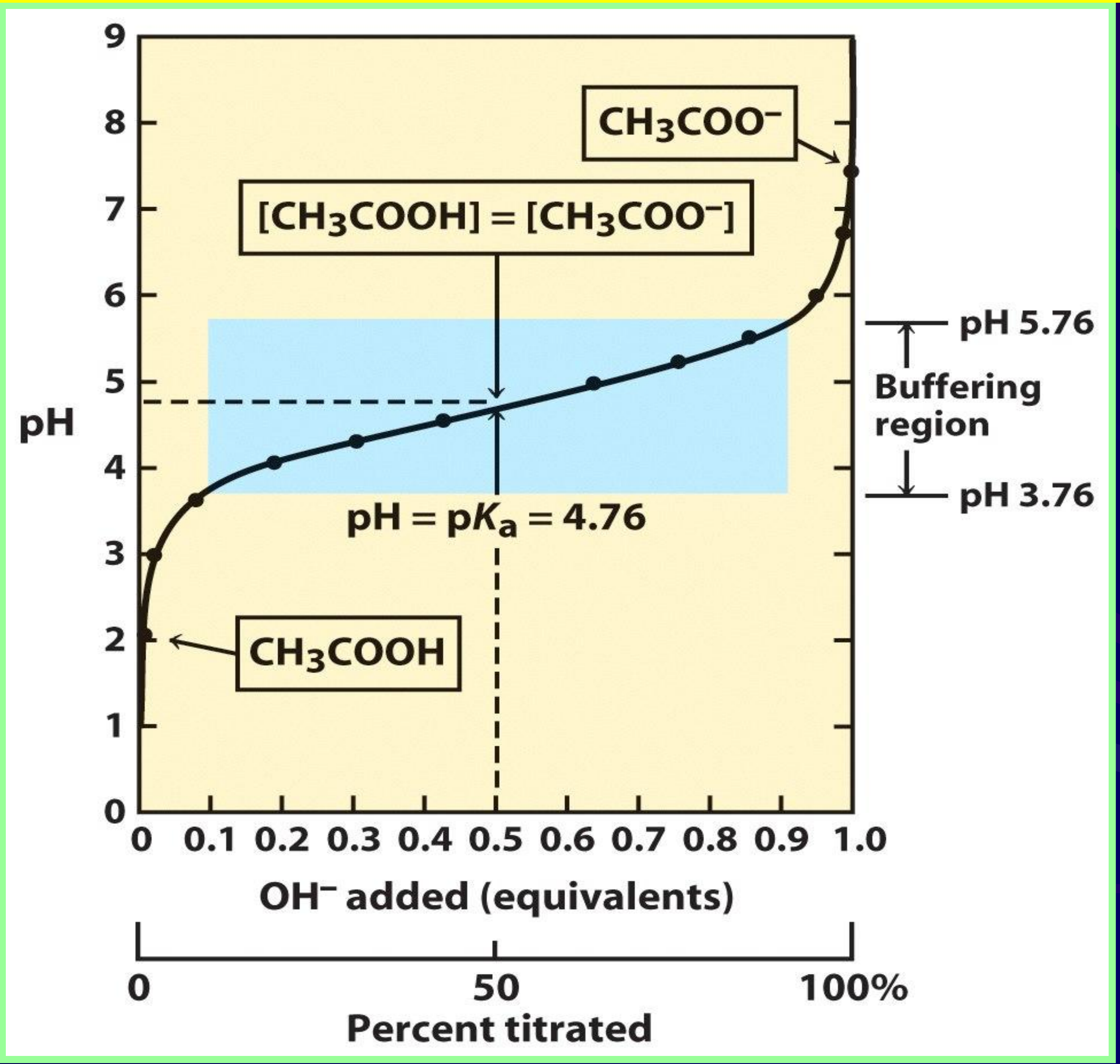
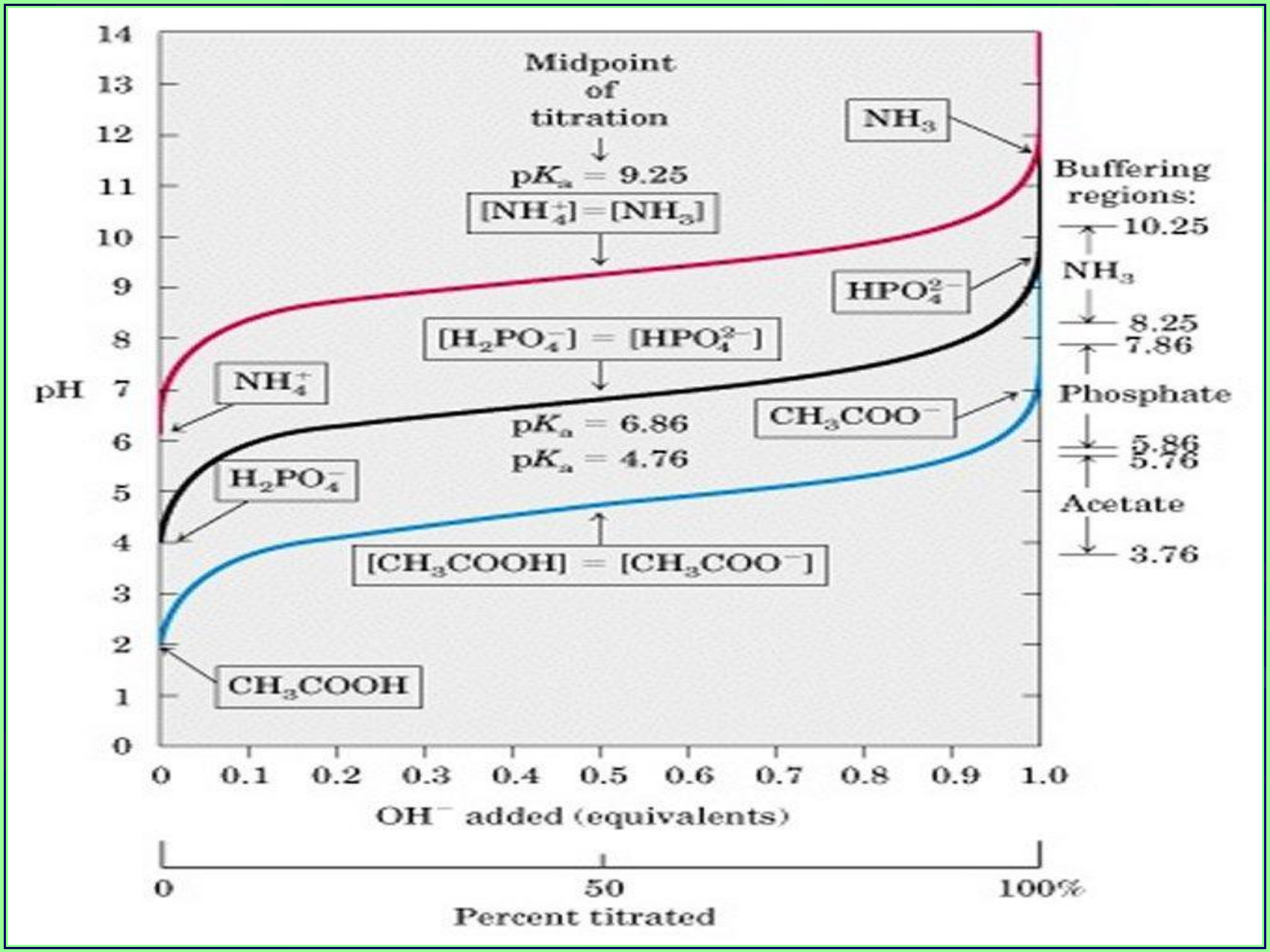
Dr.Sulieman Al-Khalil

- pH = -log [H^+] .
- Strong Acid .
- Strong base .
- Weak Acid.
- Weak Base.
- "**Conjugate Acid – Base Pair**" → a proton donor & its corresponding proton acceptor .
- Stronger acid → lower pH.
- Stronger Base → higher pH.
- pKa can be determined experimentally ; it is the midpoint of the titration curve for the acid or base.

Titration of acetic acid



Comparison of the Titration curves of Three weak acids.



Buffering Against pH Changes on Biological Systems

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

which is the same as:

$$[H^+] = (k_A)\left(\frac{1}{[A^-]}\right)([HA])$$

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

or

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

or

$$pOH = pk_B + \log \frac{[\text{conjugate cation}]}{[\text{base}]}$$

or

$$pOH = pk_B - \log \frac{[\text{base}]}{[\text{conjugate cation}]}$$

- Handerson-Hasselbach Equation

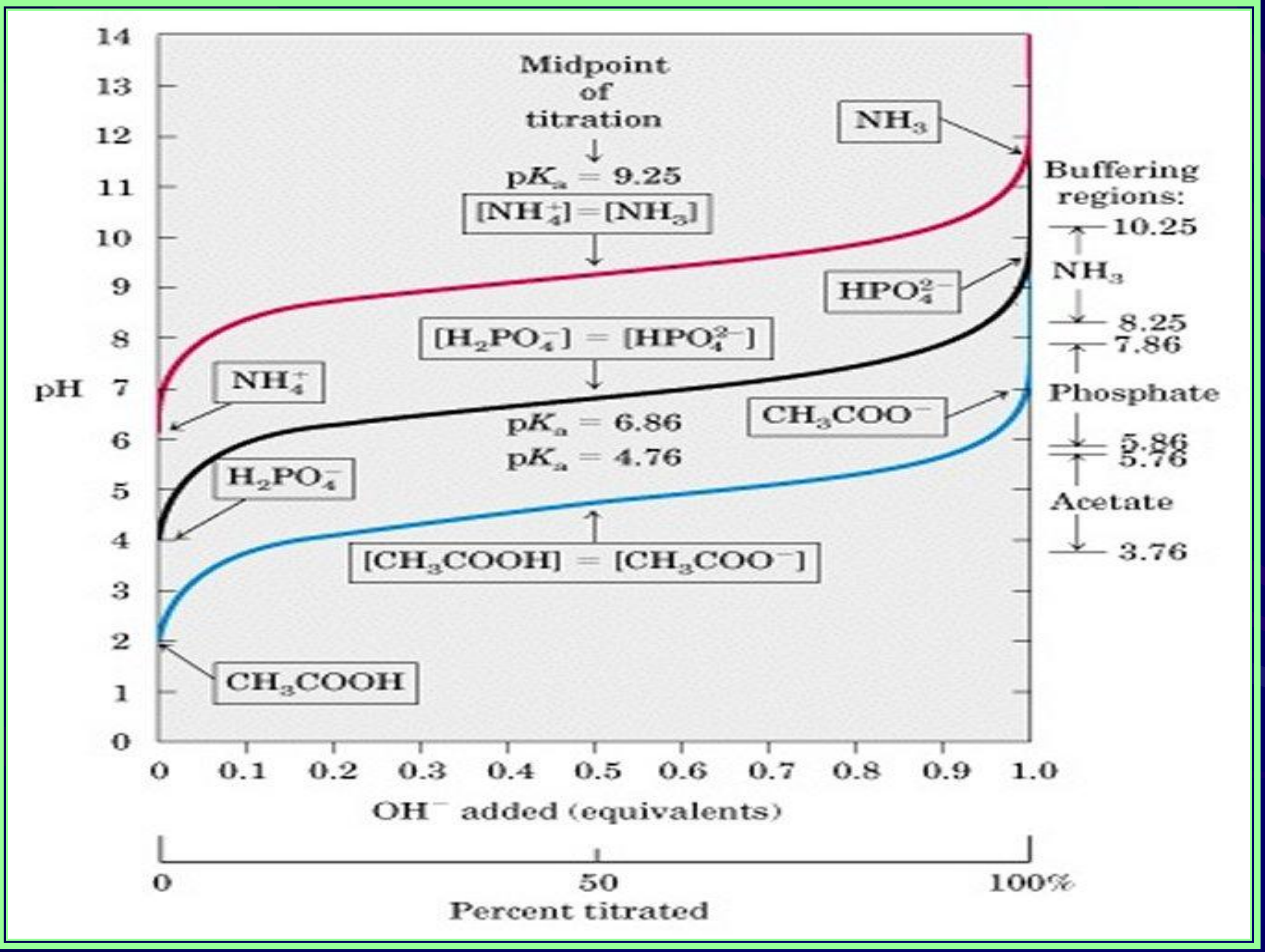
- Calculate the pKd of lactic acid , given that when [Lactic acid] = 0.010 M & [Lactate] = 0.087 M , pH = 4.80 ?

- Calculate pH mixture of 0.1 M lactic acid & 0.20 M Sodium acetate . The pka of Acetic acid = 4.76 ?
- Calculate the ratio of conc. of acetate & acetic acid required in a buffer system of pH = 5.30 ?

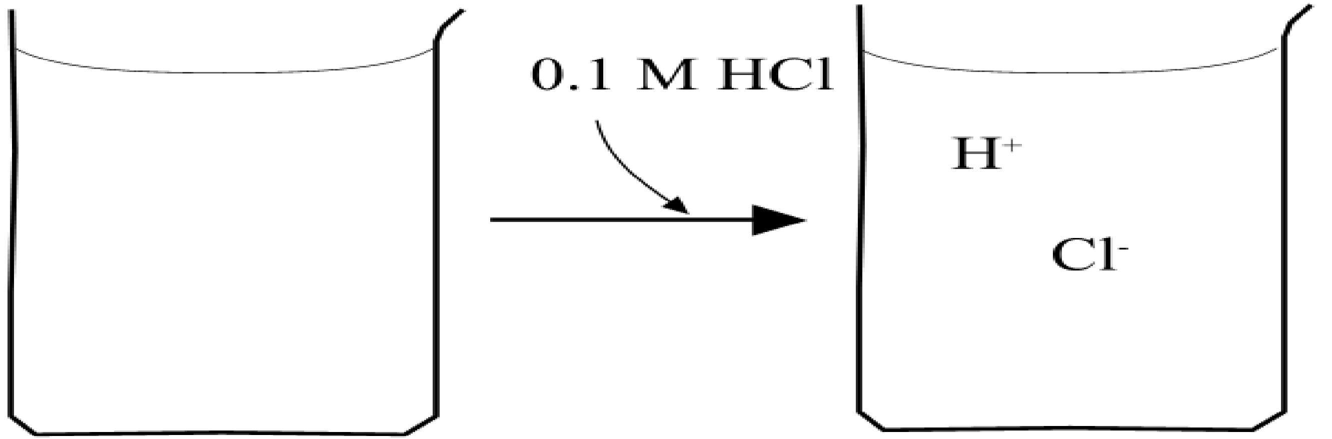
Buffers

- Buffer : is a tendency of solution to resist effectively any changes in pH in the body following the addition of strong acid or strong base .
- Most important buffers are :
  - Phosphate buffer.
  - Bicarbonate buffer .

Comparison of the Titration curves of Three weak acids.



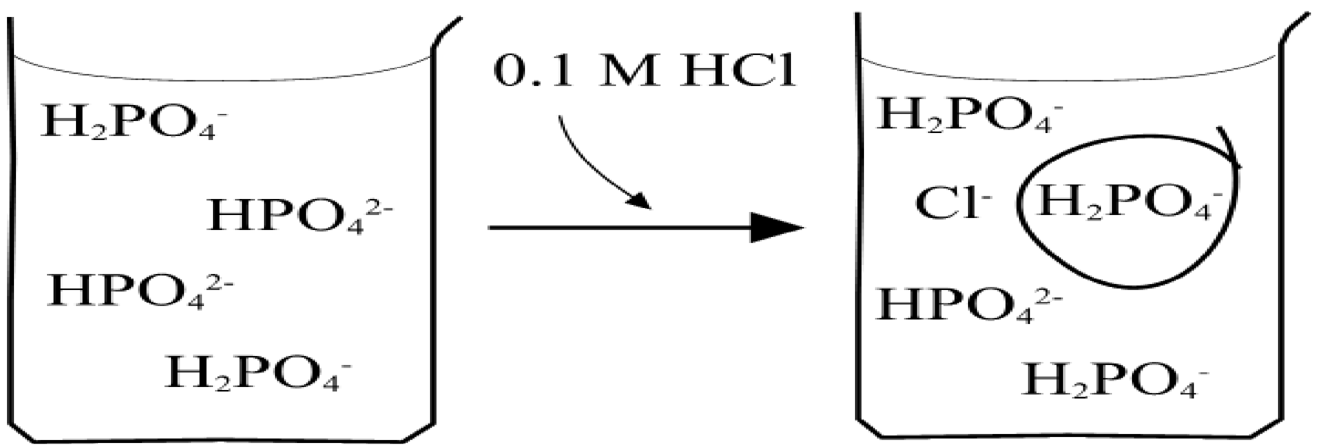
Buffers in Action--A Simple Example



pH = 7.0

pH = -log (0.1) = 1

Prepare 0.4 M phosphate buffer, pH 6.8, by making the solution 0.2 M $H_2PO_4^-$ and 0.2 M HPO_4^{2-}



$$pH = 6.8 + \log(0.2/0.2) = 6.8 + 0 = 6.8$$

$$pH = 6.8 + \log(0.1/0.3) = 6.8 - 0.5 = 6.3$$

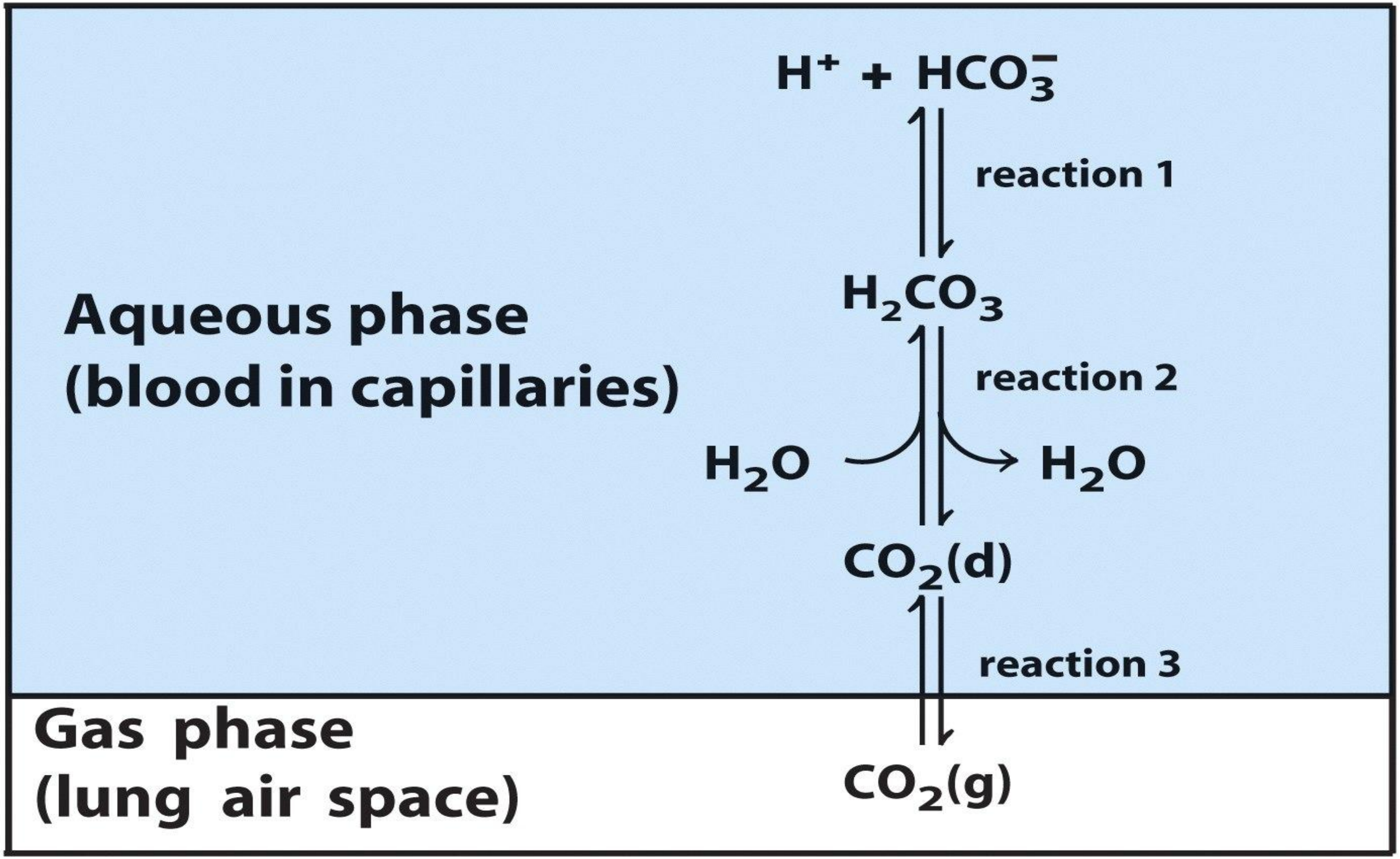
Blood Buffering: the Carbonate System

- The pH of human blood is about 7.4
- If pH falls outside range 6.8 to 7.8, you are dead.
- The main blood buffer is carbonate system (but phosphate and protein systems help)
- The pKa for carbonic acid is 6.1
- $H_2CO_3 \Leftrightarrow HCO_3^- + H^+$
- At physiologic pH, find ratio of carbonate to carbonic acid
- pH = pKa + log [A-]/[HA]
- 7.4 = 6.1 + log [A-]/[HA]
- 1.3 = log [A-]/[HA]   Therefore: [A-]/[HA] = 20.

- Given total buffer is about 28 mM, [HCO_3^-] = 26.6 mM and [H_2CO_3] = 1.4 mM
- This seems like a poor buffer BUT infinite CO_2 reservoir gives huge buffering capacity and MOST assaults on system are metabolic acid production.
- CO_2 (gas) ⇒ CO_2(aqueous) + H_2O ⇒ H_2CO_3 ⇒ + H^+ + HCO_3^-
- In clinic, conditions that effect MAINLY [HCO_3^-] are "metabolic", and conditions that effect MAINLY [H_2CO_3] are "respiratory"
- Metabolic acidosis means excess H^+: often seen with diabetic conditions
- Metabolic alkalosis is loss of H^+: may indicate vomiting or poisoning with base

- Respiratory acidosis: High CO_2 yields high $[H_2CO_3]$. May indicate poor CO_2 clearance from lungs.
- Respiratory alkalosis arises from very rapid removal of CO_2, which is indicative of hyperventilation.
- Ex. A patient comes in comatose. Blood gas analyzer shows $[HCO_3^-] = 50$ mM and $[H_2CO_3] = 1.4$ mM.
- $pH = 6.1 + \log 50/1.4 = 7.67$.
- metabolic alkalosis, maybe ingested too much bicarbonate (a Tums addict!)

Bicarbonate buffering system of blood

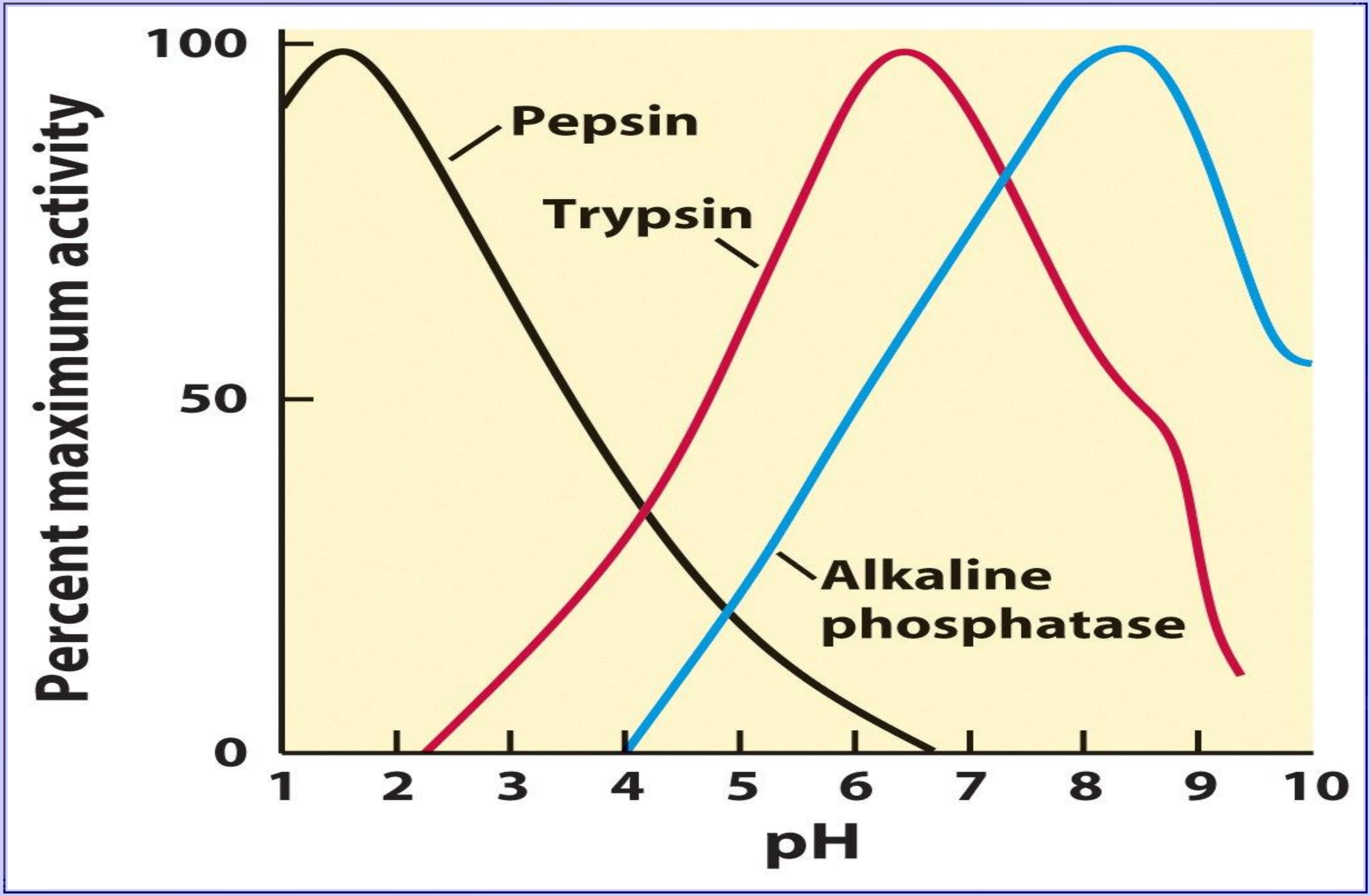


- Weak acids or base buffer cells & tissues against pH changes :-

1. The organism 1st line of defense against changes in internal pH is provided by buffer system.
2. The cytoplasm of most cells contains high [Protein] , which contain many amino acids with functional groups that are weak acids or weak bases:-

i. Side chain of histidine → has pka = 6.0 . So, protein containing histidine residues buffer effectively near neutral pH.

Many enzymes are sensitive to pH



ii. Nucleotides such as ATP & metabolites with low Mwt have ionizable groups contribute buffering power to cytoplasm.

3. Extracellular buffers have high conc. of :

- Organic acids buffer the vaccules of plant cells.
- Ammonia buffers urine.

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