

Ionization of Water, Weak acids & Weak bases:-

Ionization of $H_2O \rightarrow$ very small degree

Equilibrium Constt.



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



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

$$K_{eq} \times 55.5 = [H^+][OH^-] = K_w \quad \text{Ion product Constant for water}$$

$$K_w = 55.5 \times (1.8 \times 10^{-16} M) = 1.0 \times 10^{-14} M^2$$

$$[H^+][OH^-] = 10^{-14} M^2 \quad \text{true for pure water and solutions}$$

In Pure Water \rightarrow

$$[H^+] = [OH^-] = 10^{-7} M$$

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

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

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

$$\text{e.g. } 0.1 M HCl \quad [OH^-] = \frac{10^{-14}}{10^{-1}} = 10^{-13} M$$

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Acidic solution $[H_3O^+] > 10^{-7} M$ or $pH < 7$

Neutral solution $[H_3O^+] = 10^{-7} = [OH^-]$

Basic solution $[H_3O^+] < 10^{-7}$ or $[OH^-] > 10^{-7}$

pH scale: Measuring acidity in aqueous solutions

$$pH = -\log[H_3O^+]$$

$$\text{or } [H_3O^+] = 10^{-pH}$$

$[H_3O^+]$ for coffee is $10^{-5} M$
 $pH = -\log[10^{-5}] = 5$

Lemon juice has a $pH = 2$
 $2 = -\log[H_3O^+]$

$$[H_3O^+] = 10^{-2}$$

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

$$[H_3O^+] = \frac{10^{-14}}{10^{-3}} = 10^{-11} M$$

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

The pH Scale:

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

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

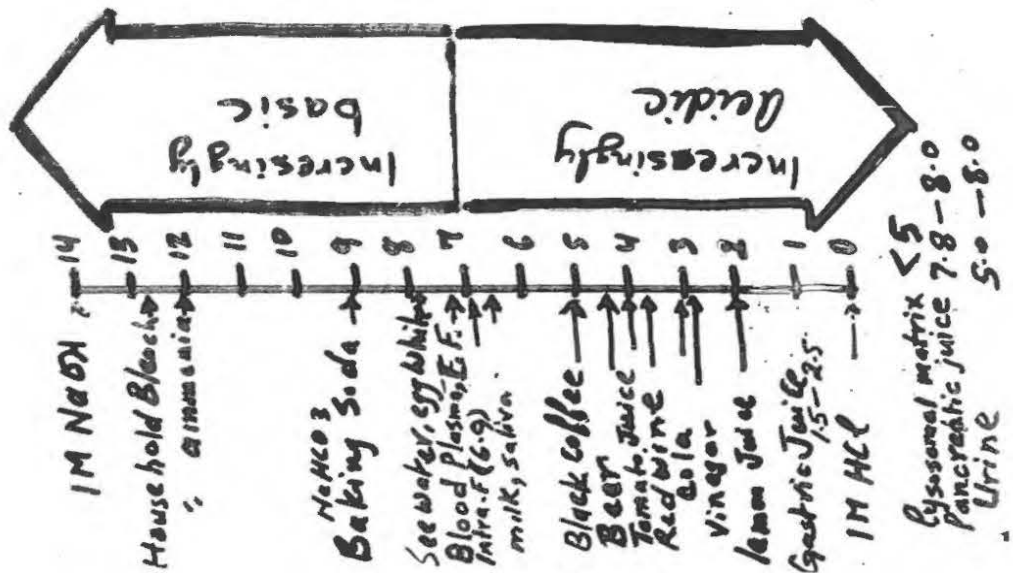
$$-\log [H^+] \times [OH^-] = -\log 10^{-14}$$

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

$$pH + pOH = 14$$

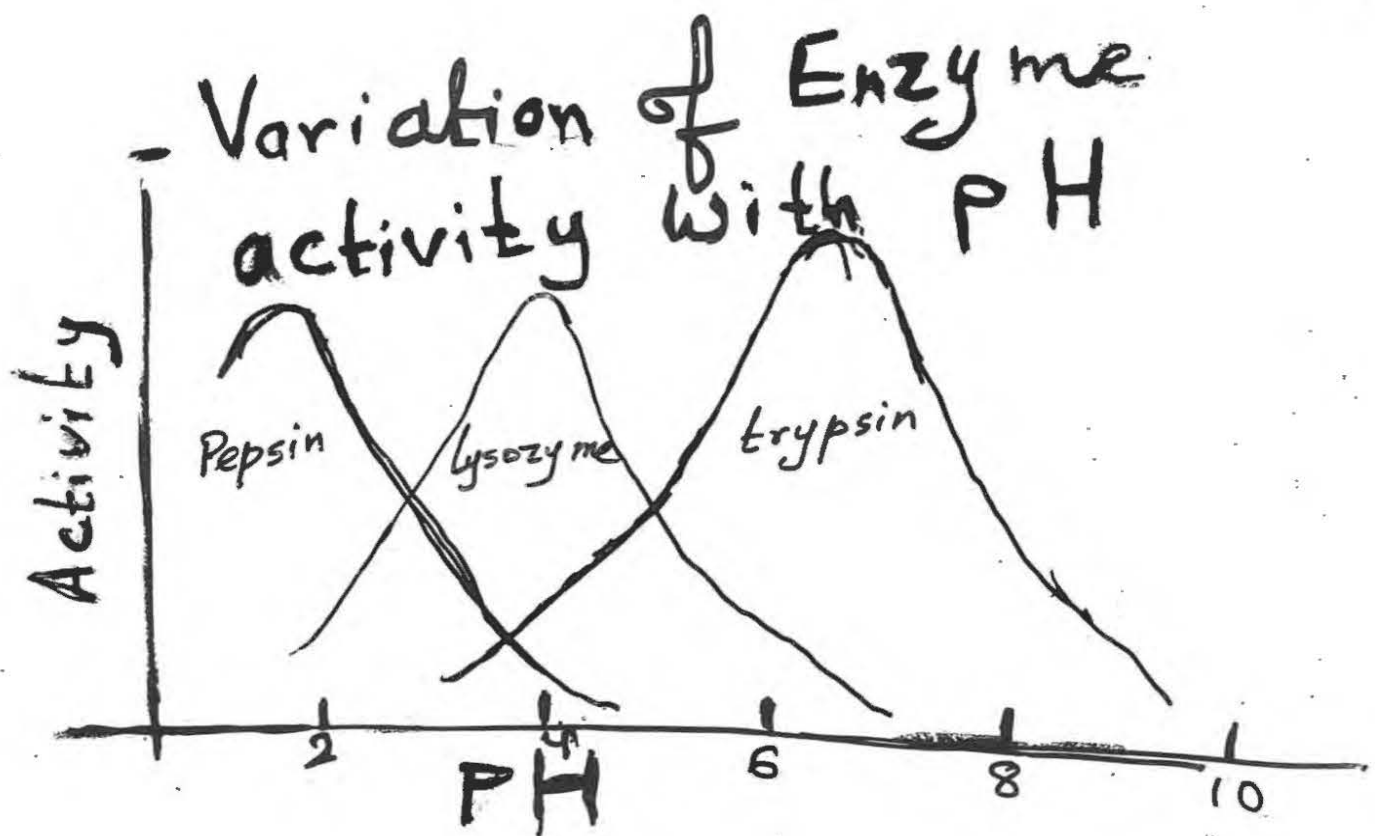
$[H^+]$ M	pH	$[OH^-]$ M	pOH
$10^0 \rightarrow 1.0$	0	10^{-14}	14
$10^{-1} \rightarrow 0.1$	1	10^{-13}	13
$10^{-2} \rightarrow 0.01$	2	10^{-12}	12
$10^{-3} \rightarrow 0.001$	3	10^{-11}	11
$10^{-4} \rightarrow 0.0001$	4	10^{-10}	10
10^{-7}	7	10^{-7}	7
10^{-12}	12	10^{-2}	2
10^{-14}	14	1	0

pH of Some Fluids



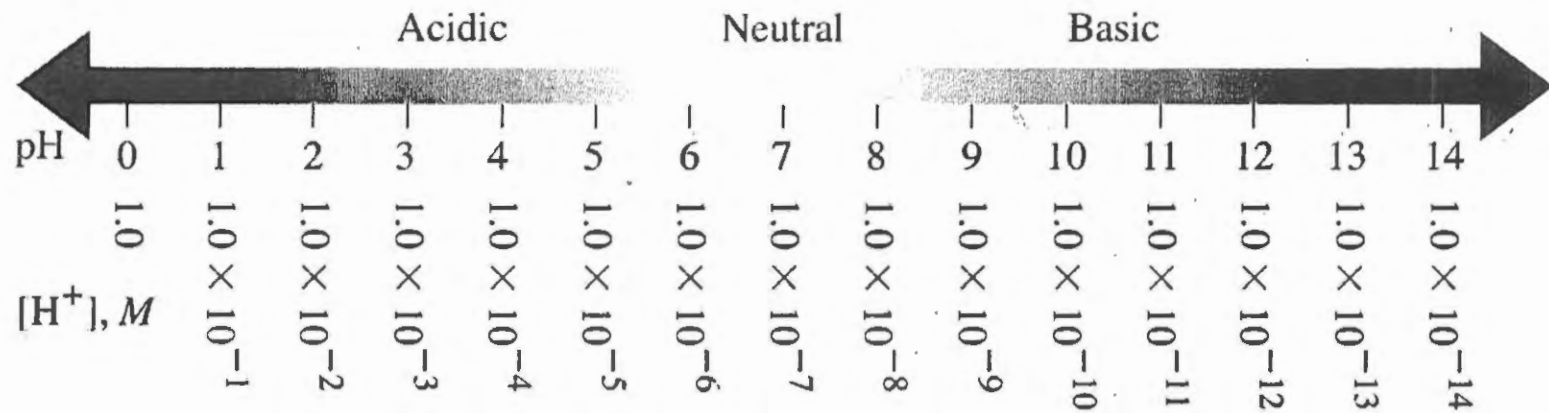
Importance of pH Regulation

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



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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

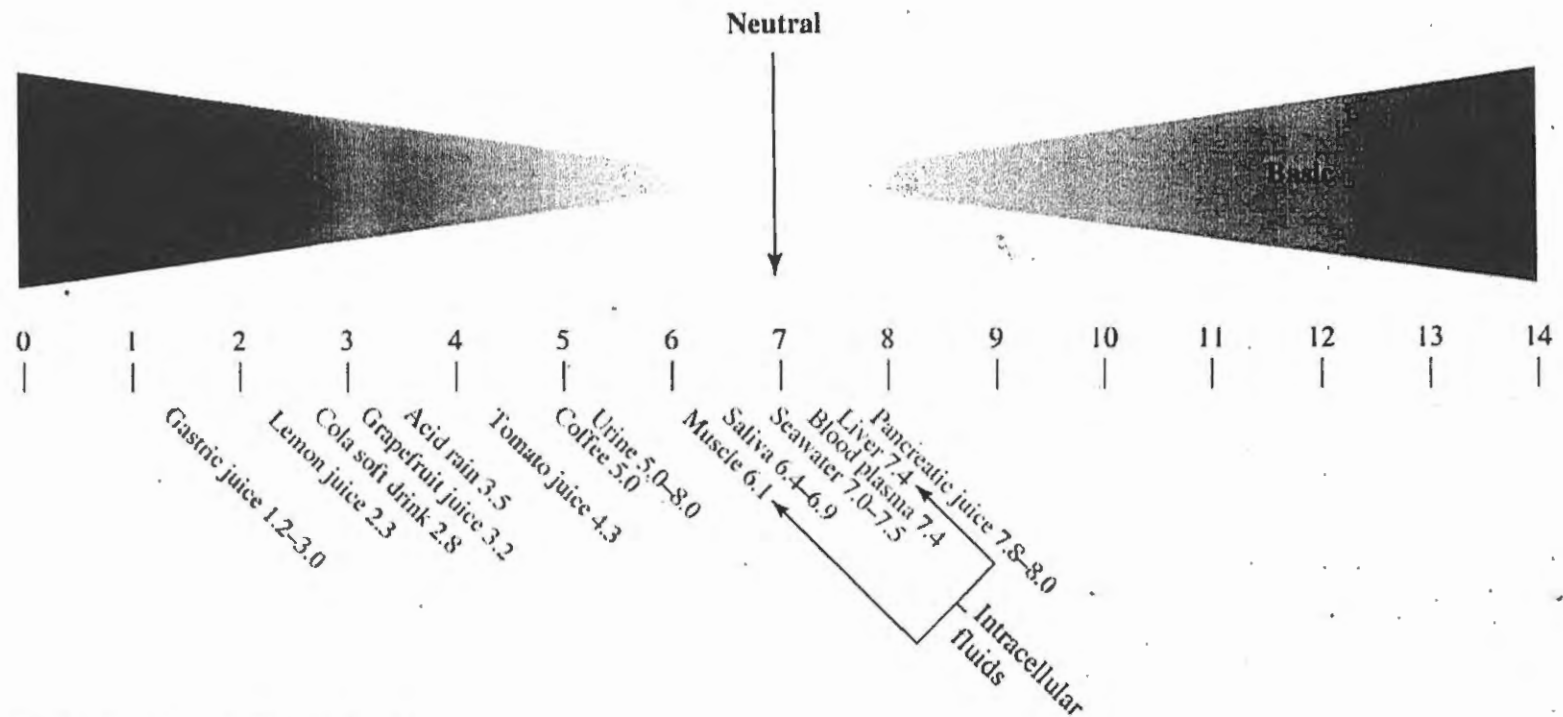
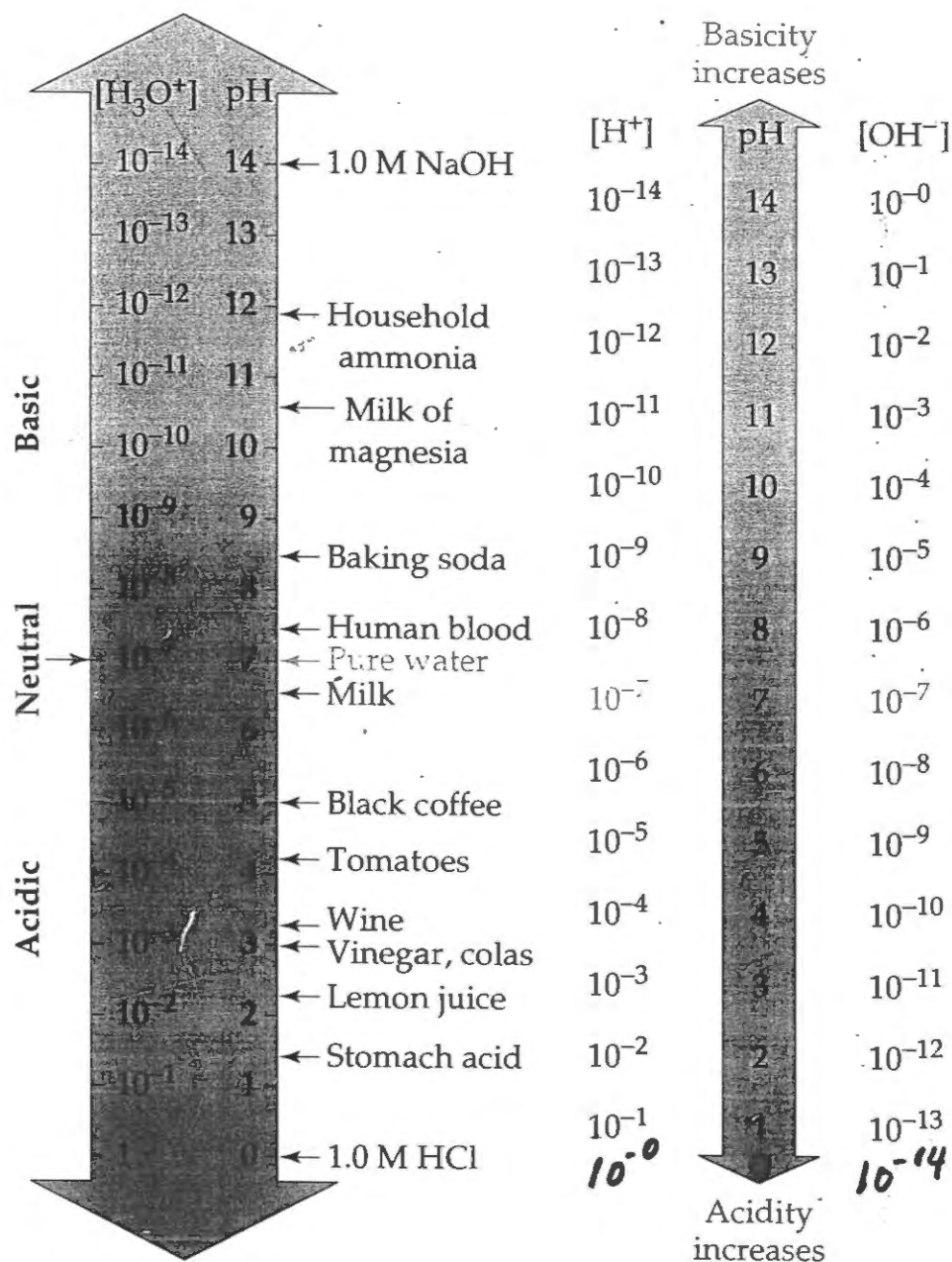


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The relationship of pH scale to H^+ and OH^- concentrations



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Dissociation of Acid in Aqueous solution



Acid

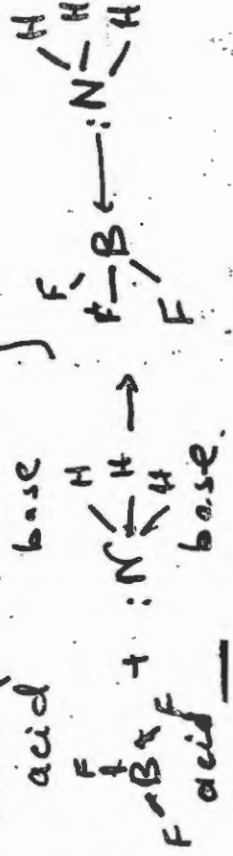
Base

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Bronsted-Lowry Acids and Bases



Lewis Acids and Bases



H₂O is amphiprotic

ACIDS & BASES

Acid - proton donor

Base - Proton acceptor

Acid dissociation constant



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

Dissociation constant = $K_a = K [H_2O]$

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

Acid	K_a (M)	pK_a
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

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

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

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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 COCO ₂ H	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	$ \begin{array}{c} \text{OH} \\ \\ \text{HOOC}-\text{CH}_2-\text{C}-\text{CH}_2-\text{COOH} \\ \\ \text{COOH} \end{array} $	(1) 8.14×10^{-4} (2) 1.78×10^{-5} (3) 3.92×10^{-6}	3.09 4.75 5.41
-Carbonic acid	$ \begin{array}{c} \text{O} \\ \\ \text{HO}-\text{C}-\text{OH} \end{array} $	(1) 4.31×10^{-7} (2) 5.62×10^{-11}	6.37 10.26

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Acid	Structure ^a	K _a	pK _a
- Phosphoric acid	$\begin{array}{c} \text{O} \\ \parallel \\ \text{HO}-\text{P}-\text{OH} \\ \\ \text{OH} \end{array}$	(1) 7.25×10^{-3} (2) 6.31×10^{-8} (3) 3.98×10^{-13}	2.14 → 7.20 12.40
- Ammonium ion	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{N}^+-\text{H} \\ \\ \text{H} \end{array}$	5.62×10^{-10}	9.25
- H_2SO_4		Large	
- HSO_4^-		1.2×10^{-2}	
- HCl		Large	
- HNO_3		Large	

Examples:

- pH of 0.04 M $\text{Ba}(\text{OH})_2$

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

$$\text{pOH} = 1.1$$

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

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

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

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

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

Weak bases:-



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

$$K_a = \frac{[\text{B}] \times [\text{H}^+]}{[\text{BH}^+]}$$

Reverse reaction for $\text{BH}^+ \rightleftharpoons \text{B} + \text{H}^+$

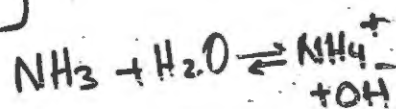
$$K_a \times K_b = [\text{H}^+] \times [\text{OH}^-] = K_w = 10^{-14}$$

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

Example

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

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



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

$$[OH^-] = \sqrt{1.8 \times 10^{-5}} = 4.24 \times 10^{-4} M$$

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

$$pH = 14 - 3.37 = 10.6$$

Example

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

$$pOH = 14 - 10 = 4$$

$$[OH^-] = 10^{-4}$$



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

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

V. Imp.

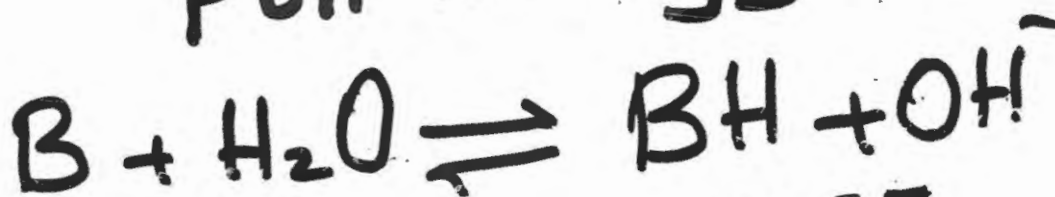
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$$pH = -\log[H^+]$$



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

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



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

MOST IMPORTANT FORMULAS
TO USE

Relation of pH, pK & buffer concentration 10

"Henderson - Hasselbalch Equation"

$$K_a = \frac{[H^+][A^-]}{[HA]} \quad \begin{matrix} d & H^+ & A^- \\ a-x & x & x \end{matrix}$$

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

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

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

$pH = pK_a$ at mid point $\Rightarrow [A^-] = [HA]$

$$pH = pK_a + \log \frac{[\text{conjugate base}]}{[\text{conjugate acid}]}$$

e.g. $pH = pK_a - 1$
base/acid = 1/10

$pH = pK_a - 2$; Ratio = 1/100

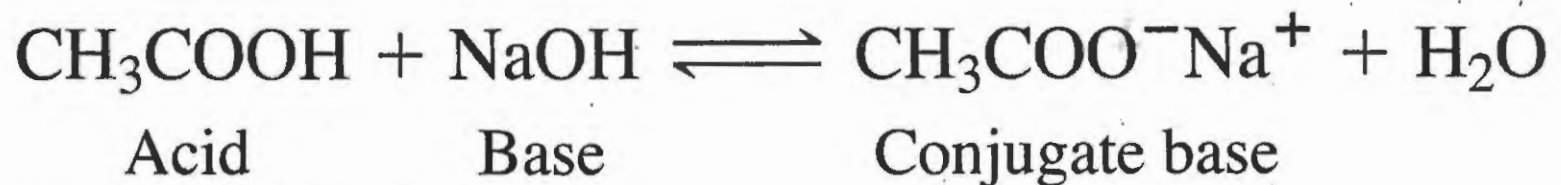
.....
 $pH = pK_a + 1$
base/acid = 10/1

$pH = pK_a + 2$
Ratio = 100/1

Titration Curves

1. Monoprotic acids

weak acid strong base



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strong acids have weak conjugate bases
weak acids have strong conjugate bases

acid strength ↑	<u>Acid</u> CH_3COOH $\text{C}_6\text{H}_5\text{OH}$ phenol H_2O $\text{C}_2\text{H}_5\text{OH}$ ethanol	<u>conjugate base</u> CH_3COO^- $\text{C}_6\text{H}_5\text{O}^-$ phenoxide OH^- hydroxide $\text{C}_2\text{H}_5\text{O}^-$ ethoxide	↓ Base strength
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Titration Curve

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

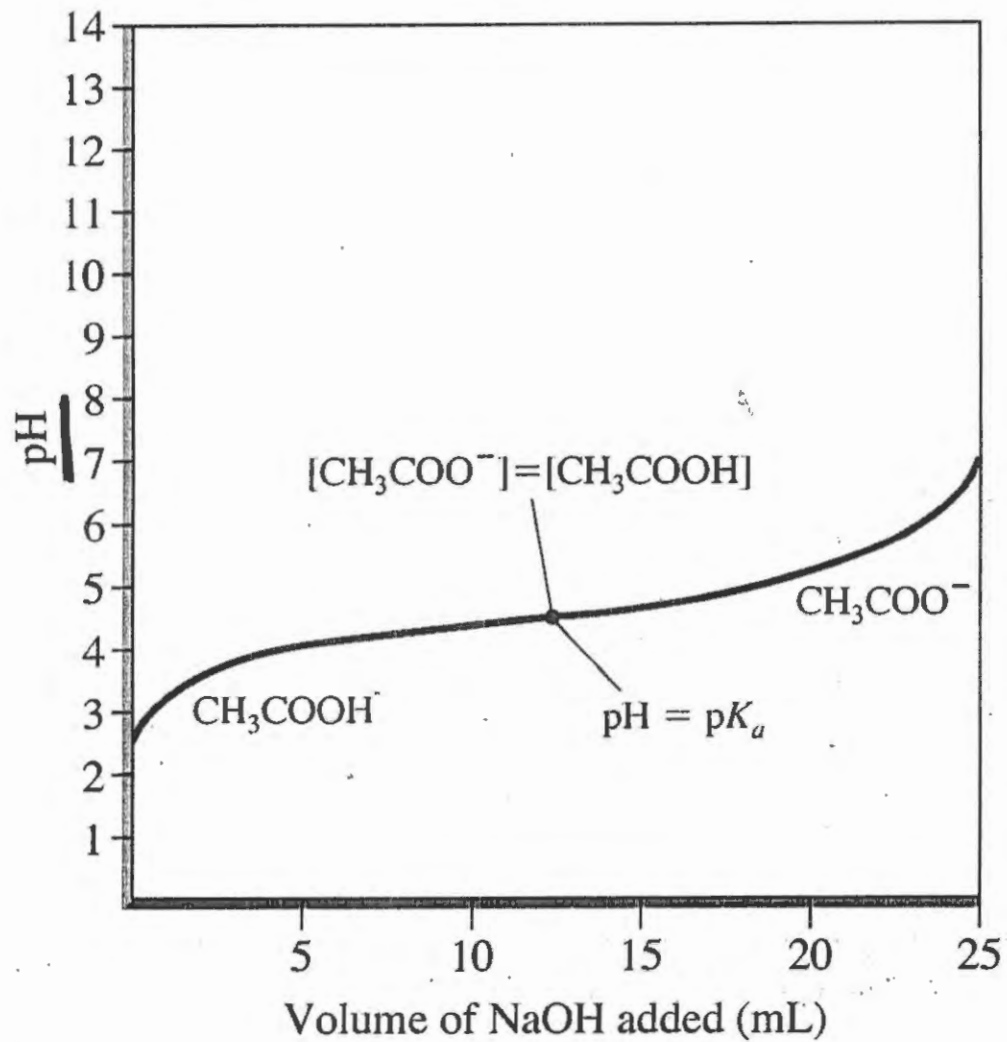
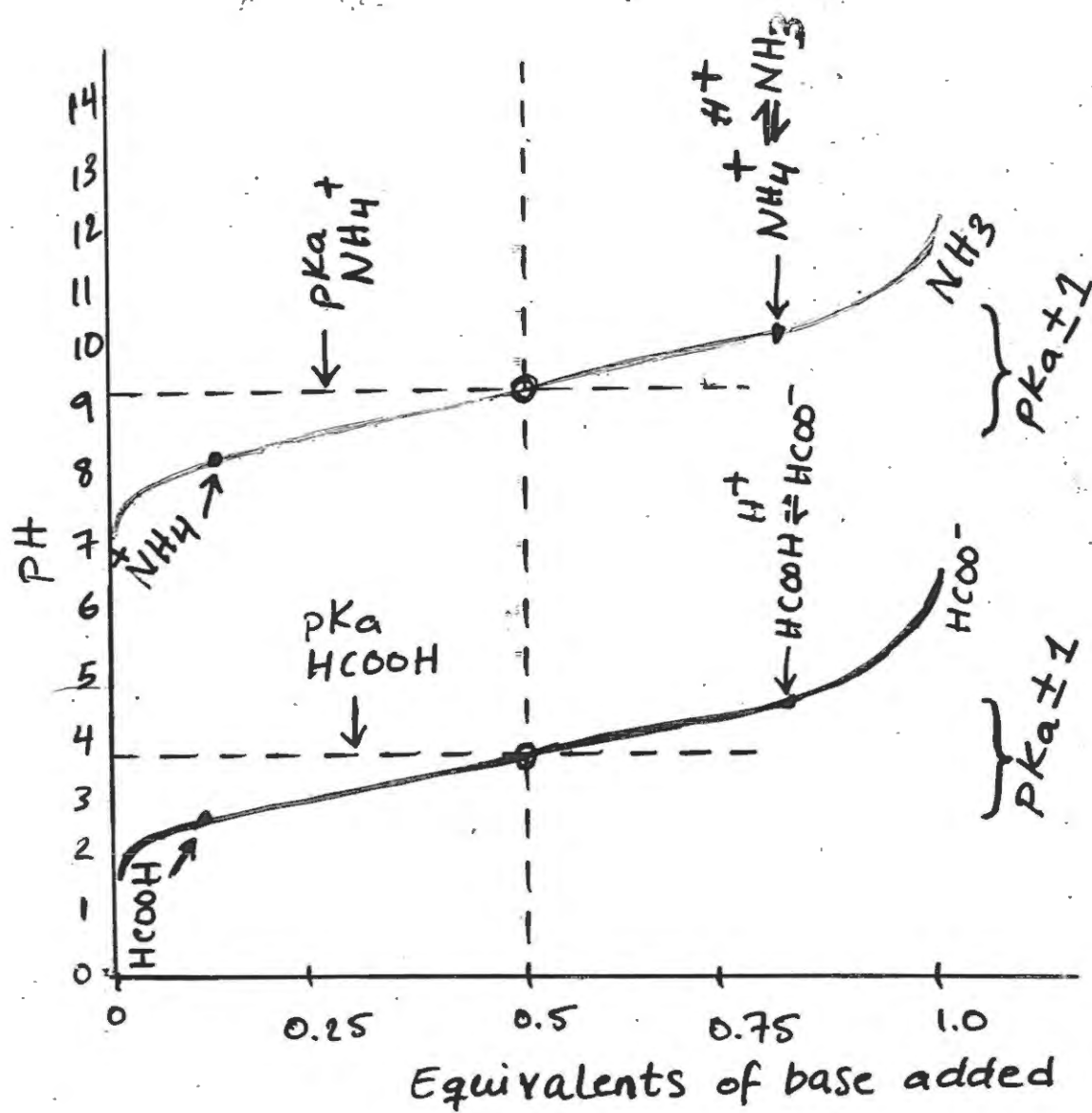
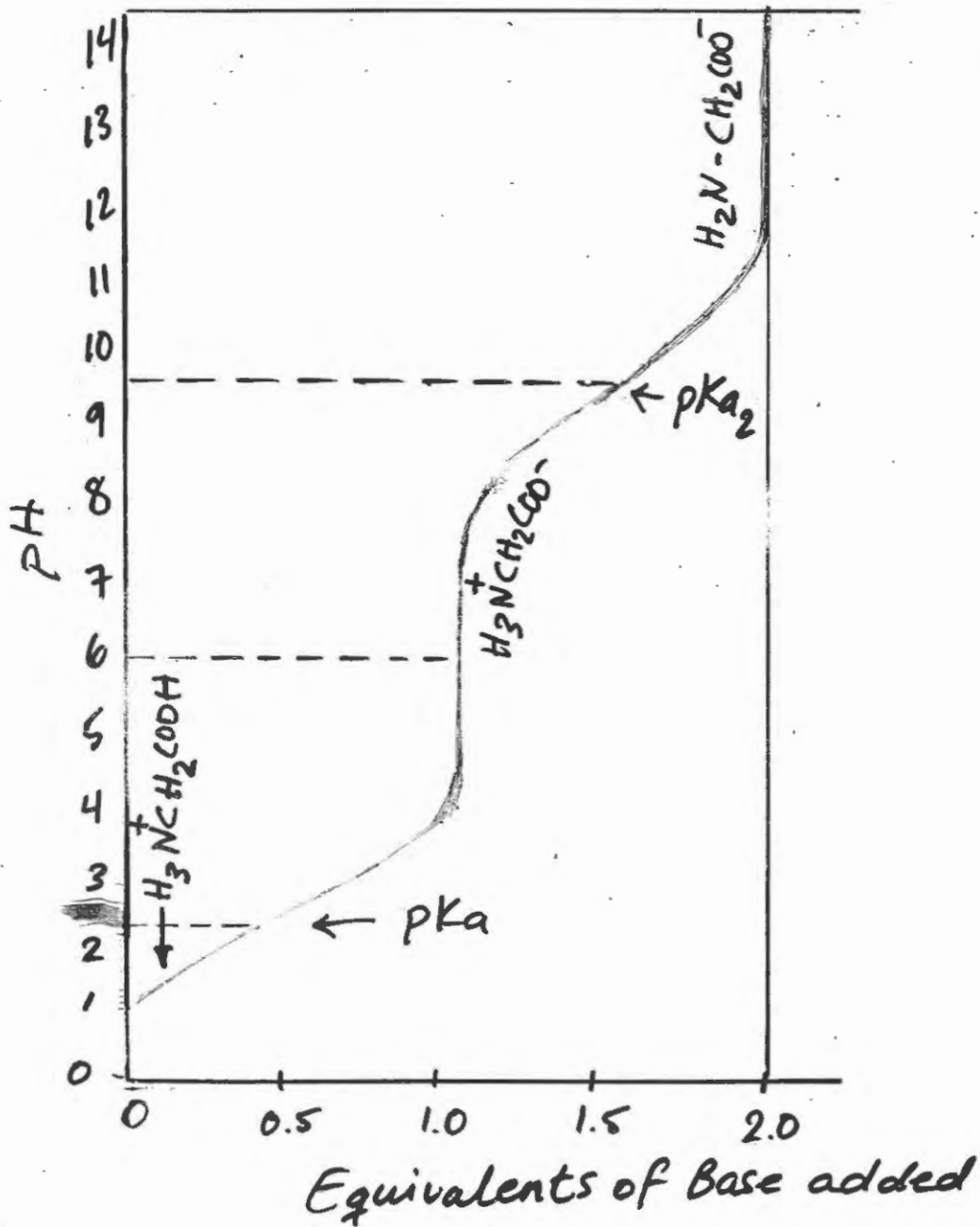


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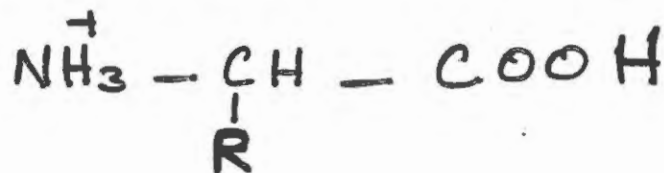


Titration Curves of Glycine

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Amino acids General Formula



13C

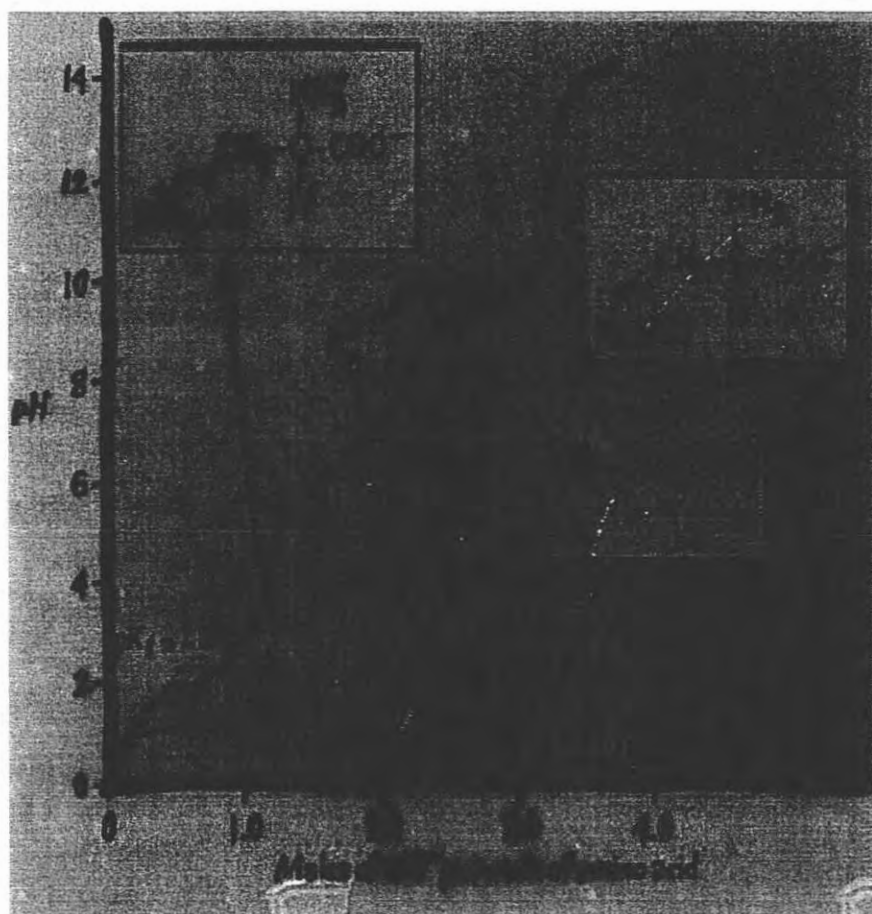
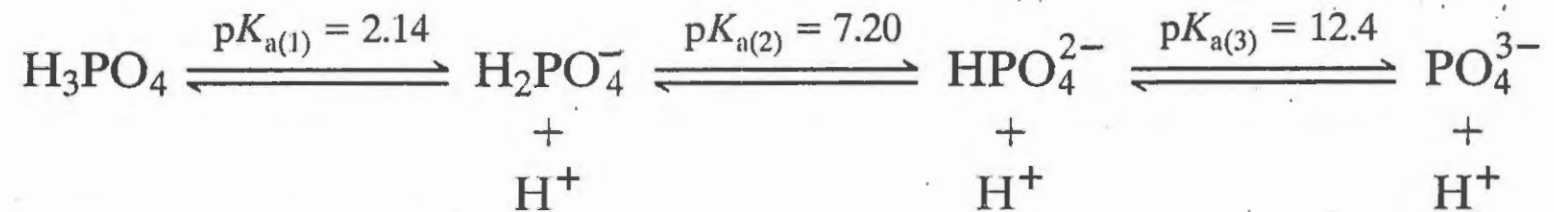


Fig 3.7 : The titration curve of Histidine .

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

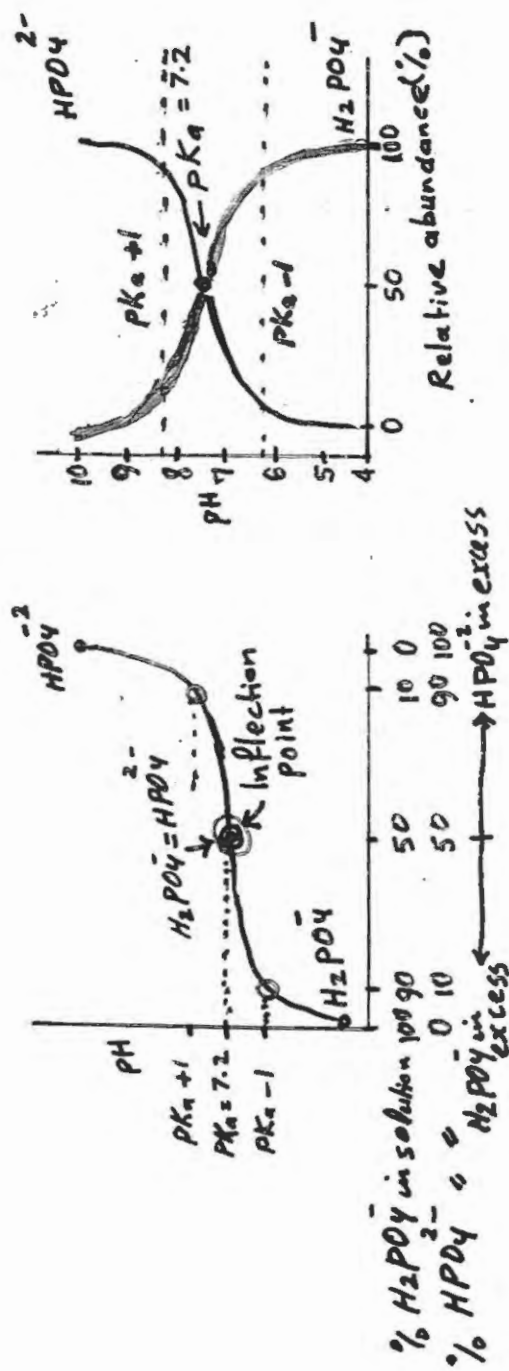
Polyprotic Acids

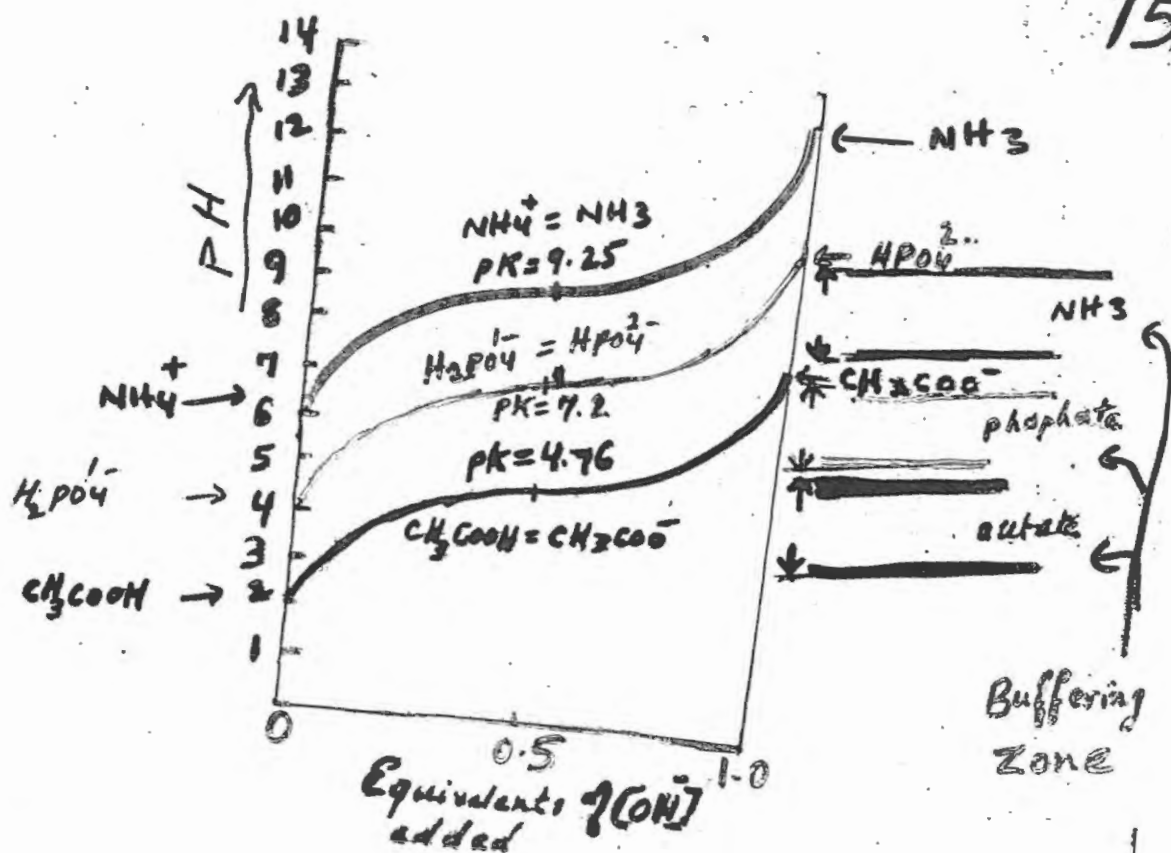
e.g. malic acid, citric acid, carbonic acid, phosphoric acid



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The titration curve of H_2PO_4^- , showing the buffer region for the $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$ pair





$$pH = pK_a + \log \frac{[conjugate\ base]}{[conjugate\ acid]}$$
 Henderson-Hasselbalch Equation

Buffer Capacity

Effective buffering range

at a pH range = $pK_a \pm 1$

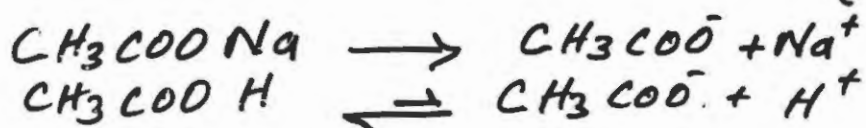
but better at $pK_a \pm 0.5$ only

- Ratio of $[A^-]/[HA]$
- Concentration of $[A^-] + [HA]$

Buffer Action:-

Resists changes in pH when small amount of acid or alkali are added to it.

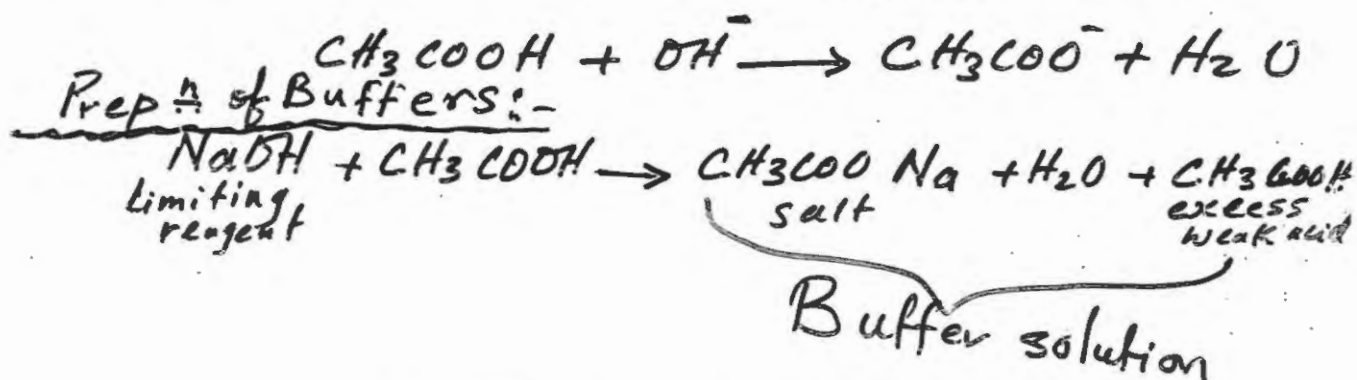
Acidic buffer = weak acid + salt of the acid
(Acid + strong base)



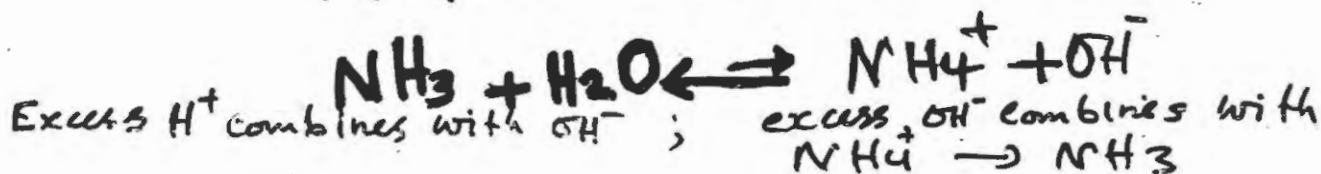
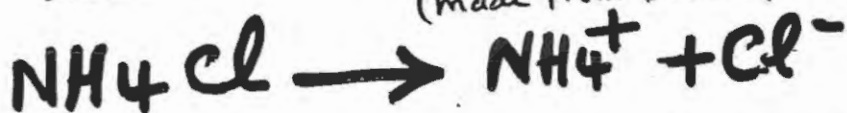
- if an acid is added, the extra H^+ combines with acetate \rightarrow undissociated **HAC**, so $[\text{H}^+]$ remains constt.



- if an alkali is added, the OH^- ions is removed by reaction with undissociated acid to form water — so $[\text{H}^+]$ remains constt.



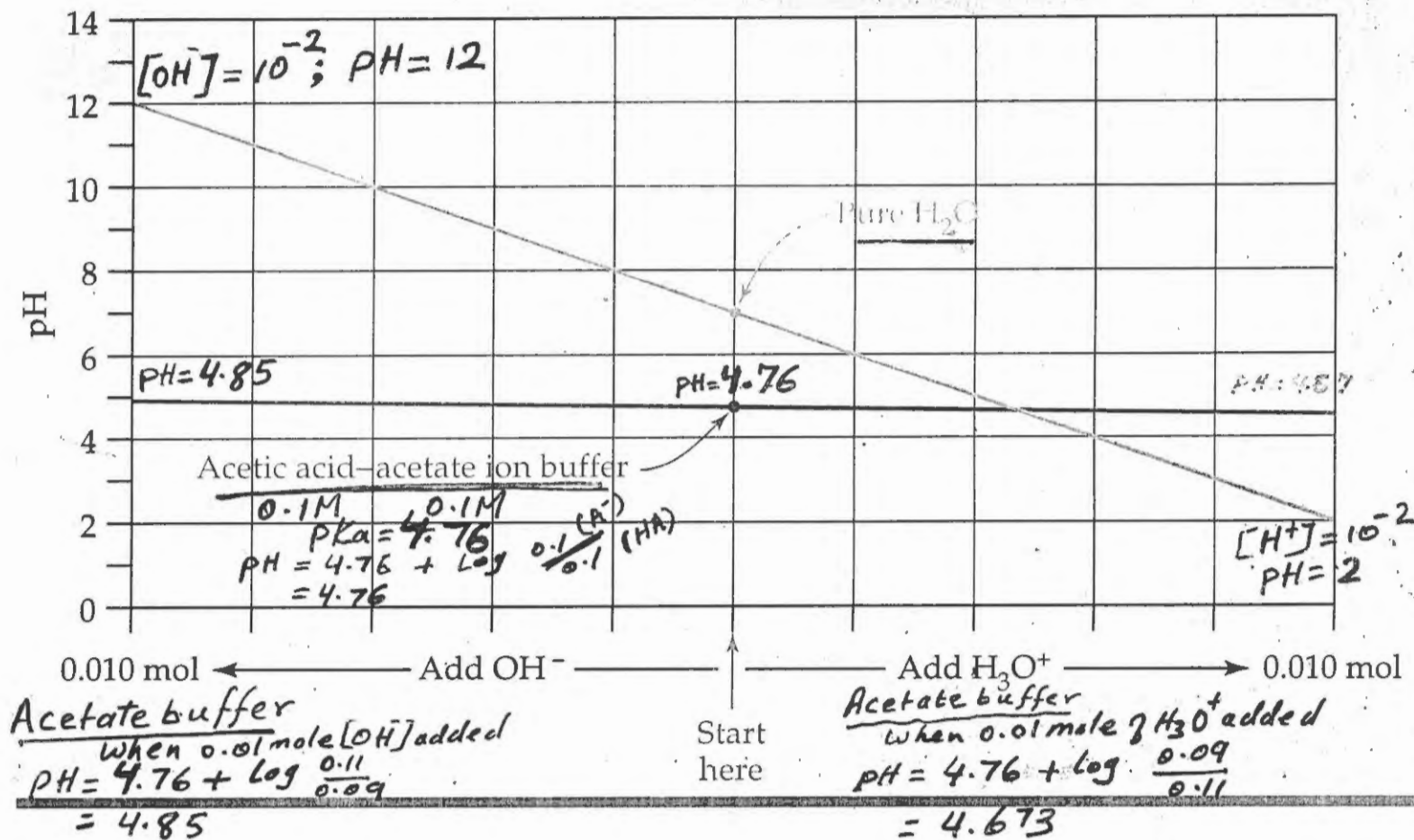
Alkali buffer :- $\text{pH} > 7.0$
Weak base + salt of the base
(made from base + strong acid)



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Figure 10.5 The effect of a buffer solution on pH

$$pH = pK_a + \log \frac{[\text{conjugate base}]}{[\text{conjugate acid}]}$$



Buffer Calculations:-

- 1) Buffer consists of 0.2 mole CH_3COONa in 500ml of 0.1M CH_3COOH ($K_a = 1.8 \times 10^{-5}$)
What is its pH?

$$\text{pH} = \text{pK}_a + \log \frac{\text{Base}}{\text{Acid}}$$

$$\text{pH} = \underbrace{-\log 1.8 \times 10^{-5}}_{\text{pK}_a} + \log \frac{2 \times 0.2}{0.1}$$

OR

$$\left[\begin{array}{l} K_a = \frac{[\text{H}^+] \times 0.4}{\text{CH}_3\text{COOH}} = 1.8 \times 10^{-5} \\ [\text{H}^+] = 4.5 \times 10^{-6} \text{ M} \\ \text{pH} = -\log 4.5 \times 10^{-6} = 5.4 \end{array} \right.$$

- 2) Calculate mass of sod. propionate to be dissolved in 1L of 1M propionic acid (M.W.=96) to give buffer of pH=4.5 ($\text{pK}_a = 4.87$)

$$\begin{array}{l} \text{pH} = 4.5 \\ \text{pK}_a = 4.87 \end{array} \quad \begin{array}{l} [\text{H}^+] = 10^{-4.5} \\ K_a = 10^{-4.87} \end{array}$$

$$K_a = \frac{[\text{H}^+] [\text{propionate}]}{\text{propionic acid}}$$

or use
Henderson-
Hasselbalch
Eq.

$$\begin{array}{l} [\text{Propionate}] = \frac{1 \times 10^{-4.5}}{10^{-4.87}} = 0.427 \text{ M} \\ 0.427 \times 96.08 (\text{M.W.}) = 41.0 \text{ gr.} \\ \hline 1.0 \end{array}$$

3) Calculate pH of a buffer when 18 ml of 0.1 M HCl is added to 32 ml of 0.1 M NH_3 ($\text{pK}_b = 4.75$)

(i) Calculate excess NH_3 and salt formed

$$\text{HCl moles} = \frac{18}{1000} \times 0.1 = 1.8 \times 10^{-3}$$

$$\text{NH}_3 \text{ moles} = \frac{32}{1000} \times 0.1 = 3.2 \times 10^{-3}$$

$$\text{Ammonium salt} = 1.8 \times 10^{-3}$$

$$\text{excess Ammonia} = 3.2 \times 10^{-3} - 1.8 \times 10^{-3} = 1.4 \times 10^{-3}$$

$$\text{Molar conc. of salt} = \frac{1.8 \times 10^{-3}}{\frac{50}{1000}} = 3.6 \times 10^{-2} \text{ M}$$

$$\text{“ “ of NH}_3 = \frac{1.4 \times 10^{-3}}{\frac{50}{1000}} = 2.8 \times 10^{-2} \text{ M}$$

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

$$10^{-4.75} = \frac{3.6 \times 10^{-2} \times \text{OH}^-}{3.2 \times 10^{-2}}$$

$$[\text{OH}^-] = 0.889 \times 10^{-5} = 8.89 \times 10^{-6} \text{ M}$$

$$\text{pOH} = -\log 8.89 \times 10^{-6} = 5.05$$

$$\text{pH} = 14 - 5.05 = 8.95$$

OR

$$\text{pOH} = \text{pK}_b + \log \frac{[\text{salt}]}{[\text{Weak base}]}$$

$$\text{pOH} = 4.75 + \log \frac{3.6 \times 10^{-2}}{3.2 \times 10^{-2}}$$

$$= 4.75 + 0.05 = 4.8; \text{ pH} = 14 - 4.8 = 9.2$$

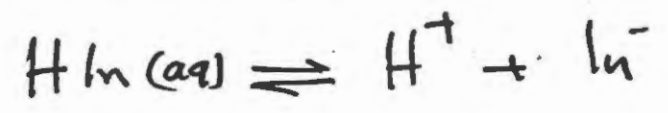
Table 2.6
Some synthetic buffers

Name (Abbreviation)	pK _a	Useful pH Range	Ionization Reaction ^a
<i>N</i> -(2-Acetamido)-2-aminoethanesulfonic acid (ACES)	6.9	6.4–7.4	$\text{H}_2\text{NCOCH}_2\text{NH}_2^+\text{CH}_2\text{CH}_2\text{SO}_3^- \rightleftharpoons \text{H}_2\text{NCOCH}_2\text{NHCH}_2\text{CH}_2\text{SO}_3^- + \text{H}^+$
3-(Cyclohexylamino)propanesulfonic acid (CHAPS)	10.5	10.0–11.0	$\text{C}_6\text{H}_{11}\text{NH}_3^+\text{CH}_2\text{CH}_2\text{CH}_2\text{SO}_3^- \rightleftharpoons \text{C}_6\text{H}_{11}\text{NHCH}_2\text{CH}_2\text{CH}_2\text{SO}_3^- + \text{H}^+$
<i>N</i> -(2-Hydroxyethyl)piperazine- <i>N'</i> -2-ethanesulfonic acid (HEPES)	7.5	7.0–8.0	$\text{HOCH}_2\text{CH}_2\text{N}^+\text{H}(\text{C}_4\text{H}_8)\text{NCH}_2\text{CH}_2\text{SO}_3^- \rightleftharpoons \text{HOCH}_2\text{CH}_2\text{N}(\text{C}_4\text{H}_8)\text{NCH}_2\text{CH}_2\text{SO}_3^- + \text{H}^+$
Tris(hydroxymethyl)aminomethane (TRIS)	8.3	7.5–9.0	$(\text{HOCH}_2)_3\text{CNH}_3^+ \rightleftharpoons (\text{HOCH}_2)_3\text{CNH}_2 + \text{H}^+$

^a Each reaction shows the two predominant forms (acid and base) present in the useful pH range.

INDICATORS

- often organic dye \rightarrow different colors in acidic and alkaline sol.
- weak acid / base



For Litmus Red Blue

In presence of acid \rightarrow HIn red

In presence of base \rightarrow In^- blue

$$K_a = \frac{[H^+][In^-]}{[HIn]}$$

$$\frac{[HIn]}{[In^-]} = \frac{[H^+]}{K_a}$$

} so, color of indicator depends on pH or $[H^+]$ and K_a .

when $pH = pK_a$, the two colors have equal conc.

So different indicators change color over different pH ranges

e.g.	pK_a	pH range	Color in acid	Color in alkali
phenolphthaleine	9.6	8.3 - 10.0	less	Pink
Methyl orange	3.7	3.1 - 4.4	Red	Yellow

In order for an indicator to be effective, color change (called the end-point) must occur rapidly at the equivalence point

e.g.

<u>Property</u>	<u>"Phenolphthalein"</u>	<u>Methylorange</u>
pKa	9.6	3.7
pH range	8.3 - 10.0	3.1 - 4.4
Color in acid	Colorless	Red
" alkali	Pink	Yellow
Useful for	Titrations involving strong bases	Titrations involving strong acids

Regulatory Mechanism against changes in $[H^+]$ of Blood

Two types of metabolic acids produced

1. Volatile acids

The physiologically most important is Carbonic acid
20,000 mEq. is produced daily from metabolism

2. Fixed Acids [60-80 mEq. per day]

• Organic acids as:

Pyruvic, lactic, keto acids (e.g. acetoacetic, β -hydroxy butyric acid), uric acid

• phosphoric and sulphuric acids produced from sulphur & phosphorus of proteins, lipoproteins and nucleoproteins

Buffering Capacity depends on:

- Conc. of buffer
- pK_a of the buffer and the desired pH

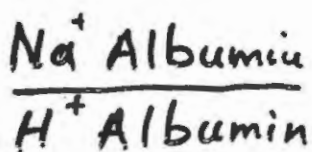
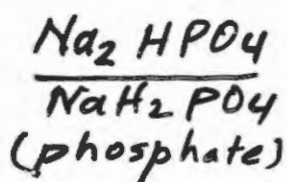
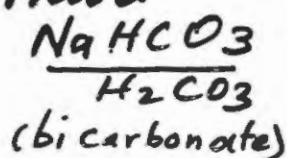
Mechanism of Regulation of pH^②

1. Buffer mechanism - First defence
2. Respiratory mechanism - second line of defence
3. Renal Mechanism - 3rd line of defence

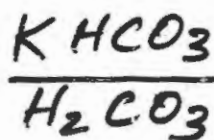
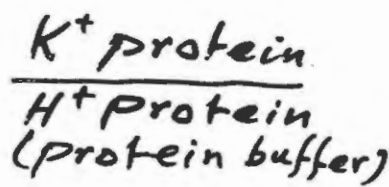
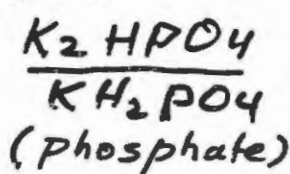
The first two lines of defence keep the $[H^+]$ from changing too much until the more slowly responding third line of defence, the kidneys, can eliminate the excess acid or base from the body

Buffer systems of the body

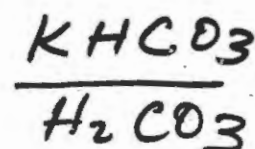
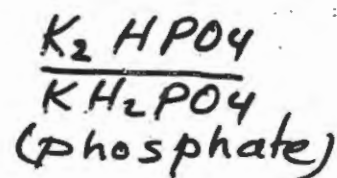
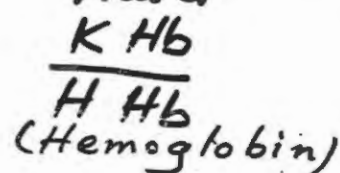
Extracellular fluid



Intracellular fluid



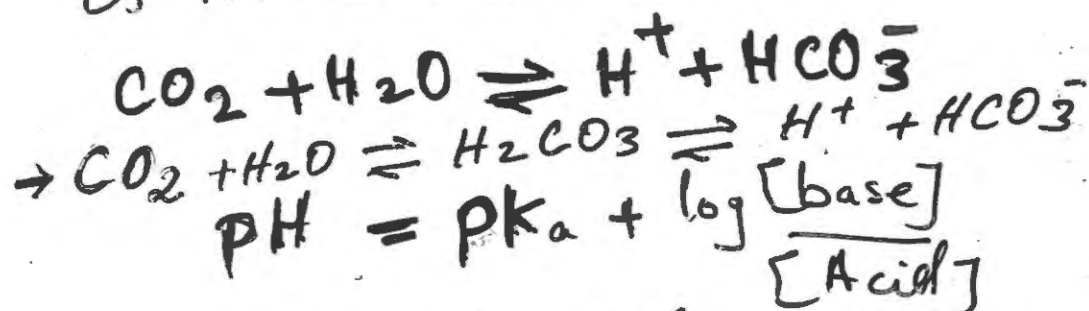
Erythrocyte fluid



(3)

Regulation of Blood pH Bicarbonate Buffer

When CO_2 is dissolved in H_2O , there is little H_2CO_3



Under physiological conditions

$$7.4 = 6.1 + \log \frac{[\text{HCO}_3^-]}{[\text{CO}_2]}$$

$$1.3 = \log \frac{[\text{HCO}_3^-]}{[\text{CO}_2]}$$

$$\frac{[\text{HCO}_3^-]}{[\text{CO}_2]} = \frac{20}{1}$$

} solubility
coefficient of CO_2
 0.03 mM/mmHg

Normal values

$$\text{pH} = 7.4$$

$$\text{PCO}_2 = 40 \text{ mm Hg } (\sim 1.2 \text{ mM})$$

$$[\text{HCO}_3^-] = 24 \text{ mM}$$

The Bicarbonate Buffer System ^(4a)

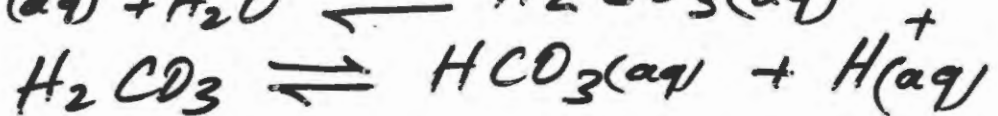
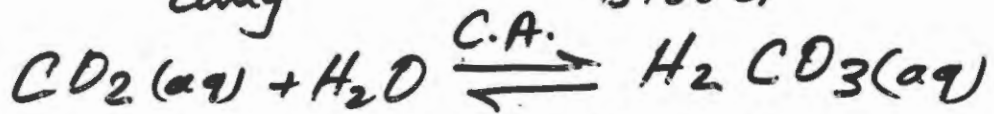
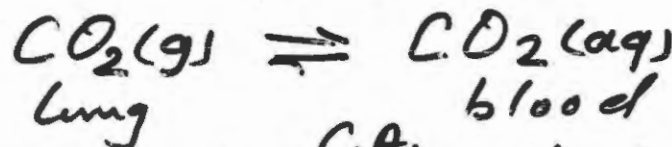
Under physiological conditions with plasma $\text{pH} = 7.4$; $\text{pK}_a = 6.1$

$$\frac{\text{HCO}_3^-}{\text{H}_2\text{CO}_3} = 20/1$$

Actual pH of blood is at the upper limit of buffering range of carbonic-bicarbonate buffer

$$6.1 \pm 1 = 5.1 - 7.1$$

Inefficiency is replenished by the reserve supply of gaseous CO_2 in lungs



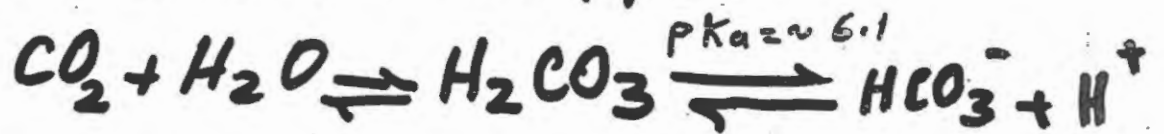
These reactions also work in reverse

Excess H^+ is removed by HCO_3^-

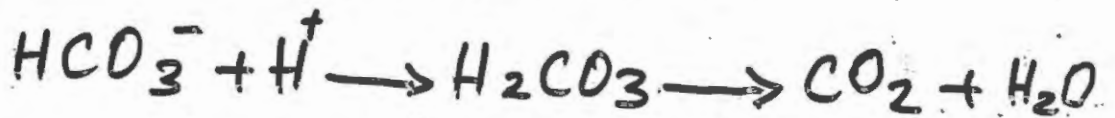


expired by
lung

Mechanism of Action of 46 Carbonate Buffer

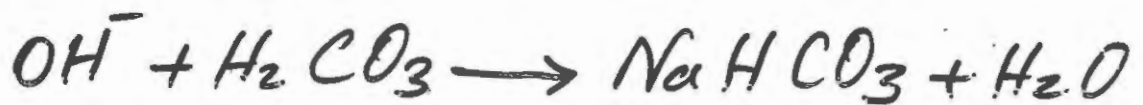


- Acid added to blood



Excess CO_2 greatly stimulate respiration which eliminate CO_2 from extracellular fluid

- Base added to the blood



$\text{CO}_2 + \text{H}_2\text{O} \xrightarrow{\text{C.A.}} \text{H}_2\text{CO}_3$ to replace H_2CO_3 in the above.

Decrease $[\text{CO}_2]$ decreases respiration rate to decrease rate of CO_2 expiration

HCO_3^- alkali reserve

$$\text{HCO}_3^- / \text{CO}_2 = \frac{25 \text{ mmole/l}}{1.25 \text{ mmole}} = \frac{20}{1}$$

Phosphate Buffer Systems:-

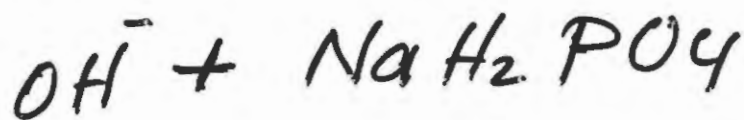
(4C)

$pK_a = 7.1 - 7.2$ very good

Considerable concentration in intracellular fluids and Tubular fluids of kidneys

In RBC, conc. of 2,3-BPG is 4-5 mmole/l is considerable conc \rightarrow 16% of non-carbonate buffer contribution

Action:



Protein Buffers

5

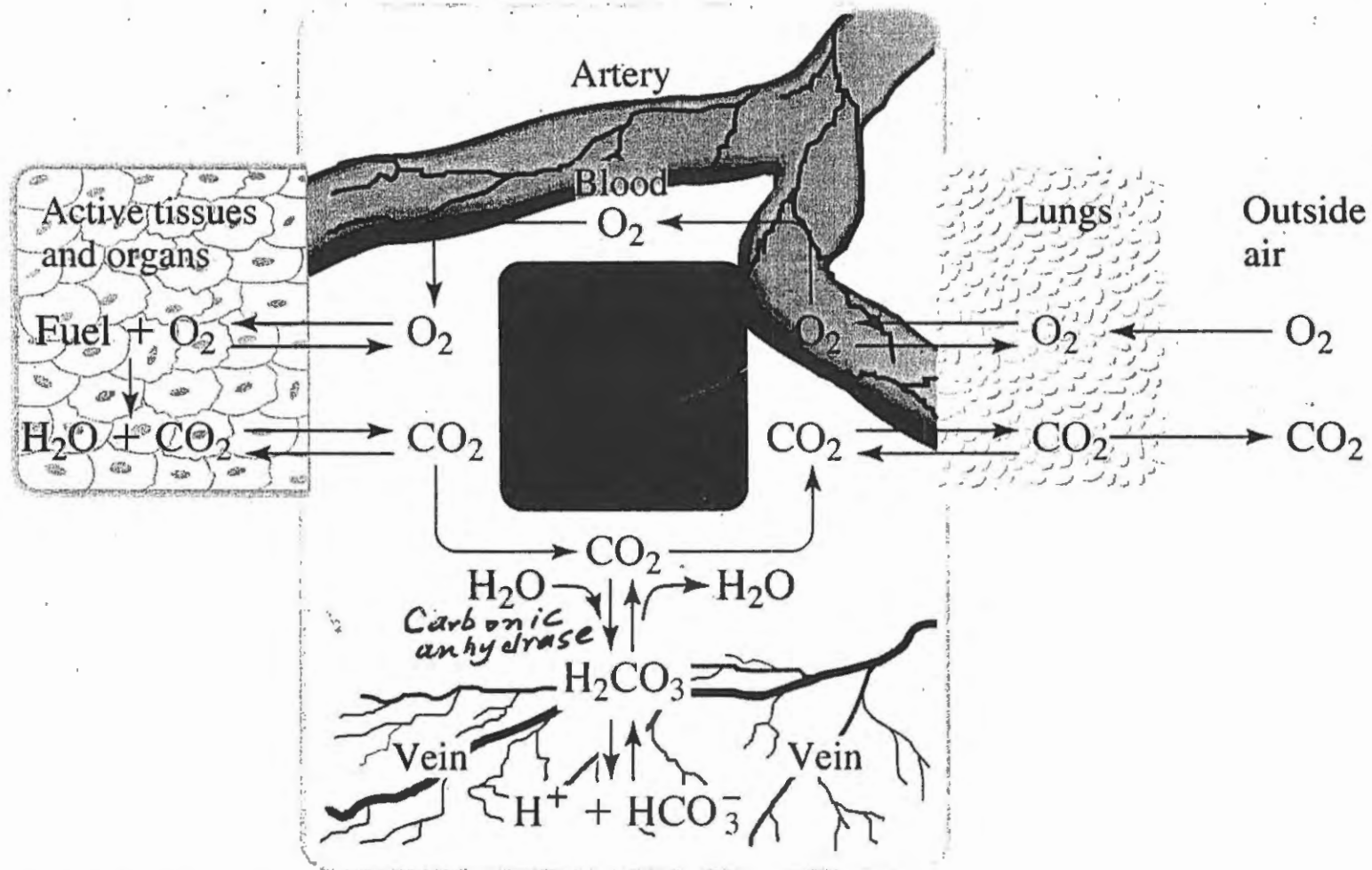
Proteins, especially ALBUMIN, accounts for 95% of non-carbonate buffering value in the Plasma

- Presence of dissociable acidic ($-COOH$) and basic ($-NH_2$) groups
- In particular the side chain of histidine (imidazole gr) having $pK_a = 7.3$
- Albumin contains 16 his/mole

Hemoglobin Buffer

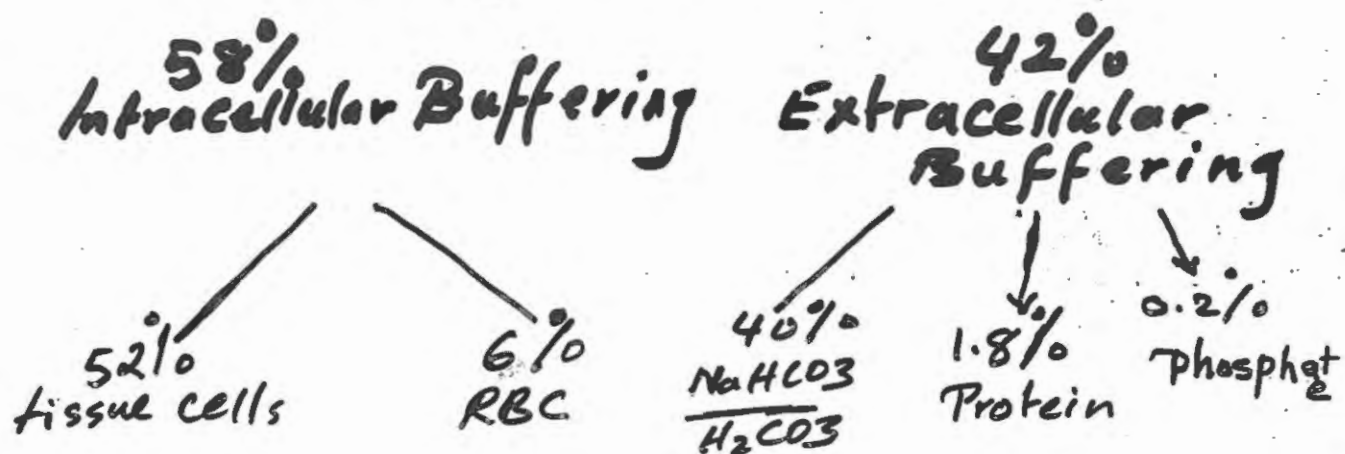
- major intracellular buffer of blood rbc
- Hb has a high conc of His (38 moles/mole of Hb)
- It buffers H_2CO_3 and CO_2
- It works in cooperation with the bicarbonate system

(more details in Blood + lymph system - 3rd yr)



Unnumbered figure pg 57c Concepts in Biochemistry, 3/e
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Relative Capacity of the Buffer Systems in the body



- Buffers act quickly but not permanently
 - They don't eliminate acids from body or replenish alkali reserve
 - Respiratory and Renal mechanism are very essential for final elimination
- Normal pH range: 7.38 — 7.42 (7.4)
- Acidosis
 - when $pH < 7.38$
 - but when $pH < 7.25 \rightarrow$ Life is threatened
 - Acidosis \rightarrow CNS depression and coma
 - when $pH < 7.0$ death occurs

Metabolic Acidosis: $\rightarrow \downarrow [\text{HCO}_3^-]$ ^⑦

Untreated diabetes } $\rightarrow \uparrow$ Ketone
Starvation diet } bodies
High protein diet } (Ketosis) $\rightarrow \uparrow [\text{H}^+]$
Low-fat diet }

Lactic acidosis, therapeutic administration of HCl
kidney disorder

Normal Metabolism

↓
Volatile acids



$\sim 20,000 \text{ mEq/day}$

excreted as CO_2 by
Lung

↓
Fixed Acids

e.g. lactic acid,
ketoacids, uric
acid, phosphoric
and sulphuric acids
 $60-80 \text{ mEq/day}$

Buffered and H^+
excreted by kidney

Respiratory Acidosis: $\uparrow [\text{CO}_2]$

Increased in $[\text{CO}_2]$ caused by

Pulmonary problems

e.g. chronic obstructive airway disease, asthma,
emphysema + pneumonia; cardiac arrest,
severe hypoxia — etc.

(8)

Alkalosis

$\text{PH} > 7.42$

$\text{PH} > 7.55$ is dangerous
 > 7.60 death

induces muscular hyperexcitability
and tetany

• Metabolic alkalosis

Results from clinical administration
in excess of alkali (e.g. NaHCO_3)

severe vomiting (loss of gastric juices)

Hypokalaemia (low cellular K^+)

→ $\uparrow [\text{HCO}_3^-]$

• Respiratory alkalosis

Hyperventilation $\downarrow \text{PCO}_2$
(heavy breathing)

Hysteria

Anxiety

altitude sickness

hot baths, working at high temp.

T-94

Figure 10.6 Breathing and the bicarbonate buffer system

