

Acids, Bases, Water & Buffers

- Weak acids and bases
- Strong acids , Conc. of acids
- The dissociation Constant K_a and pK_a
- Water, The biological Solvent:-
 - . Biological Roles of Water
 - . The structure of water
 - . Non-covalent interactions in biomolecules
 - . The importance of Hydrogen bond
 - . physical properties of water
 - . Ionization of water
- The Henderson - Hasselbalch Equation
 - . The pH of a solution is the negative log of its $[H^+]$
 - . Relationship between pH, pK_a and extent of acid dissociation
 - . Monoprotic, diprotic and polyprotic acids
 - . Selection of calculations on $[H^+]$, pH, K_a and pK_a

- The Concept of Buffers
 - Buffers composition
 - How do buffers work
 - Buffers capacity
 - How do we choose the proper buffer to use
 - How do we make a buffer in the laboratory & its calculation

- Normal metabolism generates metabolic acids, inorganic acids and CO_2 :-

Volatile acids

Non-volatile acids

- Physiological buffers:-

Bicarbonate, Phosphate & Proteins

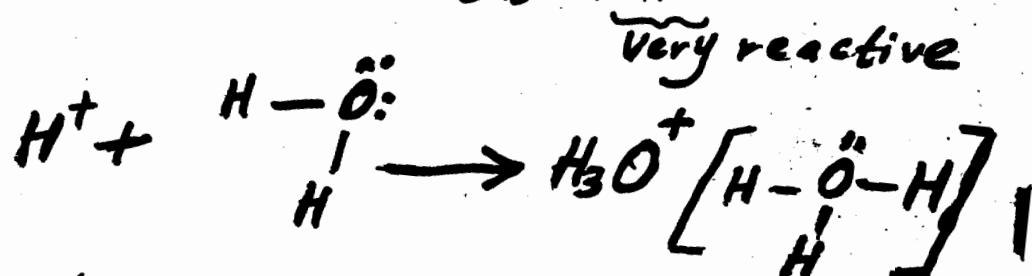
- Mechanism of action of physiological buffers

- Acid - Base Disturbances

ACIDS & BASES

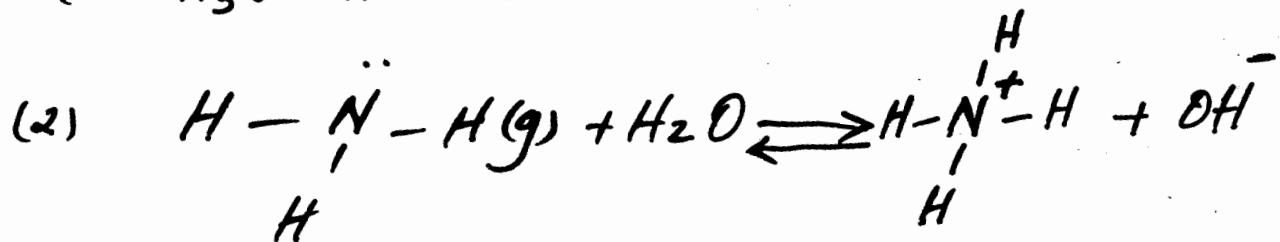
- Arrhenius Definitions of Acids & Bases

- Acids in $H_2O \rightarrow H^+$
- Bases : : $\rightarrow OH^-$
- Neutralization of acid & base \rightarrow salt + H_2O



Drawback

(1) H_3O^+ not H is released



(3) reaction is necessary in aq.

e.g. $NH_3(g) + HCl(g)$

Common Acids & Bases

Acids:- H_2SO_4 , HCl , H_3PO_4 & HNO_3 ; CH_3COOH

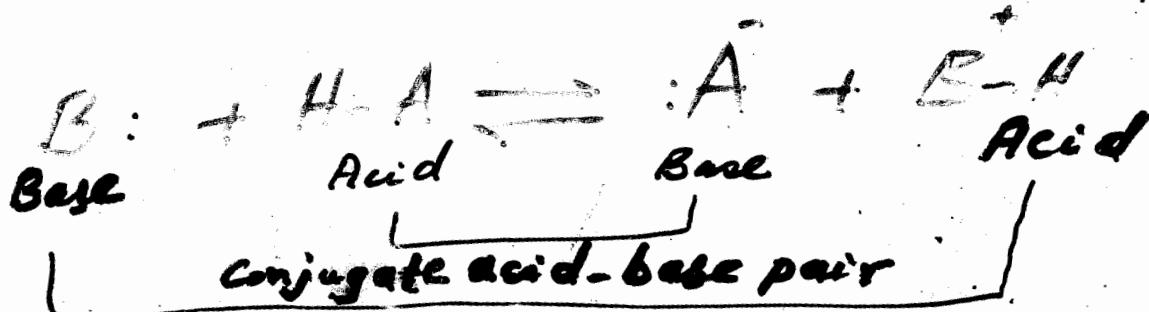
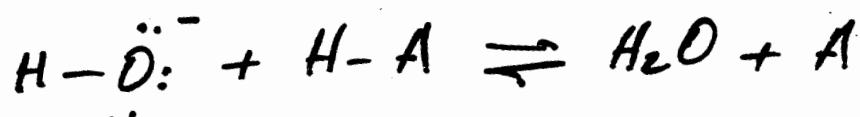
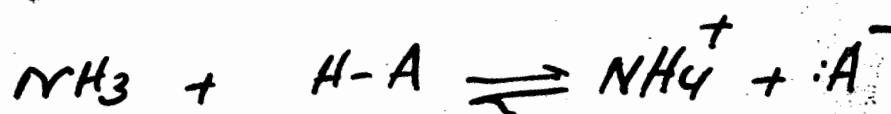
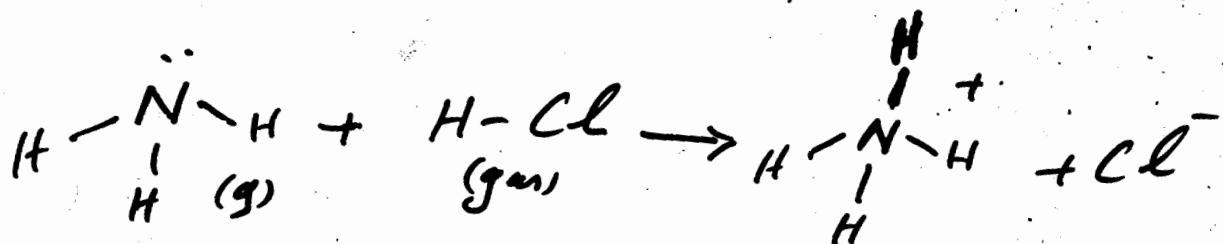
Bases:- Metal hydroxides:- $NaOH$, $Ca(OH)_2$, $Mg(OH)_2$, Ammonium (NH_4)

Bronsted-Lowry Definition of Acids & Bases

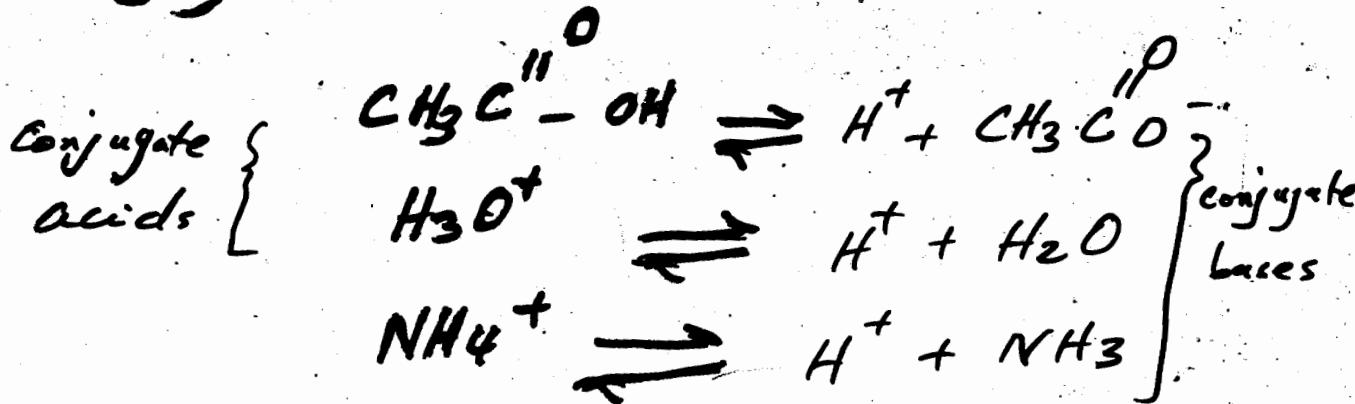
Acid : proton donor

Base : Proton acceptor

should have at least one non-bonding electron pair



Conjugate acid-base pairs:



Water as both an Acid and a Base:

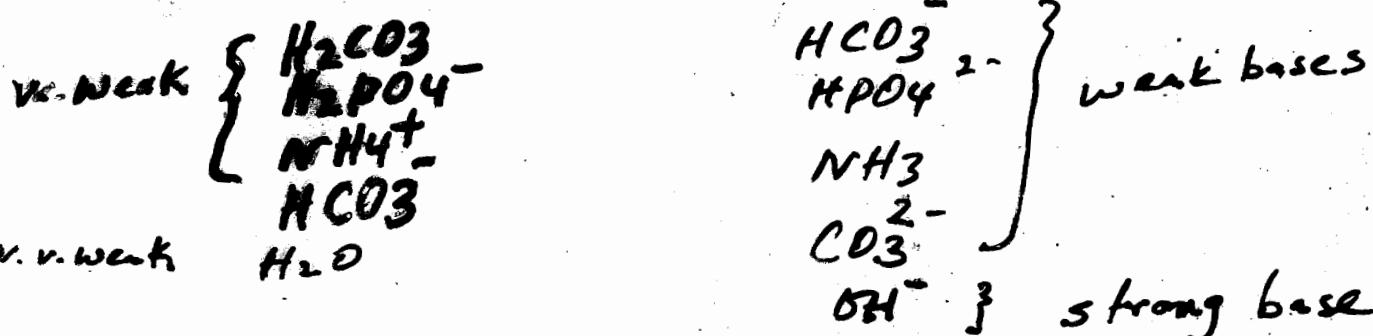
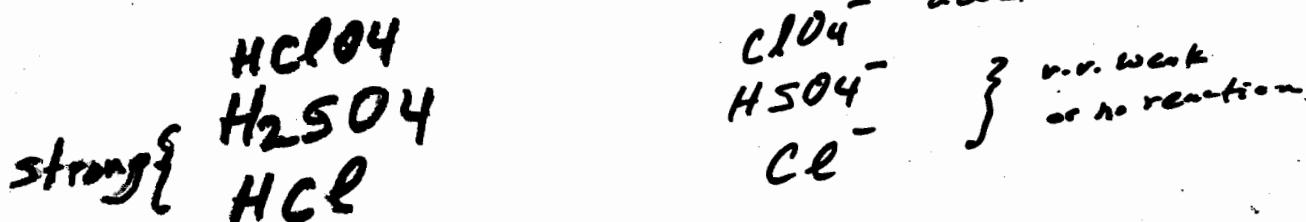
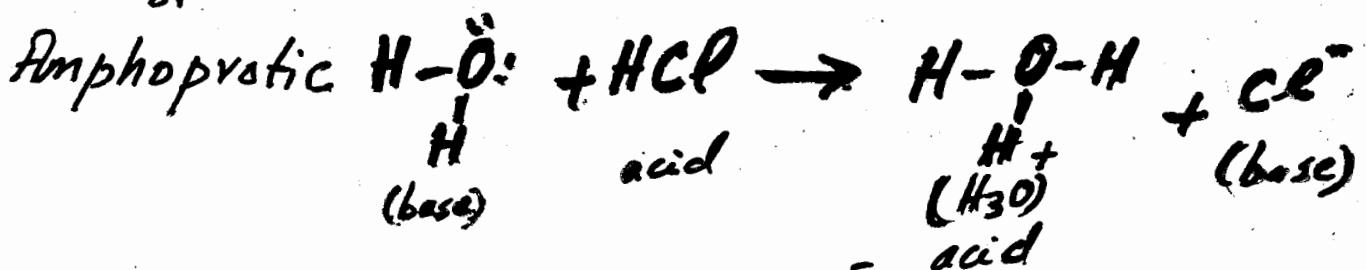
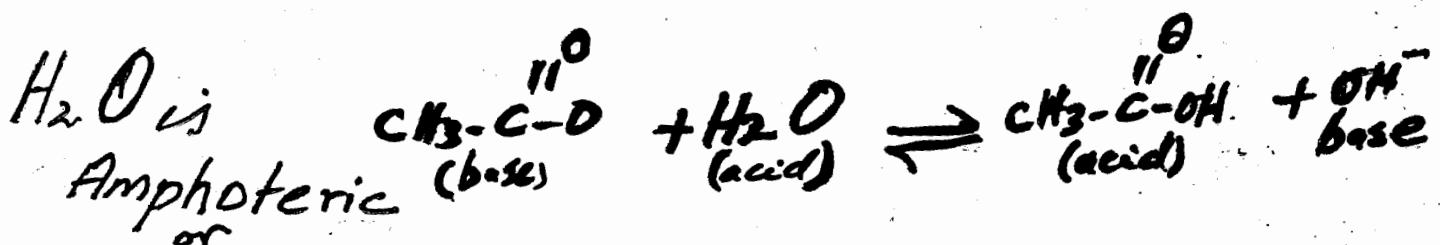
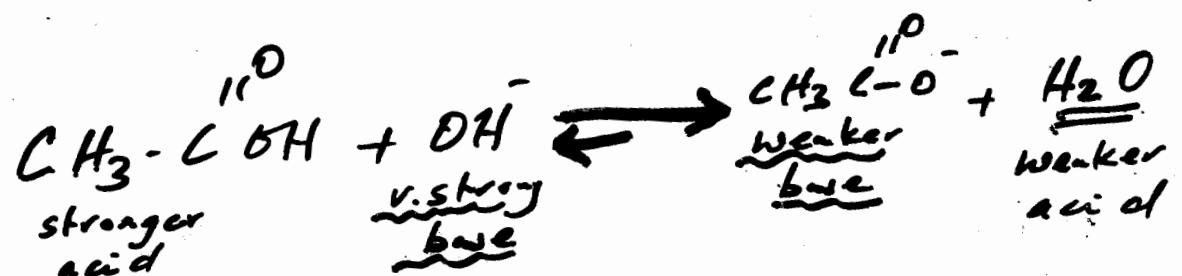
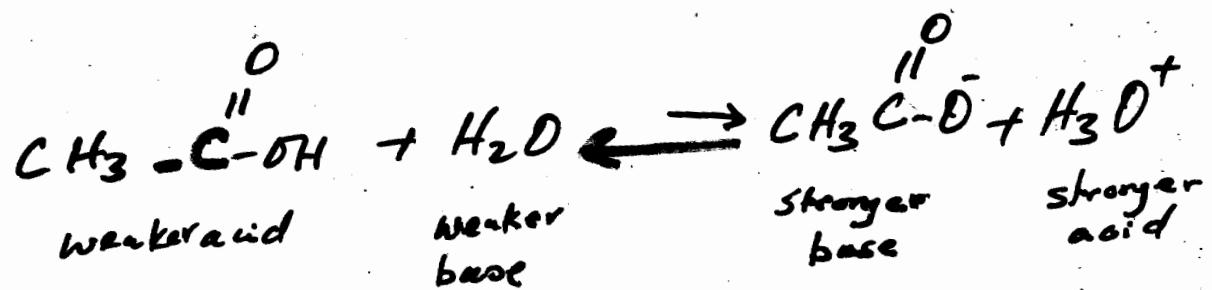
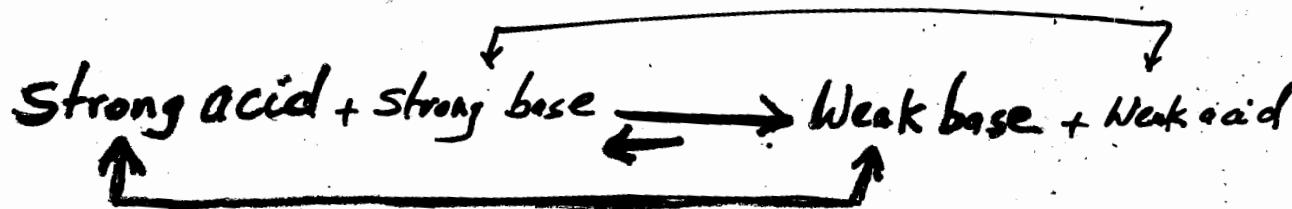
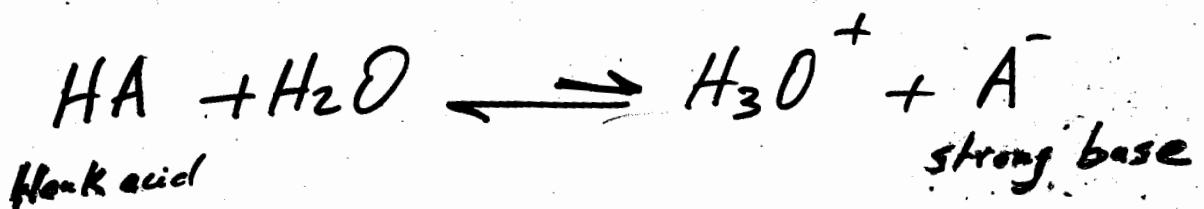
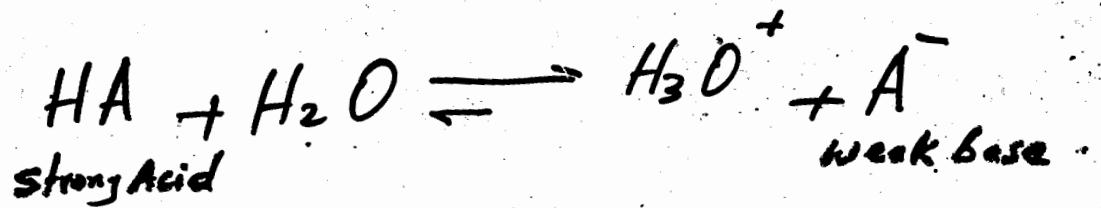


Table 10.1 Relative strengths of acids and conjugate bases

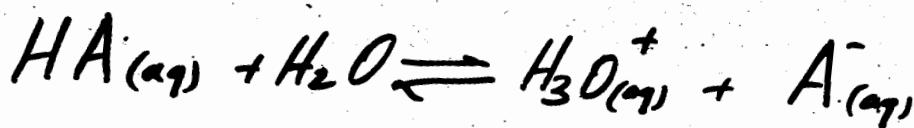
TABLE 10.1 Relative Strengths of Acids and Conjugate Bases

	ACID	CONJUGATE BASE			
Increasing acid strength					Increasing base strength
Strong acids: 100% dissociated	Perchloric acid Sulfuric acid Hydriodic acid Hydrobromic acid Hydrochloric acid Nitric acid	HClO_4 H_2SO_4 HI HBr HCl HNO_3	ClO_4^- HSO_4^- I^- Br^- Cl^- NO_3^-	Perchlorate ion Hydrogen sulfate ion Iodide ion Bromide ion Chloride ion Nitrate ion	Little or no reaction as bases
	Hydronium ion	H_3O^+	H_2O	Water	
Weak acids	Hydrogen sulfate ion Phosphoric acid Nitrous acid Hydrofluoric acid Acetic acid	HSO_4^- H_3PO_4 HNO_2 HF CH_3COOH	SO_4^{2-} H_2PO_4^- NO_2^- F^- CH_3COO^-	Sulfate ion Dihydrogen phosphate ion Nitrite ion Fluoride ion Acetate ion	Very weak bases
Very weak acids	Carbonic acid Dihydrogen phosphate ion Ammonium ion Hydrocyanic acid Bicarbonate ion Hydrogen phosphate ion	H_2CO_3 H_2PO_4^- NH_4^+ HCN HCO_3^- HPO_4^{2-}	HCO_3^- HPO_4^{2-} NH_3 CN^- CO_3^{2-} PO_4^{3-}	Bicarbonate ion Hydrogen phosphate ion Ammonia Cyanide ion Carbonate ion Phosphate ion	Weak bases
	Water	H_2O	OH^-	Hydroxide ion	Strong base



Knowing the relative strength of different acids makes it possible to predict the direction of proton-transfer reactions.

Acid Dissociation Constant



$$K = \frac{[H_3O^+][A^-]}{[HA][H_2O]} \rightarrow 55.5 M$$

$$\text{Dissociation Constt.} = K_a = K [H_2O]$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

- Strong acids have large K_a , much greater than 1
 HNO_3 , HCl , H_2SO_4

- Weak acids have K_a much less than 1

NH_4^+ , CH_3COOH , $H_2CO_3 4.3 \times 10^{-7}$, $HCO_3^- 5.6 \times 10^{-11}$
 6×10^{-10} , $HF 3.5 \times 10^{-4}$, $HSO_4^- 1.3 \times 10^{-2}$

- Donation of each successive H^+ from Polyprotic acid is more difficult than the one before -
 K_a values becomes lower

- Most organic acid, containing $-CO_2H$ group have K_a values near 10^{-5}

T-91

Table 10.2 Some acid dissociation constants, K_a , at 25 °CTABLE 10.2 Some Acid Dissociation Constants, K_a , at 25 °C

ACID	K_a	ACID	K_a
Hydrofluoric acid (HF)	3.5×10^{-4}	<i>Polyprotic acids</i>	
Hydrocyanic acid (HCN)	4.9×10^{-10}	Sulfuric acid	
Ammonium ion (NH_4^+)	5.6×10^{-10}	H_2SO_4	Large
<i>Organic acids</i>		HSO_4^-	1.2×10^{-2}
Formic acid (HCOOH)	1.8×10^{-4}	Phosphoric acid	
Acetic acid (CH_3COOH)	1.8×10^{-5}	H_3PO_4	7.5×10^{-3}
Propanoic acid ($\text{CH}_3\text{CH}_2\text{COOH}$)	1.3×10^{-5}	H_2PO_4^-	6.2×10^{-8}
Ascorbic acid (vitamin C)	7.9×10^{-5}	HPO_4^{2-}	2.2×10^{-13}
		Carbonic acid	
		H_2CO_3	4.3×10^{-7}
		HCO_3^-	5.6×10^{-11}

$$\text{p}K_a = -\log K_a$$

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

WATER or PH

- Simple and abundant
- Extraordinary physical, chemical and biological properties
- Vital to all forms of life
70% to 85% the wt. of typical cell

Biological Roles of Water

- Biological Solvent

- Water serves as an essential buffer to regulate Temp. and pH
High specific heat capacity
- Water is a participant in many biochemical reactions
 - Hydrolysis
 - photosynthesis
- $$6 \text{CO}_2 + 6 \text{H}_2\text{O} \xrightarrow{\text{light}} \text{C}_6\text{H}_{12}\text{O}_6 + 6 \text{O}_2$$
- Respiration - aerobic metabolism
- others

Noncovalent Interactions in Biomolecules

- Ionic bonds
between oppositely charged atoms or groups
Energy 20-30 kJ/mole
- Hydrogen bonds
between H atom linked to electronegative atom (O, N or F) and electronegative atom
10 - 30 kJ/mole
- van der Waals interactions
1-5 kJ/mole
- Hydrophobic interactions
5-30 kJ/mole

Characteristics of Non covalent Interactions:

- relatively weak
1-30 kJ/mole as compared to 350 kJ/mole in C-C
- Reversible
- Binding between molecules
is specific

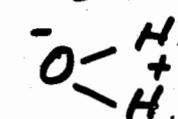
Weak Interactions in an Aqueous Environment

The nature of non-covalent interactions (Essentially Electrostatics)

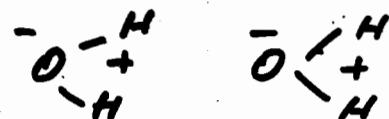
- Charge - Charge. $-NH_3^+$



- Charge - dipole $-NH_3^+$



- Dipole - dipole



- charge - induced dipole $-NH_3^+$

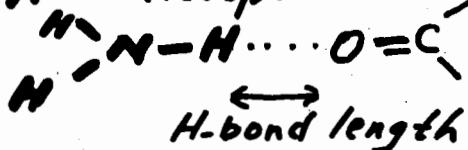


- Dipole - induced dipole

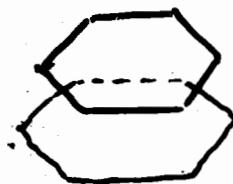
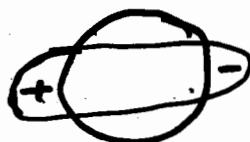


- Dispersion

- Hydrogen bond Donor-H.....Acceptor

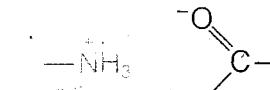
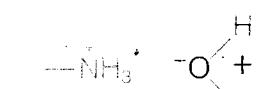
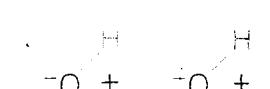
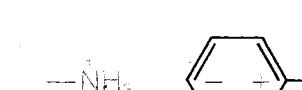
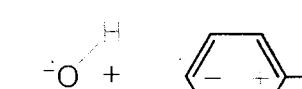
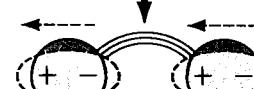
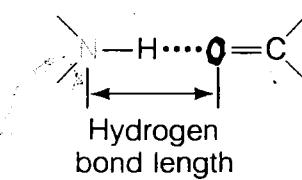


- Dispersion



Types of noncovalent interactions

Figure 2.1

TYPE OF INTERACTION	MODEL	EXAMPLE
(a) Charge-charge		
(b) Charge-dipole		
(c) Dipole-dipole		
(d) Charge-induced dipole		
(e) Dipole-induced dipole		
(f) Dispersion	 	
(g) Hydrogen bond	Donor—H....Acceptor	

From Mathews and van Holde: *Biochemistry*
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Structure of Water

• electronegativity of O atom is 3.5
 • H = 2.1

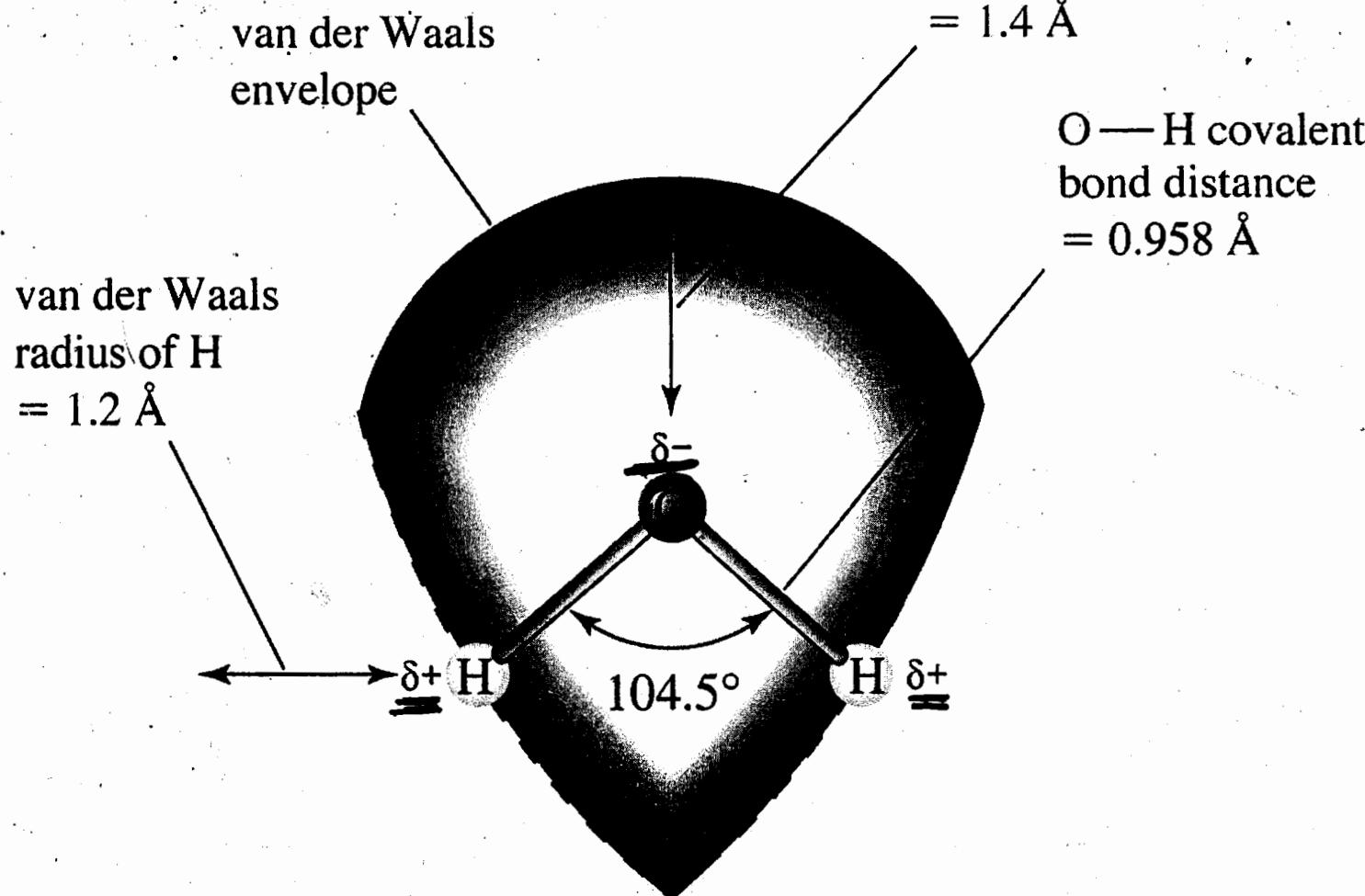


Figure 2-1a Concepts in Biochemistry, 3/e
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Water is electrically neutral (no net charge) but has relatively large dipole moment because of its bent geometry

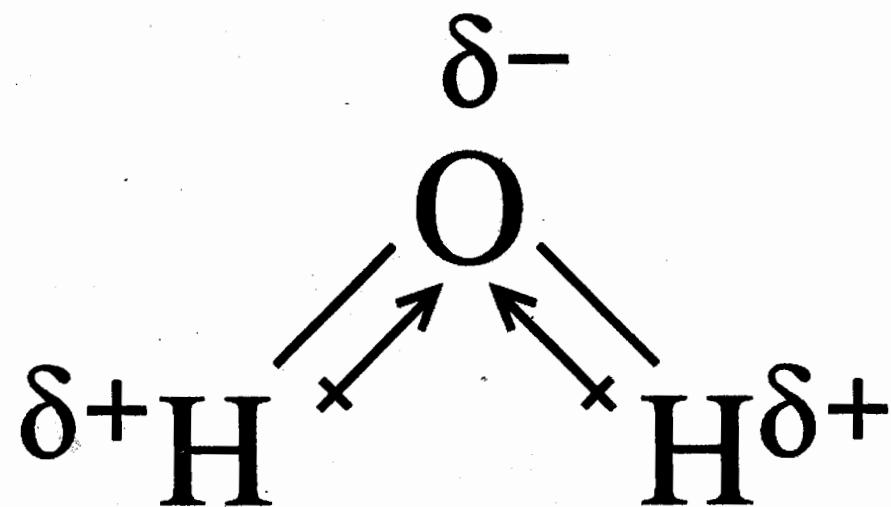
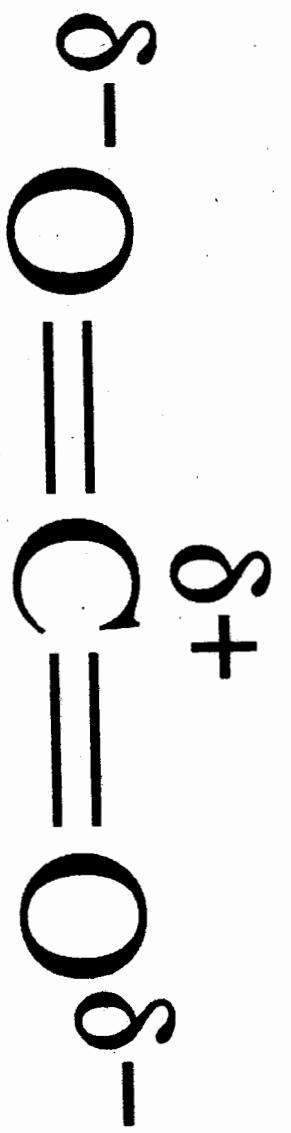


Figure 2-1b Concepts in Biochemistry, 3/e
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C₂O₂ has polar bonds caused by electronegativity between C and O atoms but no dipole moment because it is linear



Hydrogen bond between two water molecules

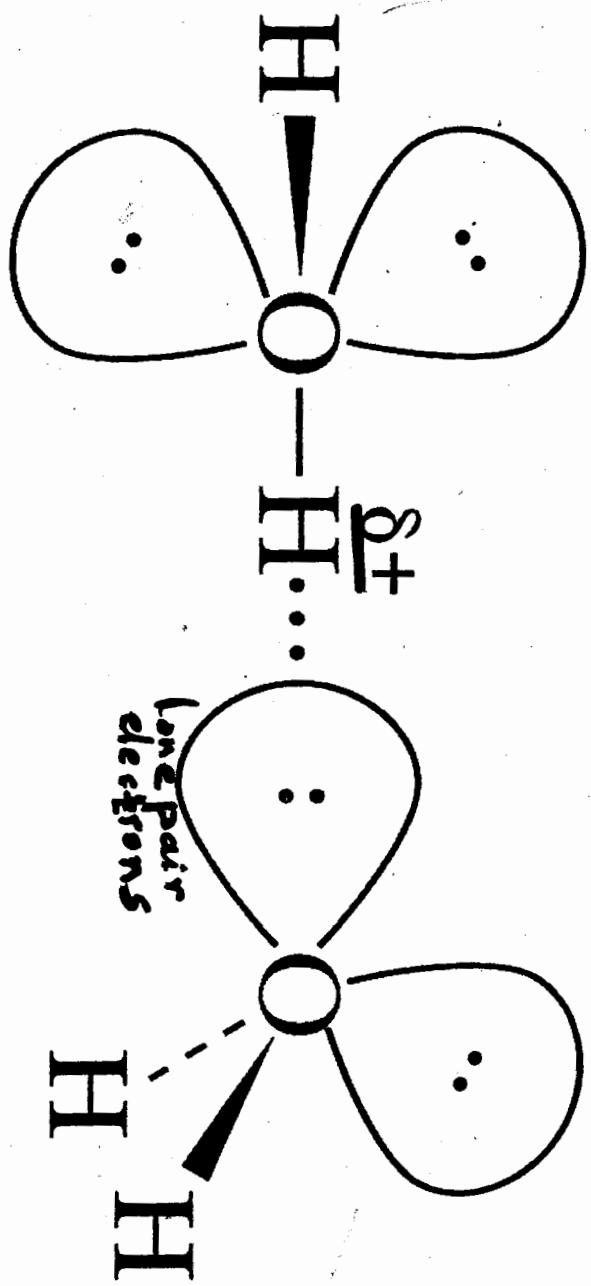
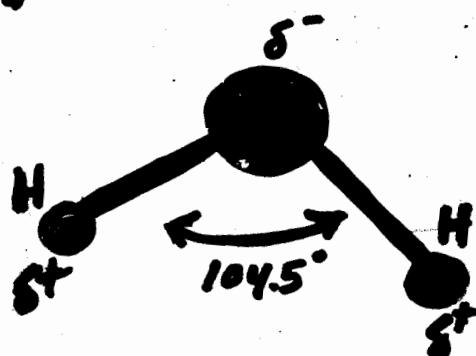


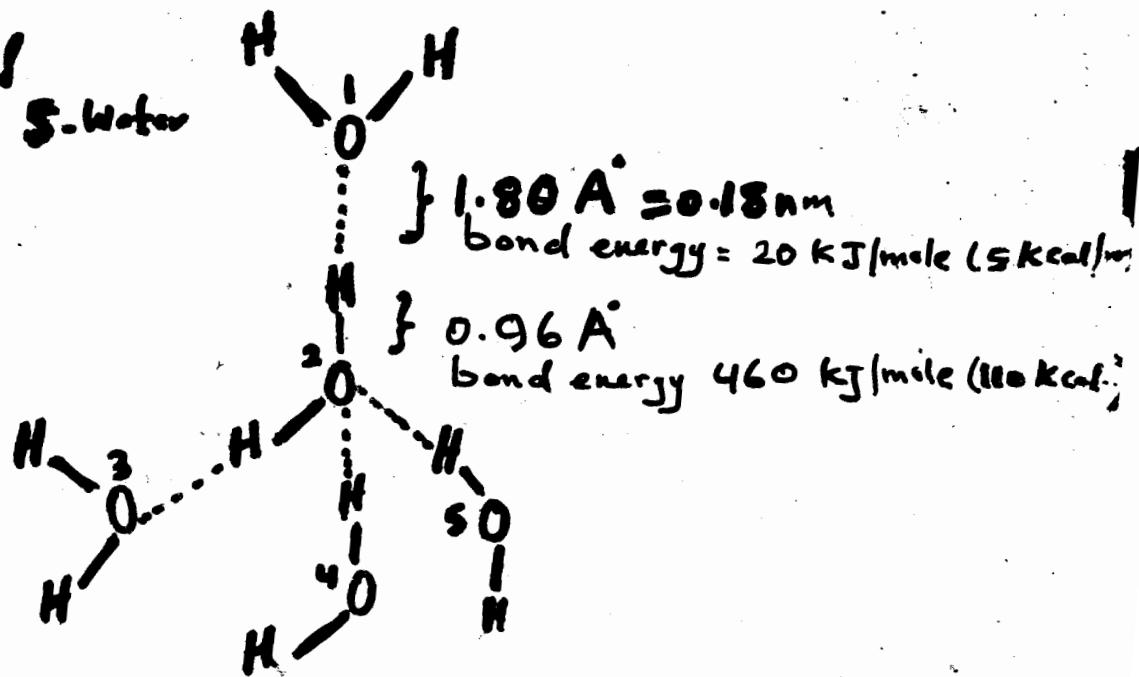
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Hydrogen Bonds Between Water Molecules

The H-O-H bond angle



Tetrahedral H-bonding of S. Water molecules



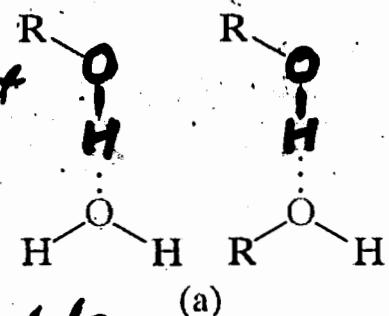
Hydrogen bond is strongest when the three atoms $X-\text{H}-\cdots A$ are linear.

A can be oxygen, nitrogen or fluorine

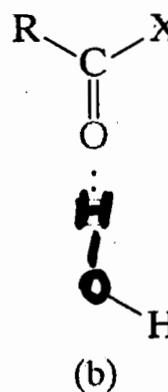
X can be oxygen, nitrogen or fluorine

I - Many biomolecules have atoms that can hydrogen bond with water, themselves & other molecules

Functional groups that participate in H-bonding include:



(a)

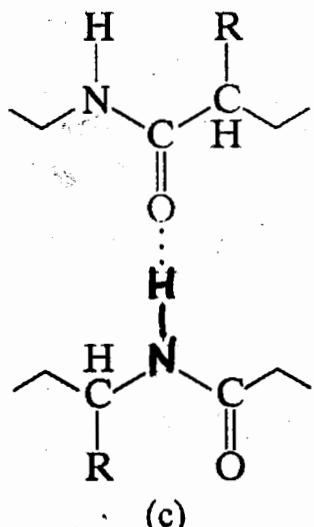


(b)

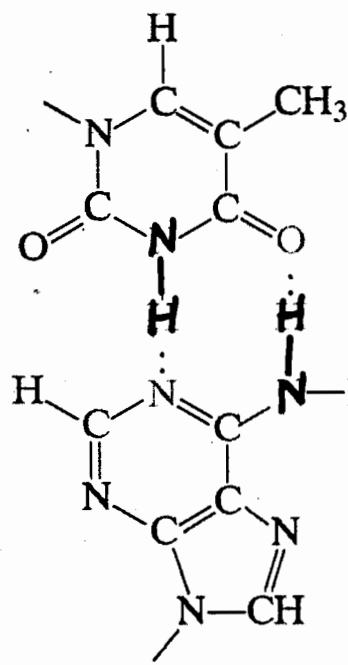
- $-\text{OH}$ gr in alcohols, org. acids & carbohydrate

- Carbonyl groups in aldehydes, Ketones, acids, amides & esters

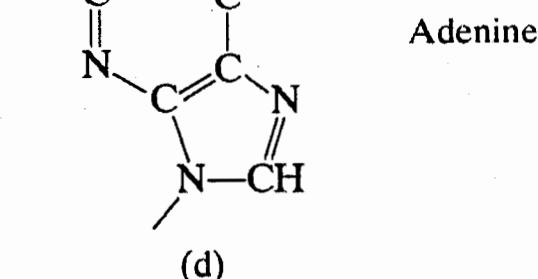
- N-H groups in Amines & Amides



(c)



Thymine



Adenine

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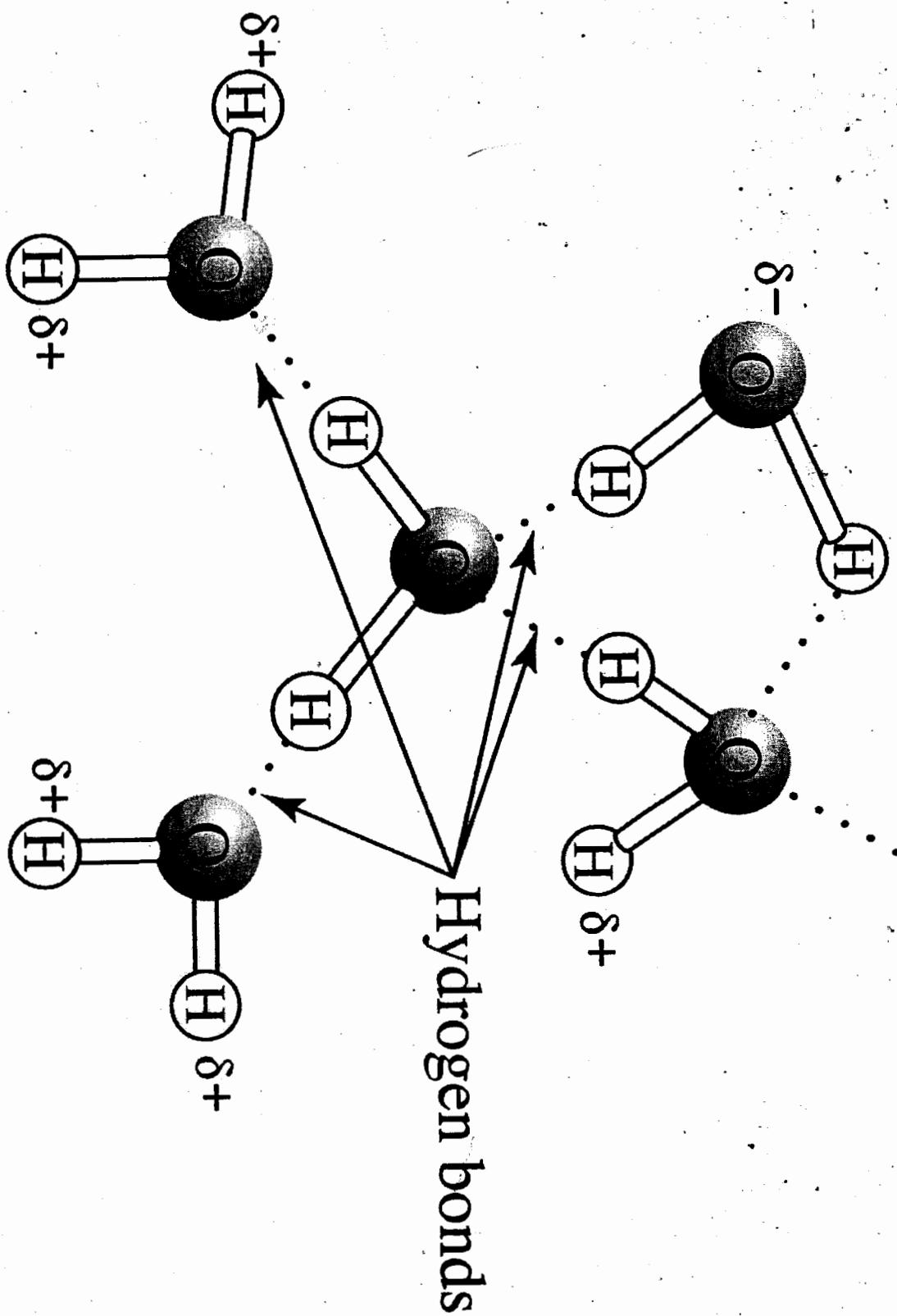


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The network of Potential H-bonds in Water

- Average number of H-bonds to each molecule in liquid water at 10°C is ~ 3.0
- Number of H-bonds decrease with increasing temp.

In crystalline ice, the number approaches four

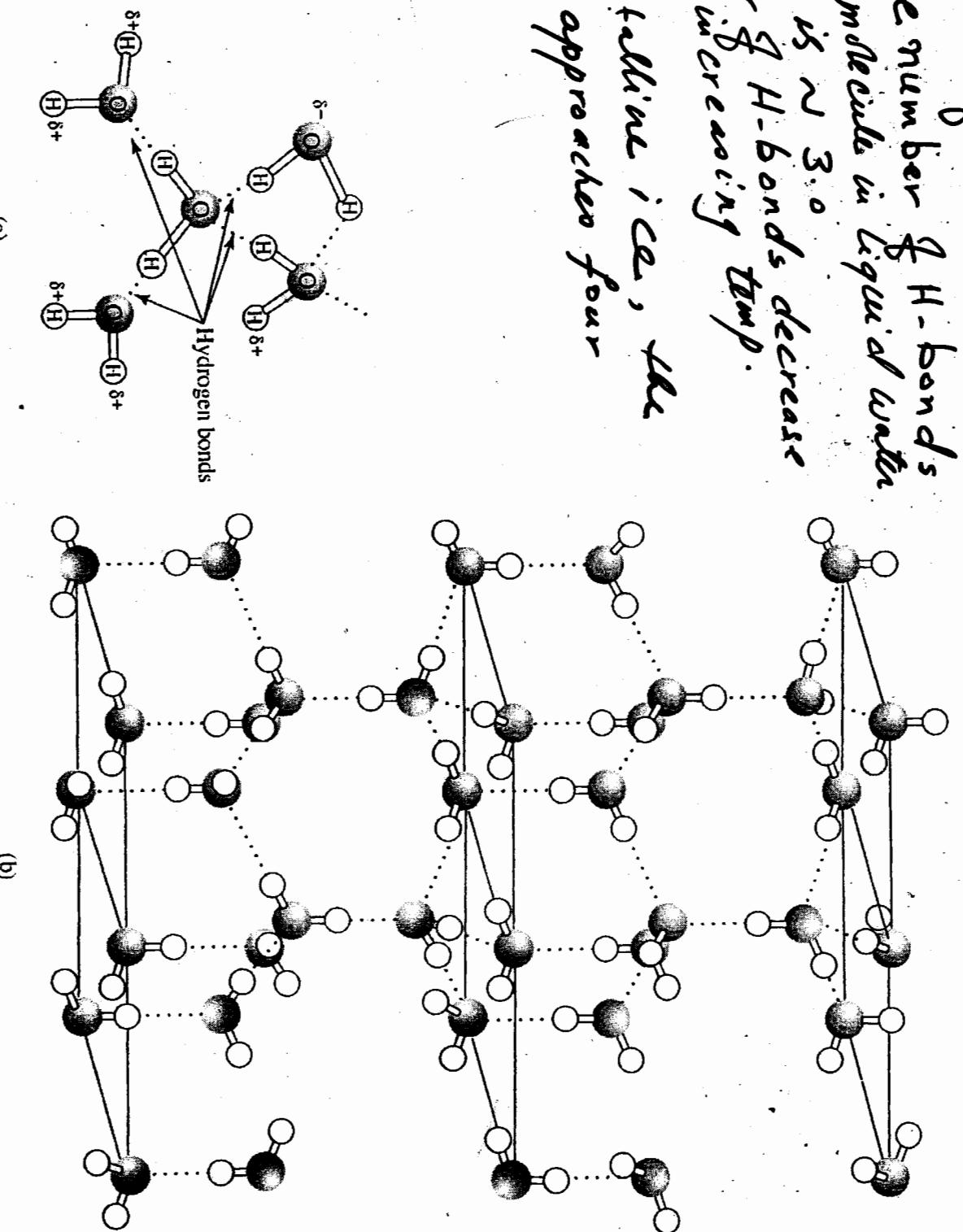


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Table 2.3
A comparison of some physical properties of water with hydrides of other nonmetallic elements:
N, C, and S

Property	H ₂ O	NH ₃	CH ₄	H ₂ S
Molecular weight	18	17	16	34
Boiling point (°C)	100	-33	-161	-60.7
Freezing point (°C)	0	-78	-183	-85.5
Viscosity ^a	1.01	0.25	0.10	0.15

^aUnits are centipoise.

Table 2-3 Concepts in Biochemistry, 3/e
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Physical Properties of Water:- 13

- Hydrogen bonding Gives Water its unusual Properties

- higher m.P.; B.P.; heat of vaporization,
higher freezing, surface tension

H-O-H bond angle is 104.5°

Bond energy of H-bond is 20 kJ/mole

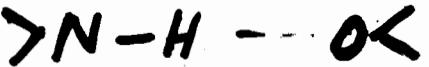
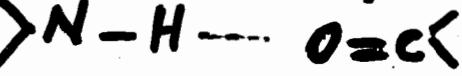
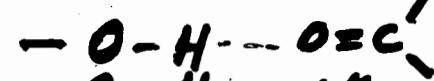
life-time $1 \times 10^9 \text{ s}$ O-H 460 kJ/mole bond energy

in liquid state each water molecule forms hydrogen bond with another 3-4 H_2O molecules

In Ice - 4 H_2O molecules
larger vol. & less dense the ice-lattice

- Water forms Hydrogen bonds with solutes
H-bonds are not unique only to Water

Hydrogen atoms covalently bonded to carbon atoms, which are not electronegative, do not participate in H-bonding



e.g.

- B.P. for butanol = 117°C



- B.P. for butane = -0.5°C



Water as a Solvent
Important Solvent and Transporter

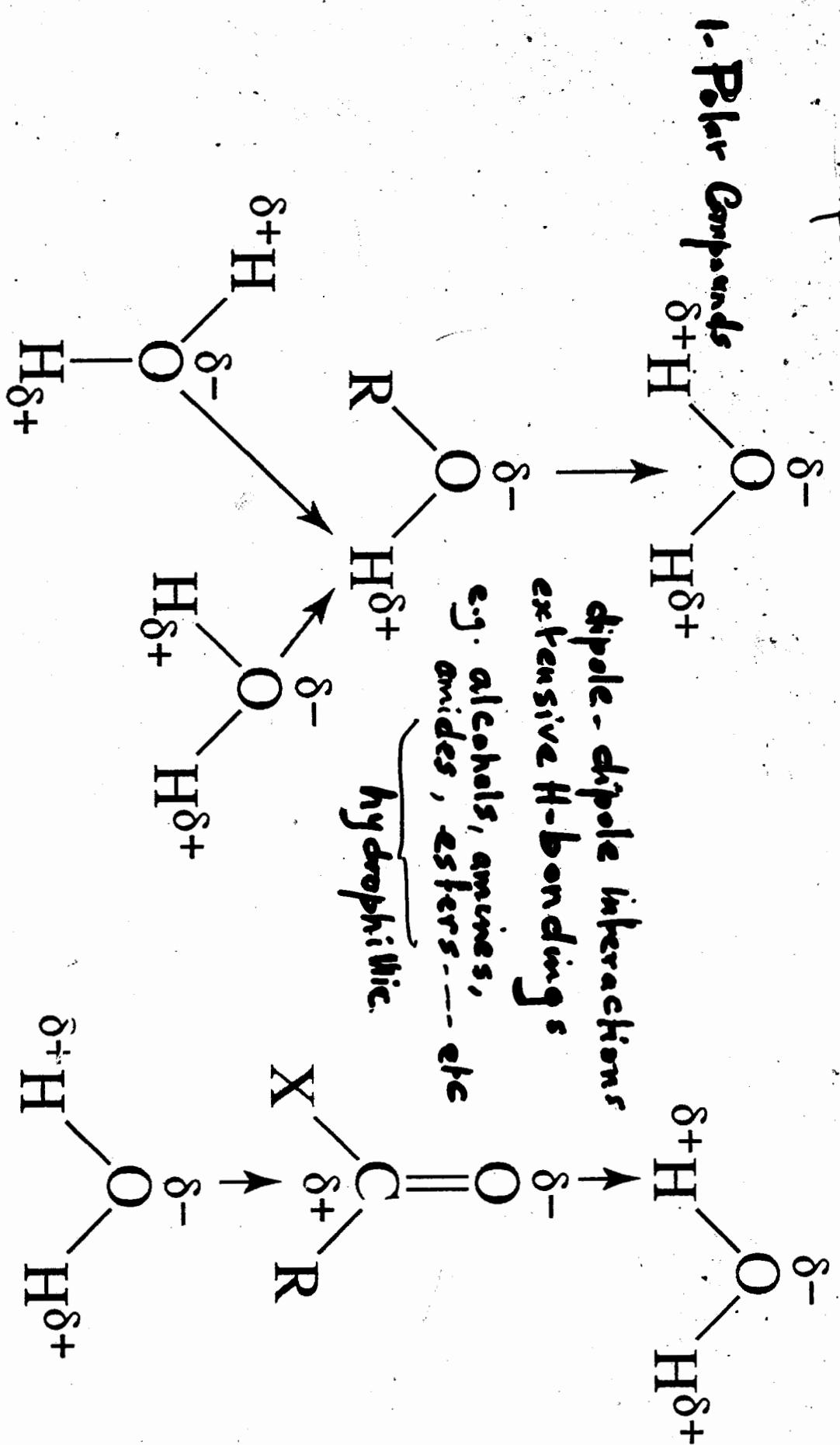


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2. Ionic Compounds

Individual ions are hydrated (solvated) by polar water molecules

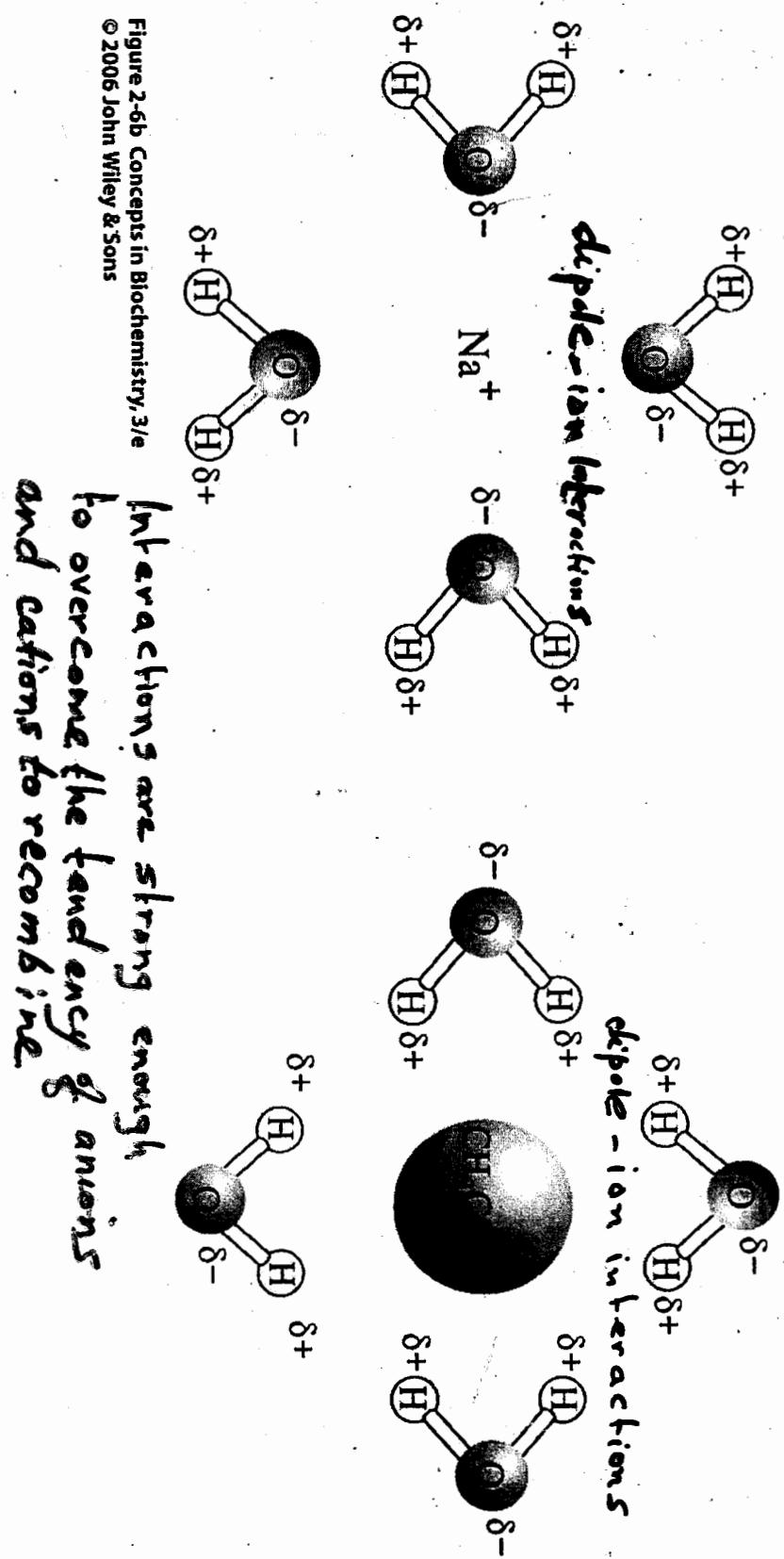
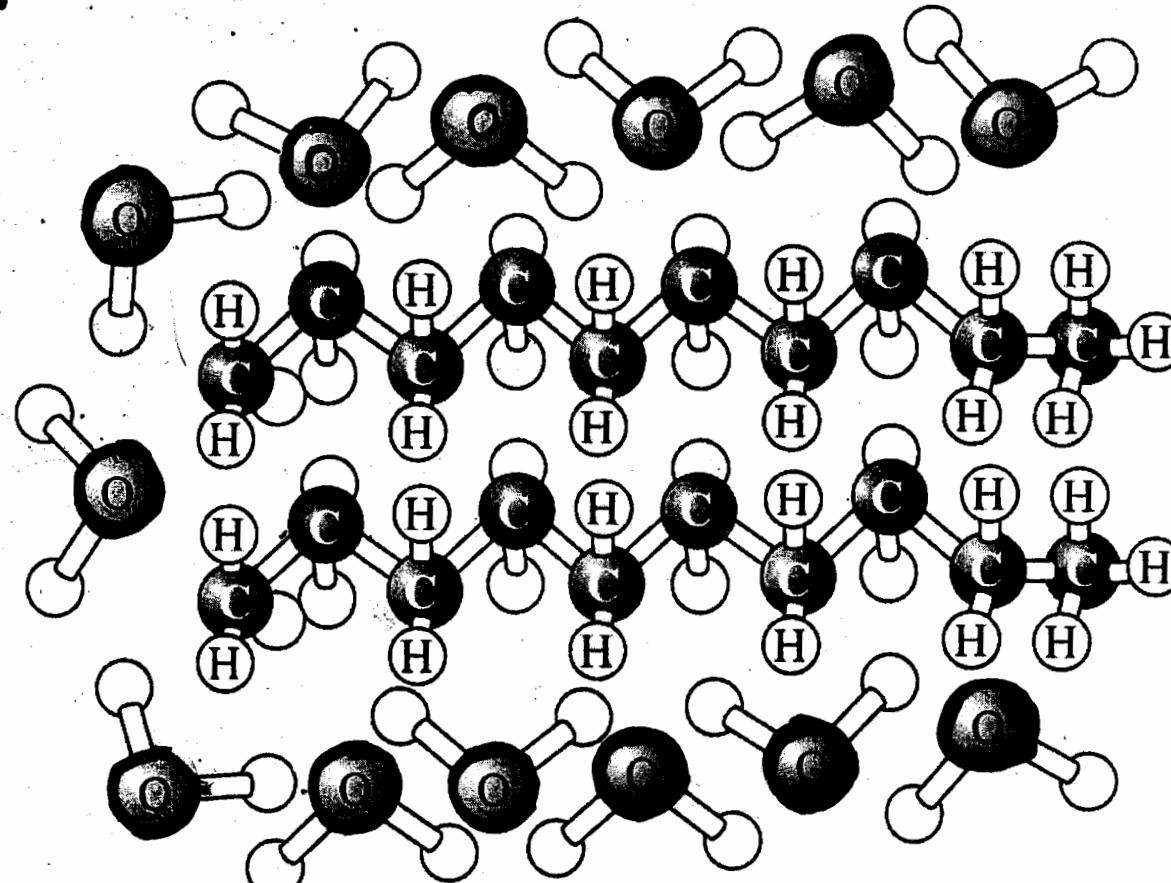


Figure 2-6b Concepts in Biochemistry, 3/e
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Interactions are strong enough to overcome the tendency of anions and cations to recombine.

3- Nonpolar Compounds :-
They do not contain ions or polar functional groups - Hydrophobic.
e.g. decane, benzene...etc 16
. amphiphilic e.g. sodium stearate



Water molecules
in cage around
hydrocarbon chain

Formation of this
highly ordered cage
of water requires
much energy, which
comes from hydrophobic
interactions

Figure 2-7 Concepts in Biochemistry, 3/e
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Amphiphilic Molecules

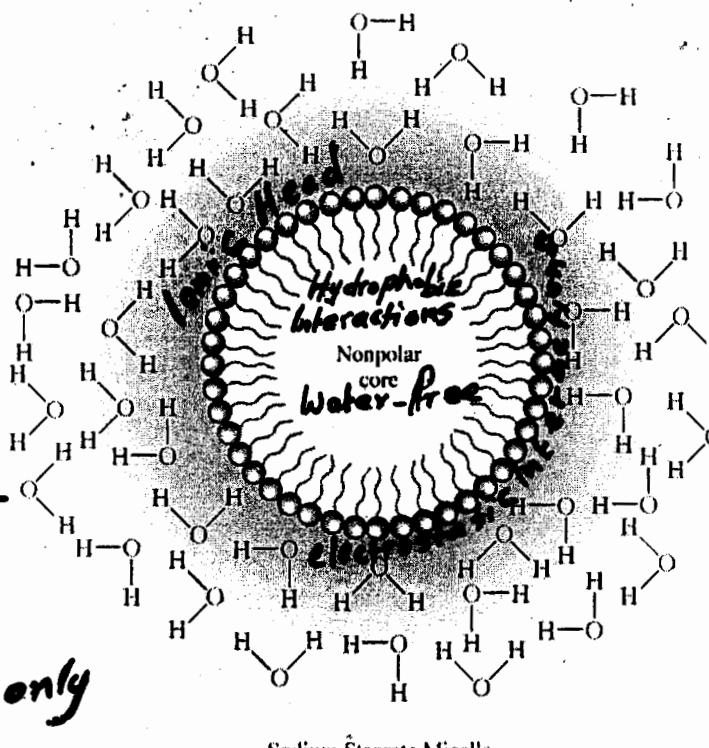
- Soap action

- Changes in water structure by solutes

- Changes in solutes structure by water

- Nucleic acids, proteins and some lipids are amphiphilic

- Ordered arrangement only are often associated with biological activity

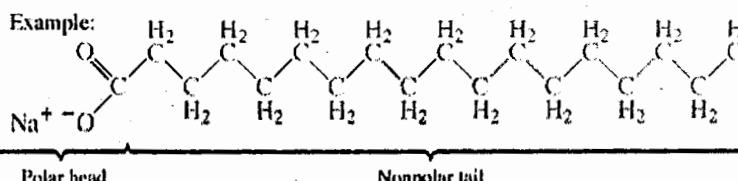


Sodium Stearate Micelle

Key: Polar head of sodium stearate



Nonpolar tail of sodium stearate



Amphiphilic Compound

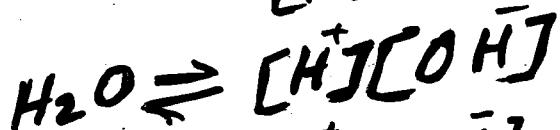
Figure 2-8 Concepts in Biochemistry, 3/e
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Ionization of Water, Weak acids & Weak bases:-

Ionization of $H_2O \rightarrow$ very small degree

Equilibrium constt.
 $A + B \rightleftharpoons C + D$

$$K_{eq} = \frac{[C][D]}{[A][B]}$$



$$K_{eq} = \frac{[H^+][OH^-]}{[H_2O]} = 55.5 M$$

$$K_{eq} \times 55.5 = [H^+][OH^-] = K_w \quad \begin{matrix} \text{ion product} \\ \text{constant for water} \end{matrix}$$

$$K_w = 55.5 \times (1.8 \times 10^{-16} M) = 1.0 \times 10^{-14} M^2 \quad \begin{matrix} \text{true for} \\ \text{pure water} \end{matrix}$$

In Pure water $\rightarrow [H^+] = [OH^-] = 10^{-7} M$ and Solutions

e.g. conc. of $[H^+]$ in a solution of 0.1M NaOH

$$K_w = [H^+][OH^-] = 10^{-1} \times [H^+] = 10^{-14}$$

$$[H^+] = \frac{10^{-14}}{10^{-1}} = 10^{-13} M$$

e.g. 0.1M HCl

$$[OH^-] = \frac{10^{-14}}{10^{-1}} = 10^{-13} M$$

I.

Acidic solution $[\text{H}_3\text{O}^+] > 10^{-7}\text{M}$ or $[\text{OH}^-] < 10^{-7}$

Neutral solution $[\text{H}_3\text{O}^+] = 10^{-7} = [\text{OH}^-]$

Basic solution $[\text{H}_3\text{O}^+] < 10^{-7}$ or $[\text{OH}^-] > 10^{-7}$

pH scale: Measuring acidity in aqueous solutions

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$\text{or } [\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

$[\text{H}_3\text{O}^+]$ for coffee is 10^{-5}M

$$\text{pH} = -\log [10^{-5}] = 5$$

Lemon juice has a pH = 2

$$2 = -\log [\text{H}_3\text{O}^+]$$

$$[\text{H}_3\text{O}^+] = 10^{-2}$$

Detergent has $[\text{OH}^-] = 10^{-3}\text{M}$

$$[\text{H}_3\text{O}^+] = \frac{10^{-14}}{10^{-3}} = 10^{-11}\text{M}$$

$$\text{pH} = -\log [10^{-11}] = 11$$

The pH Scale:

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

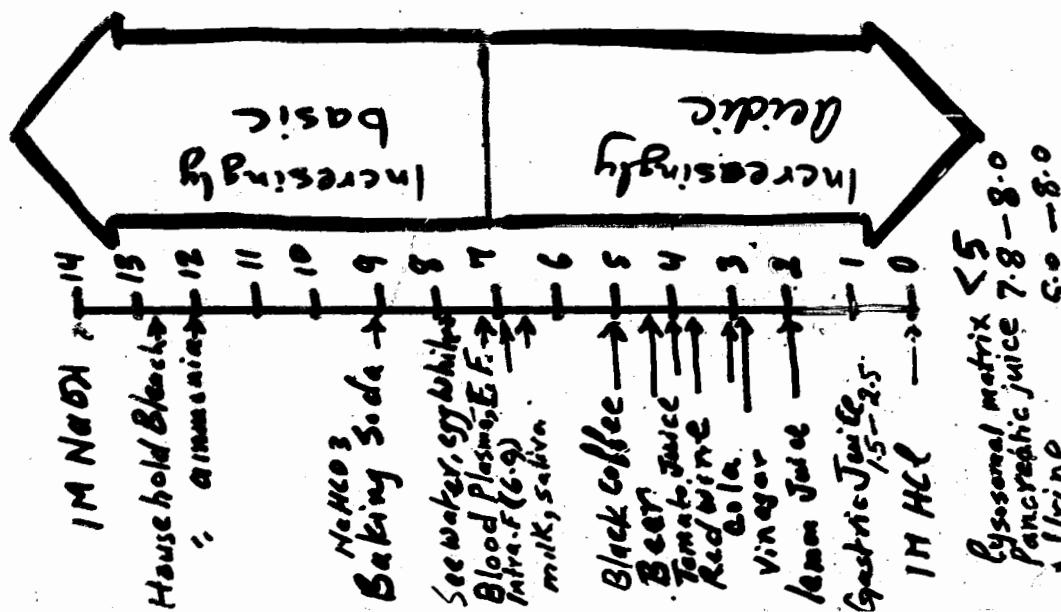
$$[H^+] \times [OH^-] = 10^{-14} \quad p[OH^-] = -\log [OH^-]$$

$$-\log [H^+] \times [OH^-] = -\log 10^{-14}$$

$$-\log [H^+] - \log [OH^-] = 14$$

$$pH + pOH = 14$$

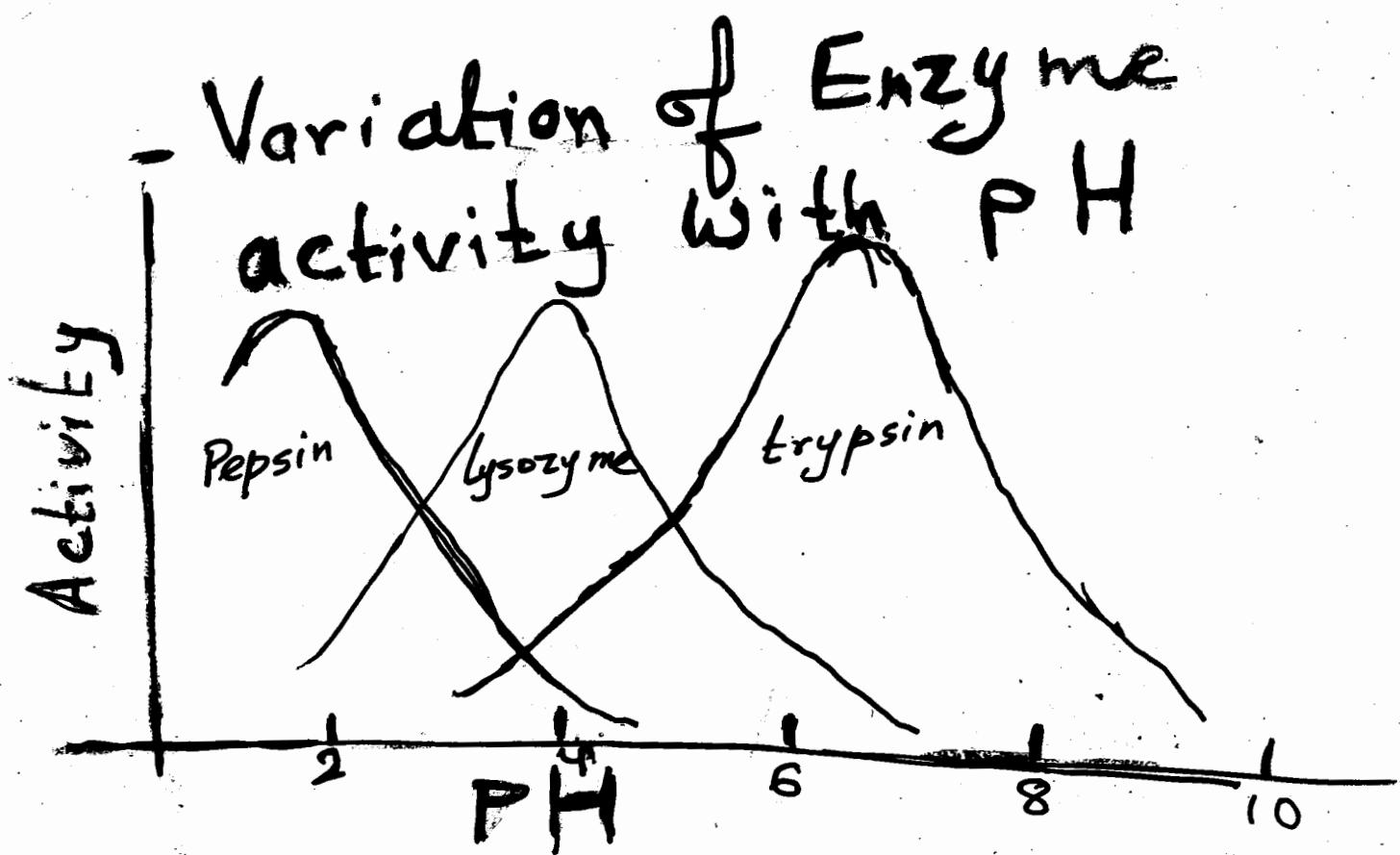
<u>$\frac{[H^+]}{M}$</u>	<u>pH</u>	<u>$\frac{[OH^-]}{M}$</u>	<u>$\frac{pOH}{M}$</u>
10^0	0	10^{-14}	14
10^{-1}	1	10^{-13}	13
10^{-2}	2	10^{-12}	12
10^{-3}	3	10^{-11}	11
10^{-4}	4	10^{-10}	10
\downarrow		10^{-7}	7
10^{-7}	7	10^{-7}	7
\downarrow		10^{-2}	2
10^{-12}	12	10^{-2}	2
10^{-14}	14	1	0



pH of Some
Fluids

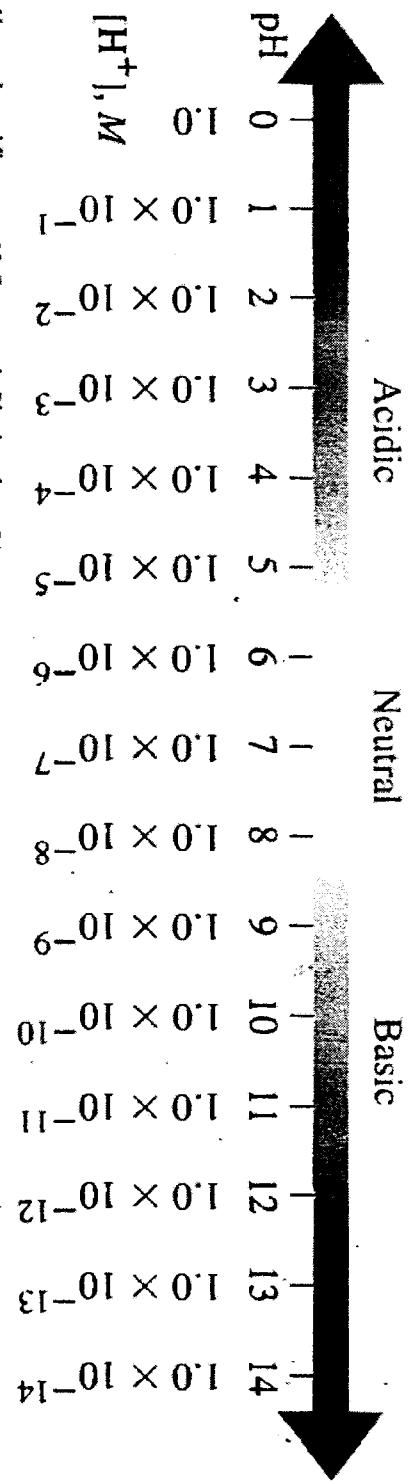
Importance of pH Regulation

- Narrow range of cellular pH in which living organism can function



4

pH scale



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- Note each digit increase or decrease represents a 10-fold change in $[H^+]$

The pH values of some substances
many natural fluids have pH around neutral pH of 7.0

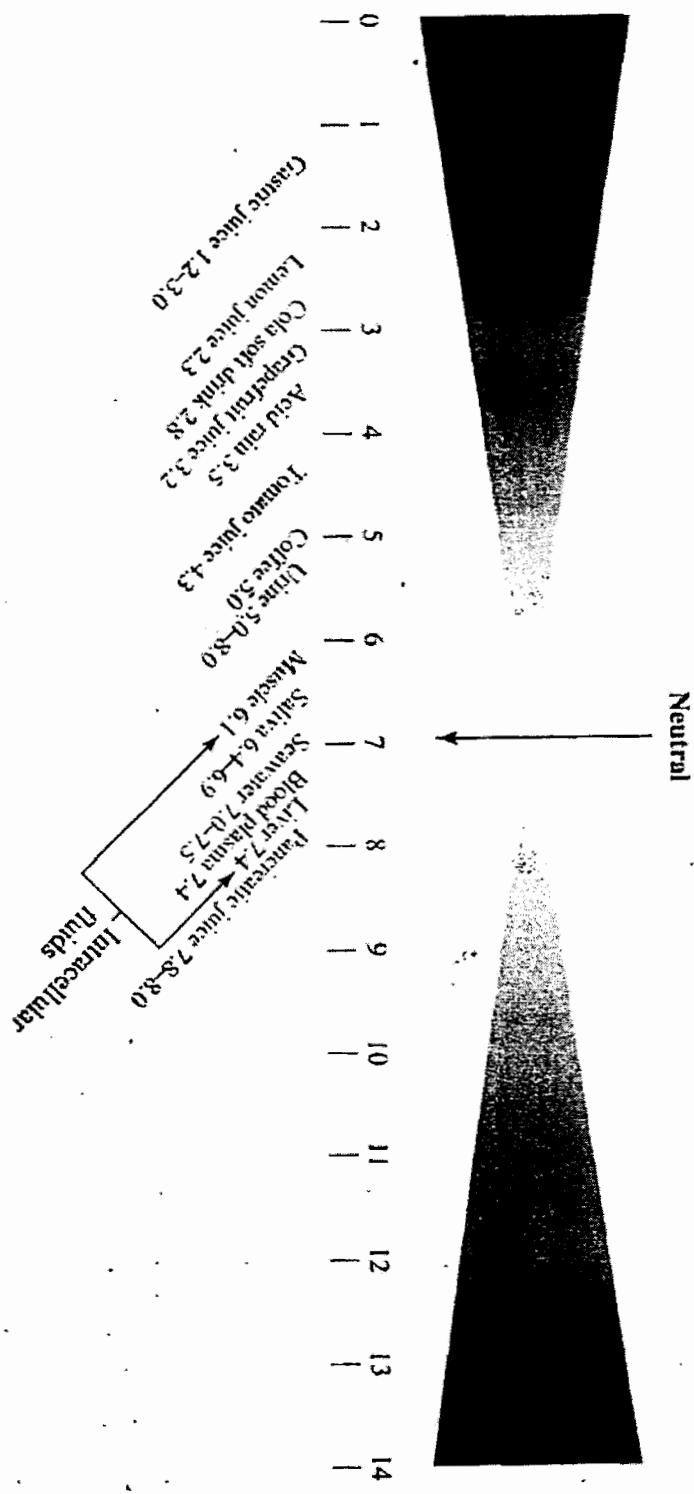
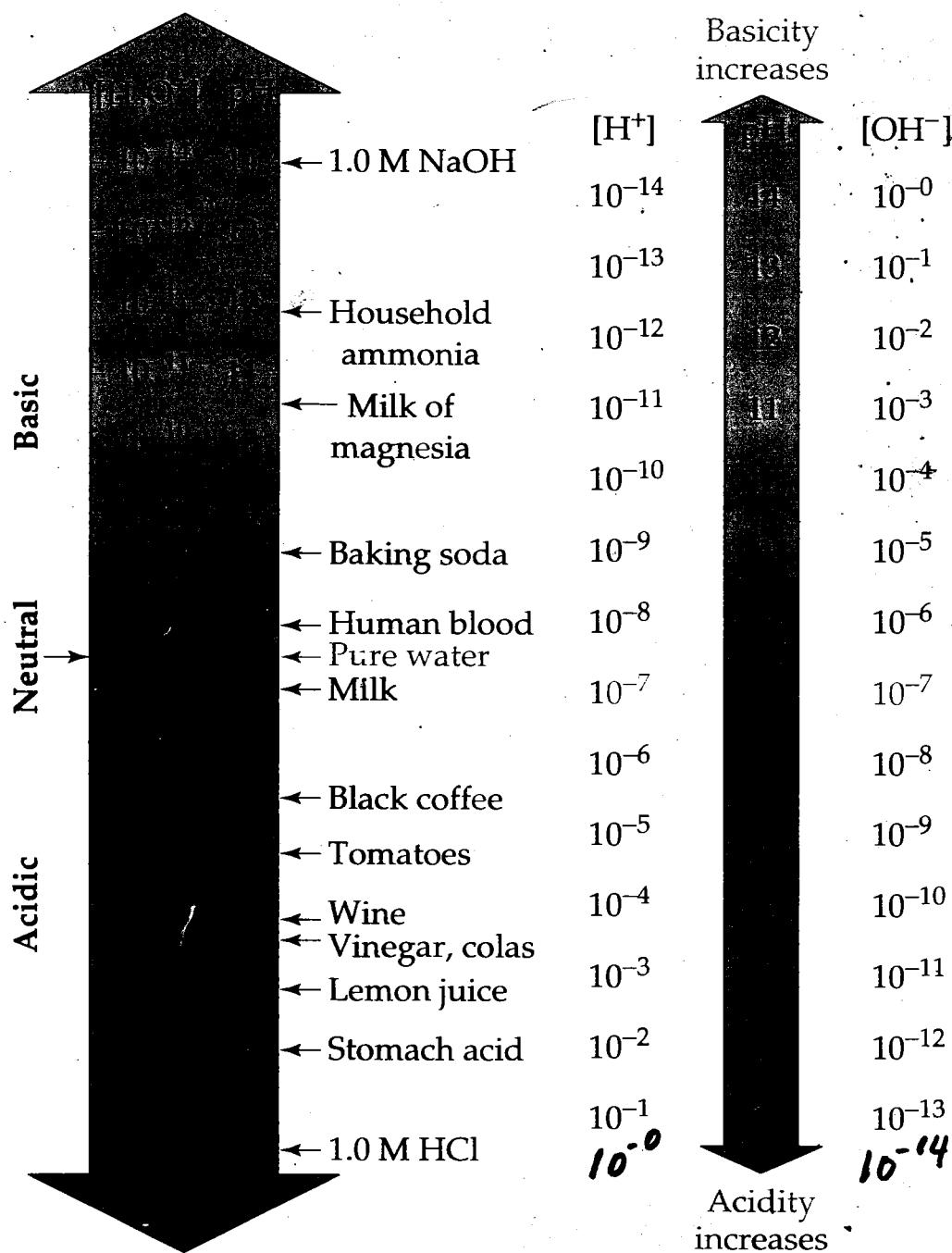


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5b

**Figures 10.1, 10.2 The pH scale and the pH of some common substances
The relationship of pH scale to H⁺ and OH⁻ concentrations**



6

Dissociation of Acid in Aqueous solution

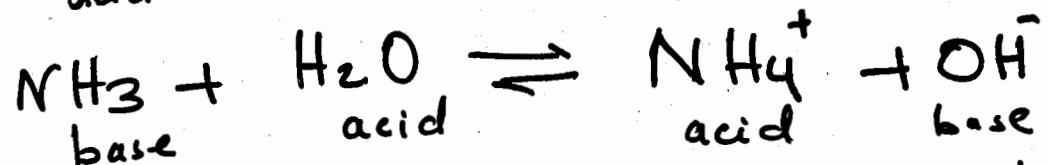
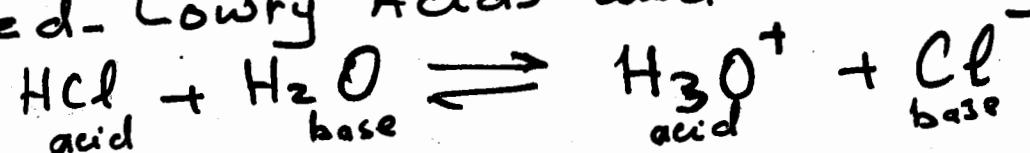


Acid

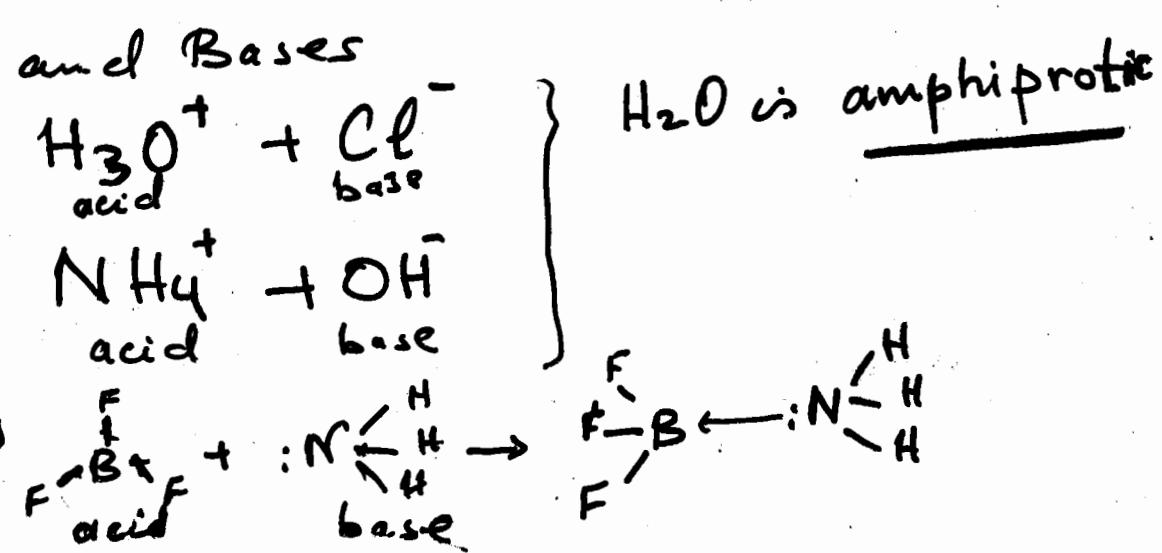
Base

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Bronsted-Lowry Acids and Bases



Lewis Acids and Bases

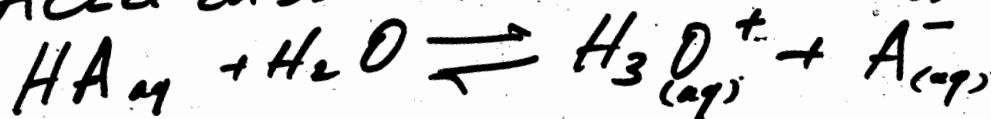


ACIDS & BASES

Acid - proton donor

Base - Proton acceptor

Acid dissociation Constant



$$K_{eq.\text{const}} = \frac{[H_3O^+][A^-]}{[HA][H_2O]}$$

$$\text{Dissociation constant} = K_a = K [H_2O]$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

<u>Acid</u>	<u>K_a (M)</u>	<u>pK_a</u>
HCOOH	1.78×10^{-4}	3.75
CH ₃ COOH	1.74×10^{-5}	4.76
H ₃ PO ₄	7.25×10^{-3}	2.14
H ₂ PO ₄ ⁻	1.38×10^{-7}	6.86
HPo ₄ ²⁻	3.98×10^{-13}	12.4
H ₂ CO ₃	1.7×10^{-4}	3.77
HCO ₃ ⁻	6.31×10^{-11}	10.2
NH ₄ ⁺	5.62×10^{-10}	9.25

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

$$PH = \log \frac{K_a}{[NH_3^+]}$$

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

Acid	Structure ^a	K_a	pK_a
-Formic acid	HCOOH	1.78×10^{-4}	3.75
-Acetic acid	CH ₃ COOH	1.74×10^{-5}	4.76
-Pyruvic acid	CH ₃ COCOOH	3.16×10^{-3}	2.50
-Lactic acid	CH ₃ CHOHCOOH	1.38×10^{-4}	3.86
-Malic acid	HOOC-CH ₂ -CHOH-COOH	(1) 3.98×10^{-4} (2) 5.50×10^{-6}	3.40 5.26
-Citric acid	HOOC-CH ₂ -C(OH)(COOH)-CH ₂ -COOH	(1) 8.14×10^{-4} (2) 1.78×10^{-5} (3) 3.92×10^{-6}	3.09 4.75 5.41
-Carbonic acid	HO-C(=O)-OH	(1) 4.31×10^{-7} (2) 5.62×10^{-11}	6.37 10.26

Acid

Structure^a

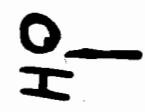
Ka

PKa

- Phosphoric acid HO-P(OH)₂

(1) 7.25×10^{-3}

2.14



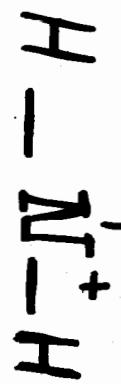
(2) 6.31×10^{-8}

7.20

(3) 3.98×10^{-13}

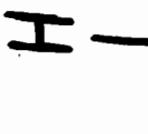
12.40

- Ammonium ion



5.62×10^{-10}

9.25



Large

1.2×10^{-2}

Large

HNO₃

- H₂SO₄

H₂SO₄

HCl

- PH of 0.04M Ba(OH)_2

$$[\text{OH}^-] = 2 \times 0.04\text{M} = 0.08\text{M}$$

$$\text{POH} = 1.1$$

$$\text{PH} = 14 - 1.1 = 12.9$$

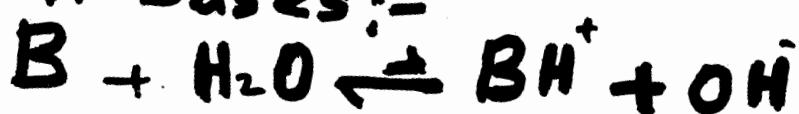
- PH of 0.02M weak acid (HA).
4.0. Find K_a .

$$K_a = \frac{[\text{H}^+] \times [\text{A}^-]}{\text{HA}} = \frac{[\text{H}^+]}{\text{HA}}^2 = \frac{10^{-4}}{10^{-4} \times 10^{-4}} = \frac{1}{0.02}$$

$$= 5 \times 10^{-7} \text{ M}$$

$$[\text{H}^+] = 10^{-\text{PH}}$$

Weak Bases:-



$$K_b = \frac{\text{BH}^+ \times \text{OH}^-}{[\text{B}]}$$

$$K_a = \frac{\text{B} \times [\text{H}^+]}{[\text{BH}^+]} \quad \begin{matrix} \text{Reverse reaction} \\ \text{for } \text{BH}^+ \rightleftharpoons \text{B} + \text{H}^+ \end{matrix}$$

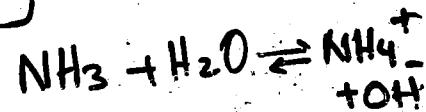
$$K_a \times K_b = [\text{H}^+] \times [\text{OH}^-] = K_w = 10^{-14}$$

$$\text{p}K_a + \text{p}K_b = 14$$

Example

K_b for ammonia is $1.8 \times 10^{-5} M$
 Find the pH of $1 \times 10^2 M$ of Ammonia

$$K_b = \frac{[NH_4^+] [OH^-]}{[NH_3]}$$



$$1.8 \times 10^{-5} = \frac{[OH^-]^2}{0.01}$$

$$OH^- = \sqrt{1.8 \times 10^{-7}} = 4.24 \times 10^{-4} M$$

$$pOH = -\log 4.24 \times 10^{-4} = 3.37$$

$$pH = 14 - 3.37 = 10.6$$

Example

The pH of 0.03M weak base solution
 is 10. Calculate pK_b

$$pOH = 14 - 10 = 4$$

$$[OH^-] = 10^{-4} \quad B + H_2O \rightleftharpoons BH^+ + OH^-$$

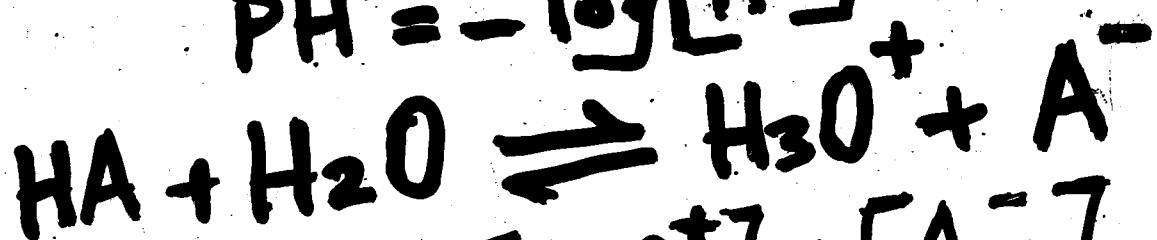
$$K_b = \frac{10^{-4} \times 10^{-4}}{0.03} = 3.33 \times 10^{-7} M$$

$$pK_b = -\log K_b = 6.48$$

V. IMP.

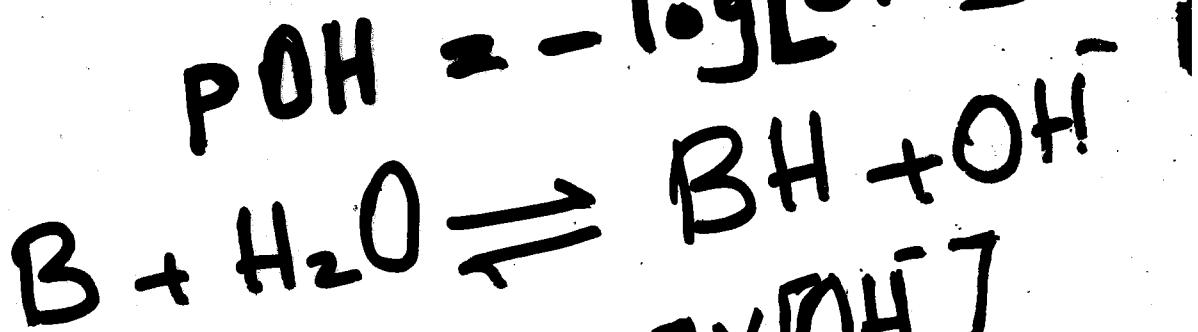
9c

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



$$\rightarrow K_a = \frac{[H_3O^+] \times [A^-]}{[HA]}$$

$$POH = -\log[OH^-]$$



$$\rightarrow K_b = \frac{[BH] \times [OH^-]}{[B]}$$

MOST IMPORTANT FORMULAS
TO USE

Relation of pH, pK & buffer Concentration
"Henderson-Hasselbach Equation"

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

a	H ⁺
a-x	x

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

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

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

$$pH = pK_a \text{ at mid point} = [A^-] = [HA]$$

$$pH = pK_a + \log \frac{\text{Conjugate base}}{\text{Conjugate acid}}$$

e.g. $pH = pK_a - 1$
 $\text{base/acid} = 1/10$

$$pH = pK_a - 2 ; \text{ Ratio} = 1/100$$

$pH = pK_a + 1$
 $\text{base/acid} = 10/1$

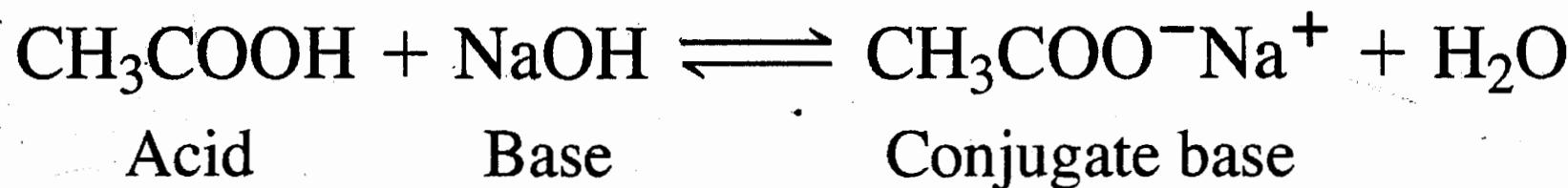
$$pH = pK_a + 2$$

Ratio = 100/1

II Titration Curves

1. Monoprotic acids

weak acid strong base



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strong acids have weak conjugate bases
Weak acid have strong conjugate bases

acid strength ↑
Acid
 CH_3COOH
 $\text{C}_6\text{H}_5\text{OH}$ phenol
 H_2O
 $\text{C}_2\text{H}_5\text{OH}$ ethanol

Conjugate base
 CH_3COO^- phenoxide
 $\text{C}_6\text{H}_5\text{O}^-$ hydroxide
 OH^- ethoxide
 $\text{C}_2\text{H}_5\text{O}^-$ ethoxide
↓
Base strength

Titration Curve

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

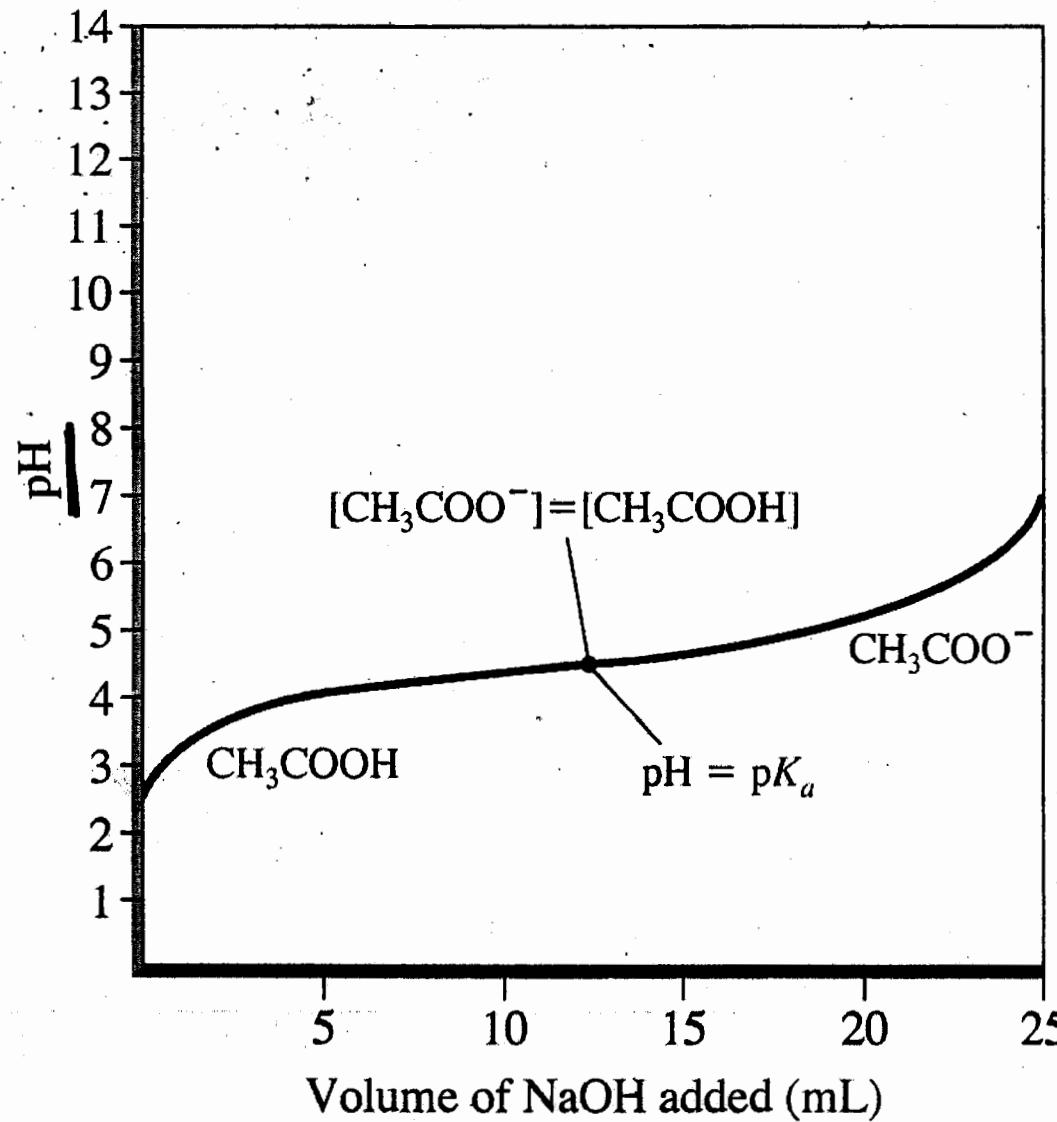
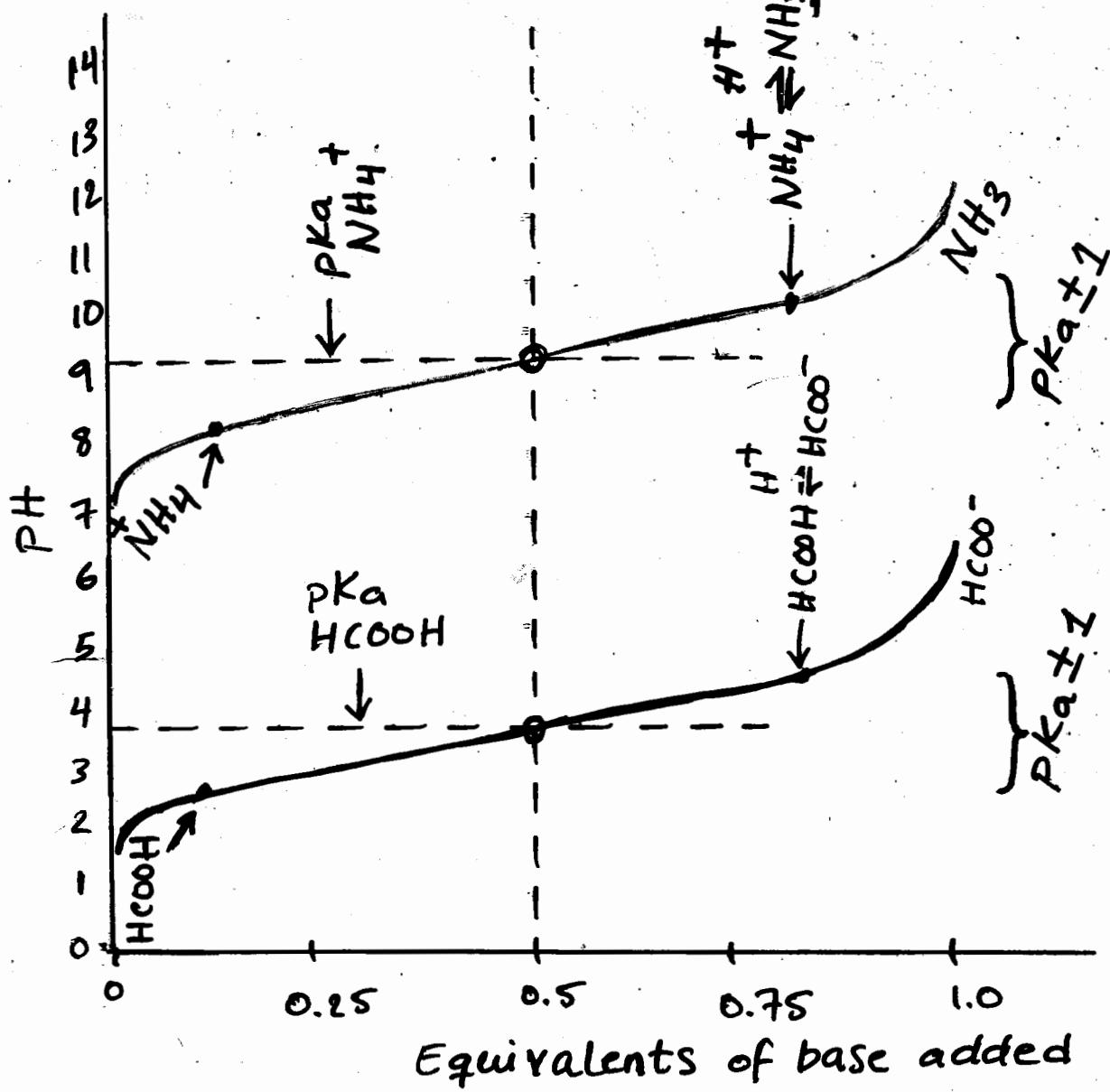
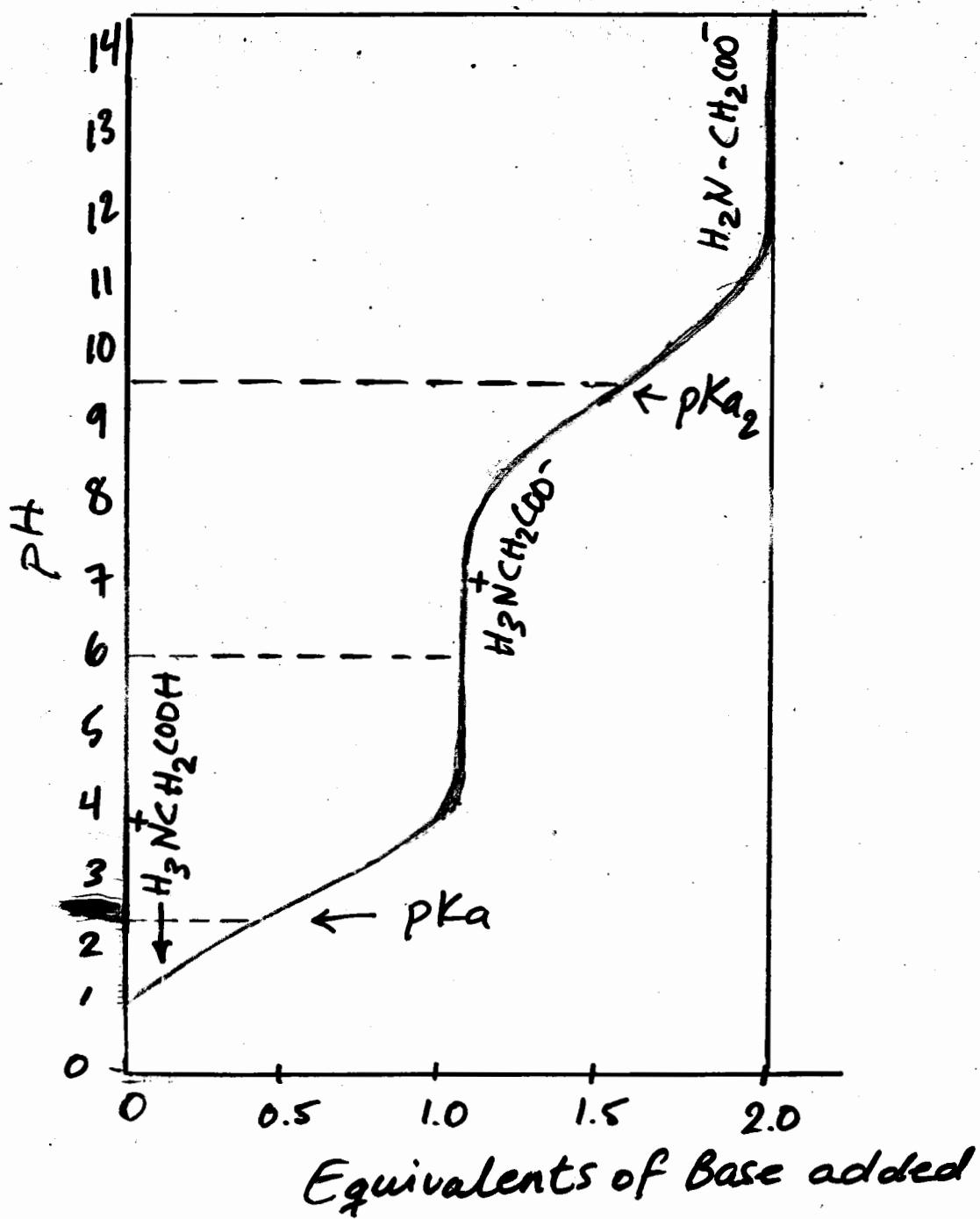


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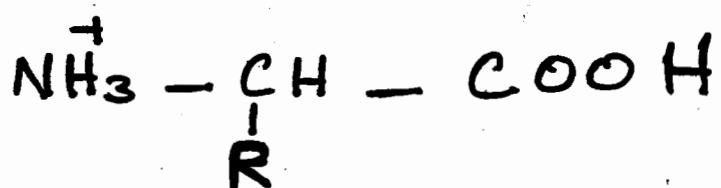


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Titration Curves of Glycine



Amino acids General Formula



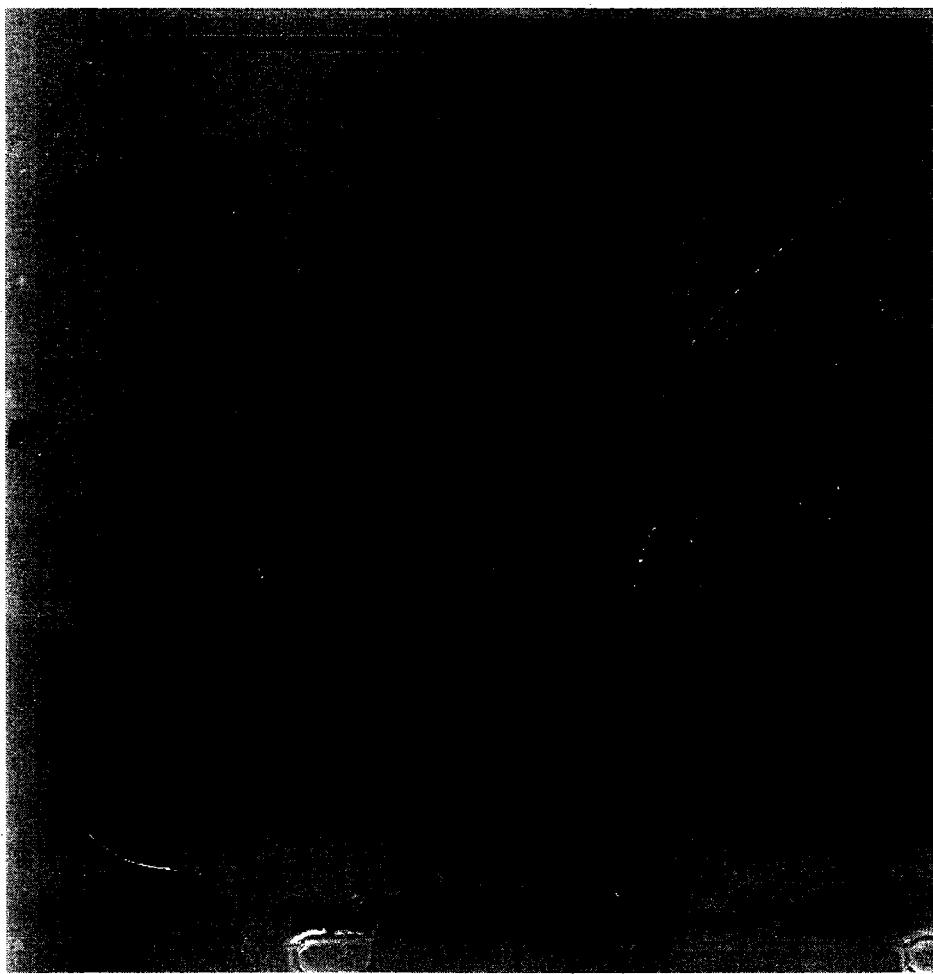
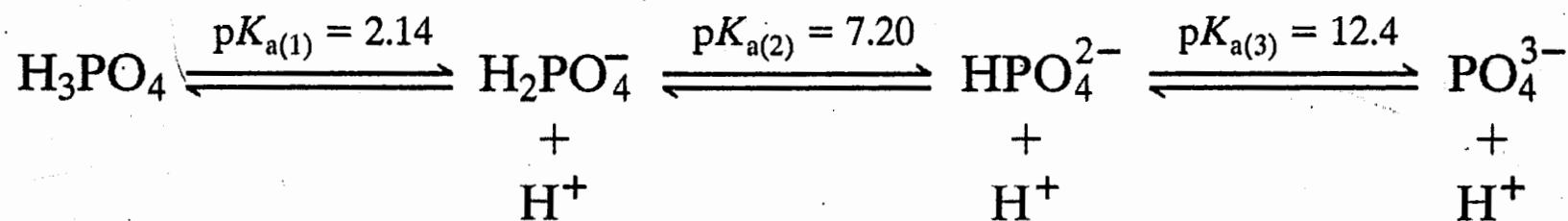


Fig 3.7 : The titration curve of Histidine .

The isoelectric pH (pI) is the value at which positives and negative charges are the same. The molecule has no net charge.

Polyprotic Acids

e.g. malic acid, citric acid, carbonic acid, phosphoric acid



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hi

The titration Curve of $H_2PO_4^-$, showing
the buffer region for the $H_2PO_4^-/HPO_4^{2-}$ pair

