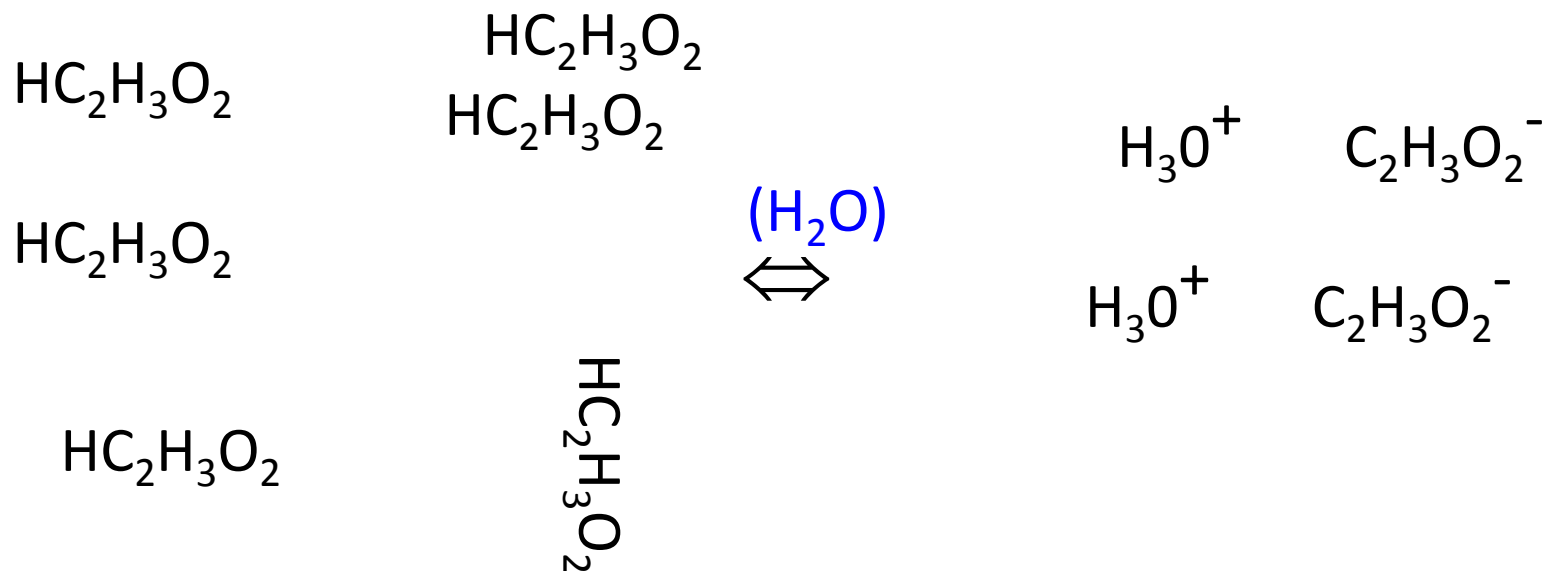


Weak Acid Ionization:



At any given time only a small portion of the acid molecules are ionized and since reactions are running in **BOTH** directions the mixture composition stays the same.

This gives rise to an Equilibrium expression, K_a

The Henderson-Hasselbalch Equation Describes the Behavior of Weak Acids & Buffers

The Henderson-Hasselbalch equation is derived below.
A weak acid, HA, ionizes as follows:



The equilibrium constant for this dissociation is

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

Cross-multiplication gives

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

Divide both sides by $[\text{A}^-]$:

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

Take the log of both sides:

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

Multiply through by -1 :

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

Substitute pH and $\text{p}K_a$ for $-\log [\text{H}^+]$ and $-\log K_a$, respectively; then:

$$\text{pH} = \text{p}K_a - \log \frac{[\text{HA}]}{[\text{A}^-]}$$

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

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

The Henderson-Hasselbalch equation has great predictive value in protonic equilibria. For example,

- (1) When an acid is exactly half-neutralized, $[A^-] = [HA]$. Under these conditions,

$$\text{pH} = \text{p}K_a + \log \frac{[A^-]}{[HA]} = \text{p}K_a + \log \frac{1}{1} = \text{p}K_a + 0$$

Therefore, at half-neutralization, $\text{pH} = \text{p}K_a$.

- (2) When the ratio $[A^-]/[HA] = 100:1$,

$$\begin{aligned}\text{pH} &= \text{p}K_a + \log \frac{[A^-]}{[HA]} \\ \text{pH} &= \text{p}K_a + \log 100 / 1 = \text{p}K_a + 2\end{aligned}$$

- (3) When the ratio $[A^-]/[HA] = 1:10$,

$$\text{pH} = \text{p}K_a + \log 1/10 = \text{p}K_a + (-1)$$

- Similarly for the weak bases :

$$\text{pOH} = \text{pKb} + \log \frac{[\text{M}^+]}{[\text{MOH}]}$$

Buffers

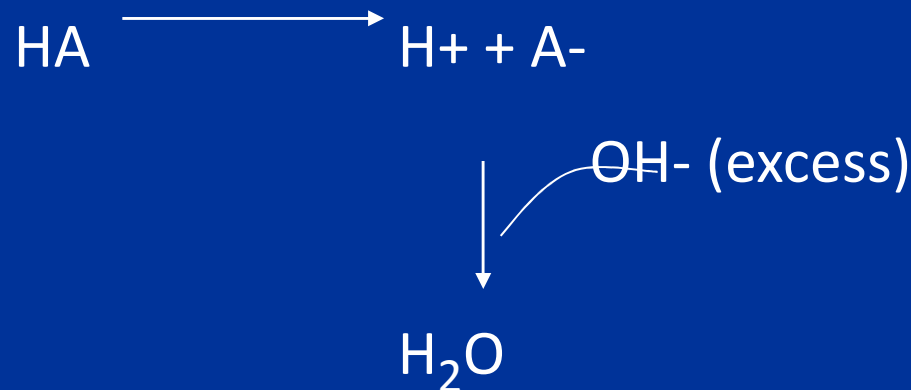
- The buffer solution has the ability to resist the changes in pH following the addition of limited amounts of acids or bases .
- Chemically buffer solutions are made up of a weak acids or bases and their conjugated salts.
- Two types :
 - 1- Acidic buffers which contain a weak acid and its salt with strong base . eg. Acetic acid +sodium acetate .
 - 2- Basic buffers which contain a weak base and its salt with a strong acid .

How buffers resist change in pH

Buffer: A chemical system that resists changes in pH when excess acid or base is added. Made usually from weak acid.

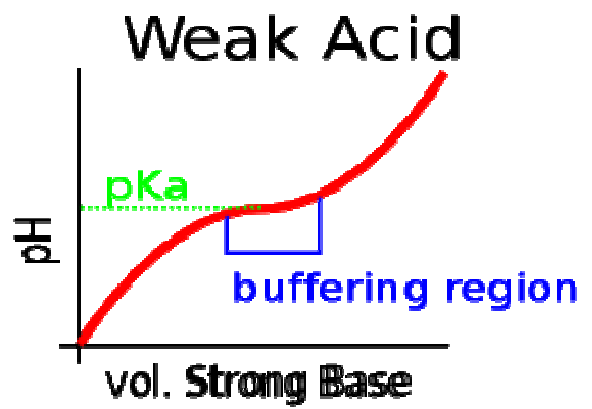
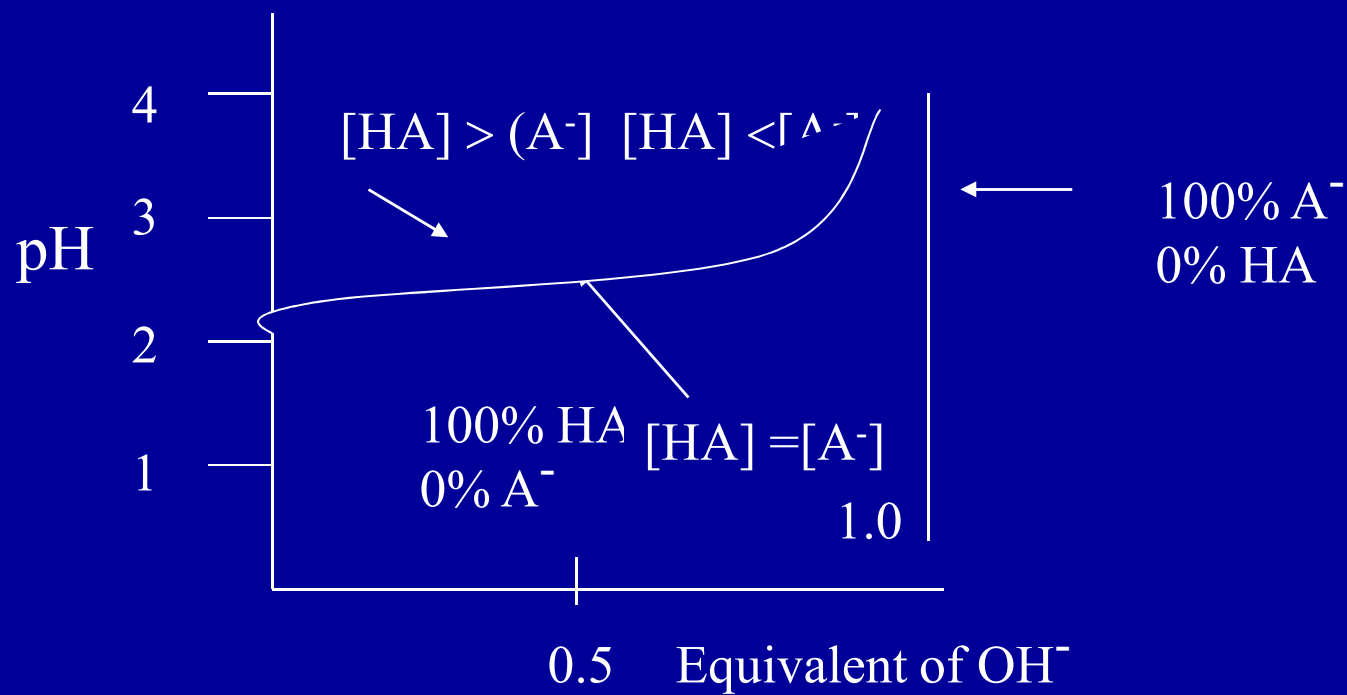
Weak acid (HA)

- Dissociate H A \rightleftharpoons H⁺ + A⁻ (slightly)
- Most HA and only small H⁺ + A⁻
- Add some OH⁻ (strong base)



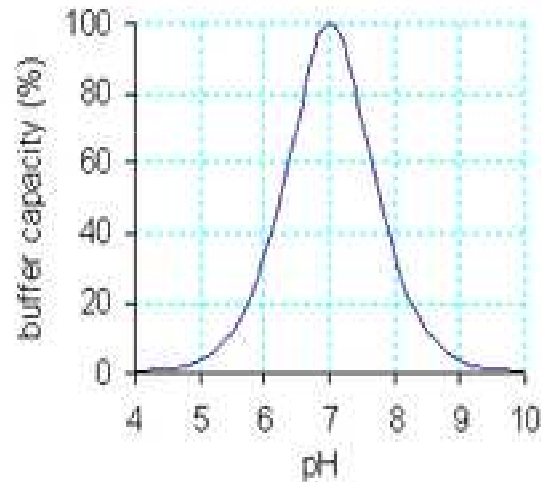
- OH^- combines with H^+ to form H_2O (Neutralization)
- This disturbs the equilibrium and causes more HA to dissociate $\text{H}^+ + \text{A}^-$.
- Newly generated H^+ reacts with OH^- to form more H_2O . Process continues till all OH^- neutralised.
- $\text{HA} \downarrow, \text{A}^- \uparrow$. If enough HA present all OH^- will be neutralised and little or no change in pH occurs.

Titration curve for a weak acid:



Buffer capacity :

Buffer capacity is a quantitative measure of the resistance of a buffer solution to pH change on addition of hydroxide ions.



Buffer capacity for $pK_a=7$ as percentage of maximum

Buffer capacity of a weak acid reaches its maximum value when $pH = pK_a$

Buffer mixtures

Components	pH range
Citric acid, Sodium citrate	2.5 - 5.6
Acetic acid, Sodium acetate	3.7 - 5.6
K_2HPO_4, KH_2PO_4	5.8 - 8
Na_2HPO_4, NaH_2PO_4	6 - 7.5
Borax, Sodium hydroxide	9.2 - 11

- The strength of weak acids is expressed by *pKa*, the negative log of the acid dissociation constant. Strong acids have low *pKa values* and weak acids have high *pKa values*.
- Buffers resist a change in pH when protons are produced or consumed. Maximum buffering capacity occurs ± 1 pH unit on either side of *pKa*.
- *Physiologic* buffers include bicarbonate, orthophosphate, and proteins.

- Since many metabolic reactions are accompanied by the release or uptake of protons, most intracellular reactions are buffered.**
- Oxidative metabolism produces CO₂, the anhydride of carbonic acid, which if not buffered would produce severe acidosis.**
- Maintenance of a constant pH involves buffering by phosphate, bicarbonate, and proteins, which accept or release protons to resist a change in pH .**
- For experiments using tissue extracts or enzymes, constant pH is maintained by the addition of buffers such as *inorganic orthophosphate (pKa2 7.2)*, or *Tris (tris[hydroxymethyl] aminomethane, pKa 8.3)*. The value of pKa relative to the *desired* pH is the major determinant of which buffer is selected.**

Functional Groups That Are Weak Acids Have Great Physiologic Significance

Many biomolecules possess functional groups that are weak acids or bases.

Carboxyl groups, amino groups, and the second phosphate dissociation of phosphate esters present in proteins and nucleic acids, most coenzymes, and most intermediary metabolites.

Knowledge of the dissociation of weak acids and bases thus is basic to understanding the influence of intracellular pH on structure and biologic activity.

The protonated species (eg, HA or R—NH₃⁺) are **acid and the unprotonated species (eg, A⁻ or R—NH₂) its conjugate base. Similarly, we may refer to a base (eg, A⁻ or R—NH₂) and its conjugate acid (eg, HA or R—NH₃⁺).**

Representative weak acids (left), their conjugate bases (center), and the *pK_a values* (right) include the following :



$$pK_a = 4 - 5$$

$$pK_a = 9 - 10$$

$$pK_a = 6.4$$

$$pK_a = 7.2$$

Their resistance to changes in pH makes buffer solutions very useful for chemical manufacturing and essential for many biochemical processes.

The ideal buffer for a particular pH has a pK_a equal to that pH, since such a solution has maximum buffer capacity. Buffer solutions are necessary to keep the correct pH for enzymes to work. Many enzymes work only under very precise conditions; if the pH changes too far out of the margin, the enzymes slow or stop working and can denature, thus permanently disabling its catalytic activity. A buffer of carbonic acid (H_2CO_3) and bicarbonate (HCO_3^-) is present in blood plasma, to maintain a pH between 7.35 and 7.45.

Relationship between $\text{HCO}_3^- / \text{H}_2\text{CO}_3$ ratio and pH

$\text{HCO}_3^- / \text{H}_2\text{CO}_3$	pH	State
50/1	7.8	Death
40/1	7.7	Alkalosis
32/1	7.6	“
25/1	7.5	“
20/1	7.4	Normal
16/1	7.3	Acidosis
12/1	7.2	“
10/1	7.1	“
8/1	7.0	“
6.25	6.9	“
5.1	6.8	“