



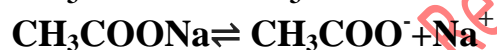
Lect.6.

Buffer Solutions

A buffer solution defined : is one in which the pH of the solution is "resistant" to small additions of either a strong acid or strong base.

Buffers usually consist of a weak acid and its conjugate base, in relatively equal and "large" quantities. Calculations are based on the equation for the ionization of the weak acid in water forming the hydronium ion and the conjugate base of the acid. "HA" represents any weak acid and "A⁻" represents the conjugate base.

A mixture of acetic acid and sodium acetate ,since acetic acid is weak.



Henderson – Hasselbalch equation

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

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

$$\text{PH} = -\text{Log}[\text{H}^+] = -\text{Log} K_a - \text{Log} \left(\frac{[\text{HA}]}{[\text{A}^-]} \right)$$

$$\text{PH} = \text{P}k_a + \text{Log} \left(\frac{[\text{A}^-]}{[\text{HA}]} \right)$$

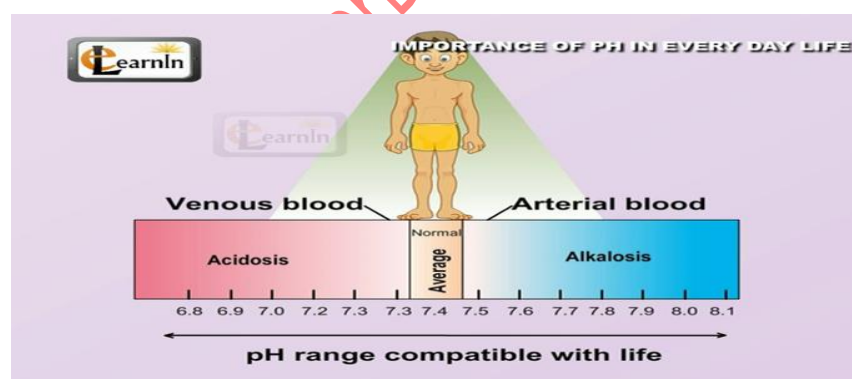
* Buffers are important in chemistry, when the PH must be used in many reactions.

In biochemistry when enzymes or proteins are being studied.

Our blood is buffered to a PH 7.4. PH units cause illness or death,

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$$

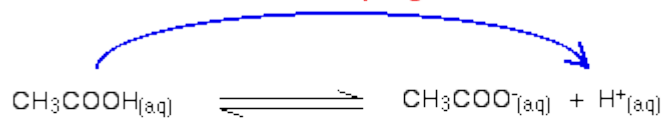
PH 7.4 in human.





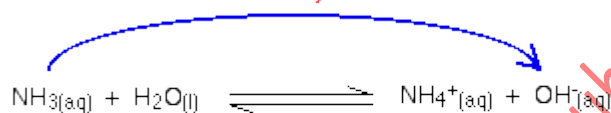
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Equilibrium moves to replace
the removed hydrogen ions.



Hydroxide ions combine
with these to make water.

Equilibrium moves to replace
the removed hydroxide ions.



Hydrogen ions combine
with these to make water.

Assist professor Dr. Reem Suhail Najm 2023/12/18



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Example 13

K_a for ethanoic acid is $1.74 \times 10^{-5} \text{ mol dm}^{-3}$, Remember that we want to calculate the pH of a buffer solution containing 0.10 mol dm^{-3} of ethanoic acid and 0.20 mol dm^{-3} of sodium ethanoate.

$$K_a = 1.74 \times 10^{-5} \text{ mol dm}^{-3}$$

pH=?

Buffer solution = 0.10 mol dm^{-3}

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$$

$$1.74 \times 10^{-5} = \frac{0.20 \times [\text{H}^+]}{0.10}$$

$$\begin{aligned} [\text{H}^+] &= 1.74 \times 10^{-5} \times \frac{0.10}{0.20} \\ &= 8.7 \times 10^{-6} \text{ mol dm}^{-3} \end{aligned}$$

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Buffer Capacity

Buffer capacity quantifies the ability of a solution to resist changes in pH by either absorbing or desorbing H⁺ and OH⁻ ions.

When an acid or base is added to a buffer system, the effect on pH change can be large or small, depending on both the initial pH and the capacity of the buffer to resist change in pH. Buffer capacity (β) is defined as the moles of an acid or base necessary to change the pH of a solution by 1, divided by the pH change and the volume of buffer in liters; it is a unitless number. A buffer resists changes in pH due to the addition of an acid or base through consumption of the buffer. As long as the buffer **has not been completely reacted, the pH will not change drastically. The pH change will increase (or decrease) more drastically as the buffer is depleted: it becomes less resistant to change.**

$([HA] \setminus [A])=1$

Number of moles of OH or H₃O-added
Buffer Capacity = $\frac{\text{Number of moles of OH or H}_3\text{O-added}}{\text{PH changes X Volume of buffer in L}}$