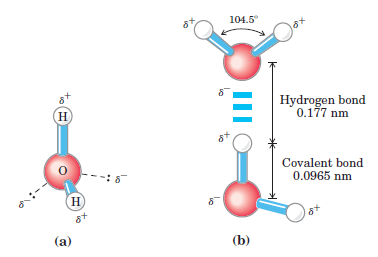
**Chapter 2 Water**

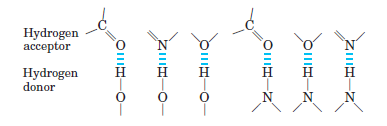
* Water is important for living systems.
* Properties of water affect the structure and function of all other cellular constituents.
* Water is the most abundant substance in living systems
* making up 70% or more of the weight of most organisms.
  1. **Weak Interactions in Aqueous Systems**

**Hydrogen Bonding Gives Water Its Unusual Properties**

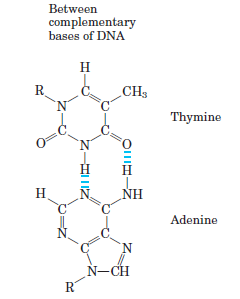
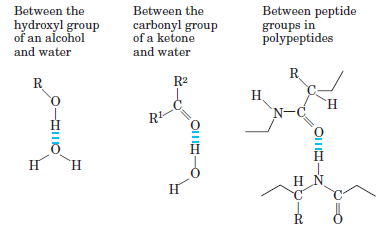
* Water has a higher
* melting point (0oC)
* boiling point (100oC)
* heat of vaporization (2,260 J/g) than most other common solvents.
* Structure of the water shows dipolar nature.
* the two H atoms have localized partial positive charges (+)
* the O atom has a partial negative charge (-) **(Fig. 2-1a)**.



* There is an electrostatic attraction between the O atom of one water molecule and the H of another, called a **hydrogen bond (Fig. 2-1b)**.
* two H2O molecules are joined by a hydrogen bond.
* H bonds are longer and weaker than covalent O-H bonds.
* Hydrogen bonds between water molecules provide the cohesive forces that make water a liquid at room temperature and a crystalline solid (ice) with a highly ordered arrangement of molecules at cold temperatures.
* Hydrogen bonds are formed between an electronegative atom (the hydrogen acceptor, usually oxygen or nitrogen) and a hydrogen atom covalently bonded to another electronegative atom (the hydrogen donor) in the same or another molecule **(Fig. 2-3).**



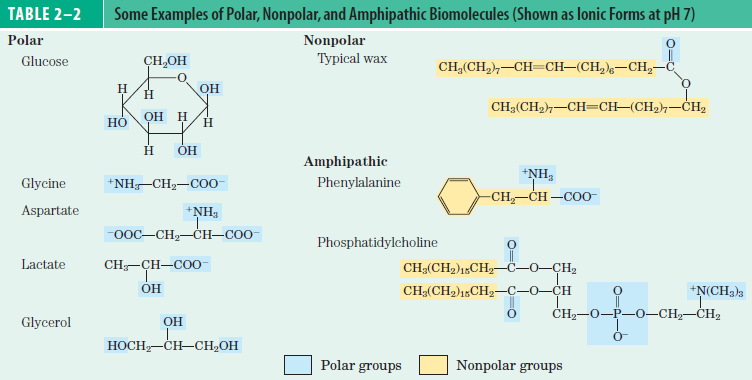
* Polar biomolecules dissolve readily in water forming hydrogen bonds
* because they can replace water-water interactions with more energetically favorable water-solute interactions.
* Some biologically important hydrogen bonds (**Fig. 2-4)**.



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**Water Interacts Electrostatically with Charged Solutes**

* Water is a polar solvent.
* It dissolves most biomolecules, which are generally charged or polar **hydrophilic** compounds **(Table 2-2)**.

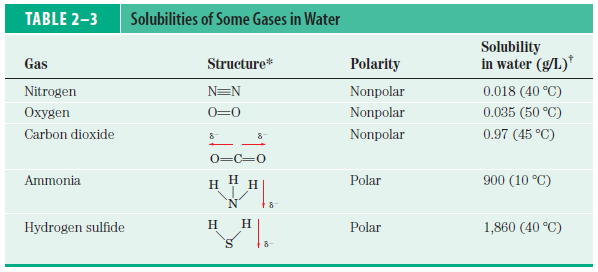


* Water dissolves salts such as NaCl.

**Nonpolar Gases Are Poorly Soluble in Water**

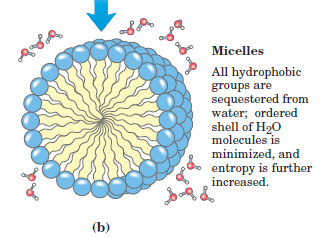
* The molecules of the biological important gases CO2, O2, and N2 are nonpolar.
* Two other gases, NH3 and H2S, also have biological roles in some organisms

- these gases are polar and dissolve in water **(Table 2-3)**.



**Nonpolar Compounds Force Energetically Unfavorable Changes in the Structure of Water**

* **Amphipathic** compounds contain regions
* polar (or charged) and nonpolar **(Table 2-2)**.
* When an amphipathic compound is mixed with water
* polar (hydrophilic) region interacts with the water and tends to dissolve.
* nonpolar **hydrophobic** region tends to avoid contact with water.
* the nonpolar regions of the molecules cluster together to present the smallest hydrophobic area.
* These stable structures of amphipathic compounds in water, called **micelles**,
* may contain hundreds or thousands of molecules.
* Structures of amphipathic compounds
* are stabilized by **hydrophobic interactions** among the nonpolar regions **(Fig. 2-7 b)**.



**Van der Waals Interactions Are Weak Interatomic Attractions**

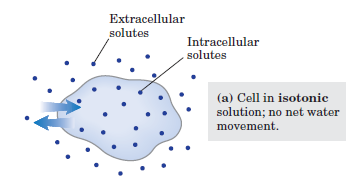
* When two uncharged atoms are brought very close together
* their surrounding electron clouds influence each other.
* Random variations in the positions of the electrons around one nucleus
* may create a transient electric dipole, which induces a transient and opposite electric dipole in the nearby atom.
* The two dipoles weakly attract each other, bringing the two nuclei closer.
* These weak interactions are called **van der Waals interactions** (also known as **London forces**).

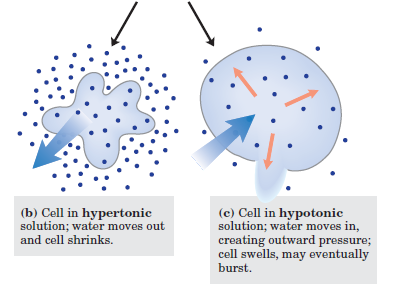
**Weak Interactions Are Crucial to Macromolecular Structure and Function**

* The noncovalent interactions
* hydrogen bonds
* ionic interactions
* hydrophobic interactions
* van der Waals interactions are individually much weaker than covalent bonds.
* The noncovalent interactions are continually formed and broken.
* The cumulative effect of many such interactions among biomolecules in aqueous solutions can be very significant.

**Solutes Affect the Colligative Properties of Aqueous Solutions**

* Solutes of all kinds alter physical properties of the solvent (water).
* These properties are vapor pressure (lowered), boiling point (increased), melting point or freezing point (decreased) and osmotic pressure.
* They are called **colligative properties**
* depends only on the number of solute particles (molecules or ions).
* When two different aqueous solutions are separated by a semipermeable membrane (allows the passage of water but not solute molecules)
* water molecules diffusing from the region of higher water concentration to the region of lower water concentration produce **osmotic pressure**.
* Plasma membranes are more permeable to water than to most other small molecules, ions, and macromolecules.







* Solutions of equal osmolarity are said to be **isotonic** (**Fig. 2-12a)**.
* Cell in a **hypertonic** solution (higher osmolarity than the cytosol) (**Fig. 2-12b)**.
* Cell in **hypotonic** solution (lower osmolarity than the cytosol) (**Fig. 2-12c)**.
* **Osmosis** is an important factor in the life of most cells.
* In natural environments of cells,
* cells generally contain higher concentrations of biomolecules and ions than their surroundings.
  1. **Ionization of Water, Weak Acids and Weak Bases**

**Pure Water Is Slightly Ionized**

* Water molecules have a slight reversible ionization to yield a hydrogen ion (proton) and a hydroxide ion

H2O H+ + OH-

* Free protons do not exist in solution
* H ions in water are immediately hydrated to **hydronium ions** (H3O+)
* Reversible ionization is crucial to the role of water in cellular function.

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

* The **equilibriım constant** for the reversible ionization of water

**Keq** = [H+] [OH-] / [H2O]

[H2O] = (1000 g/L) / (18 g/mol) = 55.5 M

Keq = [H+] [OH-] / 55.5 M

(55.5M) (Keq) = [H+] [OH-] = Kw = **ion product of water**

Keq is determined by electrical-conductivity measurements of pure water

Keq = 1.8 x 10-16 M

(55.5 M) (1.8 x 10-16 M) = 1.0 x 10-14 M2 = [H+] [OH-] = Kw

* When [H+] = [OH-], the solution is said to be at **neutral pH**

[H+] = [OH-] = 10-7 M

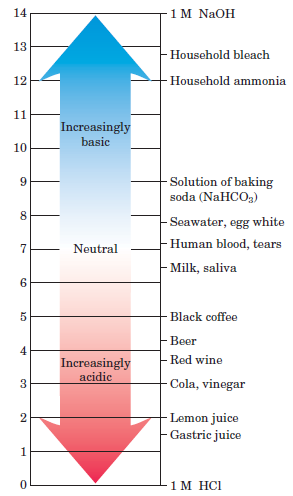
**The pH Scale Designates the H+ and OH- Concentrations**

* The term **pH** is defined by the expression

pH = - log [H+]

the symbol p denotes “negative logarithm of “

* The pH of some common aqueous fluids is given in (**Fig. 2-14)**.





* a cola drink (pH 3.0) has approximately 10,000 times higher H+ concentration than blood (pH 7.4)
* Measurement of pH is one of the most important and frequently used procedures in biochemistry.
* Measurements of the pH of blood and urine are commonly used in medical diagnoses.
* The pH affects the structure and activity of biological macromolecules.

**Weak Acids and Bases Have Characteristic Acid Dissociation Constants**

* Weak acids and bases are not completely ionized in water.
* They play important roles in metabolism and its regulation.
* The tendency of any **weak acid** (HA) to lose a proton and form its **conjugate base** (A-)
* is defined by the equilibrium constant (Keq) for the reversible reaction

HA H+ + A-

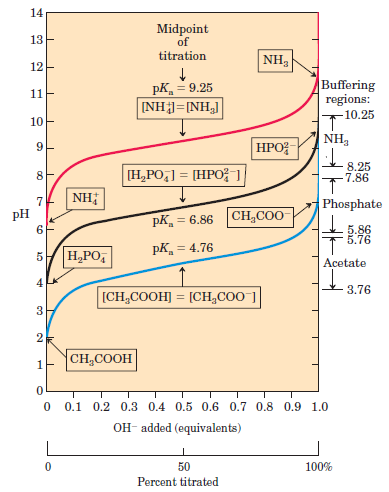
Keq = [H+] [A-] / [HA] = Ka = **acid** **dissociation constant**

pKa = - log Ka

* Stronger acid has lower pKa.

**Titration Curves Reveal the pKa of Weak Acids**

* Comparison of the **titration curves** of three weak acids**(Fig. 2-17)**.
* acetic acid (pKa = 4.76), dihydrogen phosphate (pKa = 6.86) and ammonium ion (pKa = 9.25)
* The titration curve of a weak acid shows
* a weak acid and its anion (—a conjugate acid-base pair)
* they can act as a buffer.



**2.3 Buffering against pH Changes in Biological Systems**

* Almost every biological process is pH dependent
* a small change in pH produces a large change in the rate of the process.
* Cells and organisms maintain a specific and constant cytosolic pH, usually near pH 7.
* it keeps biomolecules in their optimal ionic state.
* Constancy of pH is achieved by biological buffers.

**Buffers Are Mixtures of Weak Acids and Their Conjugate Bases**

* Buffers are aqueous systems
* tend to resist changes in pH when a small amounts of acid (H+) or base (OH-) are added.
* This change is very small change in pH compared with no buffering systems (pure water, NaCl solution).

**The Henderson-Hasselbalch Equation Relates pH, p*K*a, and Buffer Concentration**

* For the ionization of a weak acid HA, the Henderson-Hasselbalch equation can be derived

Ka = [H+] [A-] / [HA]

solve for [H+] [H+] = Ka [HA] / [A-]

the negative logarithm of both sides - log [H+] = - log Ka - log[HA] / [A-]

pH = pKa - log[HA] / [A-]

invert - log[HA] / [A-] pH = pKa + log[A-] / [HA]

generally pH = pKa + log[proton acceptor] / [proton donor]

when [HA] = [A-] pH = pKa + log 1.0 = pKa + 0 = pKa

buffer has maximum capacity

**Weak Acids or Bases Buffer Cells and Tissues against pH Changes**

* The cytoplasm of most cells contains high concentrations of proteins.
* proteins contain many amino acids with functional groups that are weak acids or weak bases.
* Two important biological buffers are the phosphate and bicarbonate systems.

H2PO4- H+ + HPO42-

* phosphate buffer works between 5.9 and 7.9.
* phosphate buffer is effective in extracellular fluids and most cytoplasmic compartments.
* bicarbonate buffer is effective in blood and lung

H2CO3 H+ + HCO3-

**Water as a Reactant**

* Water is not only the solvent
* it is a direct participant in chemical reactions of living cells.
* The addition of the elements of water is **a hydrolysis reaction**.
* for example, ATP + H2O ADP + Pi
* Hydrolysis reactions are also responsible for the enzymatic depolymerization of proteins, carbohydrates, and nucleic acids.
* H2O and CO2 are the end products of the oxidation of fuels such as glucose

C6H12O6 + 6O2 6 CO2 + 6 H2O

* **metabolic water** is formed by oxidation of foods and stored fats
  + it is enough to allow some animals in very dry habitats (camels) to survive for extended periods without drinking water.
* Green plants and algae use the energy of sunlight to split water in the process of photosynthesis
* H2O gives its electrons to electron-accepting species.