

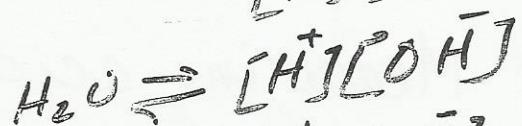
Ionization of H_2O , weak acid & its law:-

Ionization of $H_2O \rightarrow H^+ + OH^-$.

Equilibrium const.



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



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

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

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

$$[H^+] \times [OH^-] = 10^{-14} M^2$$

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

Cone. of $[H^+]$ in solution of $0.1M NaOH$

$$\text{e.g. } \text{cone. of } [H^+] = 10^{-14} M = 10^{-14} M$$

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

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

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

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

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

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

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

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

PH & P OH

$$P\text{OH} = -\log[\text{OH}^-] \quad -14$$

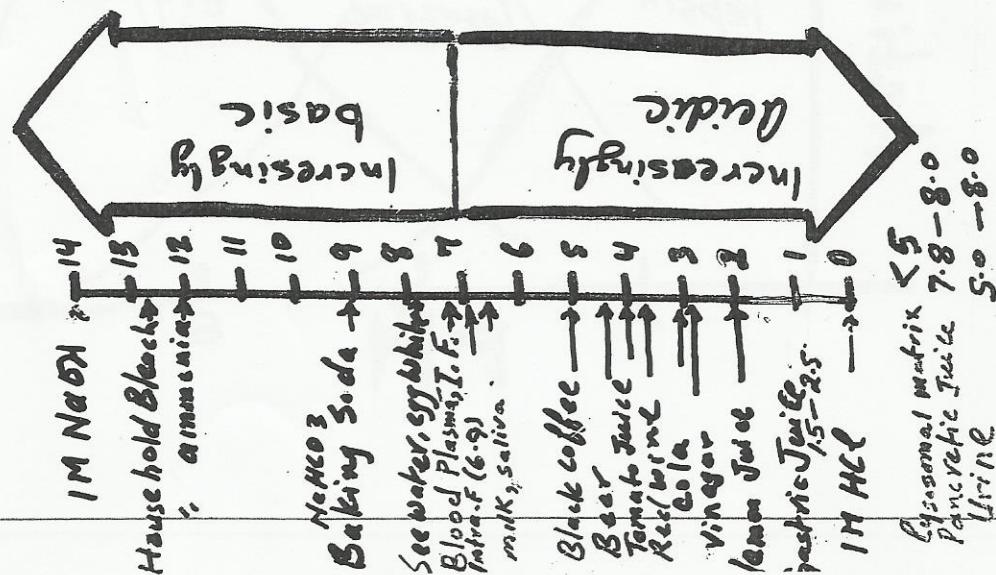
$$-\log[\text{H}^+] \times [\text{OH}^-] = -\log 10$$

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

$$\text{PH} + \text{P OH} = 14$$

$\frac{[\text{H}^+]}{M}$	<u>PH</u>	$\frac{[\text{OH}^-]}{M}$	<u>P OH</u>
1.0	0	10^{-14}	14
0.1	1	10^{-13}	13
0.01	2	10^{-12}	12
0.001	3	10^{-11}	11
10^{-4}	4	10^{-10}	10
10^{-5}	5	10^{-9}	9
10^{-6}	6	10^{-8}	8
10^{-7}	7	10^{-7}	7
10^{-8}	8	10^{-6}	6
10^{-9}	9	10^{-5}	5
10^{-10}	10	10^{-4}	4
10^{-11}	11	10^{-3}	3
10^{-12}	12	10^{-2}	2
10^{-13}	13	10^{-1}	1
10^{-14}	14	1	0

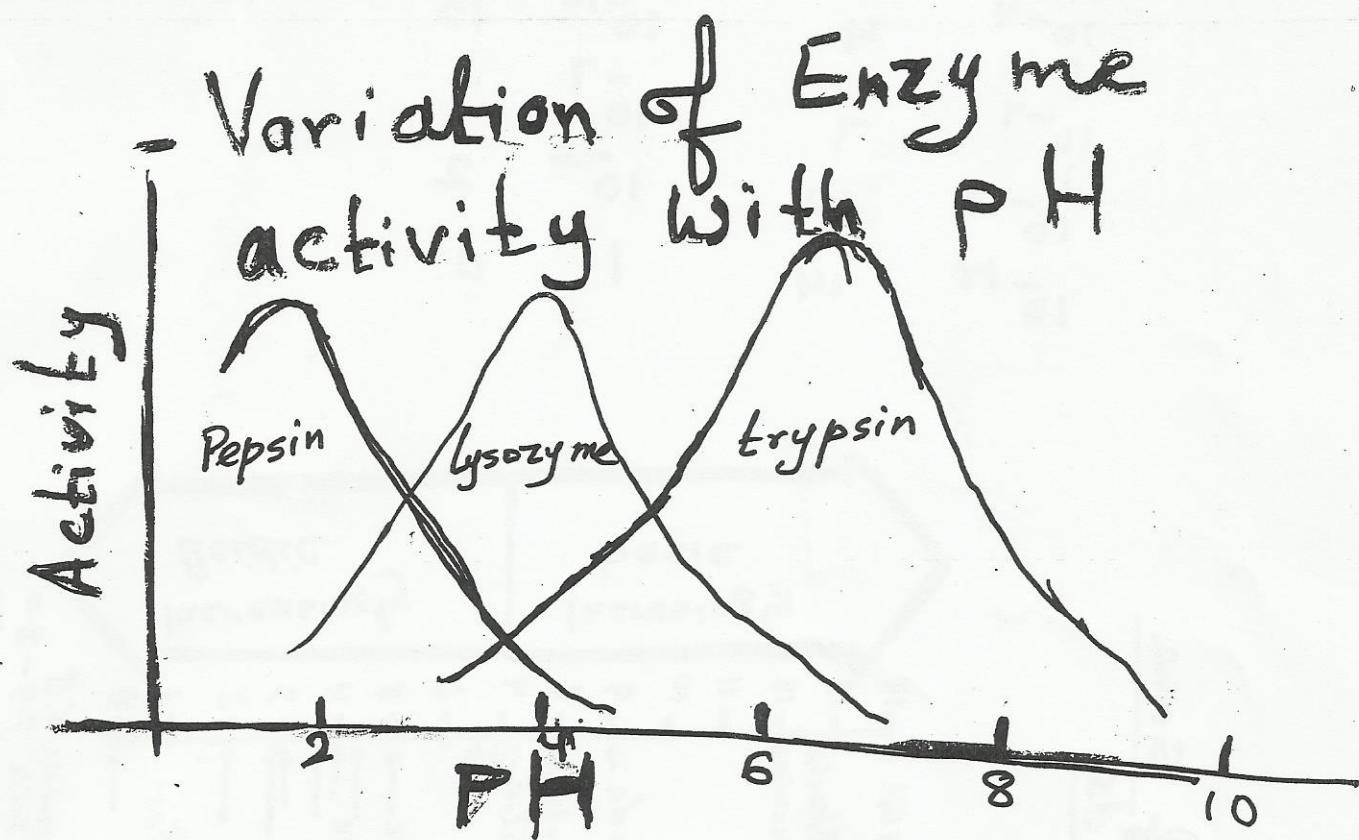
PH of Some Fluids



3.

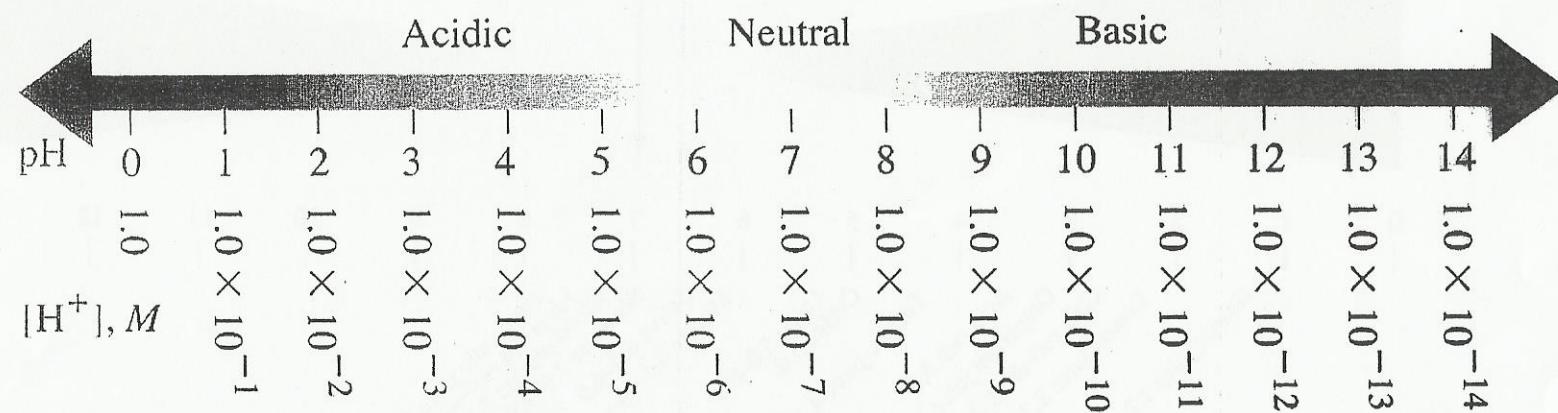
Importance of pH Regulation

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



b7

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

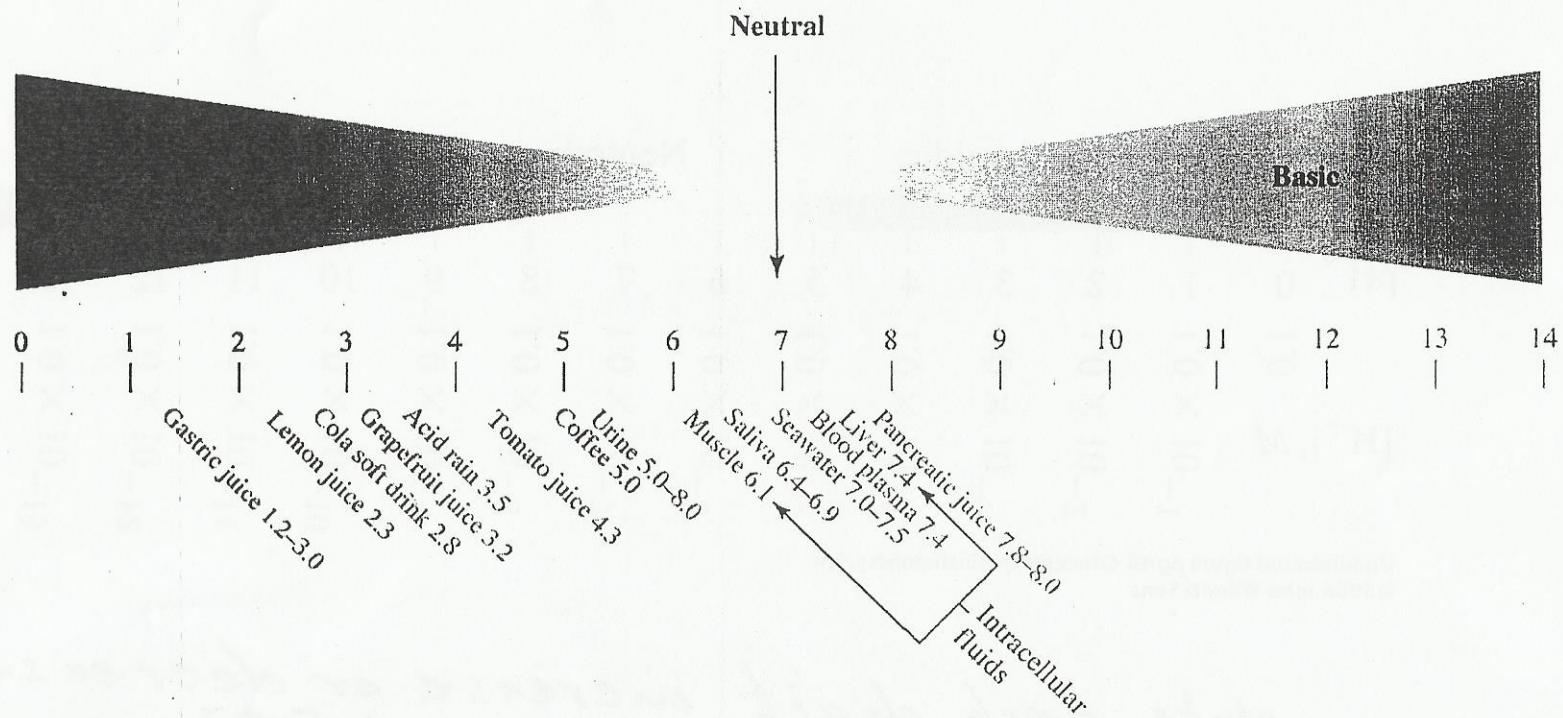
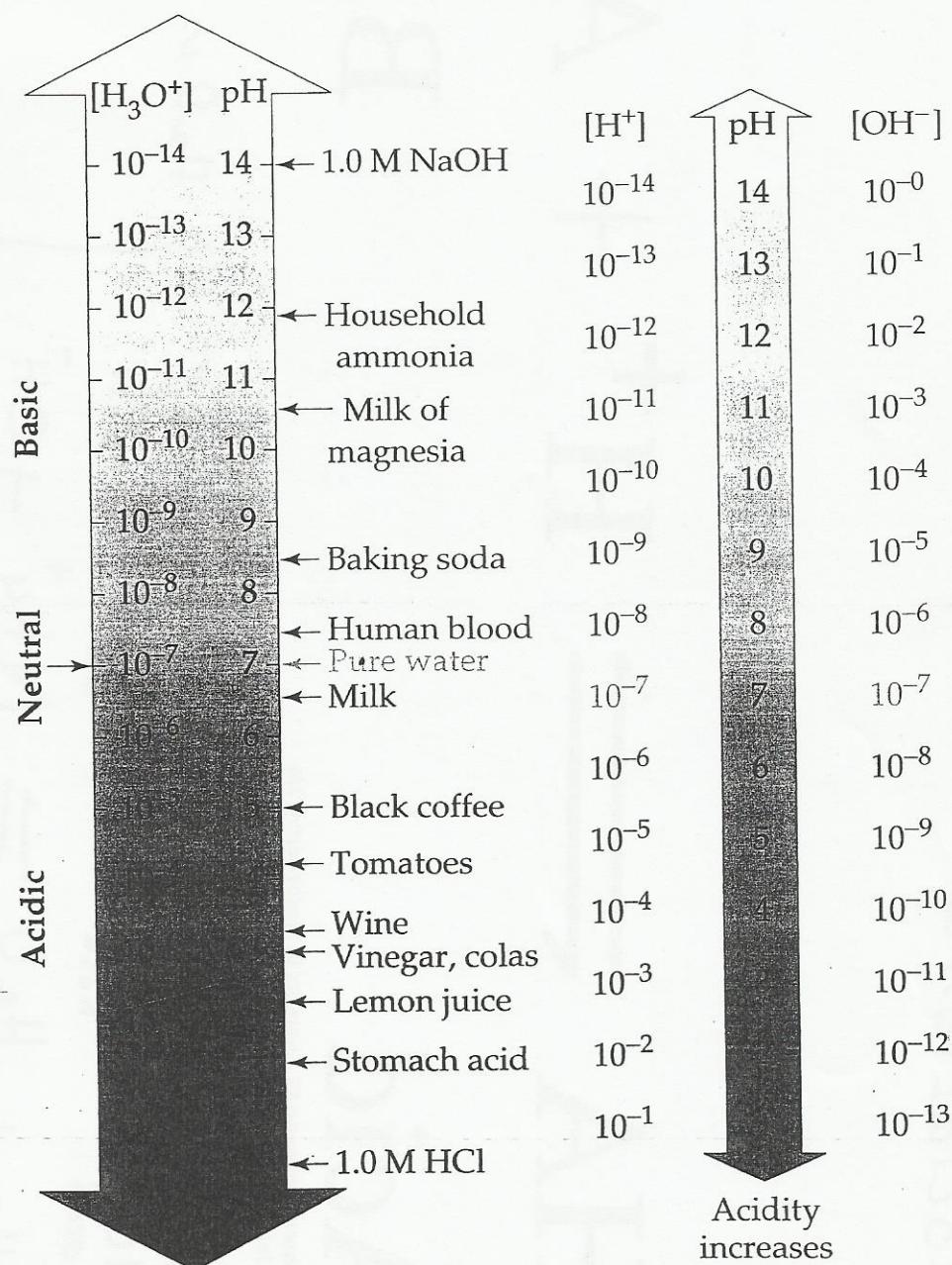


Figure 2-9 Concepts in Biochemistry, 3/e
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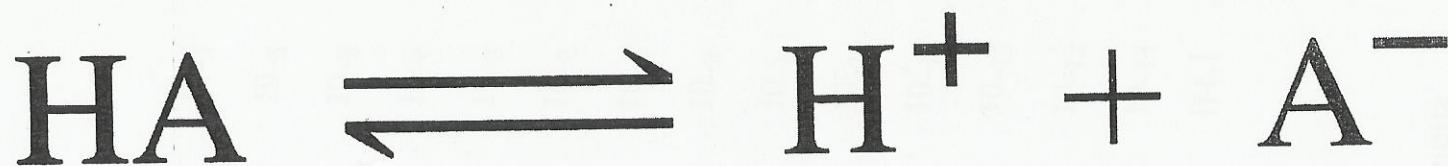
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**Figures 10.1, 10.2 The pH scale and the pH of some common substances
The relationship of pH scale to H⁺ and OH⁻ concentrations**



6

Dissociation of Acid in Aqueous Solution

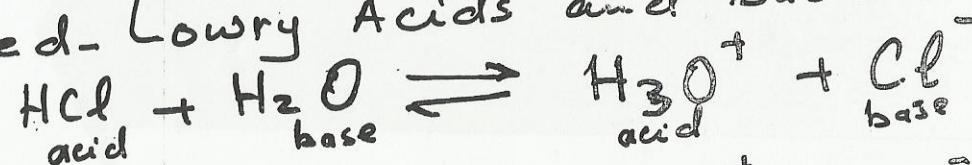


Acid

Base

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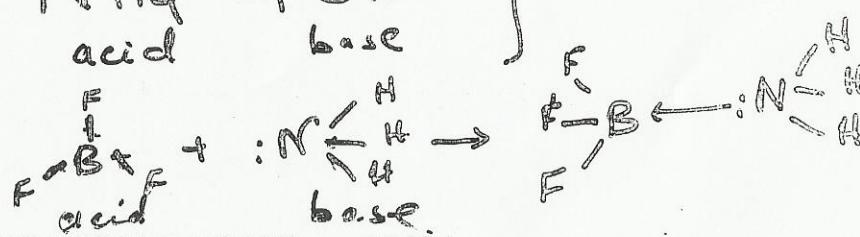
Bronsted-Lowry Acids and Bases



H_2O is amphotropic



Lewis Acids and Bases



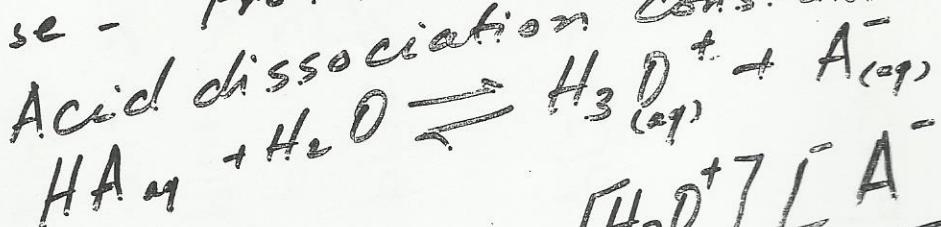
ACIDS & BASES

④

Acid - proton donor

Base - Proton acceptor

Acid dissociation constant



$$K_{\text{eq. const}} = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}][\text{H}_2\text{O}]}$$

$$\text{Dissociation constant: } K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

<u>Acid</u>	$\frac{K_a}{M}$	pKa
HCOOH	1.78×10^{-4}	3.95
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}	3.97
HCO ₃ ⁻	6.31×10^{-11}	10.2
NH ₄ ⁺	5.62×10^{-10}	9.25

$$\text{pK}_a = \log \frac{1}{K_a} = -\log K_a$$

$$\text{pH} = \log \frac{1}{[\text{H}_3\text{O}^+]} = -\log [\text{H}^+]$$

Acid	Structur ^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 -CH(OH)-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

<u>Acid</u>	<u>Structure^a</u>	<u>K_a</u>	<u>pK_a</u>
- Phosphoric acid	$\text{HO}-\overset{\text{O}}{\underset{\text{ }}{\text{P}}}-\text{OH}$ OH	(1) 7.25×10^{-3} (2) 6.31×10^{-8} (3) 3.98×10^{-13}	2.14 7.20 12.40
- Ammonium Ion	$\text{H}-\overset{\text{H}}{\underset{\text{ }}{\underset{\text{ }}{\text{N}}}}^+-\text{H}$ H	5.62×10^{-10}	9.25
- H_2SO_4		Large 1.2×10^{-2}	
- HSO_4^-			
- HCl		Large	
- HNO_3		Large	

Examples :-

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

9a

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

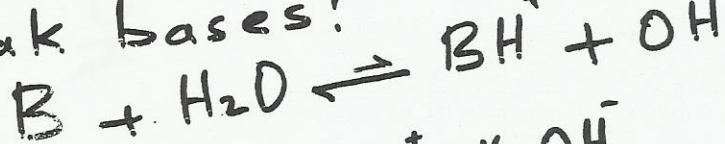
$$\text{pOH} = 1.1$$
$$\text{pH} = 14 - 1.1 = 12.9$$

- pH of 0.02 M weak acid (HA) is

4.0. Find K_a

$$K_a = \frac{[\text{H}^+] \times [\text{A}^-]}{[\text{HA}]} = \frac{[\text{H}^+]^2}{\text{HA}} = \frac{10^{-4} \times 10^{-4}}{0.02} = 10^{-8} \text{ pH}$$
$$= 5 \times 10^{-7} \text{ M} \quad [\text{H}^+] = 10^{-4}$$

Weak bases:



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

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

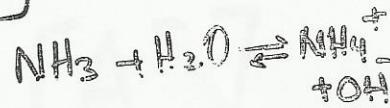
Reverse reaction
for $\text{BH}^+ \rightleftharpoons \text{B} + \text{H}^+$

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

Example

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

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



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

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

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

$$PH = 14 - 3.37 = 10.6$$

Example

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

$$POH = 14 - 10 = 4$$

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



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

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