = 2. For nonionizing solutes, i is always 1.

- **Osmosis**, water movement across a semipermeable membrane driven by differences in osmotic pressure, is an important factor in the life of most cells.

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





(a) Cell in **isotonic** solution; no net water movement.

(b) Cell in **hypertonic** solution; water moves out and cell shrinks.

(c) Cell in **hypotonic** solution; water moves in, creating outward pressure; cell swells, may eventually burst.

- Solutions of equal osmolarity are said to be **isotonic**.

- Surrounded by an isotonic solution, a cell neither gains nor loses water.

- In a **hypertonic** solution, one with higher osmolarity than the cytosol, the cell shrinks as water flows out.

- In a **hypotonic** solution, with lower osmolarity than the cytosol, the cell swells as water enters.

- In their natural environments, cells generally contain higher concentrations of biomolecules and ions than their surroundings, so osmotic pressure tends to drive water into cells. If not somehow counterbalanced, this inward movement of water would distend the plasma membrane and eventually cause bursting of the cell (osmotic lysis).

- Several mechanisms have evolved to prevent this catastrophe.
- In bacteria and plants, the plasma membrane is surrounded by a nonexpandable cell wall of sufficient rigidity and strength to resist osmotic pressure and prevent osmotic lysis.
- Certain freshwater protists that live in a highly hypotonic medium have an organelle (contractile vacuole) that pumps water out of the cell.
- In multicellular animals, blood plasma and interstitial fluid (the extracellular fluid of tissues) are maintained at an osmolarity close to that of the cytosol.
- The high concentration of albumin and other proteins in blood plasma contributes to its osmolarity.

- Cells also actively pump out ions such as Na^+ into the interstitial fluid to stay in osmotic balance with their surroundings.

- Because the effect of solutes on osmolarity depends on the number of dissolved particles, not their mass, macromolecules (proteins, nucleic acids, polysaccharides) have far less effect on the osmolarity of a solution than would an equal mass of their monomeric components.

- Plants use osmotic pressure to achieve mechanical rigidity. The very high solute concentration in the plant cell vacuole draws water into the cell.

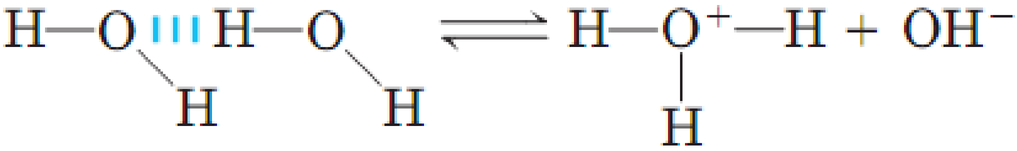
- The resulting osmotic pressure against the cell wall (turgor pressure) stiffens the cell, the tissue, and the plant body.

- Although many of the solvent properties of water can be explained in terms of the uncharged H_2O molecule, the small degree of ionization of water to hydrogen ions and hydroxide ions must also be taken into account.

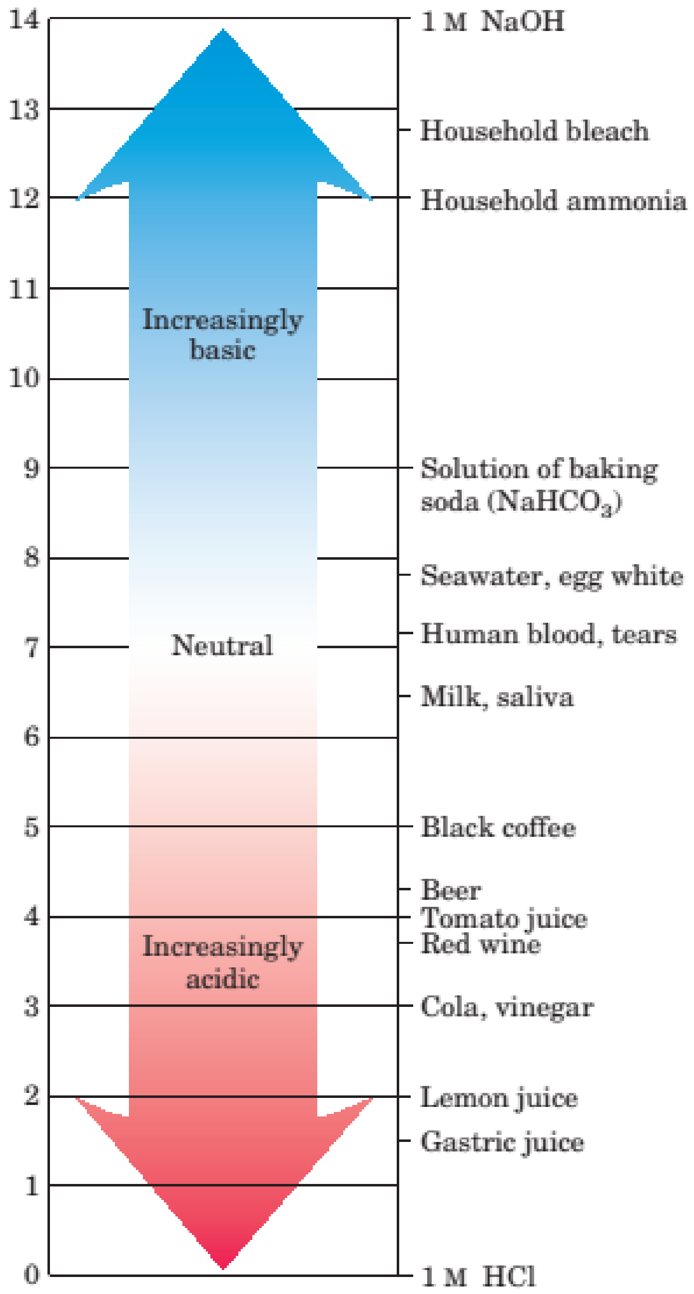
- The total hydrogen ion concentration from all sources is experimentally measurable and is expressed as the pH of the solution.

- Although we commonly show the dissociation product of water as H^+, free protons do not exist in solution; hydrogen ions formed in water are immediately hydrated to **hydronium ions** (H_3O^+).

$$H_2O \rightleftharpoons H^+ + OH^-$$



- pH is a function of H^+ ion concentration in the solution and can be defined with the equation of $pH = -\log[H^+]$.



- The pH of an aqueous solution can be approximately measured using various indicator dyes, including litmus, phenolphthalein, and phenol red, which undergo color changes as a proton dissociates from the dye molecule.

- Accurate determinations of pH in the chemical or clinical laboratory are made with a glass electrode that is selectively sensitive to H^+ concentration but insensitive to Na^+, K^+, and other cations

- Hydrochloric, sulfuric, and nitric acids, commonly called strong acids, are completely ionized in dilute aqueous solutions; the strong bases NaOH and KOH are also completely ionized.

- Of more interest to biochemists is the behavior of weak acids and bases—those not completely ionized when dissolved in water.

- Acids may be defined as proton donors and bases as proton acceptors.

$$\underset{\text{Acid}}{HA} \rightleftharpoons \underset{\text{Conjugate base}}{H^+ + A^-}$$

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

- Acetic acid (CH_3COOH), a proton donor, and the acetate anion (CH_3COO^+), the corresponding proton acceptor, constitute a conjugate acid-base pair, related by the reversible reaction

- Each acid has a characteristic tendency to lose its proton in an aqueous solution.

- The stronger the acid, the greater its tendency to lose its proton.

- The tendency of any acid (HA) to lose a proton and form its conjugate base (A^-) is defined by the equilibrium constant (K_{eq})

$$K_{\mathrm{eq}} = \frac{[\mathrm{H}^+][\mathrm{A}^-]}{[\mathrm{HA}]} = K_{\mathrm{a}}$$

- Equilibrium constants for ionization reactions are usually called ionization or acid dissociation constants, often designated K_a.

- The stronger the tendency to dissociate a proton, the stronger is the acid and the lower its **pK_a**

$$\underset{\text{Acid}}{\mathrm{HA(aq)}} + \underset{\text{Base}}{\mathrm{H_2O}(\ell)} \rightleftharpoons \underset{\text{Conjugate acid to } \mathrm{H_2O}}{\mathrm{H_3O^+(aq)}} + \underset{\text{Conjugate base to HA}}{\mathrm{A^-(aq)}}$$

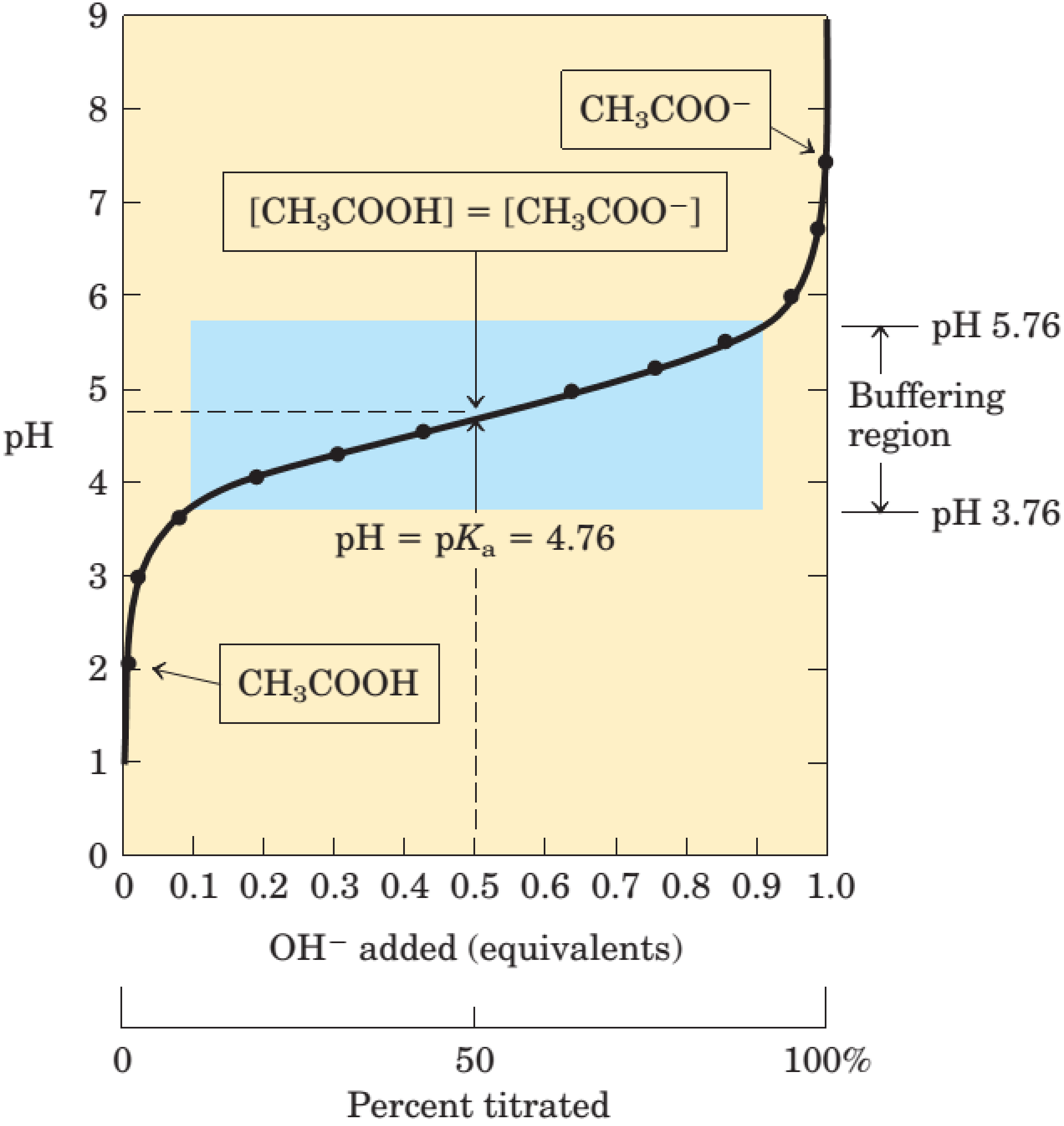
- Almost every biological process is pH dependent; a small change in pH produces a large change in the rate of the process.

- The enzymes that catalyze cellular reactions, and many of the molecules on which they act, contain ionizable groups with characteristic pKa values.

- Cells and organisms maintain a specific and constant cytosolic pH, keeping biomolecules in their optimal ionic state, usually near pH 7. In multicellular organisms, the pH of extracellular fluids is also tightly regulated.

- **Buffers** are aqueous systems that tend to resist changes in pH when small amounts of acid (H^+) or base (OH^-) are added.

- The titration curve of acetic acid has a relatively flat zone extending about 1 pH unit on either side of its midpoint pH of 4.76. In this zone, an amount of H^+ or OH^- added to the system has much less effect on pH than the same amount added outside the buffer range


pH
0
1
2
3
4
5
6
7
8
9
CH_3COO^-
$[CH_3COOH] = [CH_3COO^-]$
pH 5.76
Buffering region
pH 3.76
$pH = pK_a = 4.76$
CH_3COOH
0
0.1
0.2
0.3
0.4
0.5
0.6
0.7
0.8
0.9
1.0
OH^- added (equivalents)
0
50
100%
Percent titrated

- The shape of the titration curve of any weak acid is described by the **Henderson-Hasselbalch** equation, which is important for understanding buffer action and acid-base balance in the blood and tissues of vertebrates.

$$pH = pK_a - \log \frac{[HA]}{[A^-]} \qquad pH = pK_a + \log \frac{[A^-]}{[HA]}$$

- The cytoplasm of most cells contains high concentrations of proteins, which contain many amino acids with functional groups that are weak acids or weak bases, therefore buffer effectively near neutral pH.

- Nucleotides such as ATP, as well as many low molecular weight metabolites, contain ionizable groups that can contribute buffering power to the cytoplasm.

- Some highly specialized organelles and extracellular compartments have high concentrations of compounds that contribute buffering capacity: organic acids buffer the vacuoles of plant cells; ammonia buffers urine.

- Two especially important biological buffers are the phosphate and bicarbonate systems.

- The phosphate buffer system is maximally effective at a pH close to its pKa of 6.86 and thus tends to resist pH changes in the range between about 5.9 and 7.9.

$$H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{2-}$$

- Blood plasma is buffered in part by the bicarbonate system, consisting of carbonic acid (H_2CO_3) as proton donor and bicarbonate (HCO_3^-) as proton acceptor:

$$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$$

$$CO_2(d) + H_2O \rightleftharpoons H_2CO_3$$

$$CO_2(g) \rightleftharpoons CO_2(d)$$

- Blood plasma is buffered in part by the bicarbonate system, consisting of carbonic acid (H_2CO_3) as proton donor and bicarbonate (HCO_3^-) as proton acceptor:

- Human blood plasma normally has a pH close to 7.4. Should the pH-regulating mechanisms fail or be overwhelmed, as may happen in severe uncontrolled diabetes when an overproduction of metabolic acids causes **acidosis**, the pH of the blood can fall to 6.8 or below, leading to irreparable cell damage and death.

- Although many aspects of cell structure and function are influenced by pH, it is the catalytic activity of enzymes that is especially sensitive.

- Enzymes typically show maximal catalytic activity at a characteristic pH, called the pH optimum.

- Biological control of the pH of cells and body fluids is therefore of central importance in all aspects of metabolism and cellular activities.

- Much of biochemistry is studied by carrying out enzymatic reactions in a test tube or *in vitro*.

- Such reactions are usually buffered to maintain a constant pH. Similarly, virtually all methods for enzyme isolation, and even for growth of cells in tissue culture, use buffered solutions.

- The following criteria are typical for selecting a buffer for a biochemical reaction:

  - 1. Suitable pKa for the buffer
  - 2. No interference with the reaction or with the assay
  - 3. Suitable ionic strength of the buffer
  - 4. No precipitation of reactants or products due to presence of the buffer
  - 5. Nonbiological nature of the buffer

- Water is not just the solvent in which the chemical reactions of living cells occur; it is very often a direct participant in those reactions.

- The formation of ATP from ADP and inorganic phosphate is an example of a condensation reaction in which the elements of water are eliminated.

- The reverse of this reaction— cleavage accompanied by the addition of the elements of water—is a hydrolysis reaction.

- The high specific heat of water is useful to cells and organisms because it allows water for keeping the temperature of an organism relatively constant.

- Some vertebrates exploit the high heat of vaporization of water by using (thus losing) excess body heat to evaporate sweat.

- The high degree of internal cohesion of liquid water, due to hydrogen bonding, is exploited by plants as a means of transporting dissolved nutrients from the roots to the leaves during the process of transpiration.

- Even the density of ice, lower than that of liquid water, has important biological consequences in the life cycles of aquatic organisms.

- Ponds freeze from the top down, and the layer of ice at the top insulates the water below from frigid air, preventing the pond (and the organisms in it) from freezing solid.

- Most fundamental to all living organisms is the fact that many physical and biological properties of cell macromolecules, particularly the proteins and nucleic acids, derive from their interactions with water molecules of the surrounding medium.

- The influence of water on the course of biological evolution has been profound and determinative.

- If life forms have evolved elsewhere in the universe, they are unlikely to resemble those of Earth unless liquid water is plentiful in their planet of origin.