BIOLOGICAL PHYSICS

CHAPTER 2 –WATER: THE SOLVENT FOR BIOCHEMICAL REACTIONS

Poopalasingam Sivakumar
Department of Physics, SIUC

Medium of Life: Water

- Water constitutes ~70–90% of most cells, depending on cell type and physiological state.
- Normal metabolic activity typically requires ≥65% water content.
- (unique) Polarity and hydrogen bonding make water an exceptional biological solvent, enabling:
  - dissolution of ions and polar molecules
  - efficient molecular transport and diffusion
- Water–biomolecule interactions shape the behavior of amino acids, proteins, nucleotides, and nucleic acids, influencing their structure, dynamics, and stability.
- Water molecules and their dissociation products (H^+ and OH^-) directly participate various biochemical processes, including hydrolysis, acid–base chemistry, and redox-coupled reactions essential for metabolism.

The Unique Properties of Water

- Water exhibits unexpected physical and chemical properties compared to molecules with similar size and structure.
- These unique properties are crucial for life on Earth.

We will compare water to HF, H2S, and NH3 to understand its exceptional characteristics.

Molecule	Boiling Point	Melting Point	Heat of Vaporization	Surface Tension
H2O	100 °C	0 °C	High	High
HF	19.5 °C	-83.6 °C	Low	Low
H2S	-60.2 °C	-85.5 °C	Low	Low
NH3	-33.3 °C	-77.7 °C	Moderate	Moderate

Water has significantly higher values of physical properties despite its low molecular weight.

The Unique Properties of Water

- **High Dielectric Constant**: Water's strong polarity allows it to dissolve many polar and ionic substances.
- **Maximum Density as a Liquid**: Water is densest at 4 °C, crucial for aquatic life and global climate patterns.
- **Negative Volume of Melting**: Water expands upon freezing, which has important implications for ice formation and the environment.
- Water also demonstrates an unusually high dielectric constant, with its maximum density occurring in the liquid state rather than the solid state.
- Additionally, water exhibits a negative volume of melting.

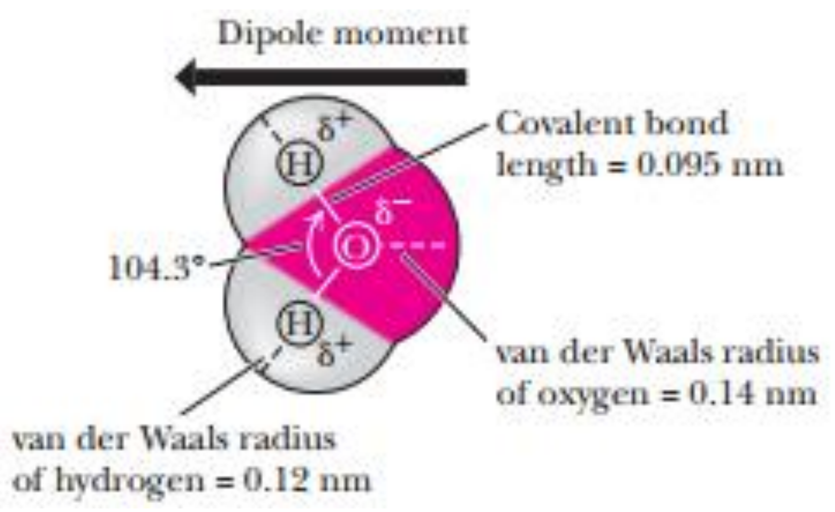
Role of Hydrogen Bonding

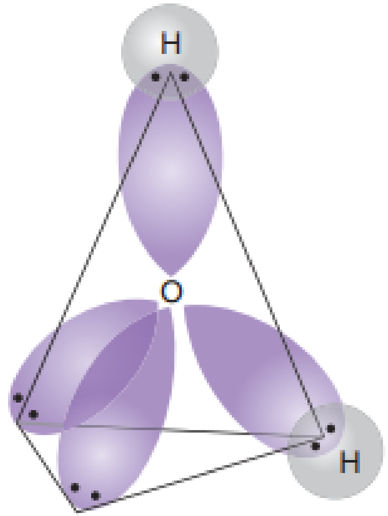
Molecules with similar size of water, CH_4, NH_3, and H_2S are gases at room temperature

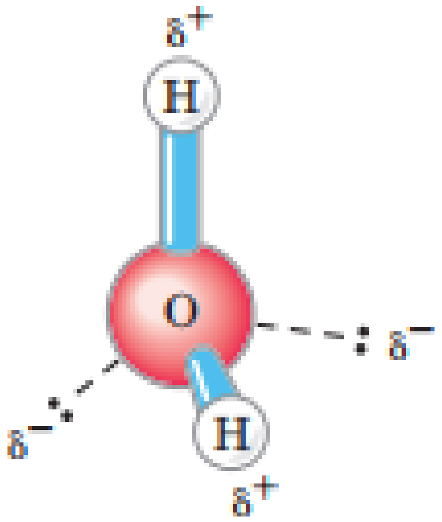
Molecule	**Molecular Weight**	**Boiling Point**	**Melting Point**	**Heat of Vaporization**	**Surface Tension**
H2O	18.01528 g/mol	100 °C	0 °C	High	High
CH4	16.04 g/mol	-161.5 °C	-182.5 °C	Low	Low
NH3	17.031 g/mol	-33.3 °C	-77.7 °C	Moderate	Moderate
H2S	34.08 g/mol	-60.2 °C	-85.5 °C	Low	Low

Unique Water Molecules

- Water is a polar molecule, meaning its electrical charge is unevenly distributed.
- Oxygen is more electronegative than hydrogen, pulling the shared electrons in the O–H covalent bonds closer to itself.
- This unequal sharing of electrons creates polar covalent bonds, where the oxygen atom carries a partial negative charge (δ^-) and each hydrogen atom bears a partial positive charge (δ^+).
- The arrangement of the H-O-H molecule is nonlinear, with a bond angle 104.5 degrees

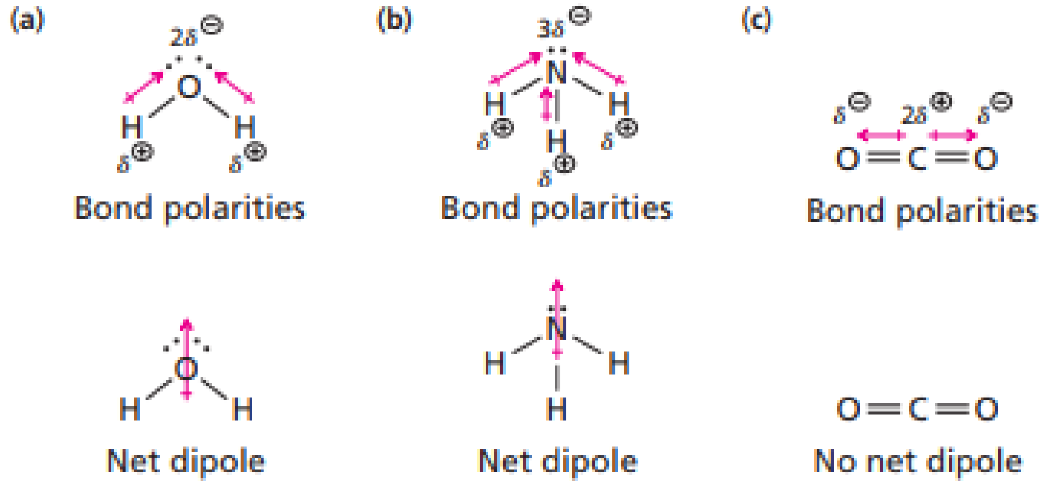






Molecular Dipoles: Water, Ammonia vs. Dioxide

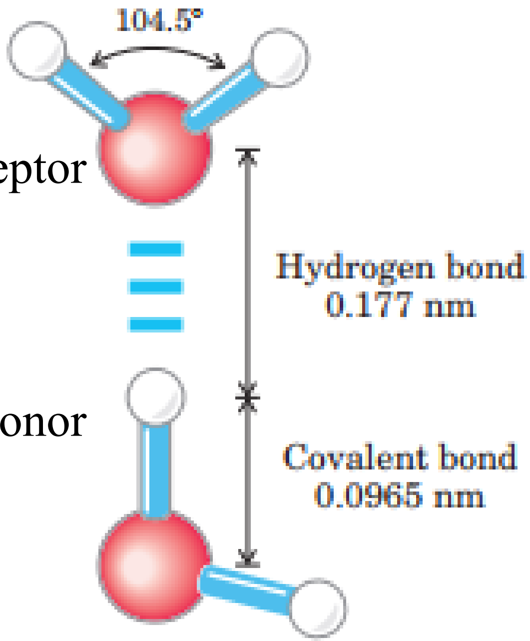
- What Creates a Molecular Dipole?
  - Polar covalent bonds + asymmetric molecular geometry
  - Net dipole depends on vector addition of bond dipoles



Molecule	**Chemical Formula**	**Molecular Geometry**	**Polar Bonds?**	**Net Dipole Moment**	**Polarity**
Water	H_2O	Bent (104.5°)	Yes (O–H)	≈ **1.85 D**	Polar
Ammonia	NH_3	Trigonal pyramidal (~107°)	Yes (N–H)	≈ **1.47 D**	Polar
Carbon Dioxide	CO_2	Linear (180°)	Yes (C=O)	**0 D**	Nonpolar

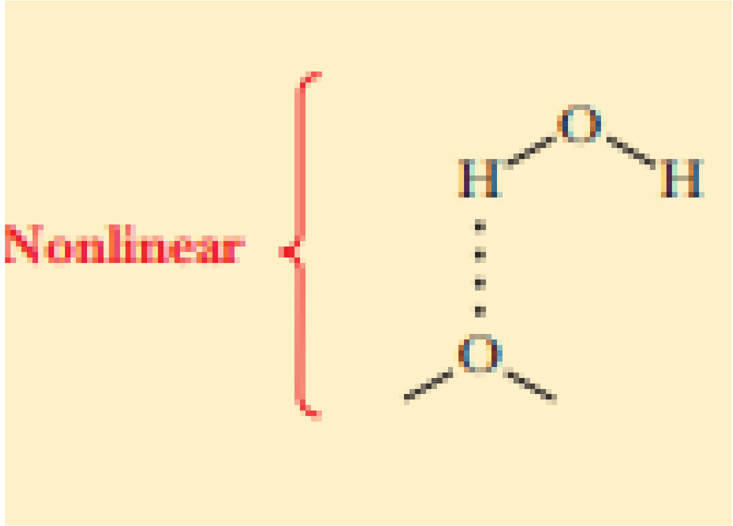
Hydrogen-bond

- Water molecules exhibit a tendency to orient themselves, aligning each partially positive hydrogen (H) with a partially negative oxygen (O), forming a **hydrogen bond**.
- Bent or nonlinear hydrogen bonds are weaker, due to reduced orbital overlap and electrostatic alignment.



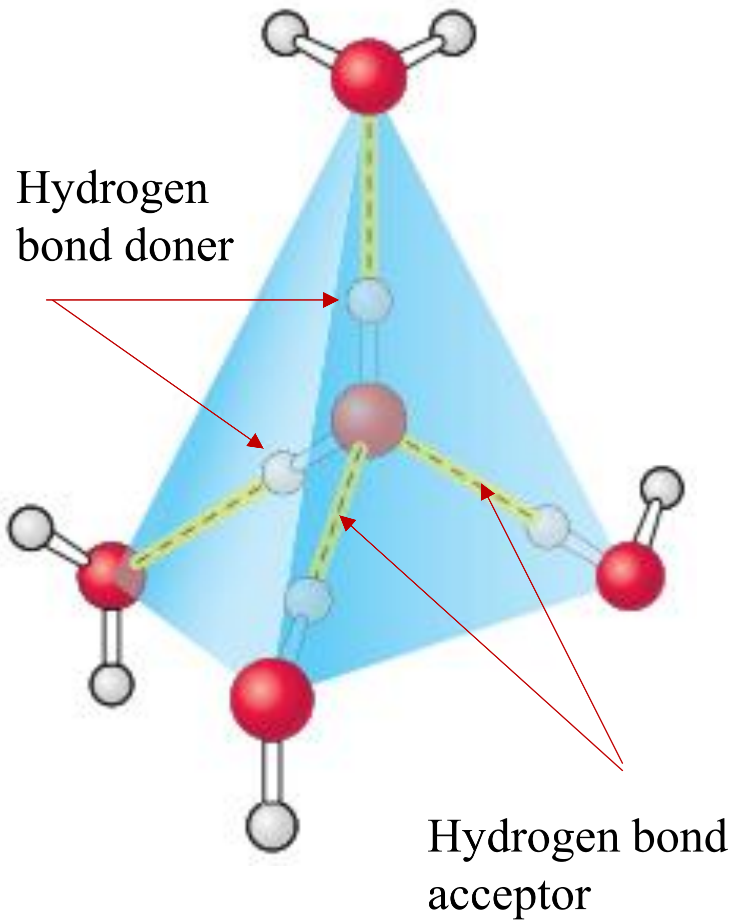
Hydrogen bond acceptor

Hydrogen bond donor

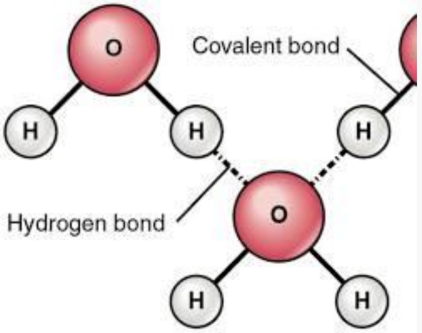


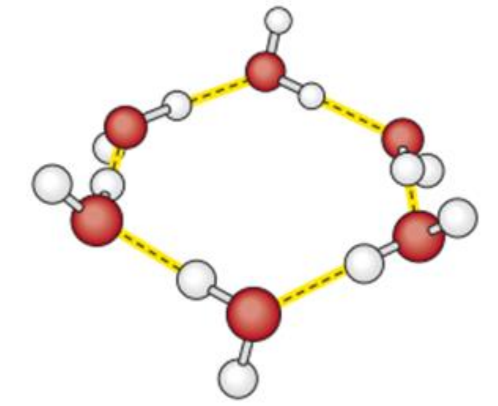
Hydrogen-bond

- A water molecule has the capability to establish up to four hydrogen bonds.
- The oxygen atom within a water molecule acts as the hydrogen acceptor for two hydrogen atoms, while each O—H group function.
- Theoretical calculations and spectroscopic data indicate that water molecules typically participate in only two strong hydrogen bonds—one as a donor and one as an acceptor.



Water Molecules and Hydrogen-bond





- Hydrogen bonds in water are transient, with typical lifetimes of approximately 10^{-11} to 10^{-12} s at 25°C, reflecting the highly dynamic nature of liquid water.
- These short-lived interactions form temporary clusters like six-membered ring structures (Fig 2), which continuously form and break
- Hydrogen bonds make water highly cohesive, meaning water molecules stick together.
- Cohesion leads to high surface tension, allowing insects to walk on water and formation of droplets.
- Hydrogen bonding is responsible for many of water’s unique life-supporting properties, including high heat capacity, solvent capability.

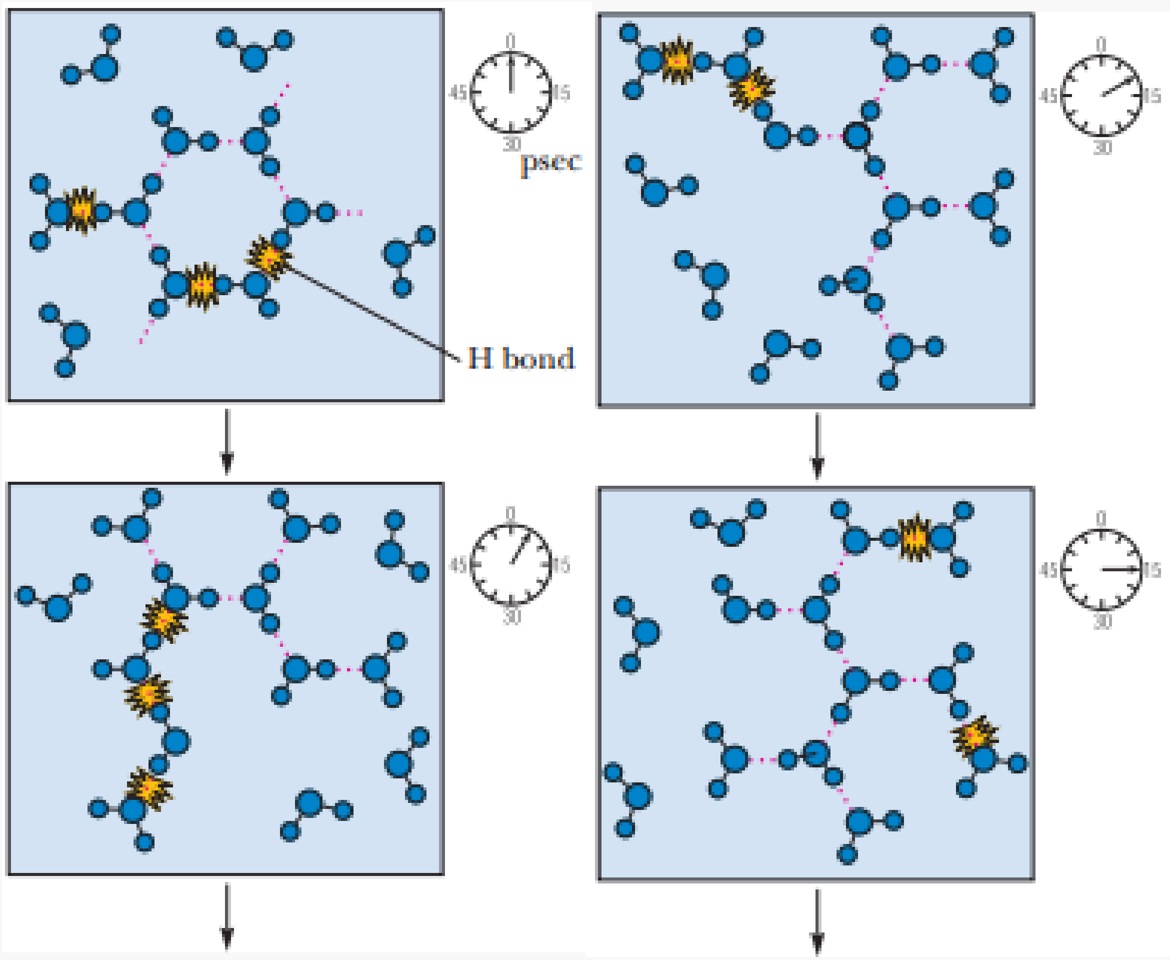
Dynamic Hydrogen-Bond Network in Liquid Water

The figure illustrates the fluid hydrogen-bond network linking water molecules in liquid water.

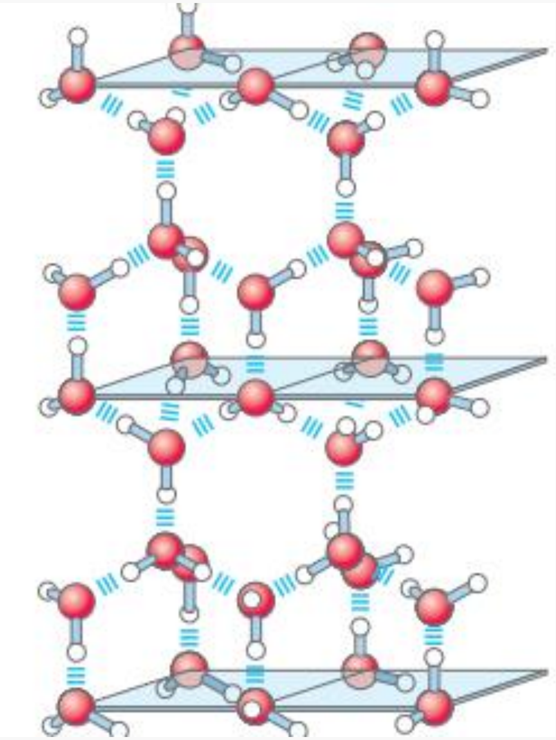
Water molecules undergo continuous vibration, rotation, and translation, maintaining a highly dynamic network.

The network is spatially heterogeneous (not uniform), with local variations in density and in the number of hydrogen bonds per molecule.

The network is influenced by temperature, higher temperatures increase bond breaking, while lower temperatures promote more extensive hydrogen bonding.



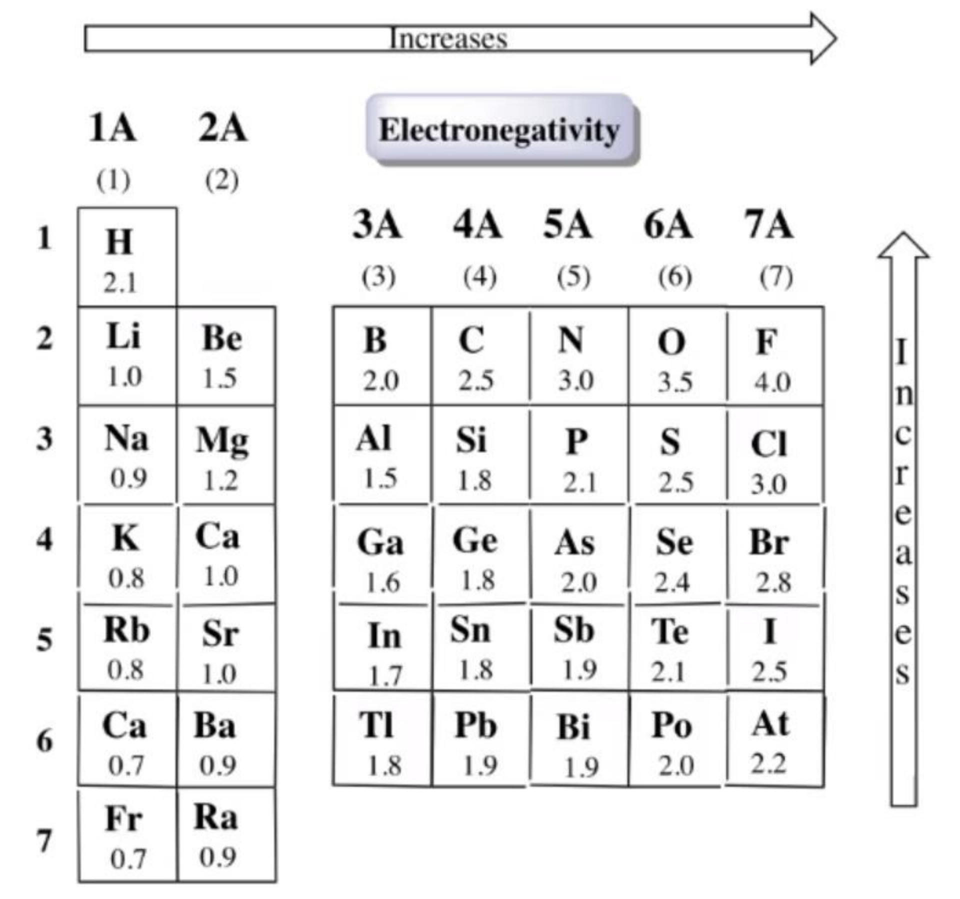
The Unique Structure of Ice



- **Four Hydrogen Bonds**: In ice, each water molecule forms four hydrogen bonds with neighboring molecules, creating a rigid, hexagonal crystalline lattice.
- **Lower Density Than Liquid Water**: This open lattice structure spaces molecules farther apart than in liquid water, making ice less dense.
- **Why Ice Floats (and Why It Matters):** Because ice is less dense, it floats on liquid water, forming an insulating layer that prevents bodies of water from freezing solid, allowing aquatic life to survive in cold climates.

Polarity

Electronegativity is the tendency of an atom to attract shared electrons toward itself in a chemical bond, resulting in a partial negative charge.



Increases

Electronegativity

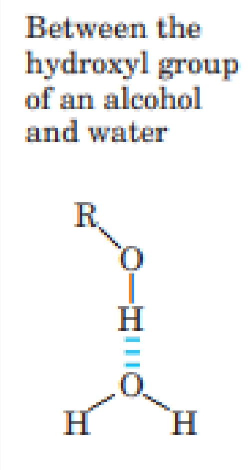
	1A (1)	2A (2)	3A (3)	4A (4)	5A (5)	6A (6)	7A (7)
1	H 2.1						
2	Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
3	Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
4	K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
5	Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
6	Ca 0.7	Ba 0.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
7	Fr 0.7	Ra 0.9					

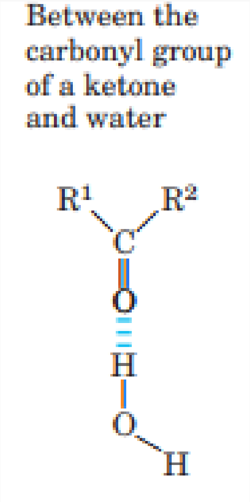
Increases

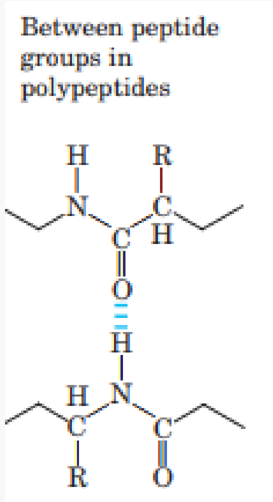
Element	**Electronegativity**
Fluorine	3.98
Oxygen	3.44
Nitrogen	3.04
Sulfur	2.6
Carbon	2.55
Phosphorus	2.20
Hydrogen	2.1

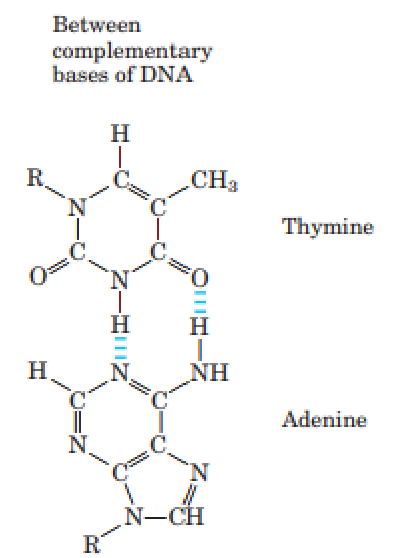
Hydrogen Bond

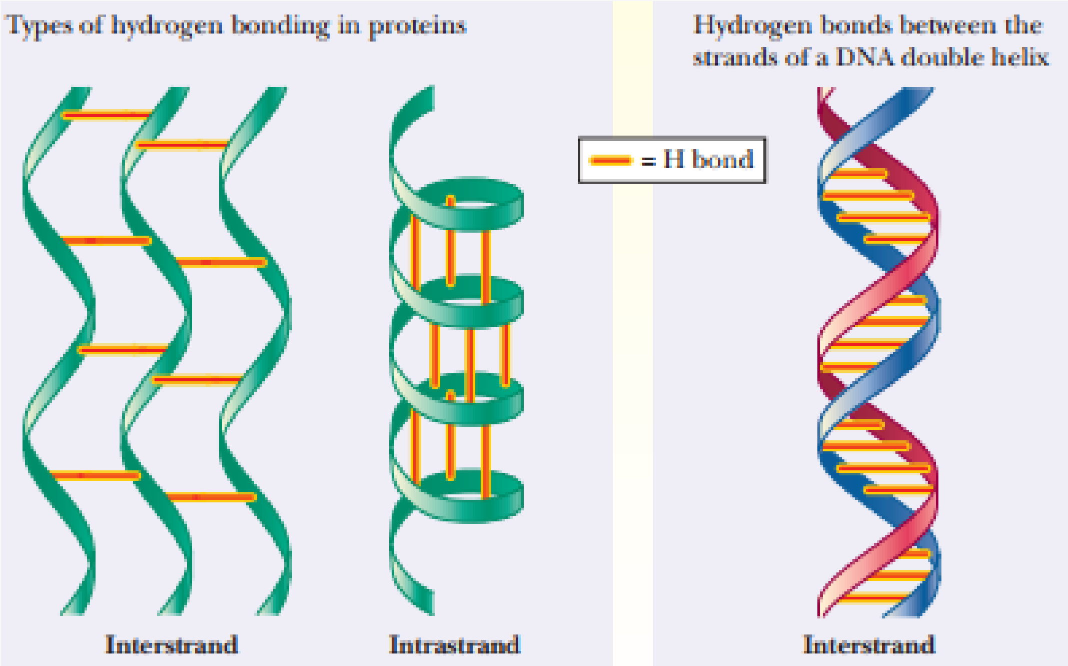
- Hydrogen bonds form between hydrogen and electronegative atoms such as N, O, or S.
- .DNA and RNA rely on hydrogen bonds to hold complementary bases together, stabilizing the DNA double helix and RNA structures.


Between the hydroxyl group of an alcohol and water

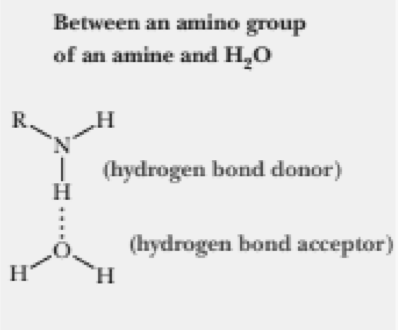

Between the carbonyl group of a ketone and water


Between peptide groups in polypeptides

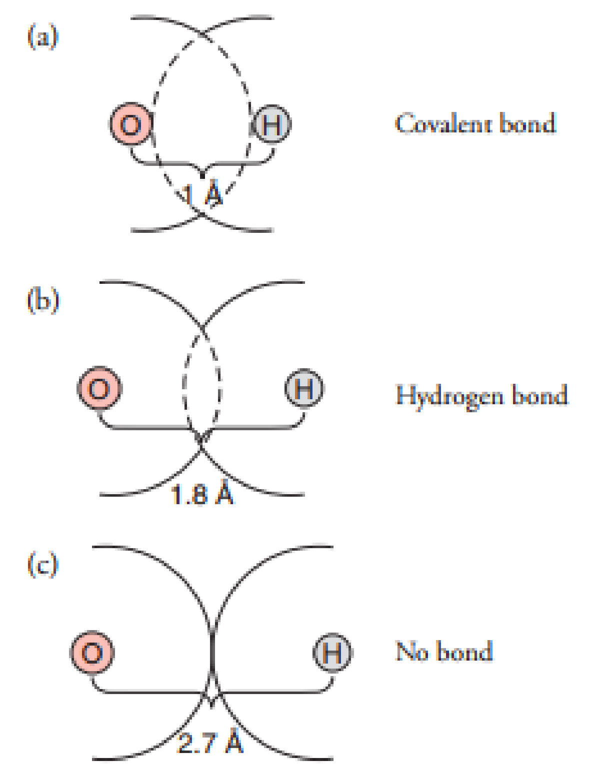

Between complementary bases of DNA


Types of hydrogen bonding in proteins

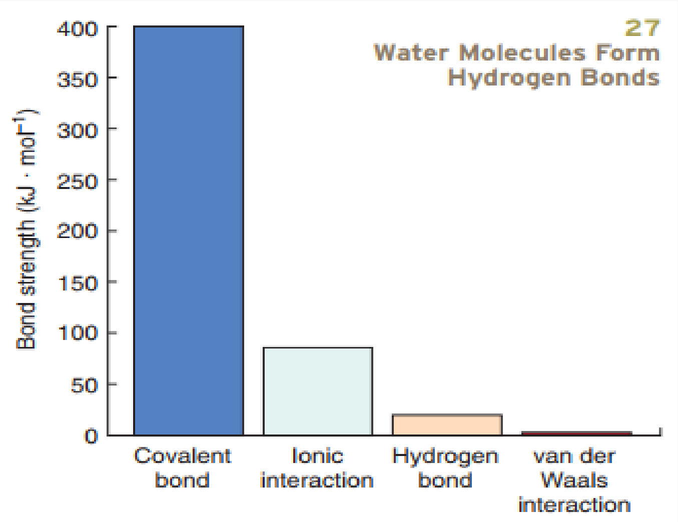
Hydrogen bonds between the strands of a DNA double helix

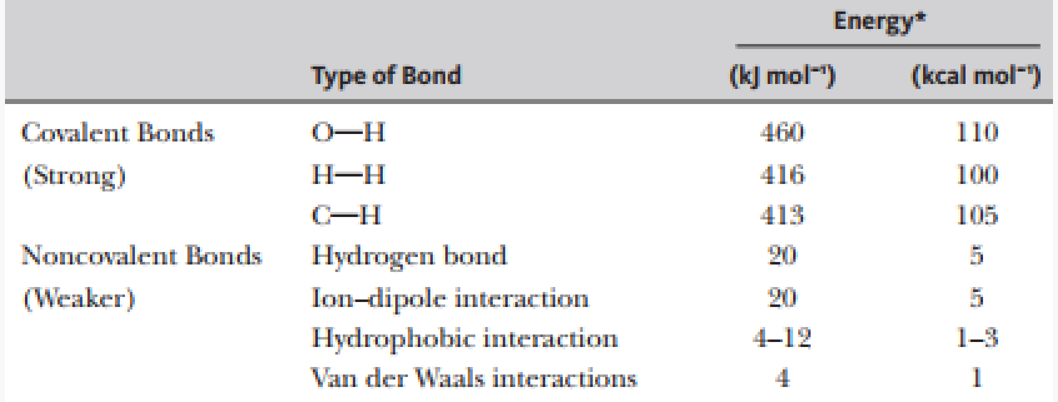

Between an amino group of an amine and H_2O

Hydrogen Bond



- Covalent bonds are strong bonds that define the primary structure of biological molecules.
- Hydrogen bonds, which are much weaker $\approx 20\ kJ.mol^{-1}$, play a crucial role in stabilizing the 3D structure of biomolecules..
- Noncovalent interactions in biological systems also include electrostatic (ionic) interactions between oppositely charged groups, such as $-COO^{-}$ and $-NH^{3+}$, which further contribute to molecular structure and stability.

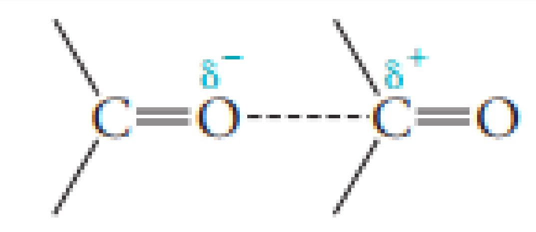


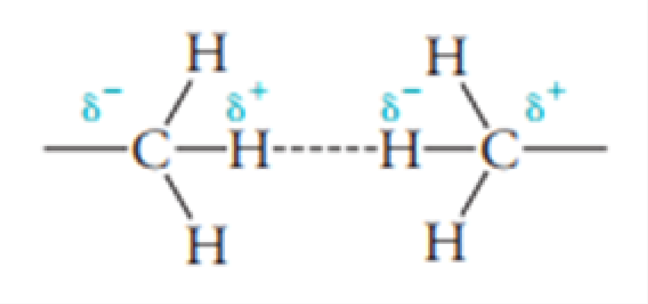


	Type of Bond	Energy* (kJ mol⁻¹)	Energy* (kcal mol⁻¹)
Covalent Bonds (Strong)	O—H	460	110
	H—H	416	100
	C—H	413	105
Noncovalent Bonds (Weaker)	Hydrogen bond	20	5
	Ion–dipole interaction	20	5
	Hydrophobic interaction	4–12	1–3
	Van der Waals interactions	4	1

Van der Waals

Van der Waals forces are weak, noncovalent intermolecular interactions that arise from temporary or permanent dipoles due to fluctuations in electron distribution between atoms and molecules. They are weaker than hydrogen bonds but become significant when many such interactions act together.





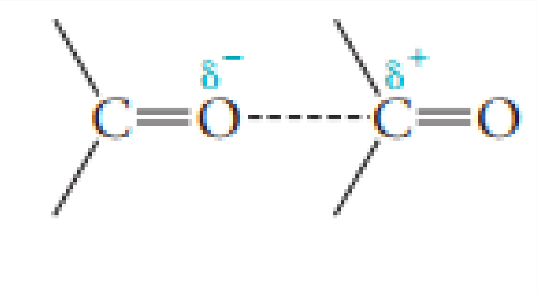
Significance of Van der Waals Forces

Biological systems: Contribute to protein folding, DNA base stacking, and enzyme–substrate recognition.

Molecular packing: Influence the structure and stability of solids and liquids, affecting melting and boiling points.

Materials science: Play a role in adhesion, lubrication, and the behavior of colloids and nanoparticles.

Types of Van der Waals

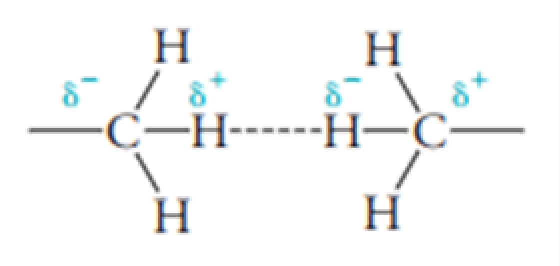


Dipole-Dipole Interactions (about 9 kJ/mol)

- Occur between molecules with permanent dipoles (i.e., polar molecules).
- The partial positive end of one molecule attracts the partial negative end of another.
- Contribute to molecular alignment and stability.
- Example: Interactions between HCl molecules.

Dispersion Forces (London Forces): very week force range (0.1-10 kJ/mol)

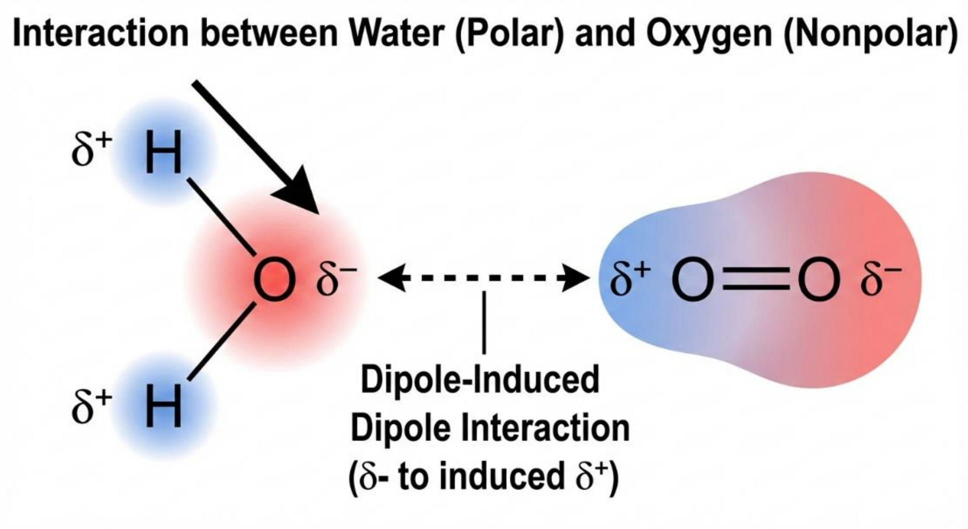
- Arise from temporary fluctuations in electron density, creating instantaneous dipoles.
- Present in all atoms and molecules, including nonpolar species.
- Strength increases with molecular size, polarizability, and surface area.
- Example: Attraction between noble gas atoms or between hydrocarbon chains..



Types of Van der Waals

Dipole-Induced Dipole Interactions

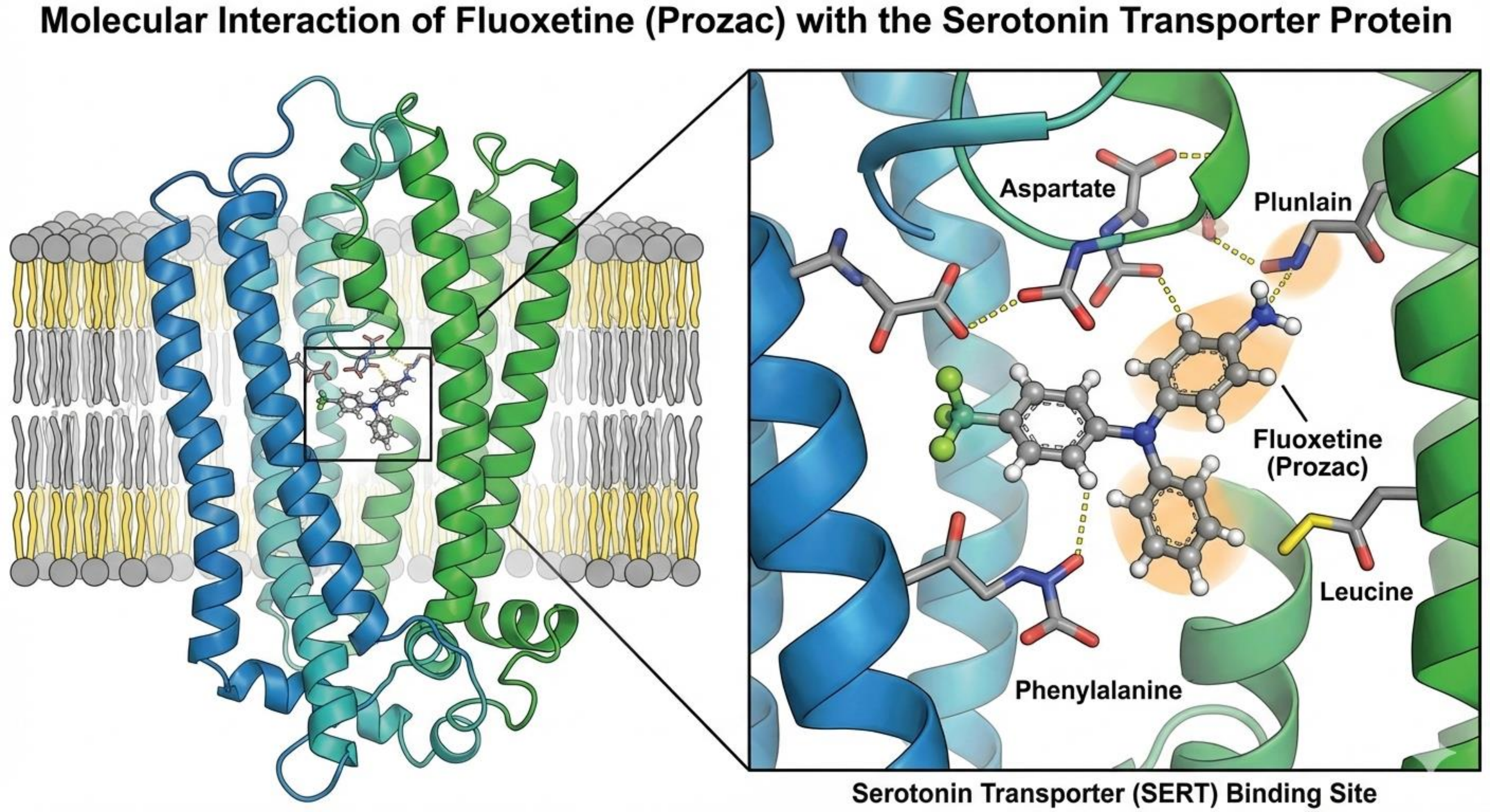
- Occur when a polar molecule induces a temporary dipole in a nearby nonpolar molecule.
- Generally weaker than dipole–dipole interactions.
- Important in solubility and weak association phenomena.
- Example: O_2 or Ar interacting with H_2O, where water induces a dipole in the nonpolar molecule.



Role of Fluorine (hydron bond) in Drug Design

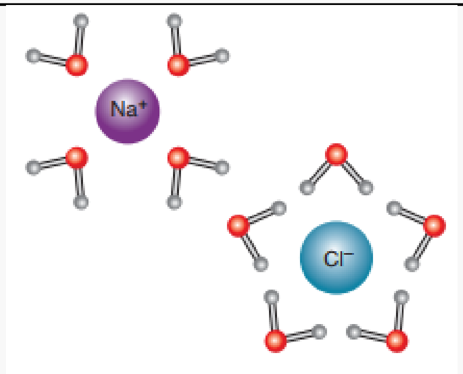
- Strategic incorporation: Drugs like Prozac (fluoxetine), fluorouracil, and Ciprofloxacin include fluorine to enhance chemical stability and biological activity without major structural changes. (see next slide).
- Electronegativity effects: Fluorine's high electronegativity reduces partial positive charges, improving membrane permeability and cellular uptake.
- Molecular interactions: Fluorine participates in hydrogen bonding (C–F…H–C) and dipole–dipole interactions (C–F…C=O), enhancing target binding.
- Size and positioning: Fluorine influences molecular shape, optimizing fit within target binding pockets.
- Therapeutic impact: Stronger binding enables higher efficacy at lower doses, improving therapeutic outcomes.

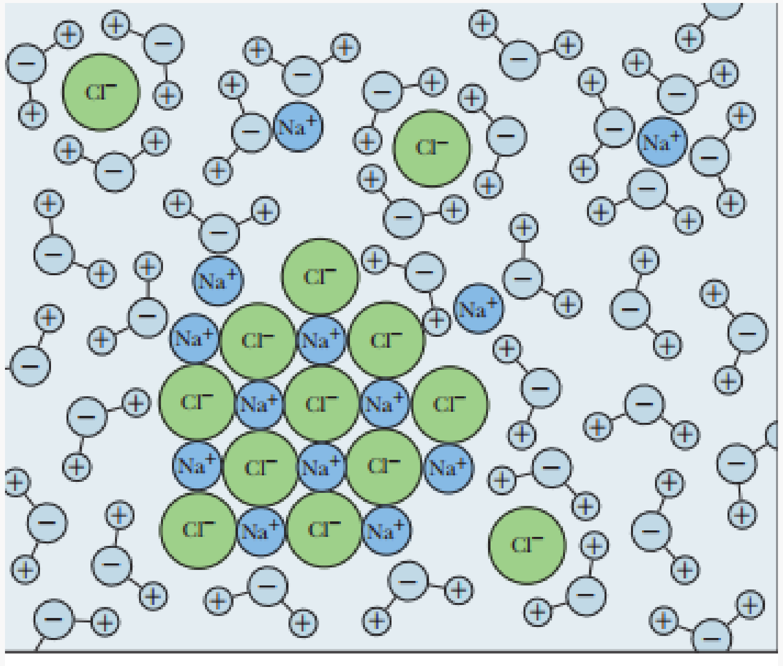
Role of Fluorine (hydron bond) in Drug Design



Water as a Universal Solvent

- Hydrogen bonding and electrostatics: Water dissolves many compounds due to its ability to form hydrogen bonds and interact electrostatically with solutes.
- High dielectric constant: Water's high relative permittivity reduces electrostatic attraction between ions, enhancing solubility.
- Ion solvation: Polar water molecules surround ions (e.g., Na^+, Cl^-), aligning their partial charges with opposite ions to stabilize them in solution..





Role of Dielectric Constant in Solubility

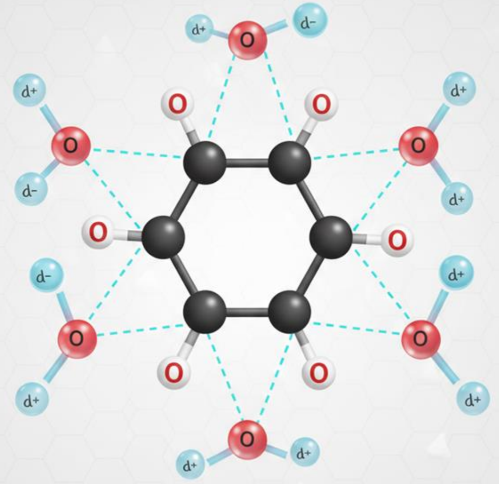
- **High dielectric constant → Greater charge separation**: Solvents with high dielectric constants weaken electrostatic attraction between oppositely charged ions, enhancing solubility.
- **Water's dielectric constant (~78.5 at 25°C)**: Allows water to stabilize ions and dissolve ionic and polar compounds effectively.
- Comparison to Other Solvents:
  - Water ($\varepsilon \approx 78.5$) – Excellent at dissolving ionic compounds.
  - Ethanol ($\varepsilon \approx 24.5$) – Moderately good at dissolving polar substances.
  - Hexane ($\varepsilon \approx 1.9$) – Poor solvent for ionic compounds.

In summary, Water's high dielectric constant is a major reason it acts as a universal solvent, crucial for biological and chemical processes.

Hydrophilic (Water-Loving) Molecules

- Definition: Hydrophilic molecules interact readily with water and often dissolve easily due to the presence of polar functional groups (e.g., –OH, –COOH, –NH_2) that can form hydrogen bonds.
- Example: Glucose has multiple hydroxyl groups capable of forming hydrogen bonds with water, making it highly soluble. Its six hydrogen-bonding oxygens enable strong interactions with solvent molecules

GLUCOSE: HIGHLY SOLUBLE DUE HYDROGEN BONDING



Six hydrogen-bonding oxygiens on glulcse form strong, multiple interactions with solvent water molecules, enabling high solubility.

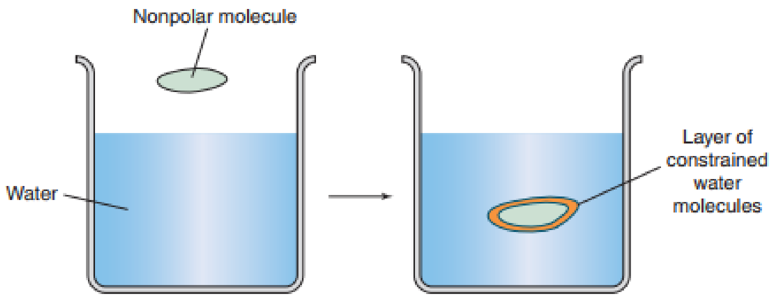
Hydrophobic Effect

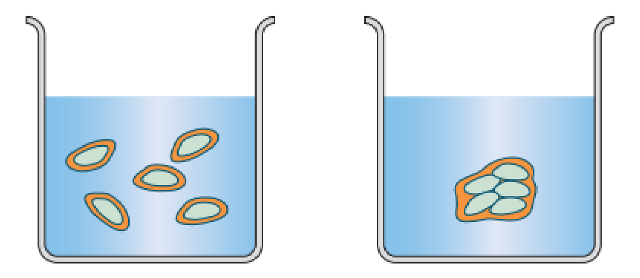
- Definition: Hydrophobic molecules do not interact favorably with water and are typically insoluble. They usually lack polar groups and are composed of nonpolar bonds (e.g., hydrocarbons).
- Examples: Oils and fats are common hydrophobic molecules.
- Hydrophobic effect: Driven by water’s tendency to maximize entropy. Water molecules organize around hydrophobic solutes, and aggregation of hydrophobic molecules allows water to adopt a more disordered, thermodynamically favorable state.

Hydrophobic Effect

The Hydrophobic Effect: Driving Biological Processes: The hydrophobic effect is a powerful force that plays a critical role in:

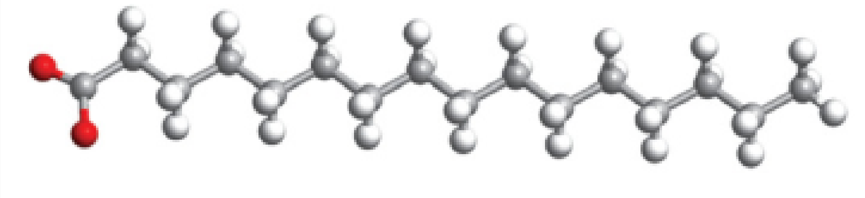
- **Protein Folding**: Hydrophobic amino acid side chains tend to cluster together in the interior of a protein, away from water, while hydrophilic side chains are exposed on the surface.
- **Membrane Formation**: Phospholipids, which make up cell membranes, have a hydrophilic "head" and hydrophobic "tails." In water, they spontaneously arrange themselves into a bilayer with the hydrophobic tails facing inward, away from water, and the hydrophilic heads facing outward, interacting with water.
- **Aggregation of Nonpolar Molecules**: Nonpolar molecules, when placed in water, tend to aggregate together.



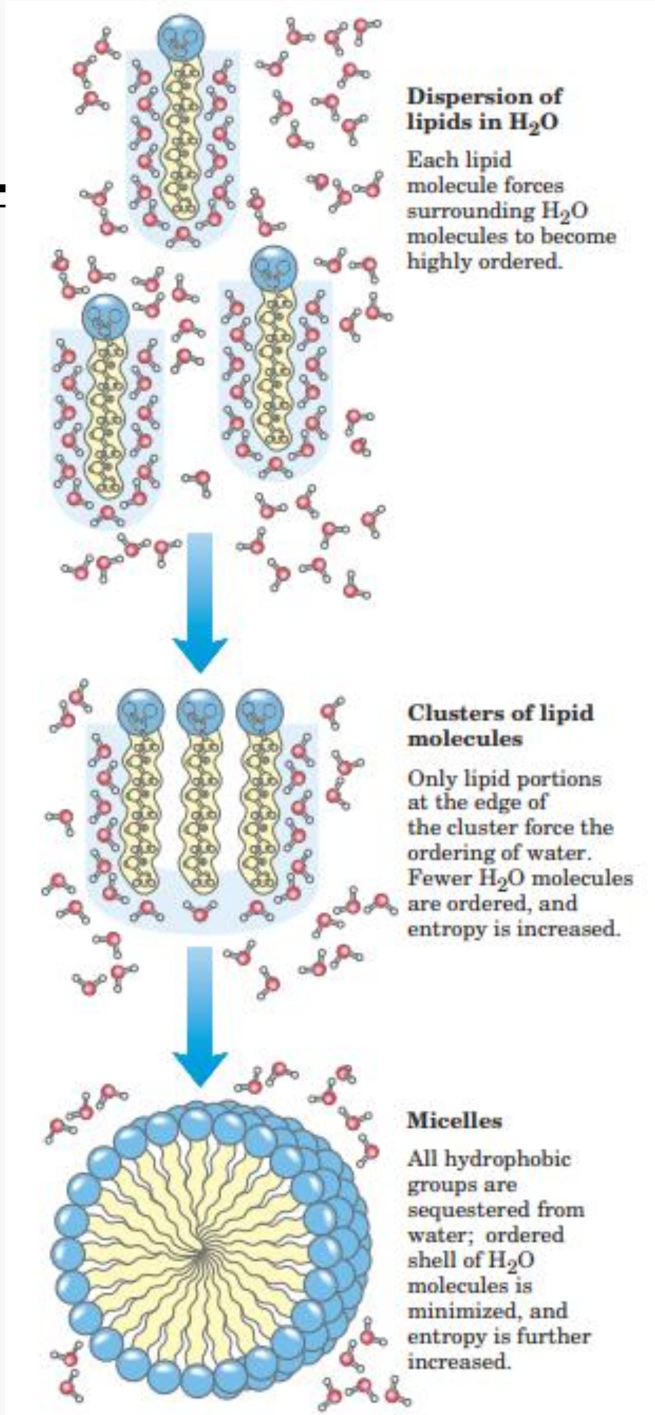


Hydrophilic and Hydrophobic Effect

- Amphiphilic molecules, like fatty acid palmitate, have both hydrophilic and hydrophobic nature.



- The polar groups of amphiphiles align towards solvent molecules, leading to hydration, while the nonpolar groups aggregate because of the hydrophobic effect.
- Amphiphilic molecules can form micelles, where the hydrophobic tails aggregate, staying away from water contact due to the hydrophobic effect.



Dispersion of lipids in H_2O

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

Clusters of lipid molecules

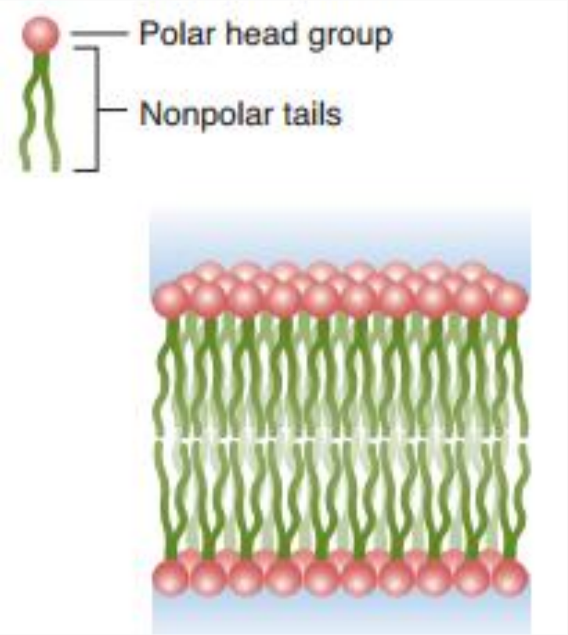
Only lipid portions at the edge of the cluster force the ordering of water. Fewer H_2O molecules are ordered, and entropy is increased.

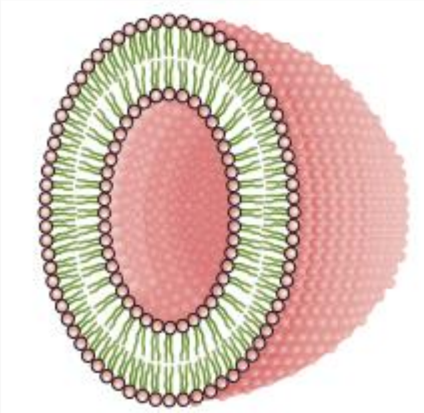
Micelles

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

Hydrophilic and Hydrophobic Effect

- Amphiphilic lipids, integral to the structural composition of biological membranes, organize into bilayers consisting of two layers.
- The hydrophobic core of a lipid bilayer acts as a barrier to diffusion.
- To minimize exposure to the solvent, a lipid bilayer tends to close up, forming a vesicle.
- Similar structural arrangements are found in many subcellular compartments (organelles) within eukaryotic cells.

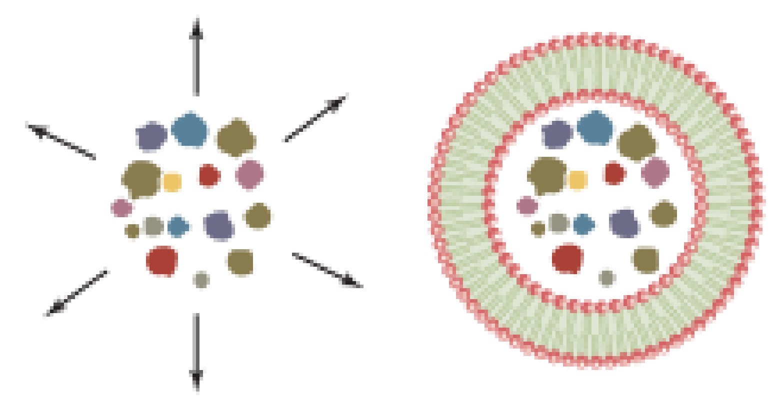


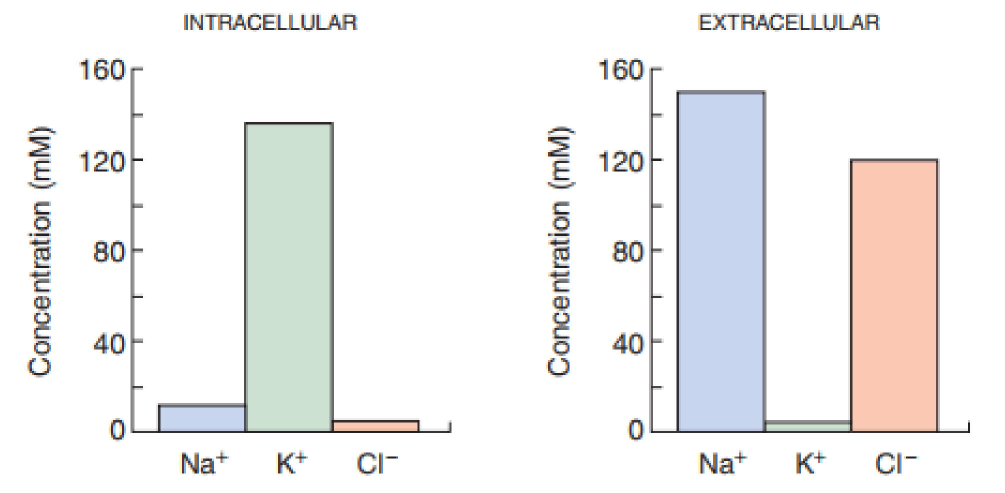


Cell Membrane and Selective Permeability

The lipid bilayer's nonpolar core acts as a high-resistance barrier.

- **Permeable**: Small, nonpolar molecules (O_2, CO_2) and some small uncharged polar molecules (H_2O, though slowly) diffuse via passive transport.
- : Large polar molecules (glucose) and all charged ions (Na^+, K^+, Cl^-) are blocked by the hydrophobic tails, requiring specific transport proteins.
- **Ion Regulation**: Human cell membranes maintain high K^+ and low Na^+/Cl^- levels relative to extracellular fluid. **K^+/Na^+ Pumps**: Active transport via Na^+/K^+ pumps maintains this balance, which is essential for nerve signaling and cellular function.





Cell Membrane and Selective Permeability

The lipid bilayer’s nonpolar core acts as a high-resistance barrier.

Feature	**Intracellular (Inside)**	**Extracellular (Outside)**	**Primary Transport Method**
Sodium ($\boldsymbol{Na^+}$)	Low	**High**	Na^+/K^+ Pump (Out)
Potassium ($\boldsymbol{K^+}$)	**High**	Low	Na^+/K^+ Pump (In)
Chloride ($\boldsymbol{Cl^-}$)	Low	**High**	Various Channels/Transporters
Oxygen ($\boldsymbol{O_2}$)	Variable	Variable	Simple Diffusion

Water Ionization

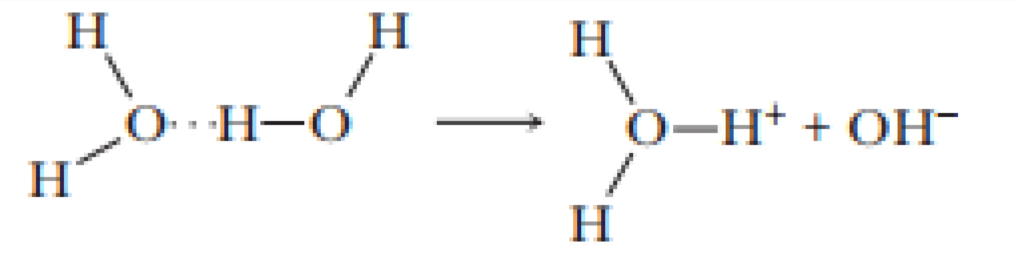
- Water can actively participate in biological process,

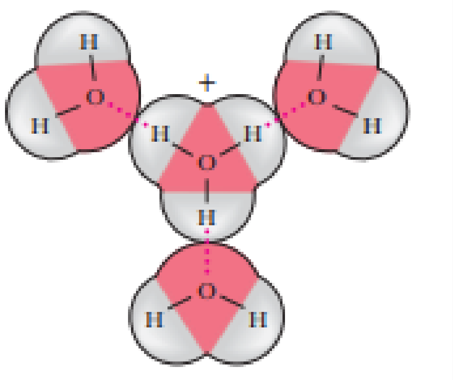
$$H_2O \rightleftharpoons H^+ + OH^-$$

- Aqueous solutions do not actually contain lone protons. Instead, the H^+ can be visualized as combining with a water molecule to produce a **hydronium ion** (H_3O^+).

$$2H_2O \rightarrow H^+ + OH^- + H_2O \rightarrow H_3O^+ + OH^-$$

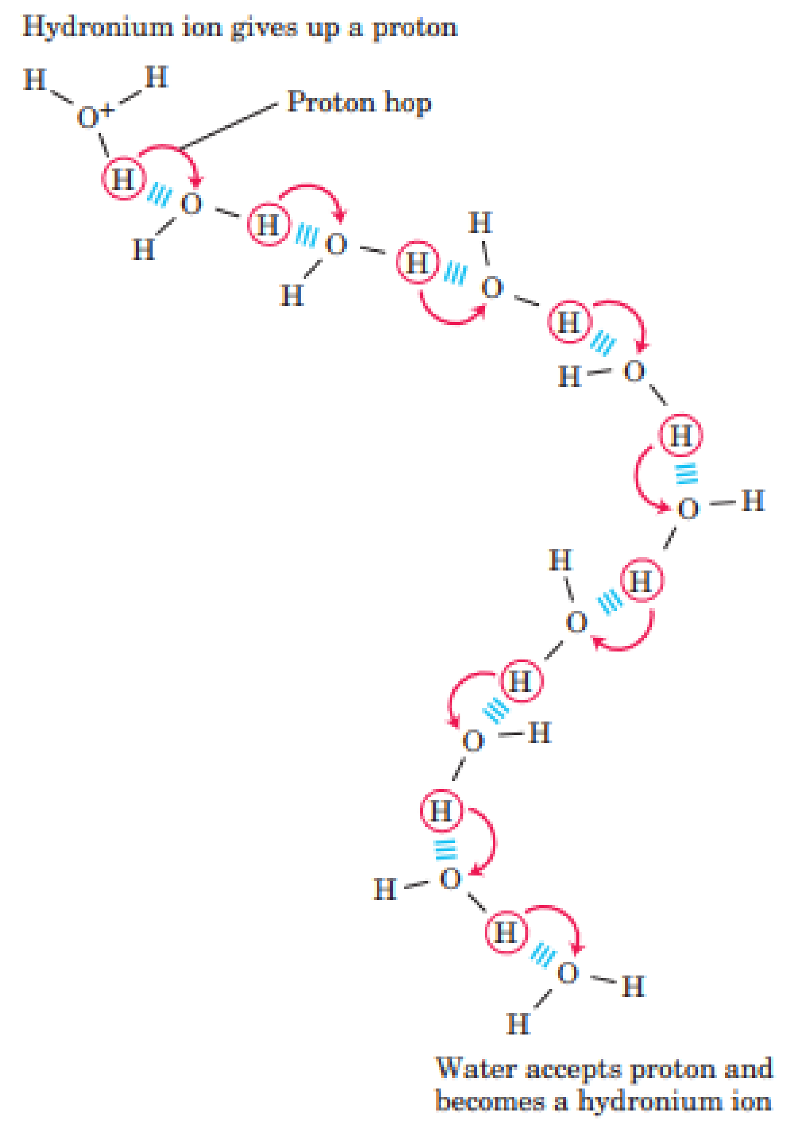
- H^+ is somewhat delocalized, so it probably exists as part of a larger, fleeting structure.

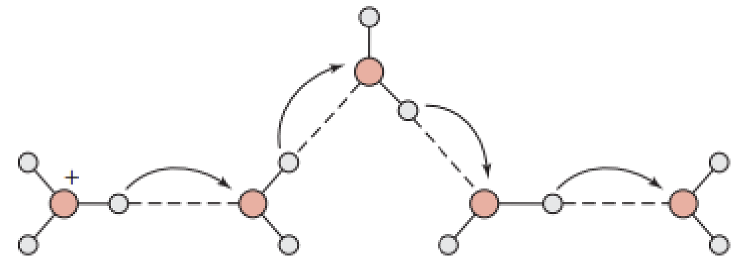




Proton Mobility in Water

- Proton hopping: A proton (H^+) does not remain bound to a single water molecule; it transfers rapidly through a network of hydrogen-bonded water molecules via the Grotthuss mechanism (also know as proton jumping).
- High effective mobility: This hopping enables H^+ to move much faster than other ions, which rely on physical diffusion.
- Biochemical significance: The rapid proton transfer makes acid–base reactions among the fastest biochemical reactions.





$[H^+]$ and $[OH^-]$ Relations

- Pure water exhibits only a slight tendency to ionize, yield $H^+$and OH^- , quite small.
- The ionization of water can be described by a dissociation constant, *K*

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

- Because $[H_2O]$ concentration is constant , we can redefine K as *K*w, **ionization constant of water**

$$K_w = [H^+][OH^-] = (10^{-7}M)(10^{-7}M) = 10^{-14}\ M^2$$

Hydrogen ion concentration is expressed as:

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

pH of water at 25^oC is

$$pH = -\log 10^{-7} = 7$$

Relationship between $[H^+]$ and $[\mathrm{OH}^-]$

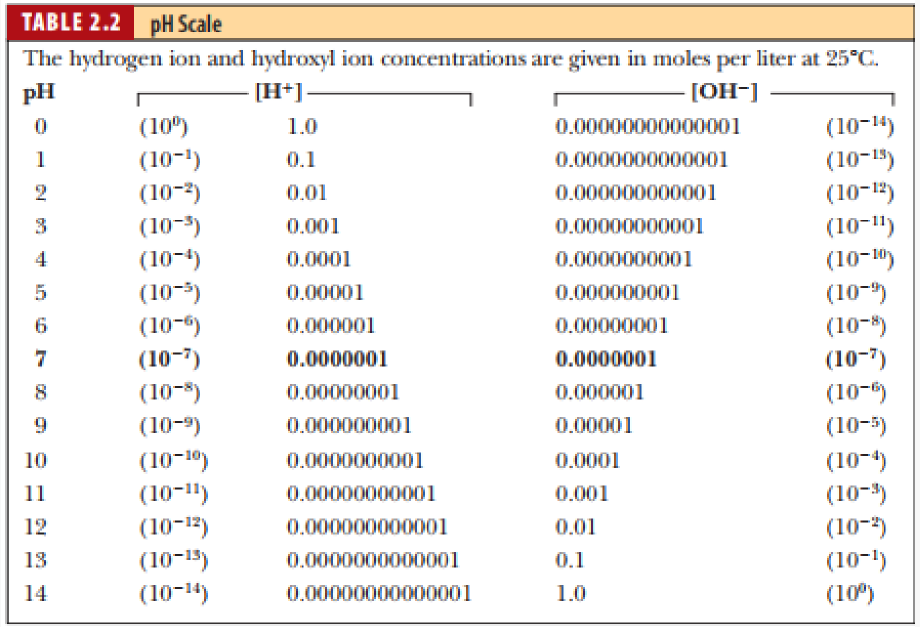
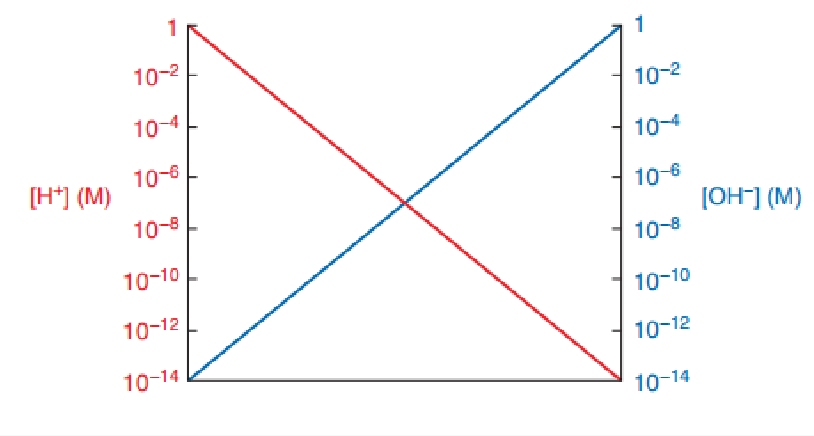
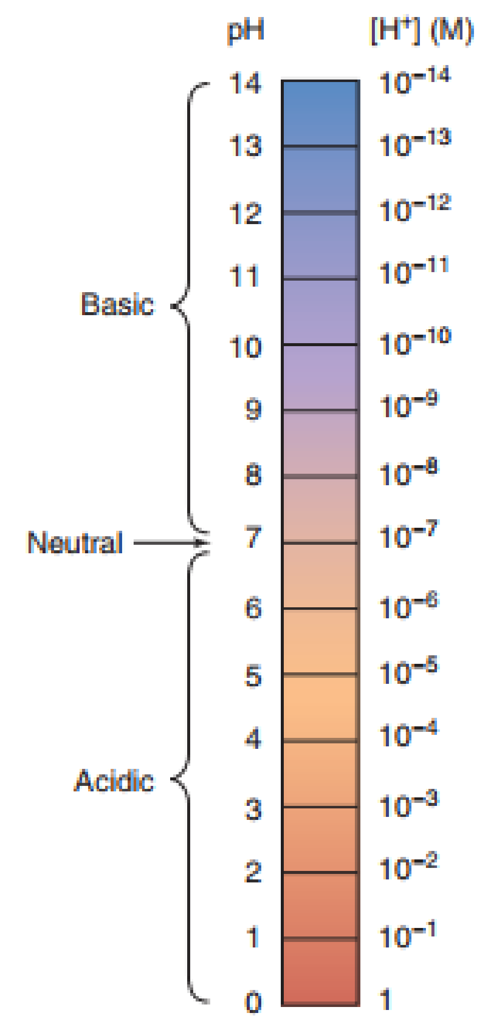


TABLE 2.2 pH Scale

The hydrogen ion and hydroxyl ion concentrations are given in moles per liter at 25°C.

pH	[H+]		[OH−]	
0	(10^{0})	1.0	0.00000000000001	(10^{-14})
1	(10^{-1})	0.1	0.0000000000001	(10^{-13})
2	(10^{-2})	0.01	0.000000000001	(10^{-12})
3	(10^{-3})	0.001	0.00000000001	(10^{-11})
4	(10^{-4})	0.0001	0.0000000001	(10^{-10})
5	(10^{-5})	0.00001	0.000000001	(10^{-9})
6	(10^{-6})	0.000001	0.00000001	(10^{-8})
7	$\mathbf{(10^{-7})}$	**0.0000001**	**0.0000001**	$\mathbf{(10^{-7})}$
8	(10^{-8})	0.00000001	0.000001	(10^{-6})
9	(10^{-9})	0.000000001	0.00001	(10^{-5})
10	(10^{-10})	0.0000000001	0.0001	(10^{-4})
11	(10^{-11})	0.00000000001	0.001	(10^{-3})
12	(10^{-12})	0.000000000001	0.01	(10^{-2})
13	(10^{-13})	0.0000000000001	0.1	(10^{-1})
14	(10^{-14})	0.00000000000001	1.0	(10^{0})





pH of a Solution

- Strong electrolytes are substances that almost entirely dissociate into ions when dissolved in a solution. The term "electrolyte" refers to substances that can produce ions in a solution, leading to an enhanced electrical conductivity of the solution.
- Examples of strong electrolytes include many salts like NaCl and K_2SO_4, strong acids such as HCl, and strong bases like NaOH.
- Dissociation of a strong acid in water

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

The equilibrium (dissociation) constant for this reaction is

$$K = \frac{[H_3O^+][Cl^-]}{[H_2O][HCl]}$$

pH of a Solution

- Because $[H_2O]$ is essentially constant in dilute aqueous solutions, $K_a = K[H_2O]$ and $[H_3O^+]$ is often replace w by H^+

$$K_a = \frac{[H^+][Cl^-]}{[HCl]}$$

For HCl, K_a is exceedingly large because vanishingly small amount of HCl in the solution.

pH of a Solution

Similarly, adding the base sodium hydroxide (NaOH) increases the pH (decreases $[H^+]$) . Here H_3O^+ is the acid that donates a proton to the added base.

$$\mathrm{NaO}H + H_3O^+ \rightarrow Na^+ + 2H_2O$$

Example-2.3

Calculate the pH of 1 L of water to which is added

(a) 10 mL of 5.0 M HCl

(b) 10 mL of 5.0 M NaOH.

Ocean Acidification and Carbonate Chemistry

- CO₂ dissolution: Atmospheric CO₂ dissolves in seawater, reacting to form carbonic acid, which partially dissociates into protons (H^+) and bicarbonate (HCO_3^-):

$$CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$$

- Ocean pH change: The average ocean pH is currently ~8.0 and is projected to decrease to ~7.8 over the next century due to increased CO₂ uptake.
- Impact on carbonate ions: Many marine organisms (e.g., corals, mollusks, plankton) rely on carbonate ions (CO_3^{2-}) to build calcium carbonate ($CaCO_3$) shells and skeletons. Increased H^+ shifts carbonate into bicarbonate:

$$CO_3^- + H^+ \rightleftharpoons HCO_3^-$$

- Shell formation challenge: Reduced availability of CO_3^{2-} makes shell formation more difficult and can slow growth of calcifying organisms.
- Dissolution of CaCO₃: Increased acidity can also dissolve existing calcium carbonate structures:

$$CaCO_3 + H^+ \rightleftharpoons HCO_3^- + Ca^{2+}$$

Weak Acid: Dissociation Constant K_a and Acid Tendency

pKa describes an acid's tendency to ionize in a solution.

- Strong acids like HCl completely dissociate in water, while weak acids like acetic acid only partially ionize.
- For example, acetic acid partially ionizes, or donates only some of its protons to water:

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

Equilibrium constant for this reaction is:

$$K = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH][H_2O]}$$

- Because the concentration of H_2O is much higher than the other concentrations, it is considered constant. incorporated into the value of K, which is then formally known as $\boldsymbol{K_a}$, the **acid dissociation constant**

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]} = 1.74 \times 10^{-5}\ M$$

pK Values and Acid Tendency

- Because K_a is typically very small number like H^+ concentration, it is convenient to use log scale, so:

$$pK = -\log K_a$$

A lower pKa value indicates a stronger acid, as it readily donates protons.

For example, pK value of acetic acid is $pK = -\log(1.74 \times 10^{-5}) = 4.76$

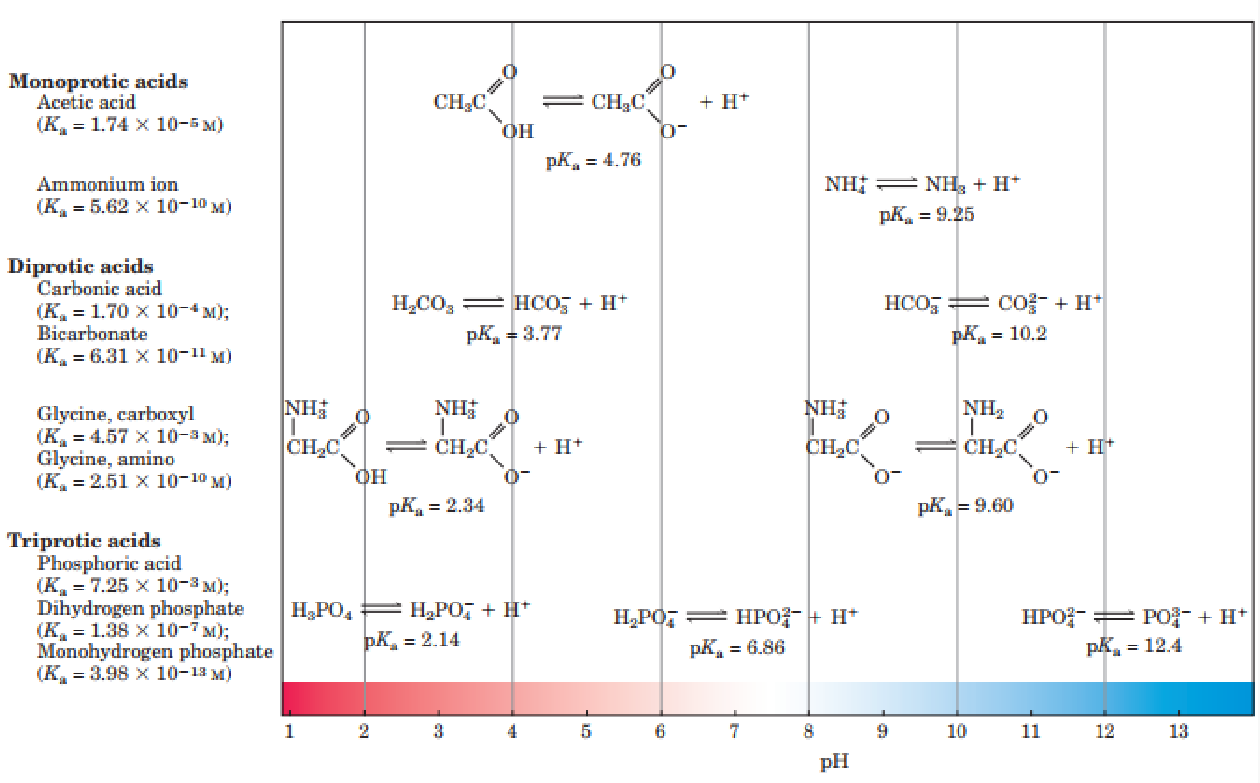
- Consider an acid such as the ammonium ion, NH_4^+:

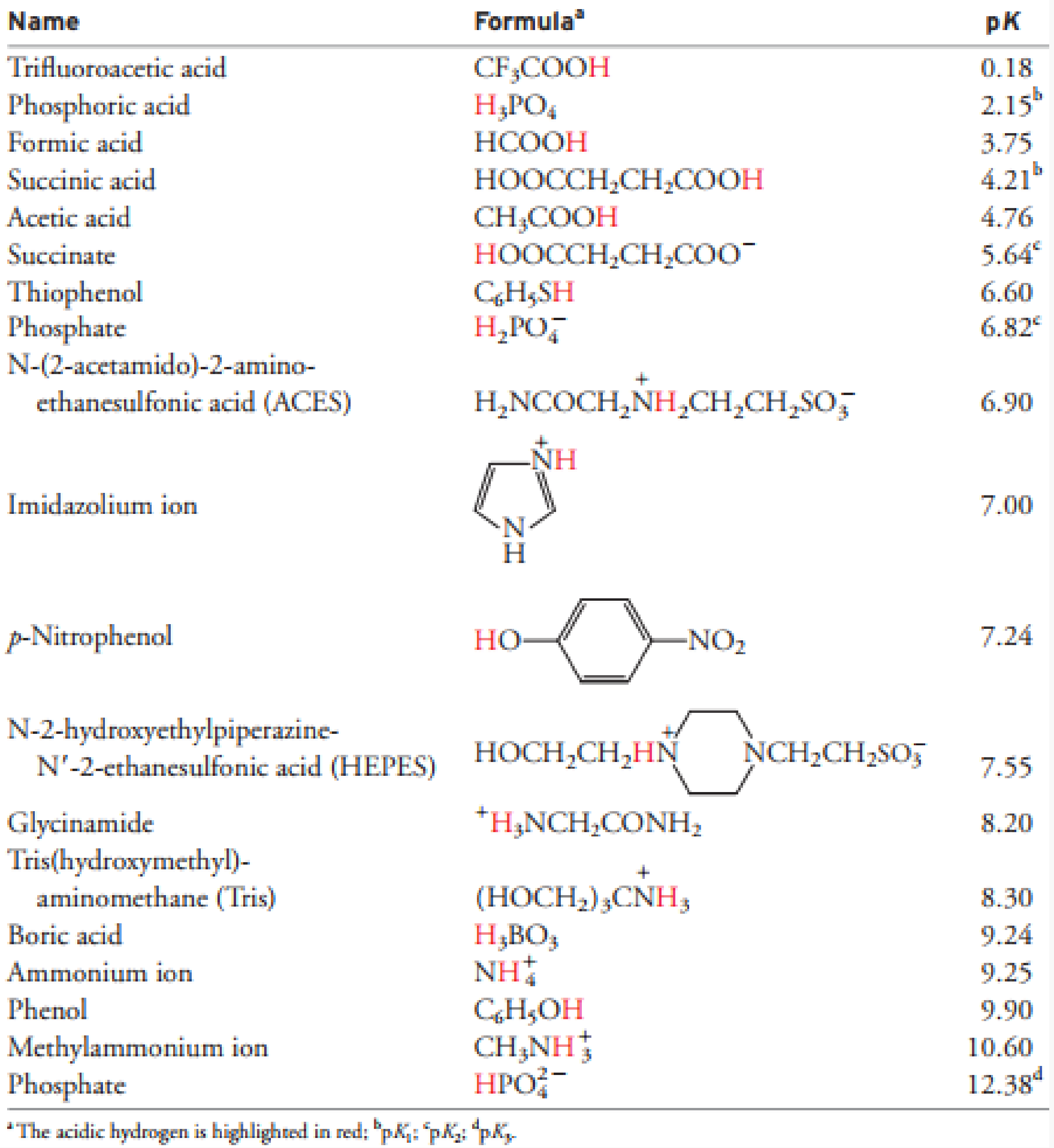
$$NH_4^+ \rightleftharpoons NH_3 + H^+$$

Where $K_a = 5.62 \times 10^{-10}$ and corresponds to a pK is 9.25. Indicate that ammonium ion is a relatively weak acid, tends not to donate a proton or ammonia (NH_3) is the **conjugate base** of the acid NH_4^+.

pH value of a Solution and pK

pK introduced as index to express the acidity of weak acids. pK values of some acids.





Name	Formula[a]	pK
Trifluoroacetic acid	CF_3COOH	0.18
Phosphoric acid	H_3PO_4	2.15[b]
Formic acid	$HCOOH$	3.75
Succinic acid	$HOOCCH_2CH_2COOH$	4.21[b]
Acetic acid	CH_3COOH	4.76
Succinate	$HOOCCH_2CH_2COO^-$	5.64[c]
Thiophenol	C_6H_5SH	6.60
Phosphate	$H_2PO_4^-$	6.82[c]
N-(2-acetamido)-2-amino-ethanesulfonic acid (ACES)	$H_2NCOCH_2\overset{+}{N}H_2CH_2CH_2SO_3^-$	6.90
Imidazolium ion	$\overset{+}{N}H$ N H	7.00
p-Nitrophenol	HO— —NO_2	7.24
N-2-hydroxyethylpiperazine-N′-2-ethanesulfonic acid (HEPES)	$HOCH_2CH_2H\overset{+}{N}$ $NCH_2CH_2SO_3^-$	7.55
Glycinamide	$^+H_3NCH_2CONH_2$	8.20
Tris(hydroxymethyl)-aminomethane (Tris)	$(HOCH_2)_3C\overset{+}{N}H_3$	8.30
Boric acid	H_3BO_3	9.24
Ammonium ion	NH_4^+	9.25
Phenol	C_6H_5OH	9.90
Methylammonium ion	$CH_3NH_3^+$	10.60
Phosphate	HPO_4^{2-}	12.38[d]

[a] The acidic hydrogen is highlighted in red; [b] pK_1; [c] pK_2; [d] pK_3.

Example

What is the final pH when 0.1 mol of acetic acid (HAc) is added to water and the volume of the solution is adjusted to equal 1 L? Hint: $K_a = 1.74 \times 10^{-5} M$ and $HAc \leftrightarrow H^+ + Ac^-$

pH of Solution and pK

pH depends on the equilibrium between HA and A^-

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

Equilibrium constant

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

And pH

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

The above equation is called **Henderson-Hasselbalch equation**. It relates the pH of the solution to the pK of an acid and the concentration of the acids (HA) and its conjugate base (A^-), useful for practical calculations to predict the pH of solution.

Example 2.3

Calculate the pH of a 1-L solution to which has been added 6.0 mL of 1.5 M acetic acid (CH_3COOH) and 5.0 mL of 0.4 M sodium acetate (CH_3COONa). [Hints: pK value of acetic acids is 4.76]

Example 2.4

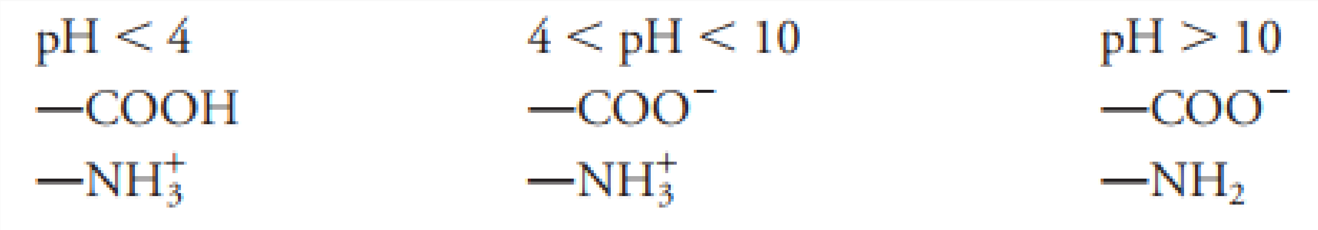
Calculate the concentration of formate ($HCOO^-$)in a 10-mM solution of formic acid at pH 4.15.
Hint: pK value of formic acid is 3.75 $HCOOH \rightleftharpoons H^+ + HCOO^-$

Example 2.5

Determine which molecular species of phosphoric acid predominates at pH values of **(a)** 1.5, **(b)** 4, **(c)** 9, and **(d)** 13.

Ionization of Biomolecules

- pH-dependent ionization of biological molecules is also key to understanding their structures and functions.
- Many of the functional groups on biological molecules act as acids and bases. Their ionization states depend on their respective p*K* values and on the pH ($[H^+]$) of their environment.



$pH < 4$	$4 < pH < 10$	$pH > 10$
$—COOH$	$—COO^-$	$—COO^-$
$—NH_3^+$	$—NH_3^+$	$—NH_2$

Tool and Techniques: Buffer

- When a strong acid such as HCl is added to pure water, all the added acid contributes directly to a decrease in pH.
- But when HCl is added to a solution containing a weak acid in equilibrium with its conjugate base (A^-), the pH does not change so dramatically.

$$HCl \rightarrow H^+ + Cl^- \; [large\ incrase\ in\ [H^+]$$

$$HCl + A^- \rightarrow HA + Cl^- \; [small\ increase\ in\ [H^+]$$

Similarly, behavior when we add a strong base.

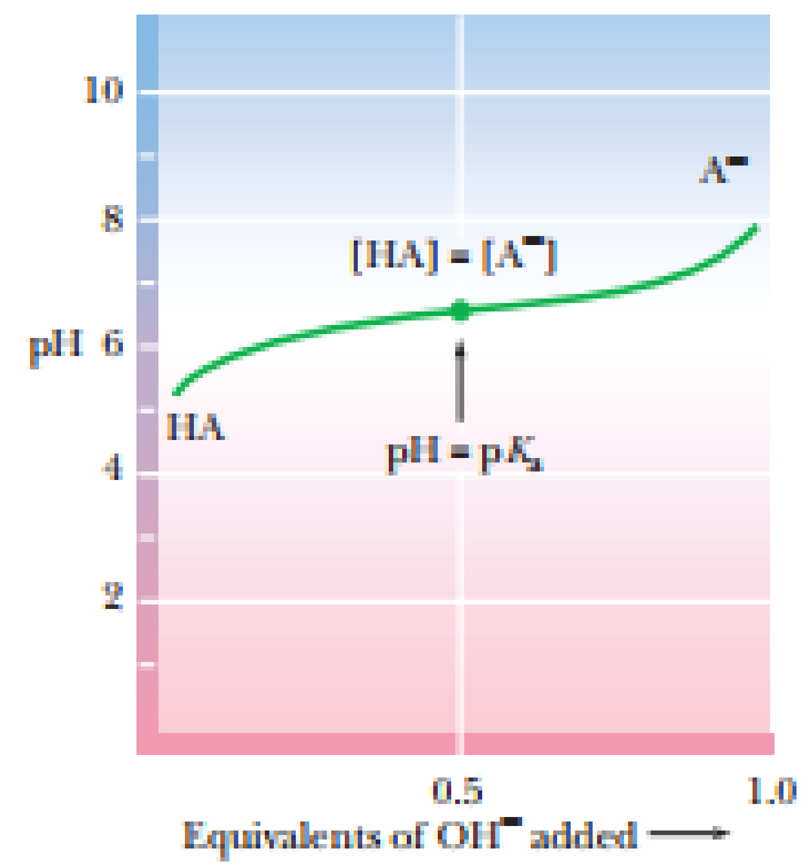
$$\text{NaOH} \rightarrow Na^+ + OH^- \; [large\ decrase\ in\ [H^+]$$

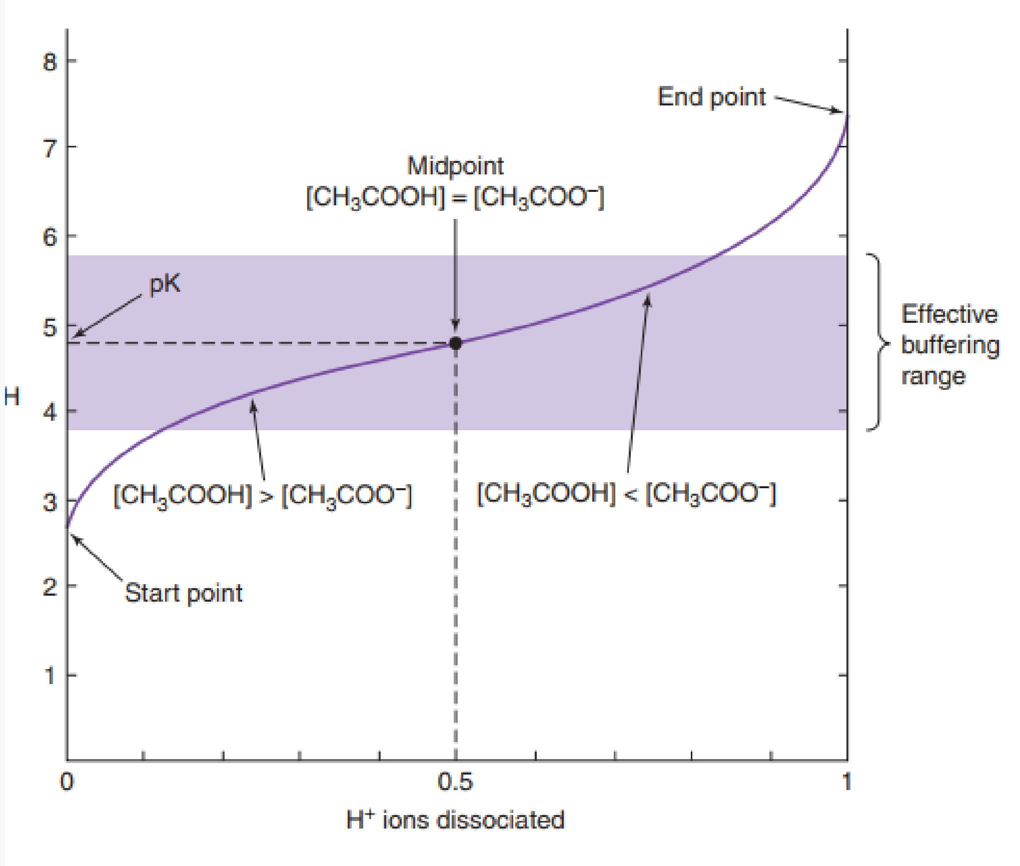
$$NaOH + HA \rightarrow A^- + H_2O \; [small\ decrease\ in\ [H^+]$$

Weak acid/conjugate bae system (HA/A^-) acts as aa **buffer** against adding acid or buffer by preventing dramatic changes in pH.

Tool and Techniques: Buffer

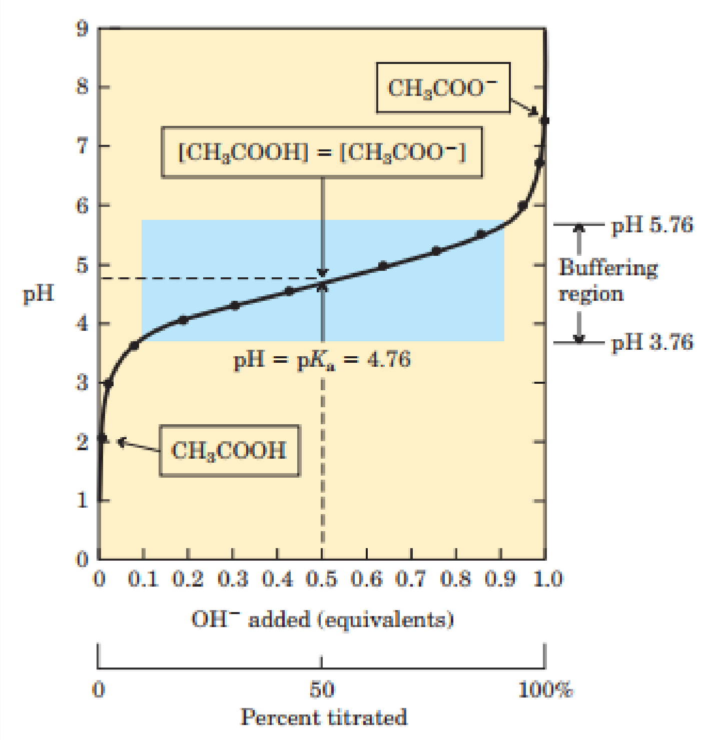
- At the midpoint, exactly half the protons have dissociated $[HA] = [A^-]$. So p$H = pK$
- pH does not change drastically with added acid or base when the pH is near the p*K*.

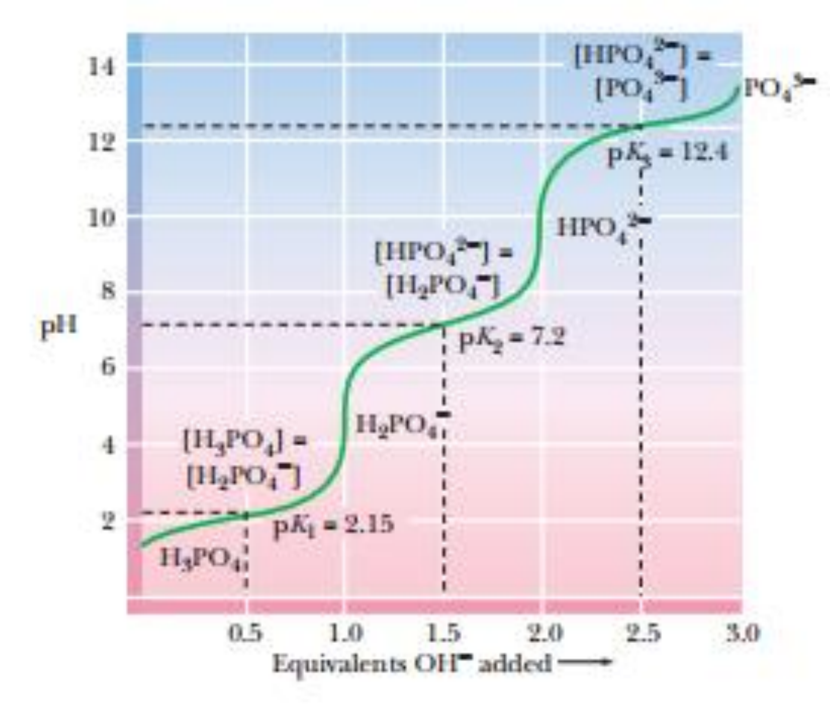




Tool and Techniques: Buffer

- The effective buffering capacity of an acid is generally taken to be within one pH unit of its pK. For acetic acid (pK= 4.76), this would be pH 3.76–5.76.
- Titration curve for phosphoric acid, H_3PO_4, three dissociable H^+ and discrete steps, each dissociate has a characteristic pK_a





Example 2.6

Acetic acid has a pK_a of 4.8. How many milliliters of 0.1 M acetic acid and 0.1 M sodium acetate are required to prepare 1 liter of 0.1 M buffer solution having a pH of 5.8?

Thank you very much for your attention