Water & pH

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BIOMEDICAL IMPORTANCE OF WATER

- Water (H_2O) is the main chemical component of living organisms.
- It is a reactant or product in various metabolic reactions.
- It has a slight tendency to dissociate into hydroxide ions (OH^-) and protons (H^+).

BIOMEDICAL IMPORTANCE OF WATER

- The concentration of protons of aqueous solutions is reported using the logarithmic pH scale.
- Bicarbonate, phosphate and other buffers function to maintain the pH of blood between 7.35 and 7.45.
- Acidosis (blood pH <7.35) vs. alkalosis (blood pH >7.45).

WATER

- $_{8}O^{16}$: $1s^2\ 2s^2\ 2p_x^{\ 2}\ 2p_y^{\ 1}\ 2p_z^{\ 1}$

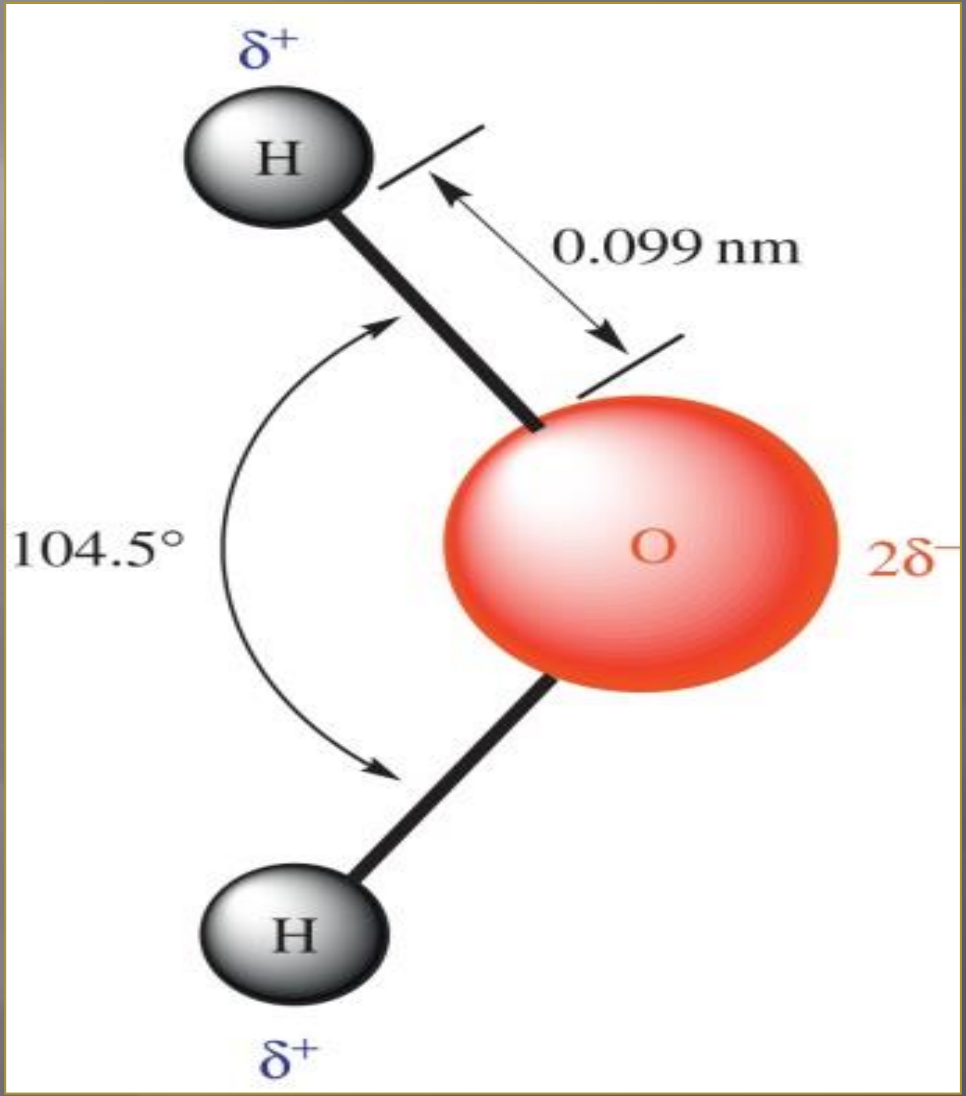
- $_{1}H^{1}$: $1s^1$

- $H \diagup \ddot{\underset{\cdot\cdot}{O}} \diagdown H$

Water Molecules Form Dipoles

- The water (H_2O) molecule has tetrahedral geometry.
- The strongly electronegative oxygen atom in a water molecule attracts electrons away from the hydrogen nuclei, leaving them with a partial positive charge, meantime its two unshared electron pairs constitute a region of local negative charge.
- A molecule with electrical charge distributed asymmetrically about its structure is referred to as a dipole.
- Its strong dipole and high dielectric constant facilitate water to dissolve large quantities of charged compounds like salts.

STRUCTURE OF WATER MOLECULE



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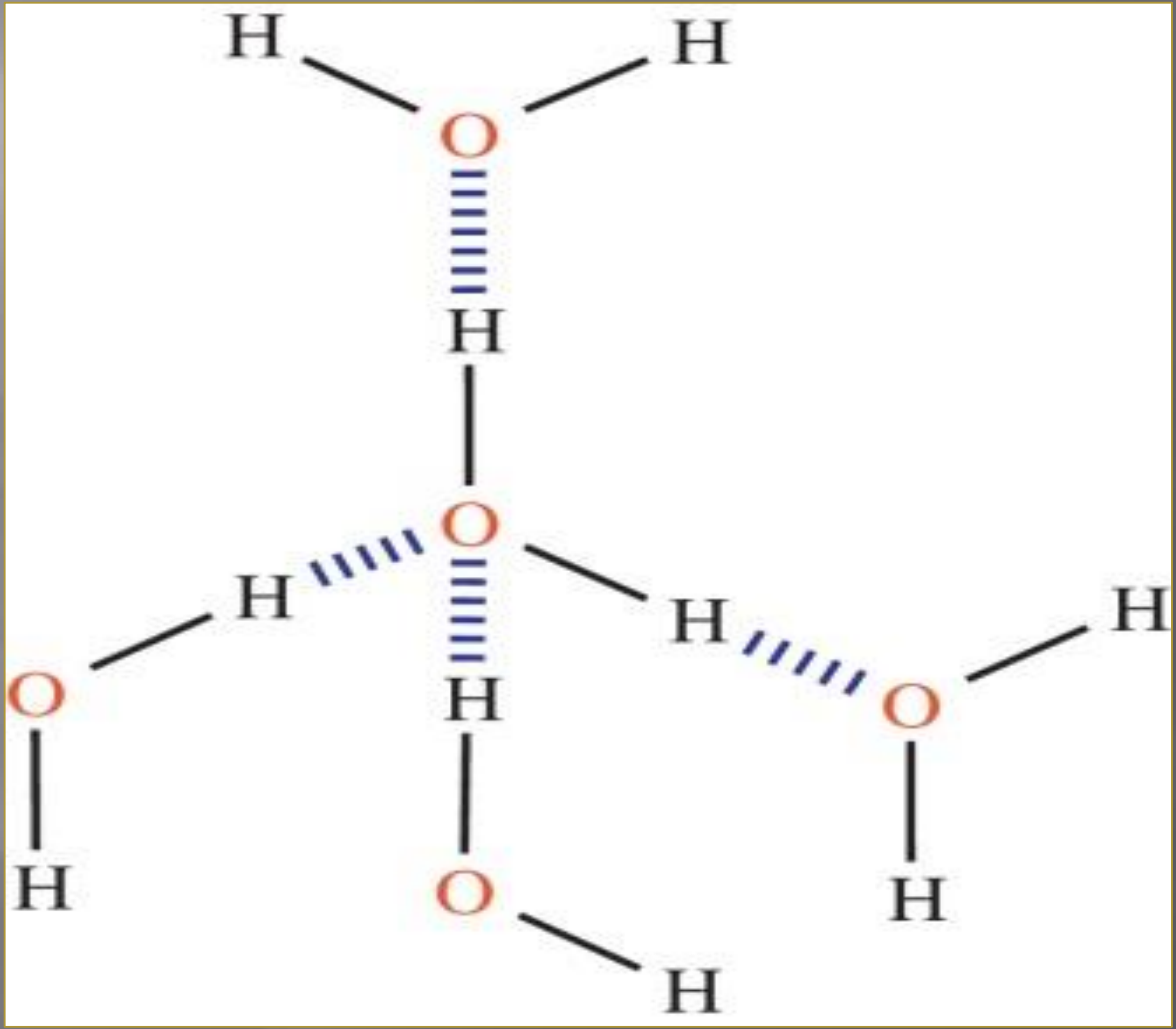
Water Molecules Form Hydrogen Bonds

- A partially unshielded hydrogen nucleus covalently bound to an electron-withdrawing oxygen or nitrogen atom can interact with an unshared electron pair on another oxygen or nitrogen atom to form a **hydrogen bond**.
- Hydrogen bonding greatly influences the physical properties of water and clarifies its relatively high viscosity, surface tension, and boiling point.
- Water molecules self-associate via hydrogen bonds.
- Hydrogen bonding enables water to dissolve many organic biomolecules that contain functional groups which can participate in hydrogen bonding.

HYDROGEN BONDS BETWEEN WATER MOLECULES



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Biologic Molecules are Stabilized by Covalent and Noncovalent Bonds

- The covalent bond is the strongest force that holds molecules together.
- Noncovalent forces, while of lesser magnitude, make significant contributions to the structure, stability, and functional competence of macromolecules in living cells.
  - Hydrogen bonds
  - Electrostatic (ionic) interactions
  - Hydrophobic interactions
  - van der Waals forces

WATER IS A GOOD NUCLEOPHILE

- Metabolic reactions often involve the attack by lone pairs of electrons residing on electron-rich molecules termed nucleophiles upon electron-poor atoms called electrophiles.
- Water, whose two lone pairs of sp^3 electrons bear a partial negative charge, is an excellent nucleophile.
- Nucleophilic attack by water typically results in the cleavage of the amide (peptide), glycoside, or ester bonds that hold biopolymers together (hydrolysis).
- When monomer units are joined together to form biopolymers, such as proteins or glycogen, water is a product.

Water Molecules Exhibit a Minor but Important Tendency to Dissociate

- The ability of water to ionize, while slight, is of central importance for life.
- Since water can act both as an acid and as a base, its ionization may be represented as an intermolecular proton transfer that forms a hydronium ion (H_3O^+; represented as H^+) and a hydroxide ion (OH^-):

$$H_2O + H_2O \rightleftarrows H_3O + OH^-$$

- Simply, water dissociates as:

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

Dissociation of Water

- For dissociation of water,

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

- K_w termed the ion product for water. The relationship between K_w and K is as shown below:

$$K = \frac{[H^+][OH^-]}{[H_2O]} = 1.8 \times 10^{-16}\,\text{mol/L}$$

$$\begin{aligned} K_w &= (K)[H_2O] = [H^+][OH^-] \\ &= (1.8 \times 10^{-16}\,\text{mol/L})(55.56\,\text{mol/L}) \\ &= 1.00 \times 10^{-14}\,(\text{mol/L})^2 \end{aligned}$$

Ion Product for Water (K_w)

- The ion product K_w is numerically equal to the product of the molar concentrations of H^+ and OH^-:

$$K_w = [H^+][OH^-]$$

- For pure water, at 25°C (room temperature) K_w equals to 10^{-14} $(mol/L)^2$, therefore $[H^+]$ in pure water is 10^{-7} mol/L.

pH IS THE NEGATIVE LOG OF THE HYDROGEN ION CONCENTRATION

$$pH = -\log[H^+]$$

For pure water at 25°C,

$$pH = -\log[H^+] = -\log 10^{-7} = -(-7) = 7.0$$

Therefore, pH of the pure water is 7.

ACIDS & BASES

- Low pH values correspond to high concentrations of H^+ and high pH values correspond to low concentrations of H^+.
- **Acids** are proton donors and **bases** are proton acceptors.
- Strong acids (e.g., HCl, H_2SO_4) completely dissociate into anions and protons even in strongly acidic solutions (low pH).
- Weak acids dissociate only partially in acidic solutions.
- Similarly, strong bases (e.g., KOH, NaOH), but not weak bases such as $Ca(OH)_2$, are also completely dissociated even at high pH.

Functional Groups That Are Weak Acids Have Remarkable Physiologic Significance

- Many biochemicals possess functional groups that are weak acids or bases.
- Knowledge of the dissociation of weak acids and bases thus is basic to understanding the influence of intracellular pH on structure and biologic activity.
- We term the protonated species (**HA** or $R\text{-}NH_3^+$) the **acid** and the unprotonated species (**A^-** or $R\text{-}NH_2$) its **conjugate base**.
- Similarly, we may refer to a **base** (A^- or $R\text{-}NH_2$) and its **conjugate acid** (HA or $R\text{-}NH_3^+$).

- The relative strengths of weak acids and bases are expressed in terms of their dissociation constants.
- Shown below is the expression for the dissociation constant (K_a) for a representative weak acid, R-COOH.

$$R{-}COOH \rightleftharpoons R{-}COO^- + H^+$$

$$K_a = \frac{[R{-}COO^-][H^+]}{[R{-}COOH]}$$

- Since the numeric values of K_a for weak acids are negative exponential numbers, we express K_a as pK_a, where

$$pK_a = -\log K_a$$

- **The stronger the acid, the lower is its pK_a value.**

- From the above equations that relate K_a to $[H^+]$ and to the concentrations of undissociated acid and its conjugate base, when

$$[R{-}COO^-] = [R{-}COOH]$$

- then,

$$K_a = [H^+]$$

- Thus, when the protonated and its conjugate base species are present at equal concentrations, the prevailing hydrogen ion concentration $[H^+]$ is numerically equal to the dissociation constant, Ka.

- Since −log K_a is defined as pK_a, and −log $[H^+]$ defines pH, the equation may be rewritten as

$$pK_a = pH$$

- **Finally, the pK_a of an acid group is the pH at which the protonated and unprotonated species are present at equal concentrations.**

The Henderson–Hasselbalch Equation Illustrates the Behavior of Weak Acids & Buffers

- A weak acid, HA, ionizes as follows:

$$HA \rightleftharpoons H^+ + A^-$$

- The equilibrium constant for this dissociation is

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

- Cross-multiplication gives

$$[H^+][A^-] = K_a[HA]$$

Derivation of The Henderson–Hasselbalch Equation

- Divide both sides by [A−]:

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

- Take the log of both sides:

$$\begin{aligned}\log[\mathrm{H^+}] &= \log\left(K_\mathrm{a}\frac{[\mathrm{HA}]}{[\mathrm{A^-}]}\right)\\ &= \log K_\mathrm{a} + \log\frac{[\mathrm{HA}]}{[\mathrm{A^-}]}\end{aligned}$$

- Multiply through by −1:

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

Derivation of The Henderson-Hasselbalch Equation

- Substitute pH and pK_a for $-\log [H^+]$ and $-\log K_a$, respectively; then

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

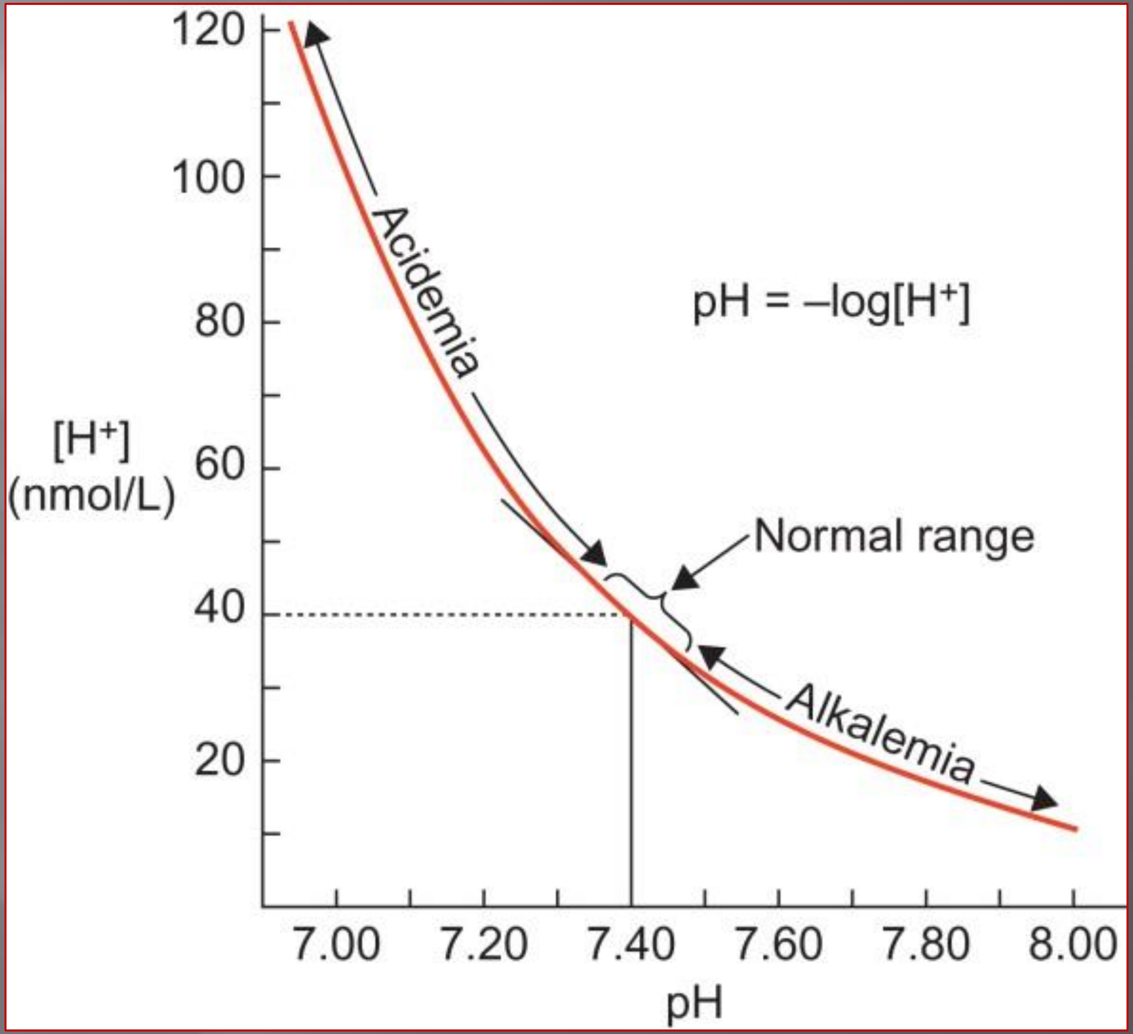
- Inversion of the last term removes the minus sign and gives the **Henderson-Hasselbalch equation**:

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

Solutions of Weak Acids & Their Salts Buffer Changes in pH

- Solutions of weak acids or bases and their conjugates exhibit **buffer**ing, the ability to resist a change in pH following addition of strong acid or base.
- In other words, **buffers** resist a change in pH when protons are produced or consumed.
- Biologic maintenance of a constant pH involves buffering by bicarbonate, orthophosphate, proteins, and hemoglobin (in red blood cells), which accept or release protons to resist a change in pH.
- A solution of a weak acid and its conjugate base buffers most effectively in the pH range $pK_a \pm 1.0$ pH unit.

Blood pH



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