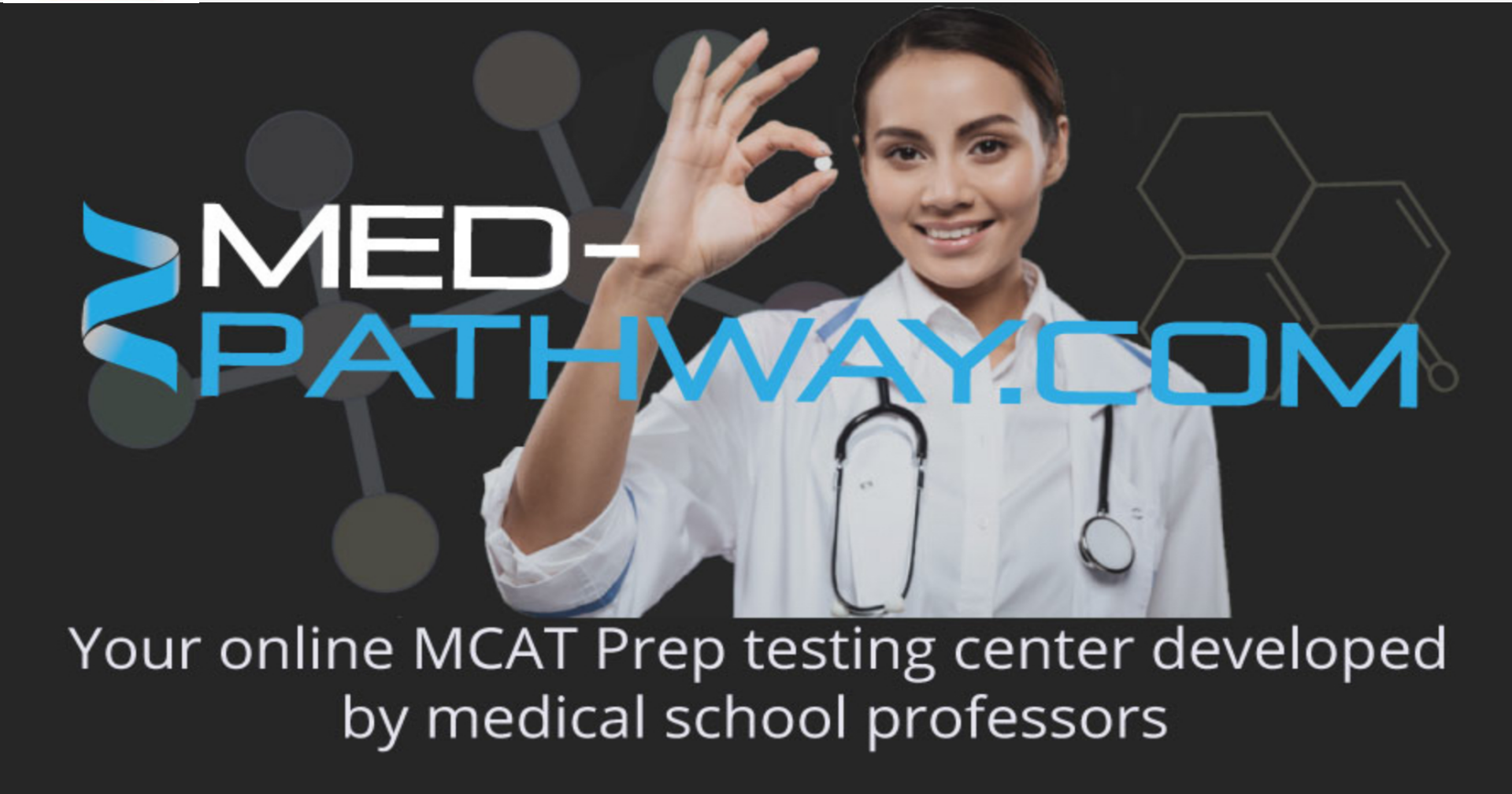
Acid Base Chemistry & Metabolism

Dr. Phillip Carpenter

medpathwaymcat

Med-pathway



Bronsted-Lowry Acids/Bases

Bronsted-Lowry Acid	Donates Protons
Bronsted-Lowry Base	Accepts Protons

Bronsted-Lowry Acids/Bases: Capable of donating and accepting protons, respectively. The system is also limited as it only focuses on protons.

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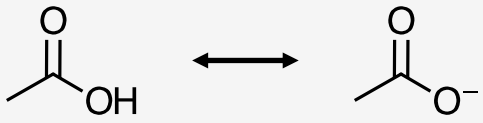


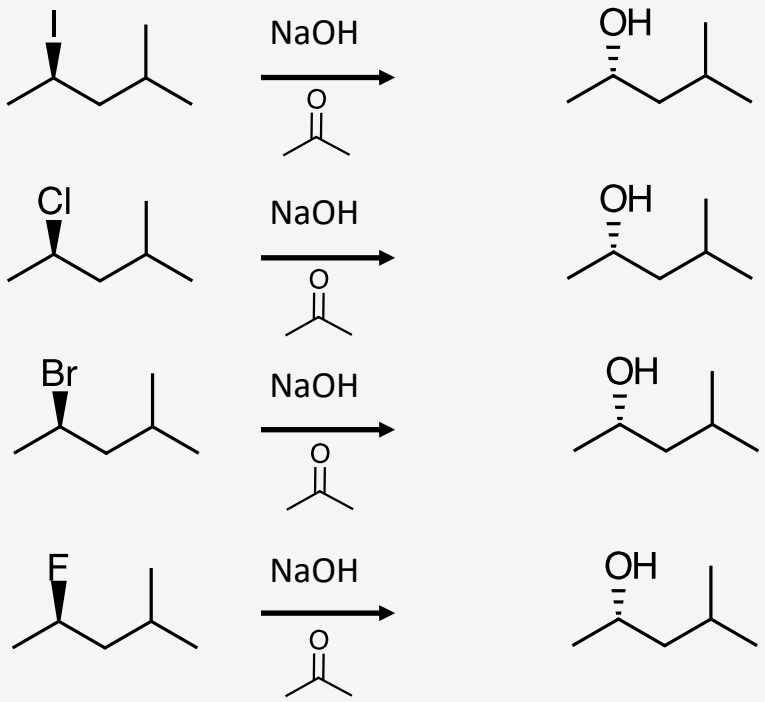
$+ H^+$ Partially dissociated $K_a = 1.8 \times 10^{-5}$ Weak

$$= \frac{[CH_3COO^-][H^+]}{CH_3COOH}$$

Strong Conjugate Base

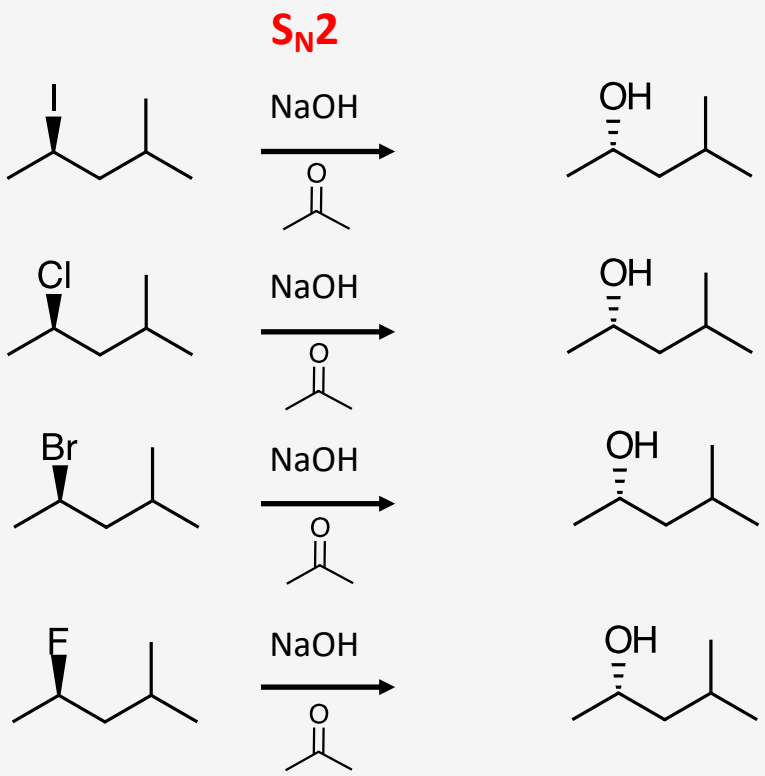
$pK_a = 4.75$

Conjugate Bases and Leaving Groups



Relative Rates of Product Formation?

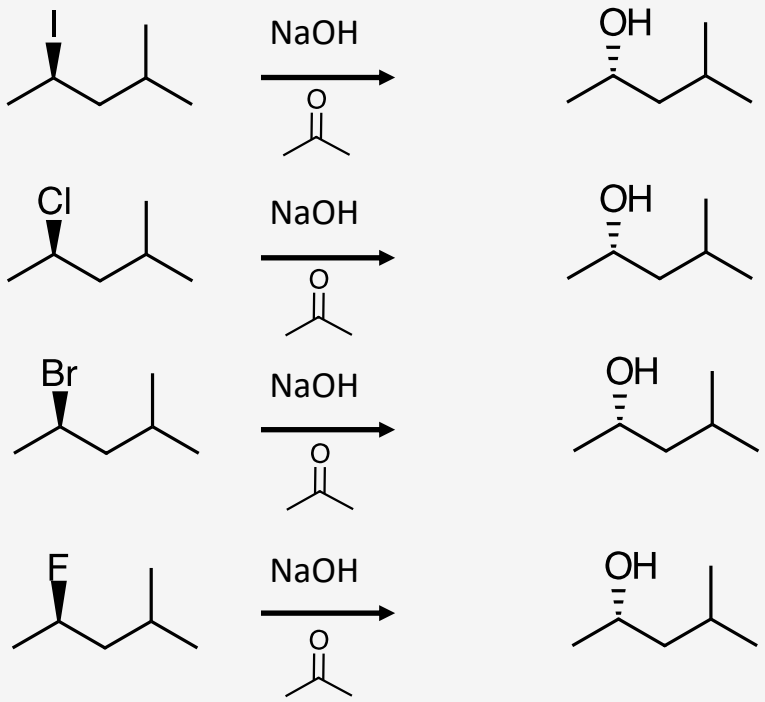
Conjugate Bases and Leaving Groups



Relative Rates of Product Formation?

HI > HBr > HCl > HF = Relative Strength of Acids

Conjugate Bases and Leaving Groups



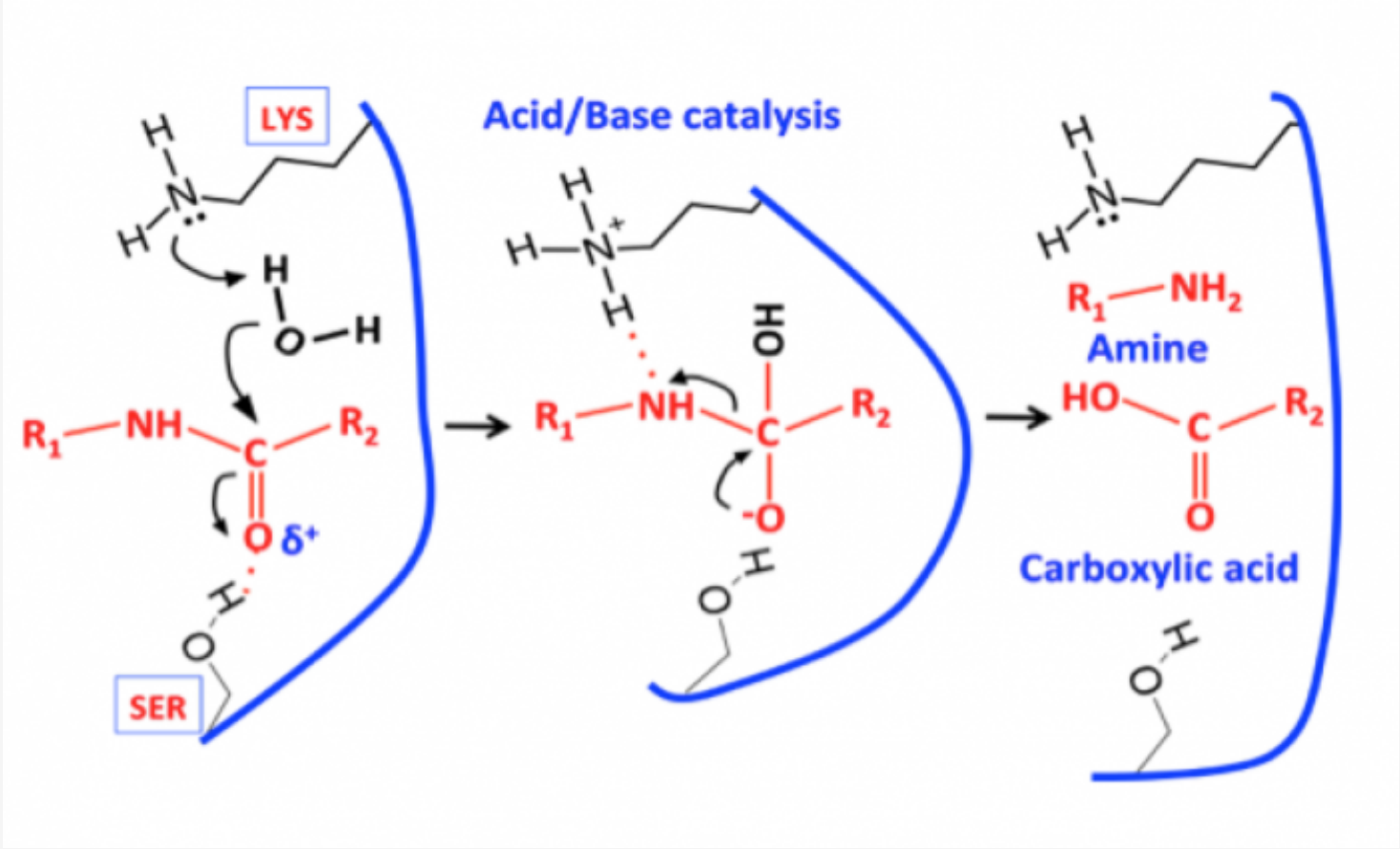
Relative Rates of Product Formation?

HI > HBr > HCl > HF = Relative Strength of Acids

Weakest conjugate base;
Best leaving group

Strongest conjugate base;
Worst leaving group

Acid/Base Catalysis as Leaving Groups



R-NH$^-$ vs R-NH$_2$ leaving groups

Acid Base Problems

What is the pH of a 0.2 M solution of a weak acid (HA) with a $K_a = 1.8 \times 10^{-4}$?

A. 1.3
B. 2.2
C. 3.4
D. 4.1

Acid Base Problems

What is the pH of a 0.2 M solution of a weak acid (HA) with a $K_a = 1.8 \times 10^{-4}$?

Weak acids do not completely dissociate into protons and their conjugate bases. Set up an ICE table describing the initial and final changes for the components of the following equilibrium:

$HA = H^+ + A^-$

ICE Table:
Initial
Change
Equilibrium

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ICE Table:
Initial
Change
Equilibrium

	[HA]	$[H^+]$	$[A^-]$
Initial	0.2M	10^{-7} M	0 M
Change	- X	+ X	+ X
Final	0.2M -X	+ X	+ X

Autoionization of water
Is negligible

$$HA = H^+ + A^-$$

$$K_A = \frac{[H^+][A^-]}{HA} = \frac{[X][X]}{0.2M - X}$$

$$1.8 \times 10^{-4} = \frac{[X][X]}{0.2M - X} = \frac{X^2}{0.2M}$$

$$X^2 = 3.6 \times 10^{-5}$$

$$X = .006$$

Acid Base Problems

pH = -log [.006] = ??

Acid Base Problems

pH = -log [.006] = ??

A. 1.3
B. 2.2
C. 3.4
D. 4.1

$[H^+]$	[pH]
0.001 M	3.0
0.006 M	2.22
0.01 M	2.0

Log [.001] = -3.0

Log [.01] = -2.0

Arrhenius Acids/Bases

Arrhenius Acid	Increases $[H^+]$ in solution
Arrhenius Base	Increases $[HO^-]$ in solution

Arrhenius Acids/Bases: The Arrhenius definition of an acid is limited as it only focuses on water as the solvent.

$$HCl \longrightarrow H^+ + Cl^-$$
$$H_2O + H^+ \longrightarrow H_3O^+$$

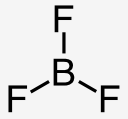
$$NaOH \longrightarrow Na^+ + OH^-$$

Lewis Acids/Bases

Lewis Acid	**Electron Pair Acceptor**
Lewis Base	**Electron Pair Donor**

Lewis Acids/Bases: Electron pair acceptors and donors, respectively. The Lewis acid/base definition focuses on what electrons are actually doing in a reaction.

:NH_3 = Lewis Base (Nonbonding electrons)



Lewis acid

Nucleophile

Electrophile

3 Acid/Base Systems

Arrhenius Acid	Increases $[H^+]$ in solution
Arrhenius Base	Increases $[HO^-]$ in solution
Bronsted-Lowry Acid	Donates Protons
Bronsted-Lowry Base	Accepts Protons
Lewis Acid	Electron Pair Acceptor
Lewis Base	Electron Pair Donor

Arrhenius Acids/Bases: The Arrhenius definition of an acid is limited as it only focuses on water as the solvent.

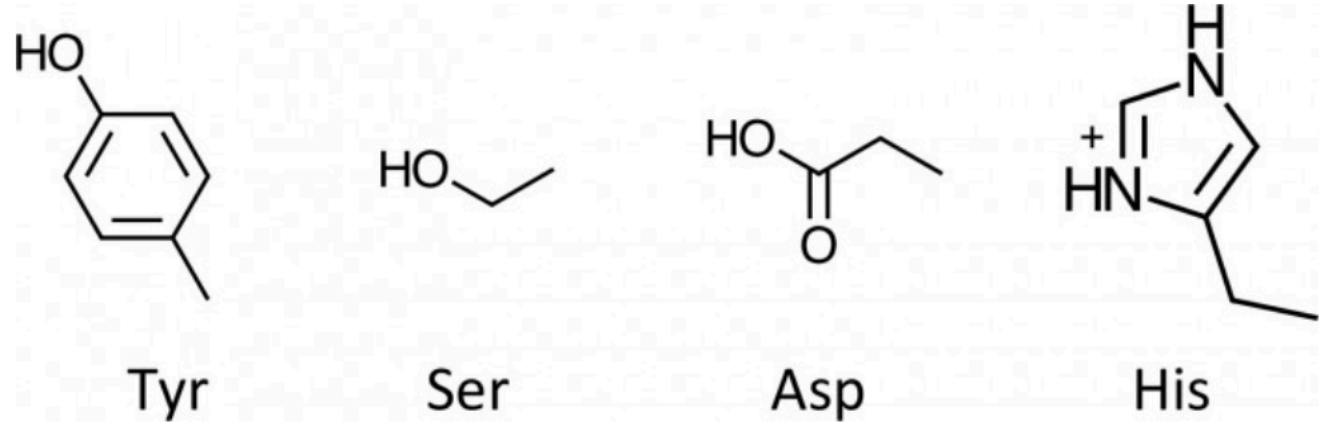
Bronsted-Lowry Acids/Bases: Capable of donating and accepting protons, respectively. The system is also limited as it only focuses on protons.

Lewis Acids/Bases: Electron pair acceptors and donors, respectively. The Lewis acid/base definition focuses on what electrons are actually doing in a reaction.

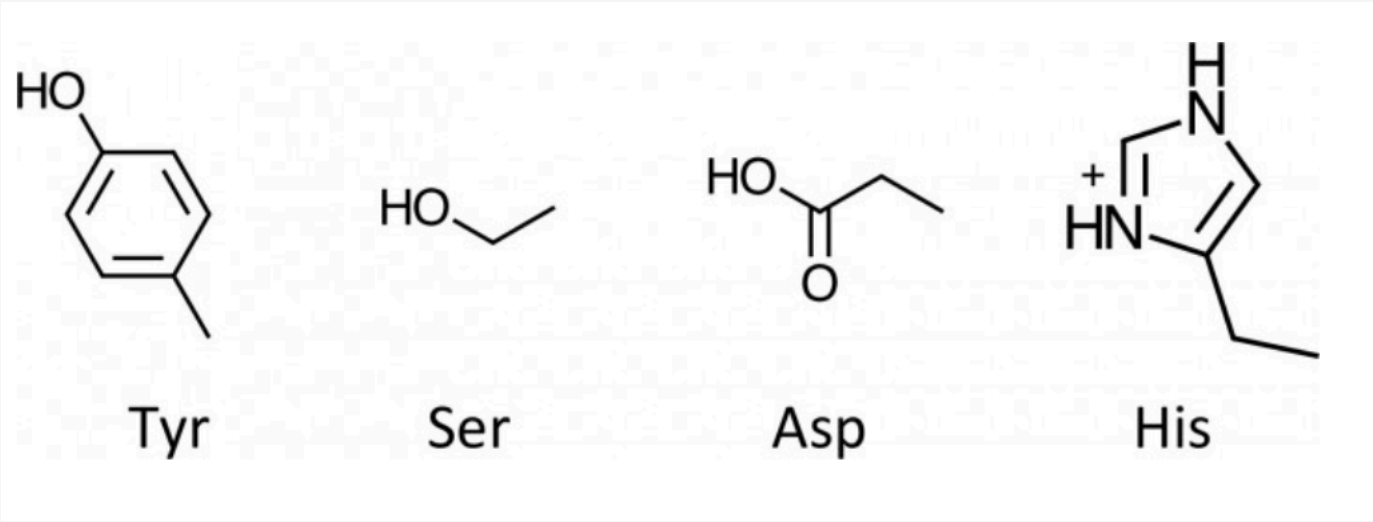
Relative Strengths of Acids

$pK_a = -\log [K_a]$ $pK_a + pK_b = 14$

$H_2O = H^+ + HO^-$ $K_w = 10^{-14}$

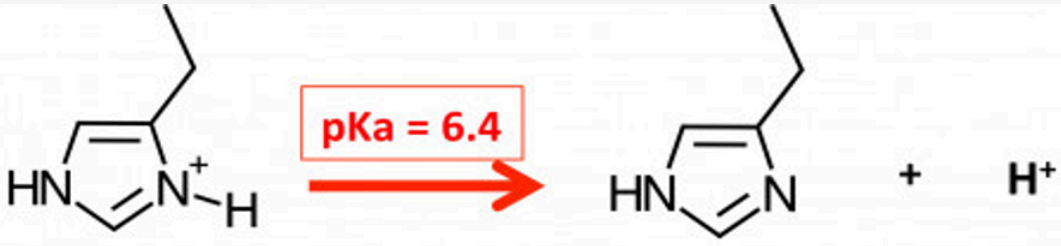


Relative pK_a values



pK_a = 10 **pK_a = 14** **pK_a = 3** **pK_a = 6.4**

Henderson Hasselbalch



HA
(Acid)

A-
(Conjugate Base)

$$pH = pKa + \log[Base]/[Acid]$$

Q: What is the fraction of histidine residues that are positively charged at pH = 7.4?

Henderson Hasselbalch

HA
(Acid)

A^-
(Conjugate Base)

$$pH = pKa + \log[Base]/[Acid]$$

Q: What is the fraction of histidine residues that are positively charged at pH = 7.4?

Solve for [Base]/[Acid] knowing that pKa = 6.4 and pH = 7.4.

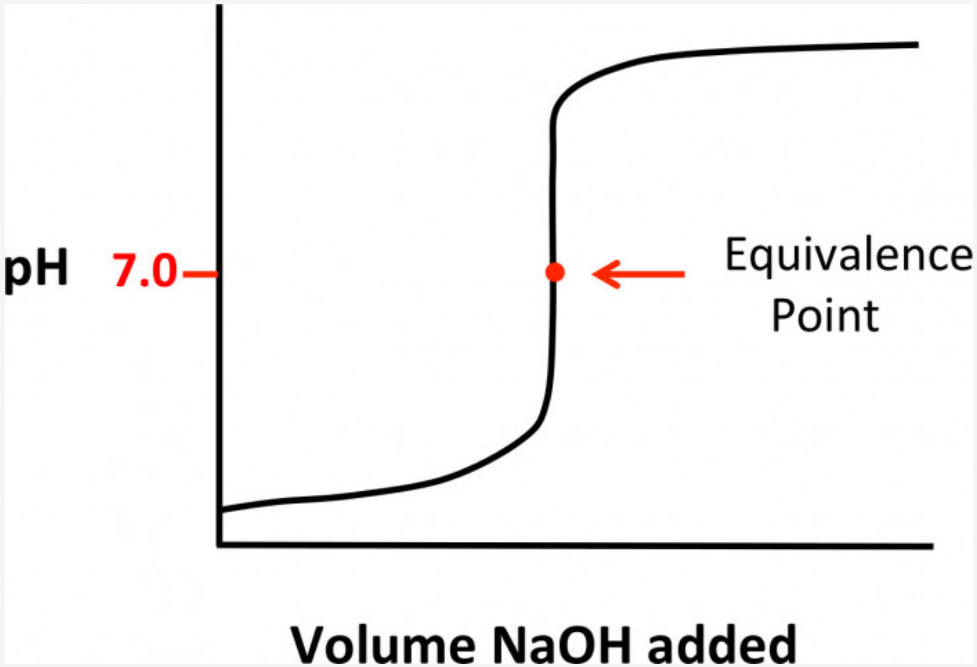
$pH - pKa = \text{Log } (X)$ $X = [A^-][HA]$

A: Antilog (pH-pKa) = antilog[log(X)] Antilog 1 = X ; X = 10
Therefore, as X= [Base][Acid], the ratio of [Base]/[Acid] = 10.

There are 10X as many His residues in conjugate base than His residues than in acidic, positively charged form.

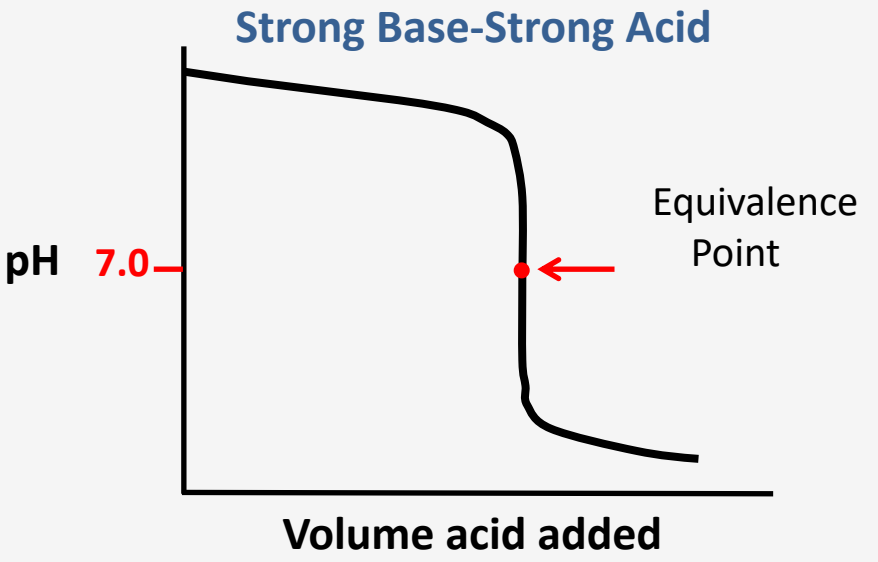
Titration of Strong Acid w/ Strong Base

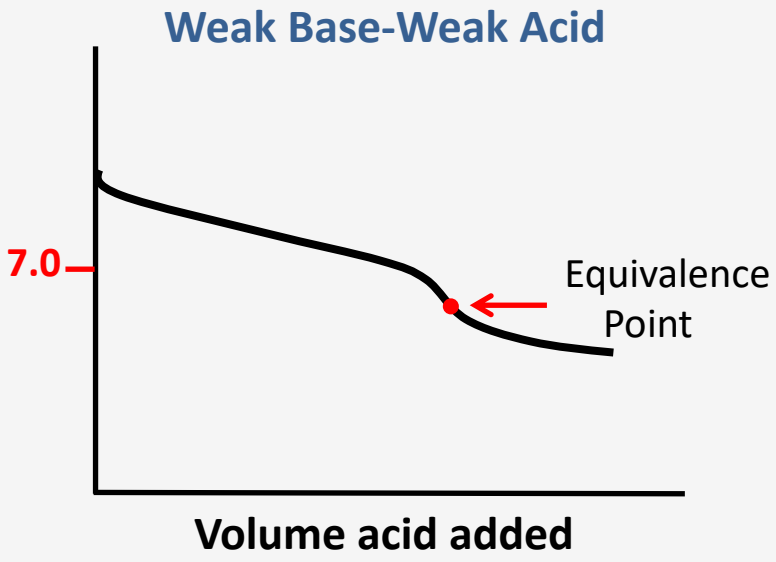
$HCl + NaOH = NaCl + H_2O$ Neutral Salt

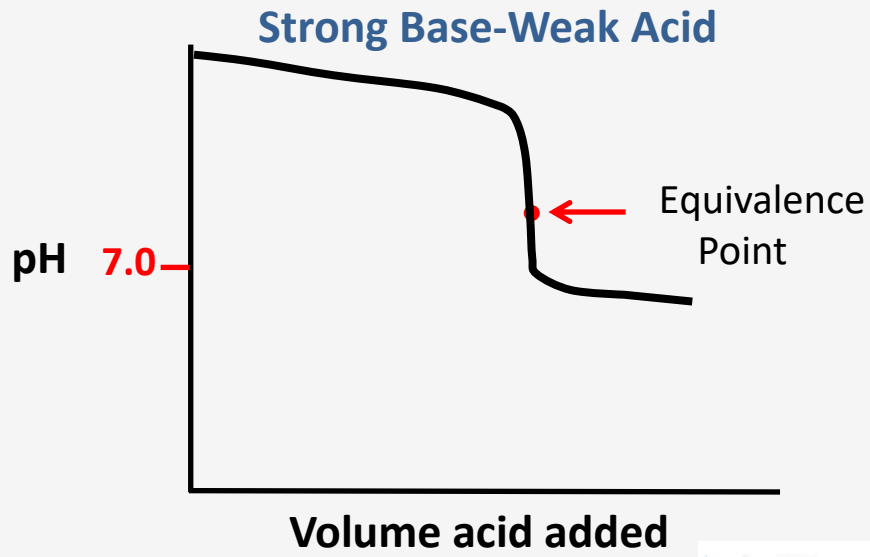


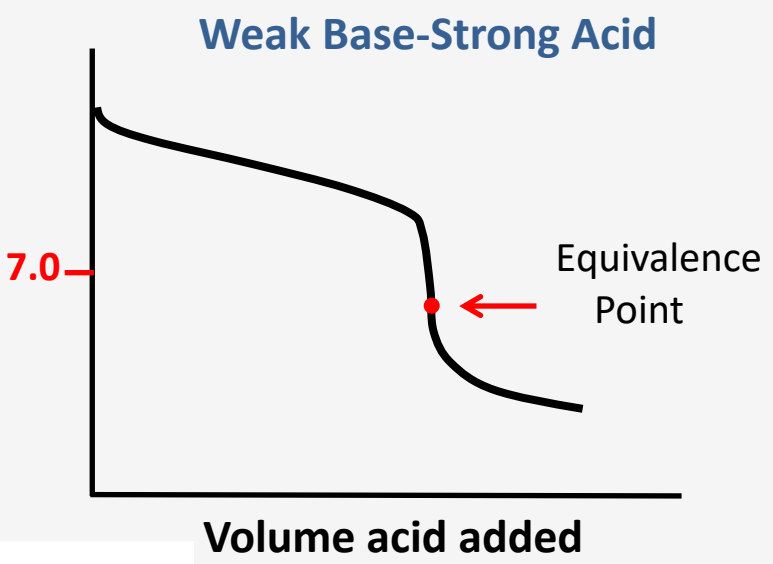
 # moles of base added = # number of moles of acid present in the solution

Titration Curves



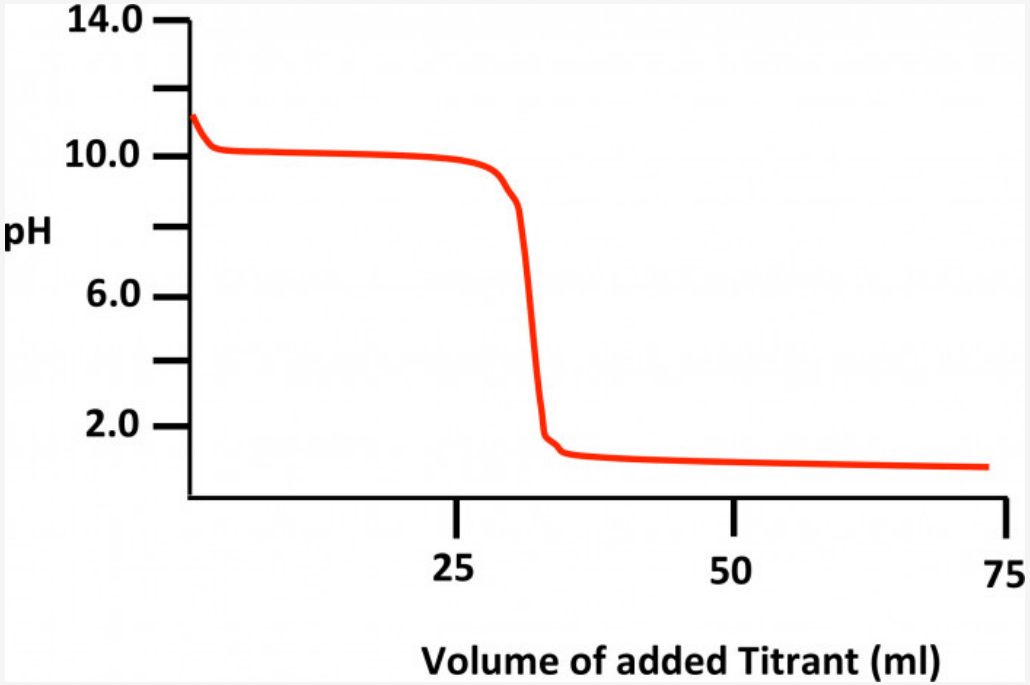






Titrations

A compound was titrated as shown. What is the K_a value of the compound?



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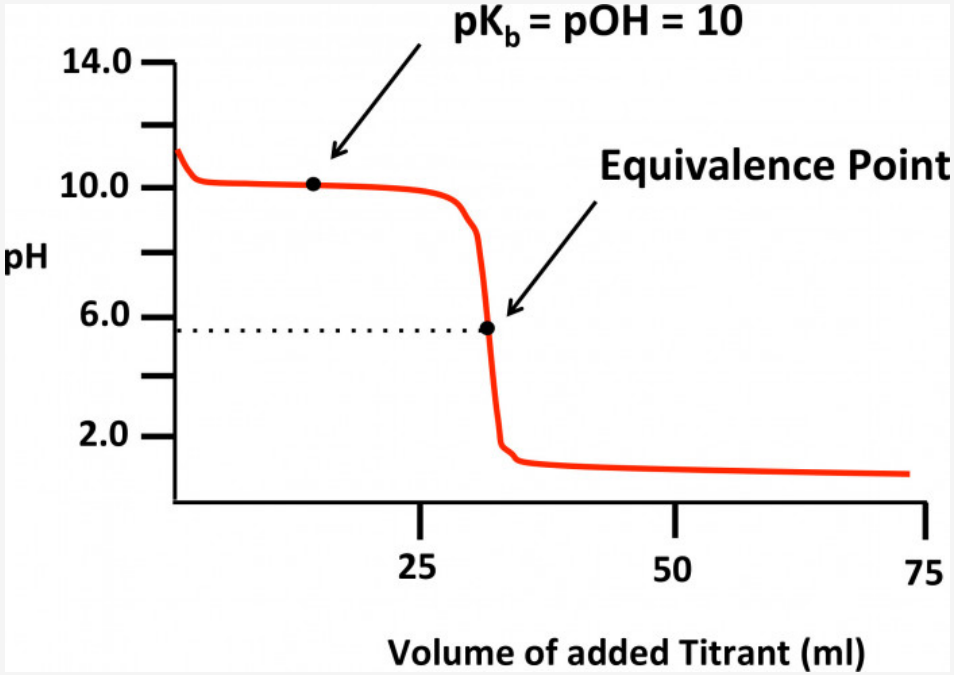
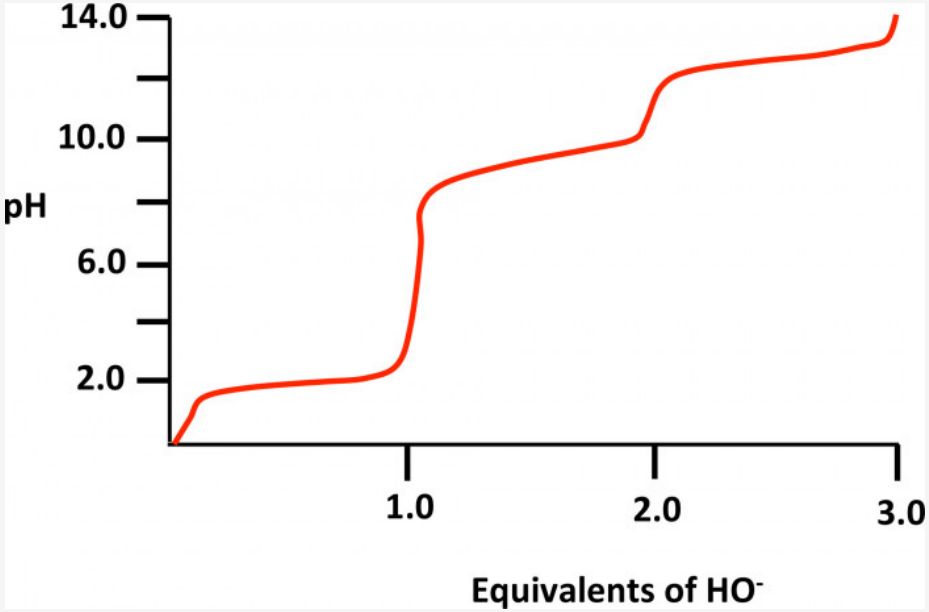


The titration is between a weak base and an acid as the pH decreases with the addition of titrant.

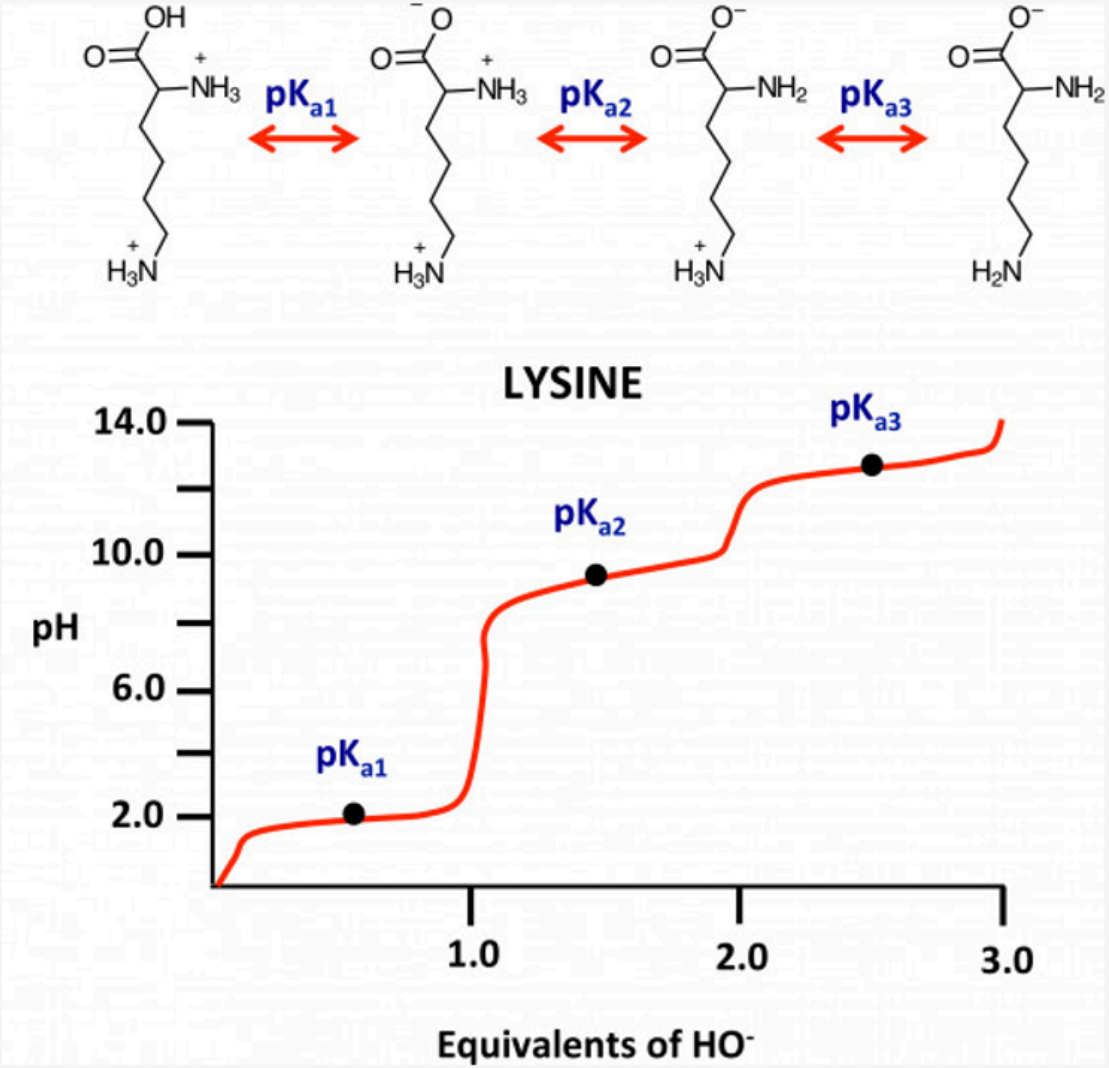
Observe from the titration that the inflection point (where the slope = 0), represents the pK$_b$. Therefore, pOH = pK$_b$ = 10. Recall that $K_b = 10^{-pKb} = 10^{-10}$ and $K_aK_b = K_w = 10^{-14}$. Therefore, $K_a = 10^{-14}/10^{-10} = 10^{-4}$.

Amino Acid Titrations

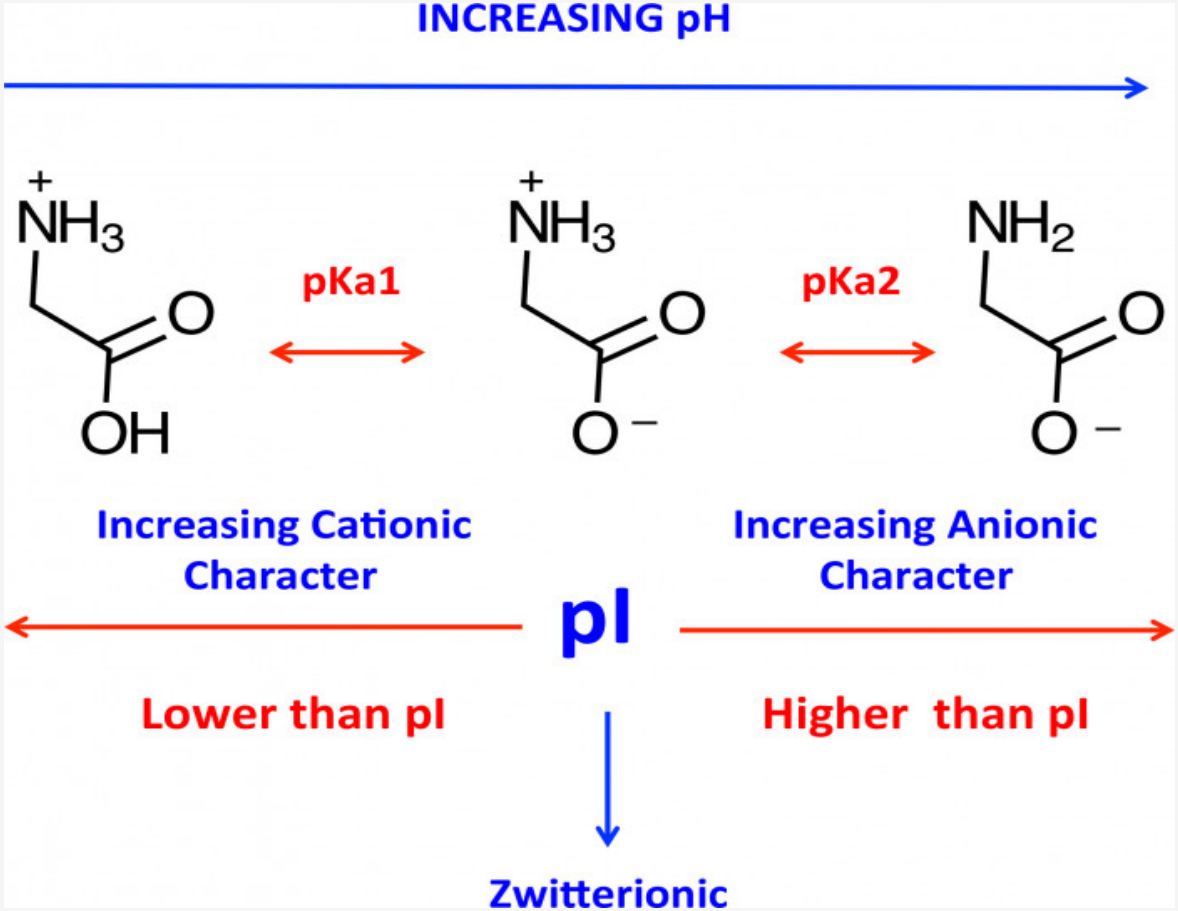
Which of the following amino acids best fits the titration profile shown below?



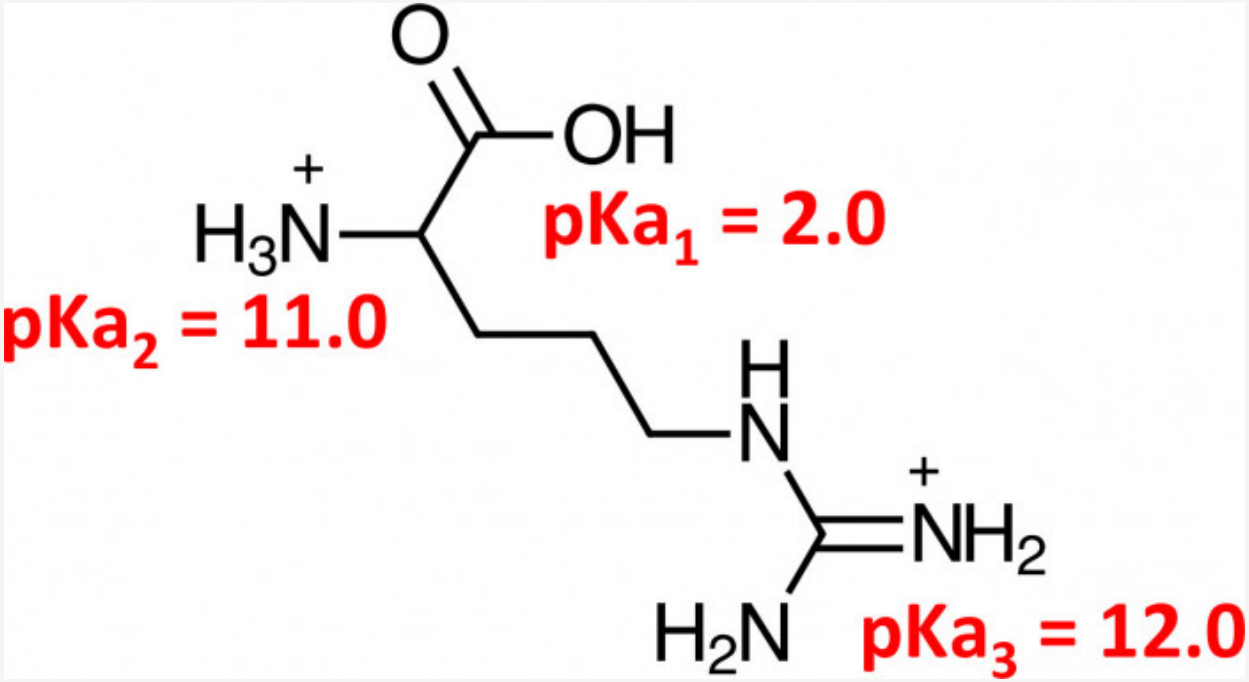
Lysine Titration



Isoelectric Point



What is the isoelectric point of arginine?



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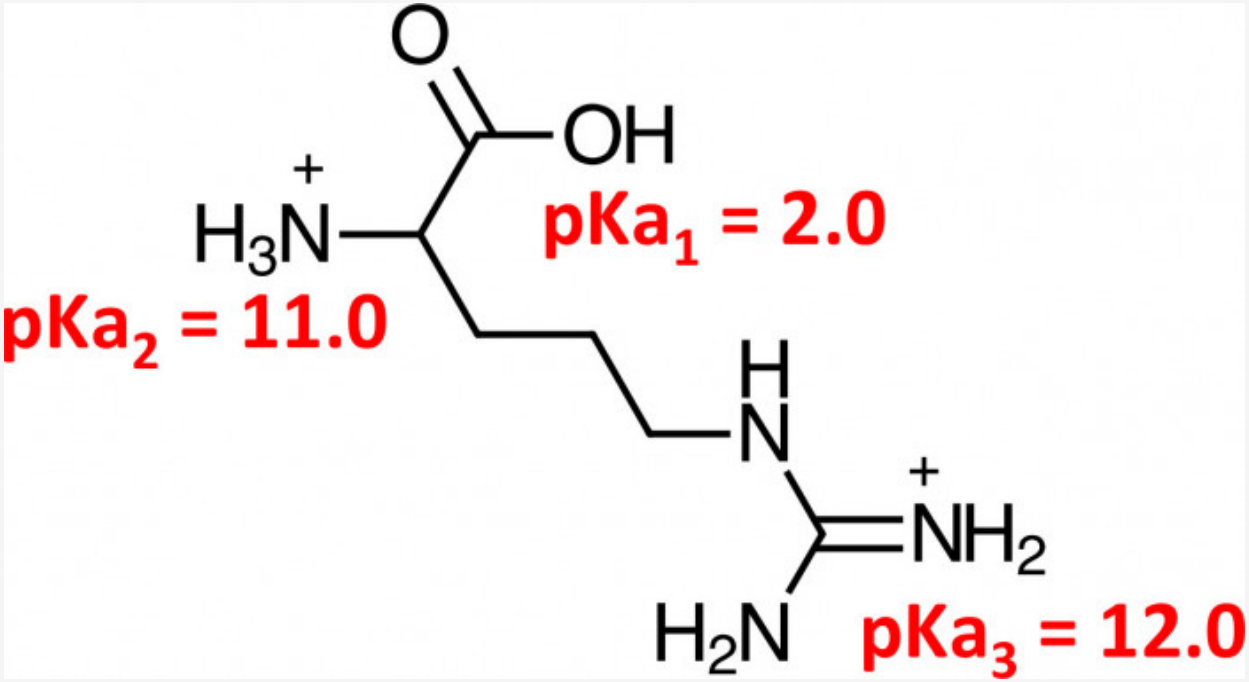
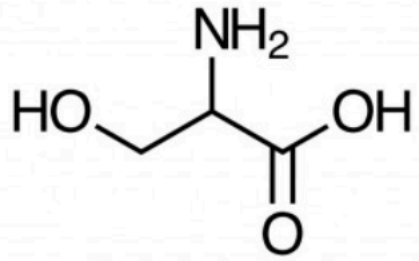


pI = (11 + 12) x 0.5 = 11.5

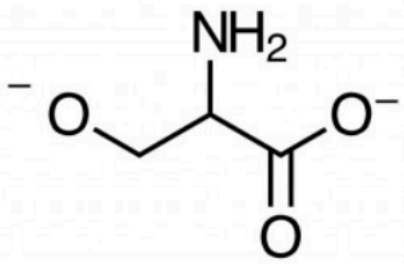
Relative pK_a Values

Which of the following forms of serine would fail to exist at any pH?

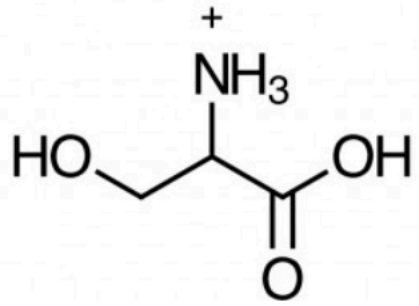
A



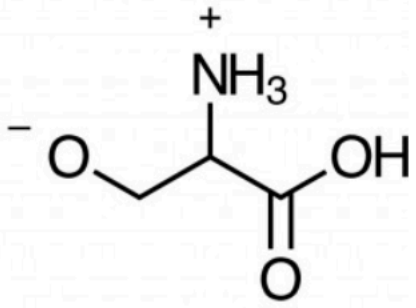
B



C



D



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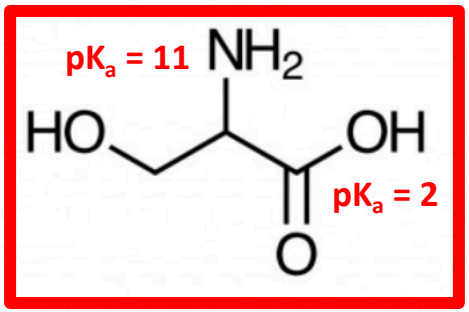
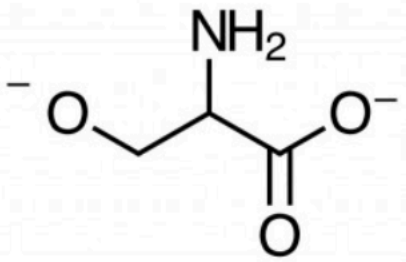


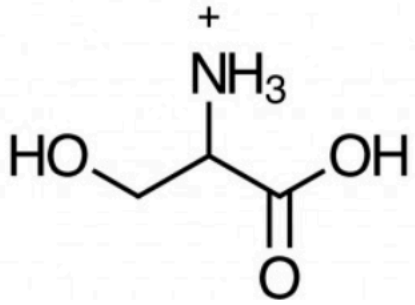
Loss of amino proton before Carboxyl Proton

B



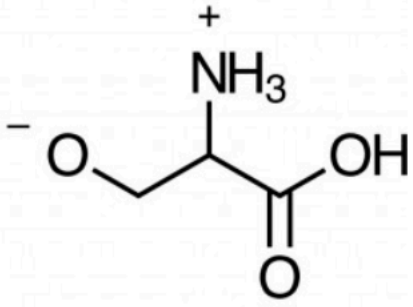
pH > pI

C



pH < pI

D



Isoelectric Point

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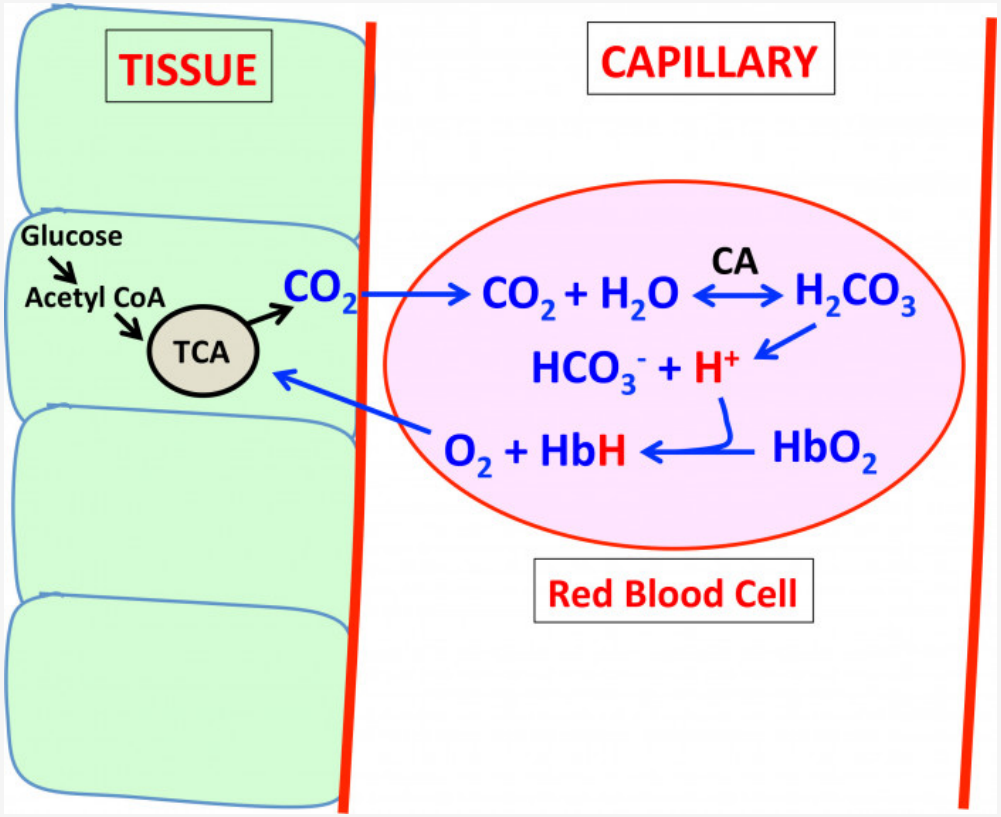
Carbonic Acid Buffering System

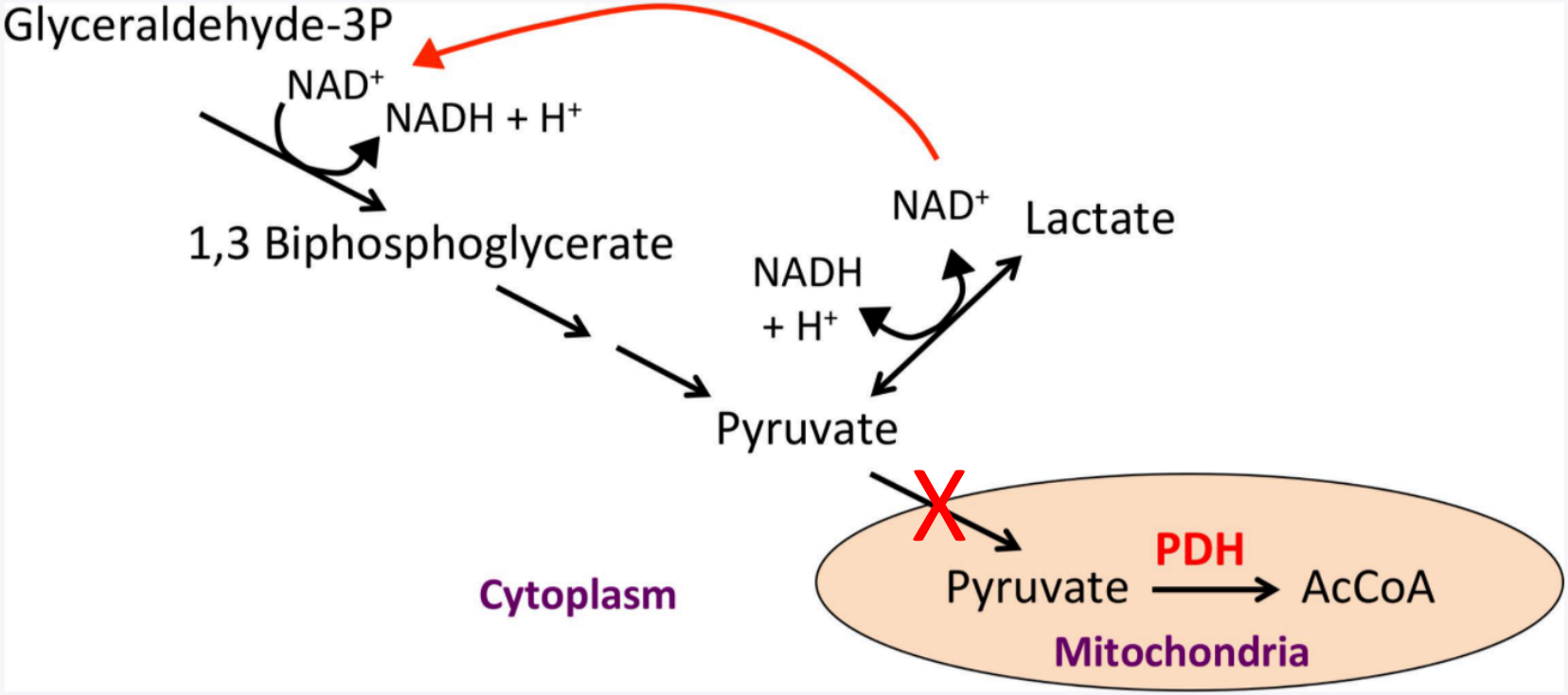
Carbonic Anhydrase

$$CO_2 + H_2O \longleftrightarrow H_2CO_3 \longleftrightarrow HCO_3^- + H^+$$



Metabolic Acidosis

Lactic Acidosis



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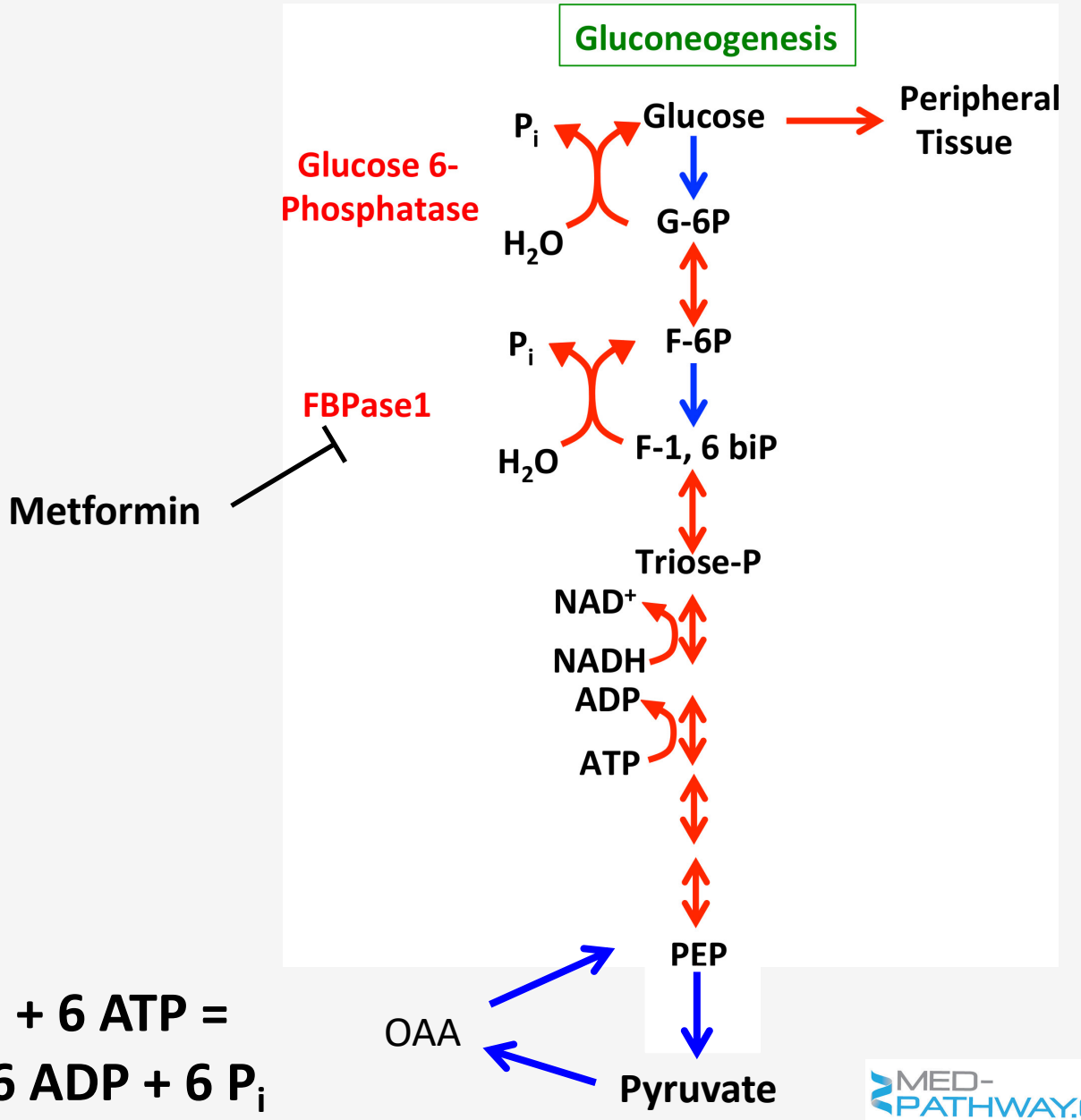
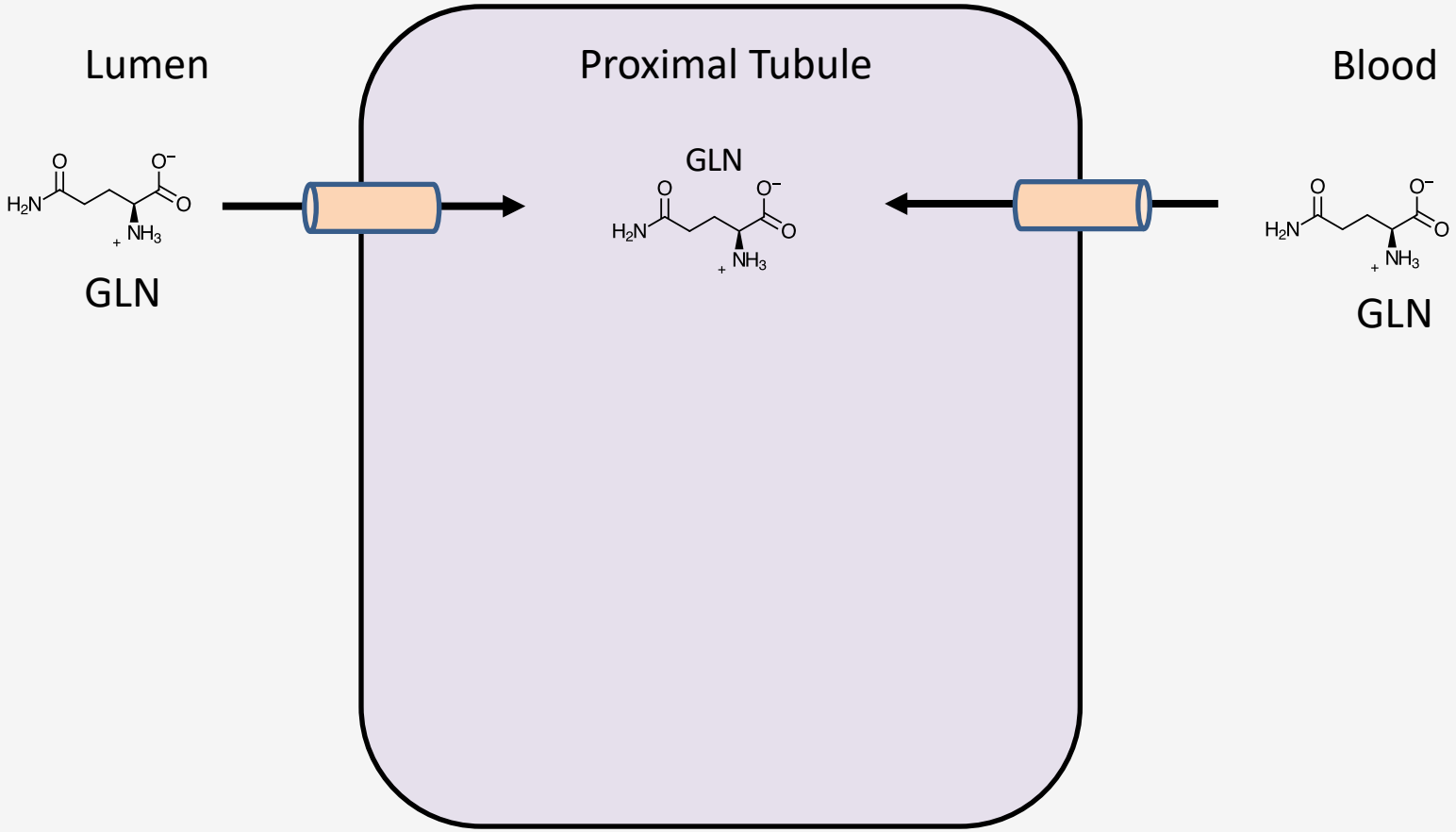
Metabolic Acidosis



2 Pyruvate + 2 NADH + 6 ATP =
Glucose + 2 NAD^+ + 6 ADP + 6 P_i

Metabolic Acidosis: The Kidney

Physiological Goal: To Raise Blood pH levels



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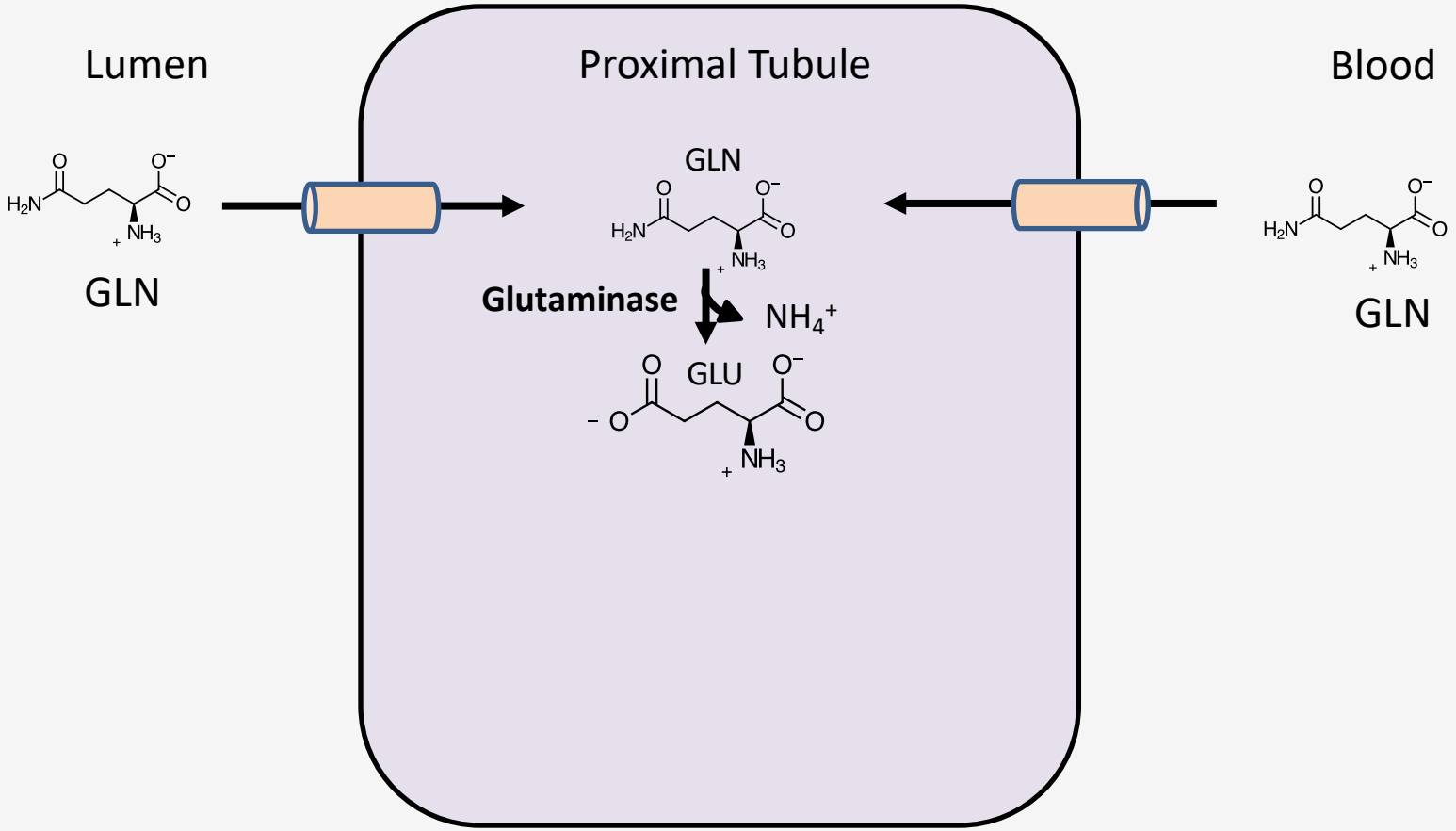

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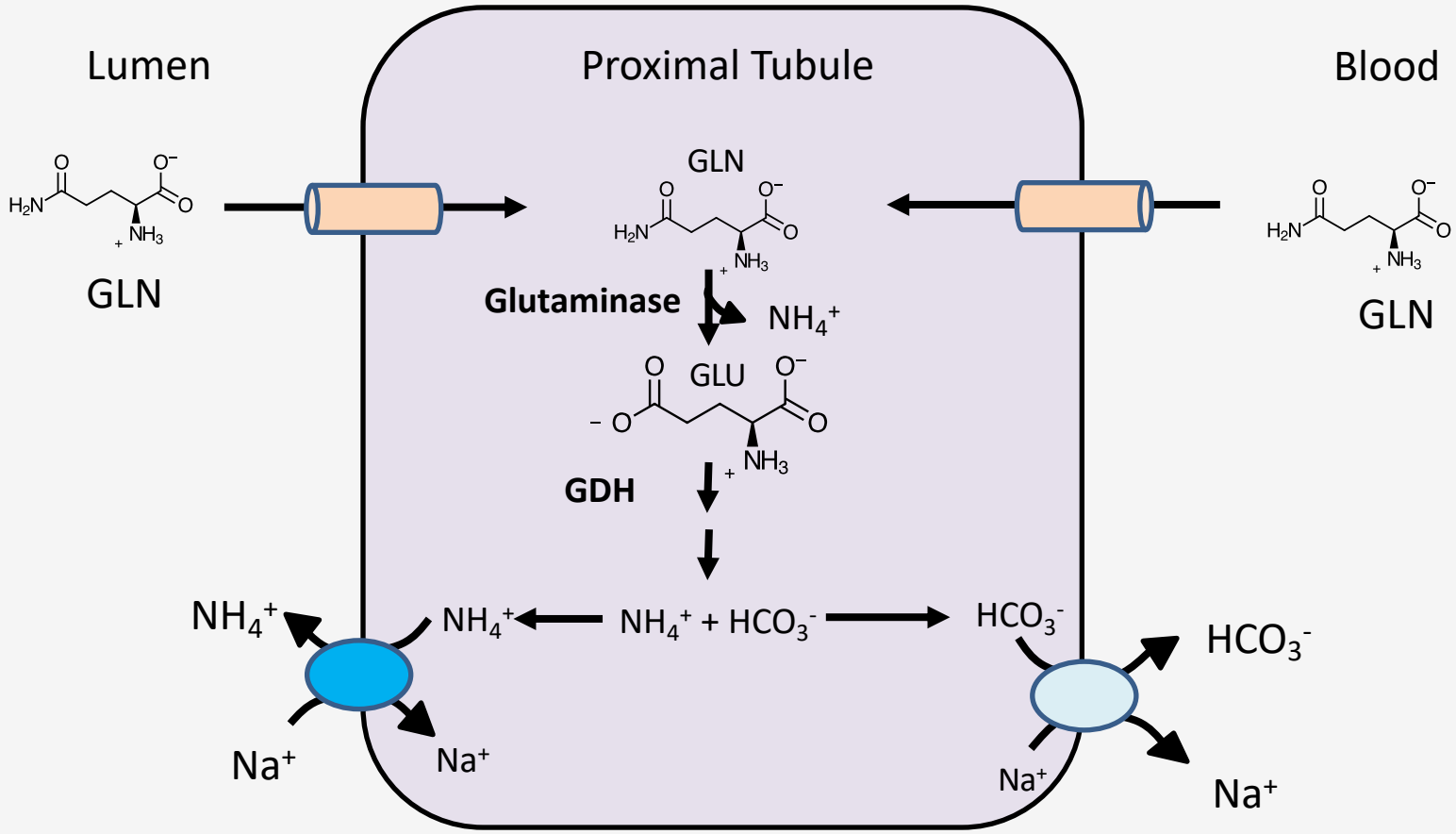

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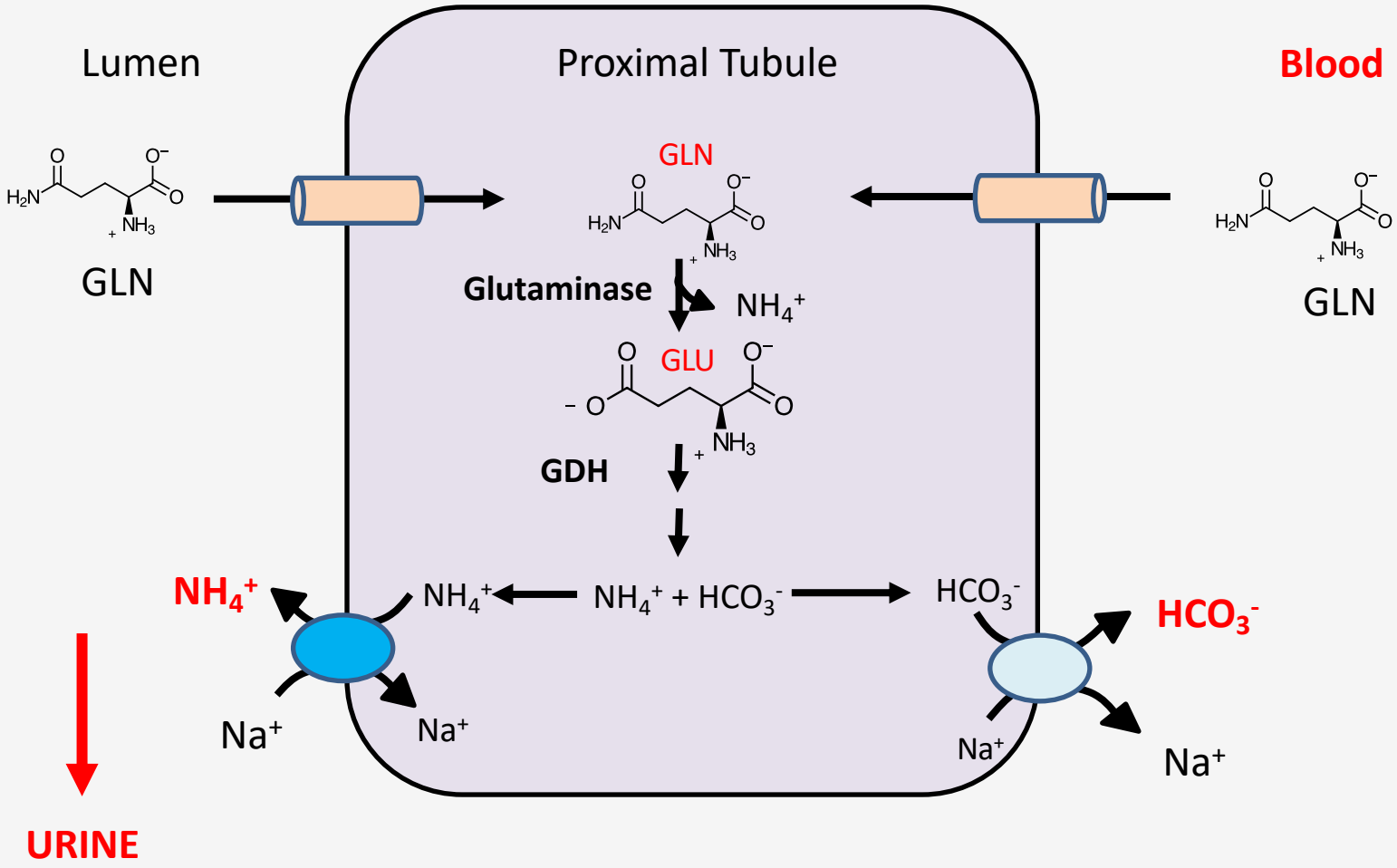
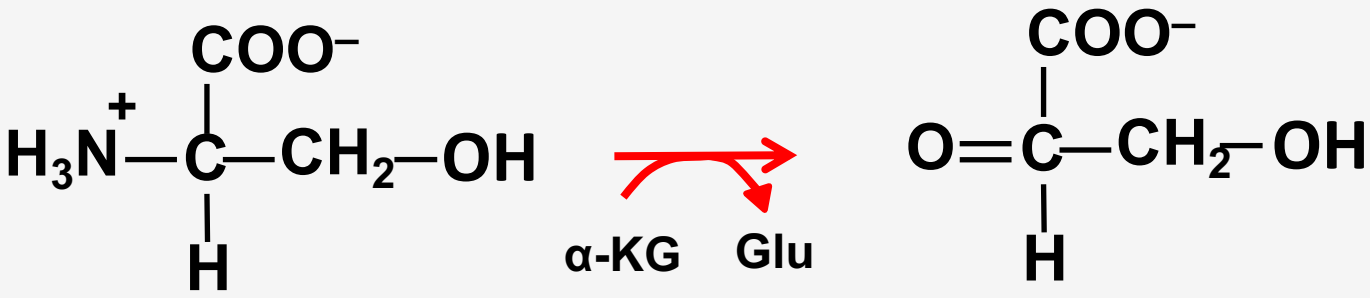
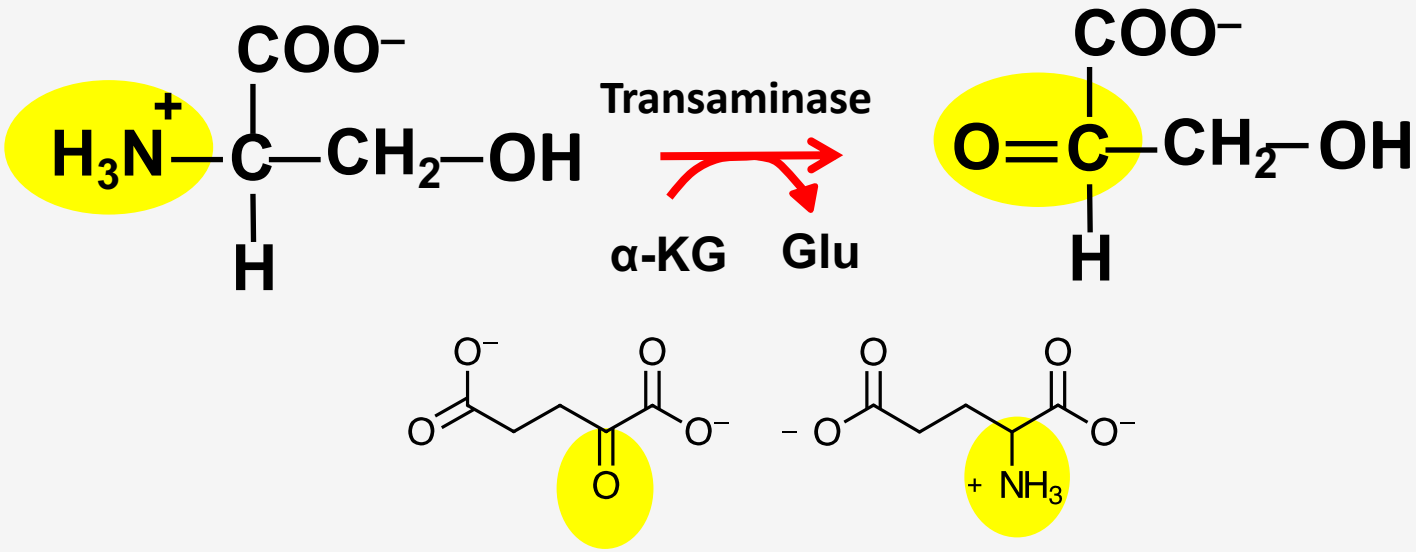


pH is linked to nitrogen balance

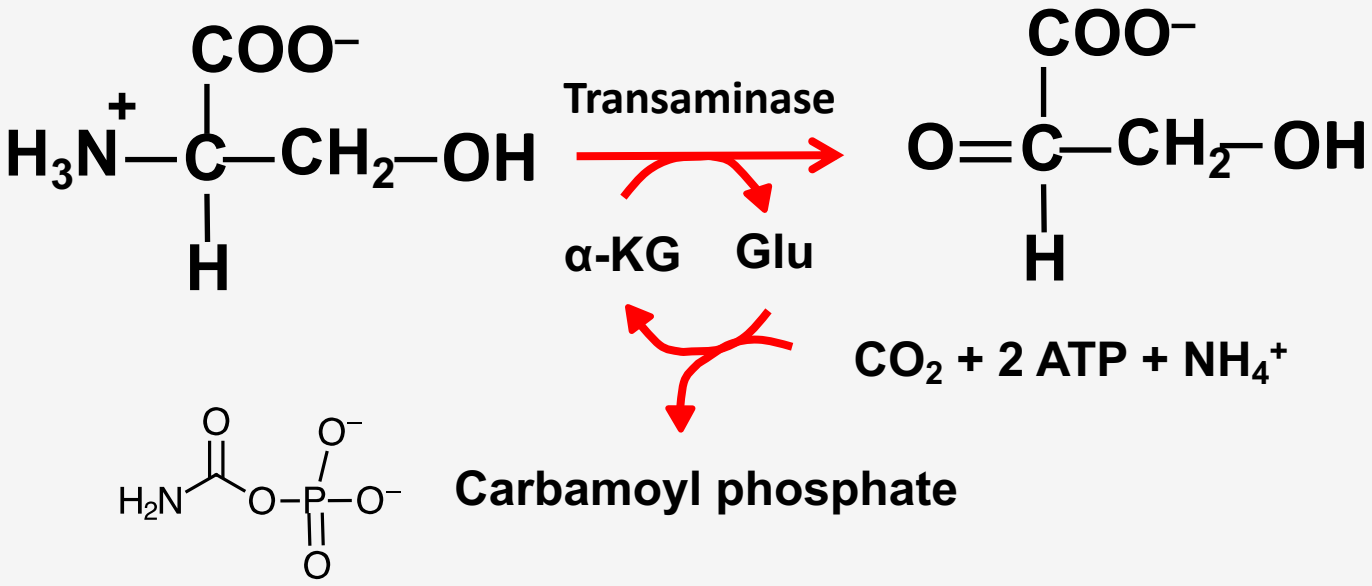
Nitrogen Balance & pH



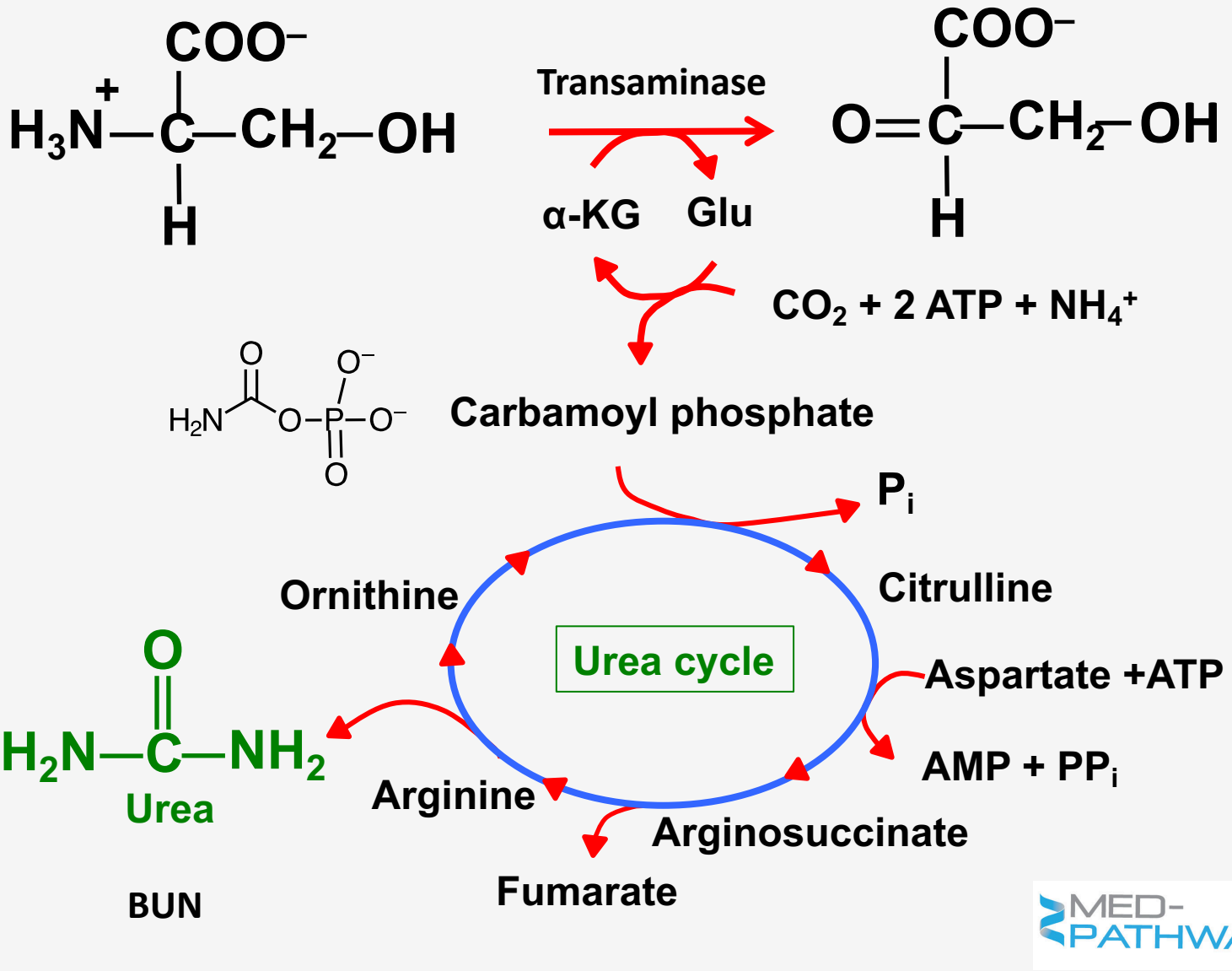
Nitrogen Balance & pH



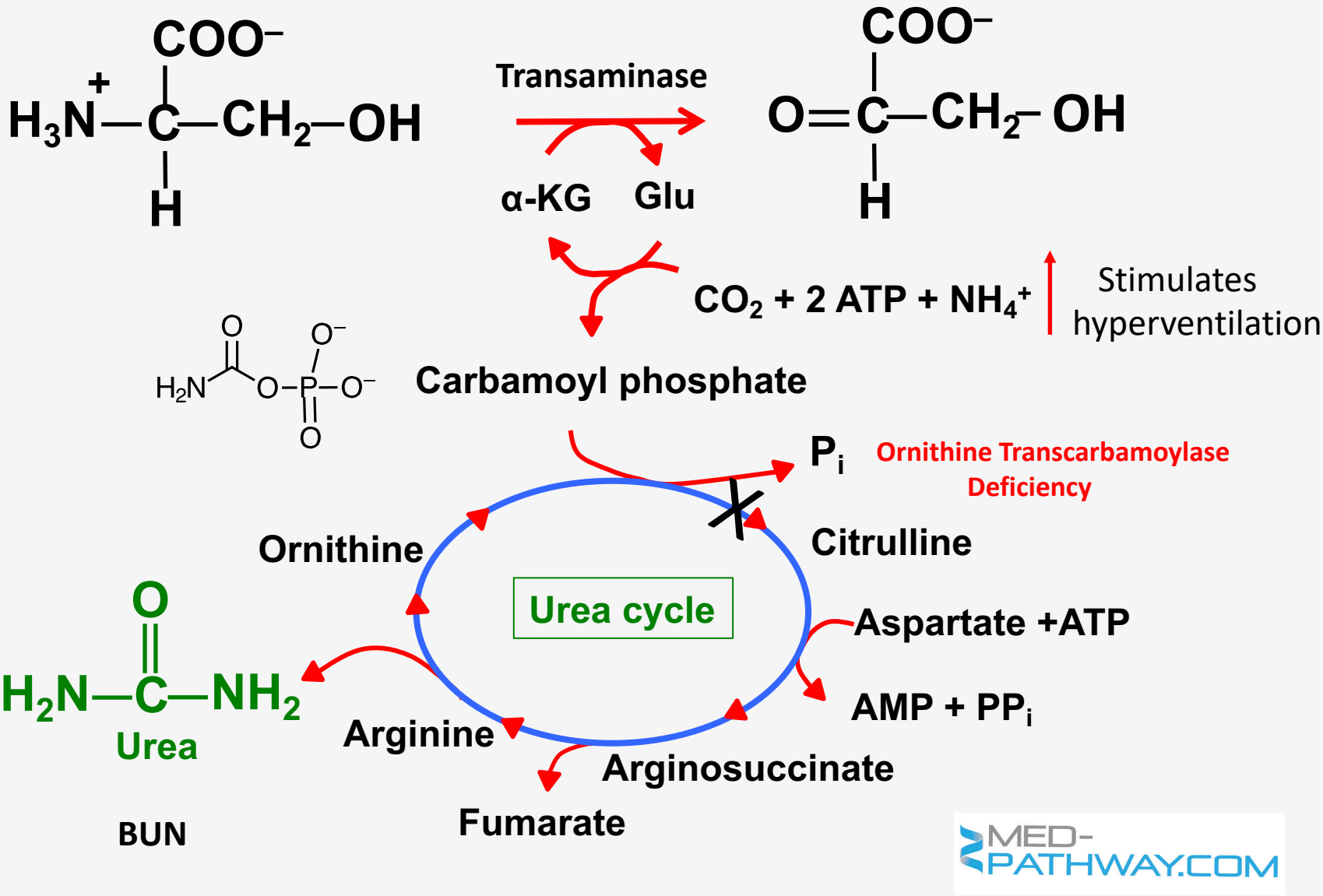
Nitrogen Balance & pH



Nitrogen Balance & pH



Nitrogen Balance & pH



Workshop Passages

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