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ACID–BASE REGULATION

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JAYPEE

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CHAPTER 3

The Regulations of Hydrogen Ion

"Every chemical reaction has a transition state"

–Derek Barton

INTRODUCTION

In the body, there is not one single receptor or center that is specifically sensitive to all aspects of hydrogen ion regulation. Many organs participate. Nevertheless, there are mechanisms that resist change in hydrogen ion concentration. They are efficient and well-coordinated. These are three in number:

1. Buffer systems within the body
2. Renal regulation of hydrogen ion excretion
3. Respiratory regulation of CO_2 excretion

BUFFER SYSTEM IN THE BODY

Buffer literally means a safeguard or resistance against change and or to prevent radical change against shift. A solution that relatively maintains a constant pH even upon the addition of an acid or base is known as a buffer solution. A buffer is a mixture of either a weak acid and its alkali salt or a weak base and its acid salt. So, it is a pair of weak acids and their conjugate base. So, the buffer solution is a weak acid plus a conjugate base. The conjugate base can be "any" negatively charged ion or molecule. Conjugate means linked to or united in function.

It can be represented as:
Buffer solution = Weak acid + conjugate base
Or,

$$HBuff = H^+ + Buff^-$$

So, the buffer is any substance that can reversibly bind with hydrogen. The weak acid is often called buffer acid and the conjugate base is the "buffer base". The buffer bases are anions, uncharged molecules with electron clouds. Functionally, the buffer systems do not eliminate or add hydrogen ions to the body, they only keep them tied up until balance can be re-established. The buffers are mainly divided into two major groups, i.e., bicarbonate and nonbicarbonate buffer systems. The only buffer system that can buffer the hydrogen ion H^+ of volatile acid is the nonbicarbonate buffer system. Both nonbicarbonate and bicarbonate buffer systems can buffer the hydrogen ion H^+ produced by fixed acids. If a strong acid or alkali is added to extracellular fluid (ECF) 30–50% is buffered by systems available in the ECF (mostly the bicarbonate-carbonic acid the first line of defense) and the remainder is buffered in cells, bones, or by release of buffers from these. This buffering is so effective that the addition of 1 mEq of strong acid per liter only increases hydrogen ion (H) by about 5 nM. That is all except that 5 of every million hydrogen ions added are buffered. The sites of these buffers govern their availability and the rate of resultant change in hydrogen ion concentration. The buffers of the blood and ECF are immediately available, those of

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intracellular are less so but still operate in a few minutes. The immense buffering capacity of bones requires hours to days to become effective. These differences are partly due to blood flow and also to the tissue rates of exchange with surrounding fluids.

Table 1 summarizes the approximate contributions of various blood buffers to the total buffer base.

The buffering process is not foolproof even if they are in equilibrium and work in cooperation. The disturbances happen in diseased states and are accentuated in many clinical situations/conditions.

The buffer systems in the body are broadly classified into two groups—intracellular buffers and extracellular buffers with respect to their cell stations.

Extracellular Buffers

- The bicarbonate and carbonic acid are quantitatively the most important buffer system in ECF. It is made up of sodium bicarbonate and carbonic acid. The normal ratio of the two components of this buffer system is 20:1.
- The phosphate buffer system consists of the mixtures of alkaline and acid phosphates of sodium, i.e., Na_2HPO_4/NaH_2PO_4. The normal ratio of base to acid is 4:1.
- The plasma proteins act as buffers. It is so because a protein molecule consists of many polypeptides containing various amino acids which behave as Bronsted acids and bases (proton donors and acceptors respectively). An amino acid has an aminoterminal and a carboxyl-terminal. The amino terminal is basic, and the carboxyl-terminal is acidic. At the isoelectric point, the substance protein has no net charge and has no tendency to move. For the proteins, this means that the sum of charges on the positive side, i.e., the aminoterminal group (R-NH_3) equals the negative side charges (R-COO^-). They can accept or donate hydrogen ions. So, all proteins have titratable groups within their molecular structure to buffer pH changes. At normal physiologic pH = 7.4, the plasma proteins are polyvalent anions. It can be symbolized as $Prot^{n-1}$. It can be written as:

$$Protein = H^+ + Prot^-$$

They react differently with acid or alkaline environments **(Figs. 1 and 2)**.

TABLE 1: Approximate contributions of various blood buffers to the total buffer base.

Buffer type	*Total buffering (%)*
Bicarbonate	
Plasma bicarbonate	35
Erythrocyte bicarbonate	18
Total bicarbonate buffering	**53**
Nonbicarbonate	
Hemoglobin	35
Organic phosphates	3
Inorganic phosphates	2
Plasma proteins	7
Total nonbicarbonate buffering	**47**
Total	**100**

Source: Beachey W. Respiratory care: anatomy and physiology. Foundations for Clinical Practice, 2nd edition. St Louis: Mosby Inc.; 2007.

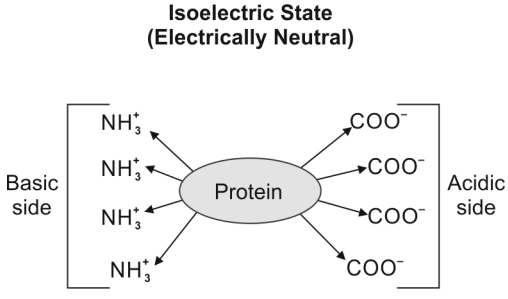

Fig. 1: Structure of a protein molecule.

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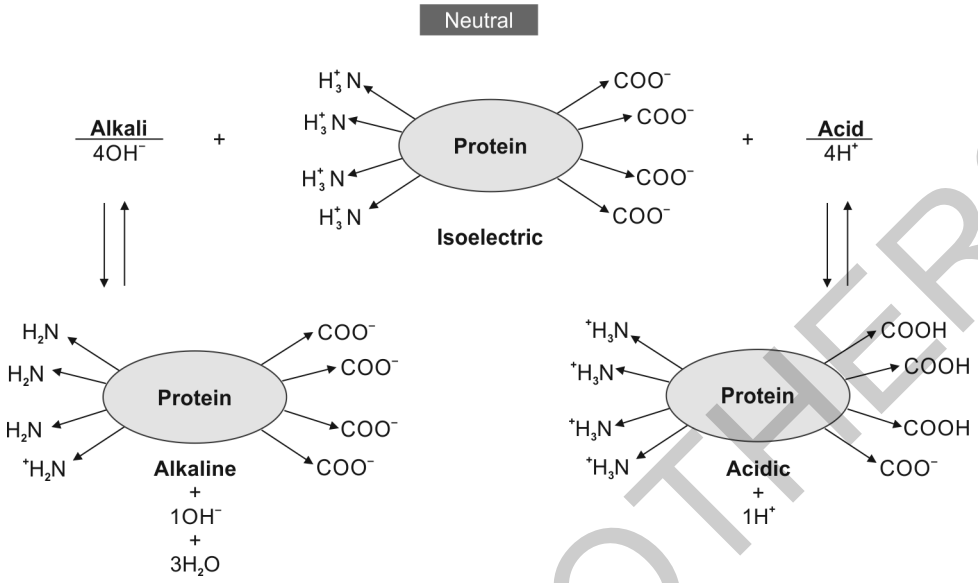

Fig. 2: Protein buffering.

Intracellular Buffers

They operate inside the cells.

- Intracellular organic phosphate complexes, adenosine diphosphate (ADP), and creatine phosphate
- Hydroxy hemoglobin/hemoglobinate molecule
- Other cellular proteins and organic-inorganic phosphates.

These are important buffer systems that are distinct and separate from those of plasma. The primary intracellular buffers are protein and organic-inorganic phosphate complexes. They operate on various tissues including bone. Hemoglobin is the other protein buffer that operates in red blood cells. Approximately 60–70% of the total chemical buffering of body fluids is inside the cells. Most of these buffering results from intracellular proteins.

Functioning of the Buffer system

From a pure chemical point of view, buffer system reactions are chemical equilibrium reactions that are governed by laws associated with them. So, it is better to repeat to reinforce the conceptualization that has been discussed in Chapter 1. The most commonly used buffer system in the body is discussed here.

The carbonic acid can dissociate into a (H^+) ion and (HCO_3^-) ions, i.e.,

$$H_2CO_3 = H^+ + HCO_3^-$$

Accordingly, the velocity of the reaction to the right is:

$$V_1 = k_1 \times (HCO_3^-)$$

The velocity of reactions which moves from right to left, is:

$$V_2 = k_2 \times (H^+) \times (HCO_3^-)$$

At equilibrium, $V_1 = V_2$. So,

$$k_1 \times (H_2CO_3) = k_2 \times (H^+) \times (HCO_3^-)$$

$$\text{Or } \frac{k_1}{k_2} = (H^+) \times \frac{(HCO_3^-)}{H_2CO_3}$$

where k_1 and k_2 are rate constants.

$$\text{Or } K = (H^+) \times \frac{(HCO_3^-)}{H_2CO_3}$$

This is the equilibrium reaction which is called the law of mass action. K is the

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equilibrium constant. Gases dissolve in water. CO_2 dissolves in water to form carbonic acid H_2CO_3. The partial pressure of CO_2 noted as the partial pressure of carbon dioxide (pCO_2) in the arterial blood is in equilibrium with that in alveolar air (So the bicarbonate buffer system is open) and normally is approximately 40 mm Hg. At 37°C (the normal body temperature), the amount of CO_2 dissolved in the plasma is:

$$CO_2 \text{ dissolved} = 0.03 \times pCO_2$$

Where 0.03 is the solubility constant for CO_2 in plasma. So, in plasma, the law of mass action can be written as:

$$CO_2 \text{ dissolved} + H_2O = H_2CO_3 = H^+ + HCO_3^-$$

Which can be simplified to:

$$CO_2 \text{ dissolved} + H_2O = H^+ + HCO_3^-$$

The law of mass action for this reaction is:

$$K = \frac{(H^+) \times (HCO_3^-)}{CO_2 \text{ dissolved in } H_2O}$$

$$\text{Or } K \times (H_2O) = (H^+) \times \frac{(HCO_3^-)}{CO_2 \text{ dissolved}}$$

Since the concentration of water is constant $K \times (H_2O)$ can be replaced by K', so the equation becomes:

Equation:

$$K' = \frac{(H^+) \times (HCO_3^-)}{CO_2 \text{ dissolved}}$$

If we solve for (H^+), the equation becomes

$$(H^+) = K' \times \frac{CO_2 \text{ dissolved}}{(HCO_3^-)}$$

In plasma at 37°C, K' is equal to 800 nmol/L, and pK equals 6.1.

So, the equation becomes

$$(H^+) = 800 \times \frac{CO_2 \text{ dissolved}}{(HCO_3^-)}$$

And if we substitute 0.03 pCO_2 for dissolved CO_2 the equation becomes:

$$(H^+) = 800 \times \frac{0.03\, pCO_2}{(HCO_3^-)}$$

$$= 24 \times \frac{pCO_2}{(HCO_3^-)}$$

This is also called the Henderson equation.

Since the normal (H^+) is 40 nMol/L, then Henderson equation becomes:

$$40 = 24 \times \frac{pCO_2}{(HCO_3^-)} = 24 \times \frac{40}{(HCO_3^-)}$$

So $HCO_3^- = 24$

Now let us calculate the pH (*See* Chapter 1). For example, the acid HA dissociates, it forms:

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

So, the equilibrium constant $Ka = H^+ \times \frac{A^-}{HA}$

Or, $Ka \times HA = (A^-) \times (H^+)$

Or, $(H^+) \times (A^-) = Ka \times (HA)$ by transposition

$$\text{Or, } (H^+) = Ka \times \frac{HA^-}{A^-}$$

By taking reciprocals of both sides, you get

$$\frac{1}{(H^+)} = \frac{1}{Ka \times \frac{HA}{(A^-)}}$$

By definition, $pH = \log \frac{1}{(H^+)}$

So, taking the logarithm of both sides, it becomes

$$\log \frac{1}{(H^+)} = \log \frac{1}{Ka} \times \frac{1}{Ka \times \frac{HA}{(A^-)}}$$

$$pH = \log \frac{1}{Ka} + \log \frac{1}{\frac{HA}{(A^-)}}$$

$$\text{Or} \quad = \log 1 - \log Ka + \log 1 - \log \frac{HA}{(A^-)}$$

$$\text{Or} \quad = 0 - \log Ka - \log \frac{HA}{(A^-)}$$

$$= pK - \log \frac{HA}{(A^-)}$$

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As log 1 is = 0

$$pH = pK + \log \frac{A^-}{HA}$$

For any molecule that can accept or donate a proton forming a conjugate base pair, the pH can be stated as

$$pH = pK + \log \frac{(Unprotonated)}{(Protonated)}$$

$$= pK + \log \frac{(Proton\ acceptor)}{(Proton\ donor)}$$

$$pH = pK + \frac{\log (HCO_3^-)}{0.03 \times pCO_2} \text{ where pK' is 6.1}$$

This is Henderson–Hasselbalch equation (HHE). Every cell respects HHE which relates the pH of a system to the degree of ionization of species under consideration. Using the Henderson equation is easy as it does not use logarithmic operations. Hasselbalch introduced logarithmic into Henderson equation.

If one studies, the HHE diligently multiple information and multiple inferences can be obtained. For better understanding, it can be stated in various formats.

The HHE statement can be rewritten as:

$$pH = pK + \log \frac{(conjugate\ base)}{(conjugate\ acid)}$$

The H_2CO_3 has been formed from the hydration of carbon dioxide CO_2 and this buffer system can be more accurately described by following a series of reactions:

$$CO_2 \text{ (gas phase)} = CO_2 \text{ (aqueous phase)}$$
$$+ H_2O = H_2CO_3 = H^+ + HCO_3^-$$

The concentration of undissociated H_2CO_3 cannot be measured in solution because it rapidly dissociates into either (CO_2 and H_2O) or (H^+ and HCO_3^-)

The HHE can again be stated as:

$$pH = pK + \log \frac{(HCO_3^-)}{CO_2 \text{ dissolved}}$$

$$= pK + \log \frac{(HCO_3^-)}{0.03 \times pCO_2}$$

The bicarbonate is the principal buffer in ECF. The reactions in these fluids can be regulated by mechanisms that control the concentration of bicarbonate and the partial pressures of carbon dioxide. Both variables are subjected to physical control each by a separate system which is briefly summarized here:

- The partial pressure of carbon dioxide is maintained at a steady level by the balance struck between the rate at which metabolic oxidations produce carbon dioxide and the rate at which lungs remove it. The rate of pulmonary ventilation (respiration rate) is regulated by respiratory centers in the brain so as to keep the partial pressure of carbon dioxide close to 40 mm Hg which is normal.
- The concentration of bicarbonate is regulated by the kidneys. The tubular epithelium reabsorbs cations and anions from the glomerular filtrate in such a way that the sum of the concentration of cations in the plasma is normally kept greater than the sum of anions other than bicarbonates by about 24–27 mmol/L maintaining electrical neutrality. So, bicarbonate concentration averaged normally as 24 mmol/L.

The steady supply of carbon dioxide from metabolic sources ensures that the balance of anions required for electroneutrality is made up by bicarbonate. The HHE can be written for better understanding as given here:

$$pH = pK + \log \frac{(kidneys)}{(lungs + respiration\ centers)}$$

Now let us discuss the other applications of HHE.

Other Applications of Henderson–Hasselbalch Equation

If the bicarbonate (HCO_3^-) and carbonic acid (H_2CO_3) exist in equal concentration the HHE becomes:

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$$pH = pK + \log \frac{(HCO_3^-)}{(H_2CO_3)}$$

$$= pK + \log \frac{(1)}{(1)}$$

As both HCO_3^- and H_2CO_3 cancel out being equal = pK + 0

So, pH = pK. Here, pK is equal to the pH of the dilute solutions in which a weak acid is half neutralized so that the concentration of the acid and of its conjugate base are equal.

Now let us consider what is pK. The pK is akin to pH. It is defined and calculated as:

$$pK = \log \frac{1}{(K)} = - \log K$$

The pH and pK are related to each other by HHE and by function too. This is more expressive if we further operationalize the HHE.

$$pH = pK + \log \frac{(salt)}{(acid)}$$

$$= pK + \log (salt) - \log (acid)$$

If we rearrange the equation it becomes:

$$0 = pK - pH + \log (salt) - \log (acid)$$

Or

$$\log (acid) - \log (salt) = pK - pH$$

$$pK - pH = \log (acid) - \log (salt)$$

This is by transposition.

The K is the equilibrium constant which speaks about the rate of a reaction. The value of pK is important and gives information and the relation between pH and pK where the solvent is water. The pK is the property of a particular solution that depends on the concentrations and identities of the components. It enables one to judge how far a reaction proceeds.

- If K is greater than 1, i.e., K >1, the equilibrium sets in when almost all reactants have entered into the reaction. The equilibrium mixture "is" or "contains" mainly products. Here the equilibrium "shifts to right."
- If K is less than 1, i.e., K <1 only a small amount of reactants are transformed into products. The equilibrium reaction is mainly reactants. Here the equilibrium "shifts to left".
- If the K = 1. The equilibrium is not static. The concentration of reactants and products remains unchanged. It reminds one of the equilibrium on balance. It should be remembered that both forward and reverse reactions proceed continuously but at the same rate. The system as a whole is in motion. Therefore, equilibrium is dynamic. The mixture contains comparable amounts of reactants and products. It means the conjugate base anion and acid cations (protons) are in equal numbers almost.

The value of K depends on the physical nature of the reactants. The electrolytes may remain as ionic solids. When it comes in contact with solvent water they are frequently broken into ions (NaCl sodium chloride, the common salt breaks into sodium Na^+ and chloride Cl^- ions). All ionic solids are strong electrolytes. The strong electrolytes in dilute aqueous solutions largely remain as ions. It is merely a way of saying that they react (dissociate) almost completely in water. For example, $HCl = H^+ + Cl^-$ and $NaCl = Na^+ + Cl^-$. It can be said that the tendency of ions to combine in dilute solutions is practically zero. In a more concentrated solution, the ions will stick together long enough to be considered an "ion pair." When an ion-pair exists, it behaves as one particle—a solid particle. This is not the case with weak electrolytes.

The value of K depends on the nature of reactants, their concentration, and prevalent temperature but does not depend on the catalyst. The presence of a catalyst in a system only cuts the time needed for equilibrium to set in. In a balanced chemical equation, it

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tells us the number of moles that react and give products. As it is learned strong acids like HCl are practically 100% dissociated in water. This means that the number of moles of H_3O^+ ($H_2O + H^+$) in the solution is the same as the number of moles of HCl present.

For example, a 0.010 M HCl solution contains 0.010 H_3O^+. The pH of this solution can be calculated:

$$pH = \log \frac{1}{H} = -\log H$$
$$= \log \frac{1}{0.010} = \log \frac{1}{10^{-2}}$$
$$= \log 1 - \log 10^{-2}$$
$$= 0 - (-2) = 2$$

So, the pH of the solution is 2. The K value of the solution is $1.3 \times 10^{+6}$

$$\text{The pK is } \log \frac{1}{K} = -\log K$$
$$= \log \frac{1}{1.3 \times 10^{+6}}$$
$$= \log 1 - \log (1.3 \times 10^{+6})$$
$$= 0 - \log 1.3 - \log 10^{+6}$$
$$= -0.113 - 6$$
$$= -6.113$$

Similarly, a strong base like NaOH is practically 100% dissociated in water, so that a 0.010M NaOH solution contains 0.010 M OH^-. So, what is the pH of the solution? First, find the H^+ concentration.

As told before in Chapter 1

$$(H^+) \times (OH^-) = 10^{-14}$$

The H concentration becomes

$$\frac{10^{-14}}{10^{-2}} = 10^{-14+2} = 10^{-12}$$

$$\text{So, pH} = -\log H = -\log 10^{-12}$$
$$= -(-12) = +12$$

Unlike strong acids and bases, the solutions of weak acids and bases contain positive and negative ions in equilibrium with molecules of undissociated acid or base. Acetic acid is a typical weak acid that exists largely as molecules in dilute aqueous solutions. Its tendency to form ions is very small, as seen from the small value of equilibrium constant ($K = 1.8 \times 10^{-5}$ very much smaller than 1.0) ($CH_3COOH = H^+ + CH_3COO^-$)

$$\text{The pK} = \log \frac{1}{K} = -\log K$$
$$= -\log 1.8 \times 10^{-5}$$
$$= -\log 1.8 + \log 10^{-5}$$
$$= -0.2553 + 5$$
$$= 4.7447$$

The equilibrium constant describes the relationship between hydrogen ion concentration and the undissociated acid concentration. In undissociated acid, the hydrogen ion is still attached. You can call it protonated. If the hydrogen ion is not associated, you call it deprotonated. Based on the relationship between pK and the pH of a solution, one can predict whether a compound is/will be protonated (undissociated) or deprotonated (ionized). If the pH is lower than the pK, then compound (acid concerned) will be protonated. There will be more unionized acid. If the pH is higher than pK then the compound will be deprotonated (ionized).

FURTHER CONSIDERATIONS ON THE CHARGE ON COMPOUND

The acids are neutral when protonated and negatively charged when deprotonated. These theoretical considerations have practical clinical importance for many organ systems of the body. Many body fluids function and respect the pK and pH values. The functions of the kidney are particularly important. It helps maintain the milieu intérieur through filtration absorption excretions and secretions. The kidney tubules handle

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water, sodium, potassium, calcium, chloride, bicarbonate, phosphate, sulfate, ammonia and its byproducts urea, etc., glucose, fructose, a large variety of amino acids, several vitamins, and hormones including various drugs and their metabolites. The drug's actions and functions also depend on local pH and pK values and functions.

In normal situations, the pH of blood is 7.4. The average value of urine pH is 6 but the kidneys have the capacity to produce urine with a pH value of 4.0, but it can range from 4.5 to 8. The different laboratories have different ranges for "normal" pH levels. One of the major factors affecting the urine pH is the food that one person eats. The foods may be acidic or alkaline. Needless here to say any acid that has a pK value higher or "not so much lower" or nearer the pH of urine can be excreted to an appreciable extent in unionized forms. Acids with pKs high enough to appear in the urine in the undissociated form are monobasic phosphate ($H_2PO_4^-$ = HPO_4^- + H) pK 6.8, β-hydroxybutyric acid pK 4.7, creatine (CrH = Cr^- + H) pK 5.0 **(Table 2)**. For most acids that enter the ECF, their excretion as undissociated acid is impossible. Hydrochloric acid and sulfuric acid are completely dissociated at all pHs as they are strong acids. They are excreted as salt. Even weak acids having pKs much below 5 must be excreted almost as salts.

The importance of pK is that it represents the pH at which maximum buffering capacity can be achieved for that particular reaction. The extent to which an acid is unionized at a given pH can be deduced. The pK – pH indicates. The buffer is efficient when the pH of the solution is within +/– 1.0 pH units of its pK. If the pH is outside these limits buffering will still occur.

The pK of common acid–base is discussed in **Table 2**.

TABLE 2: The pK of common acid–base.

Strength of acid	*Order of pK*	*Acid = Proton + conjugate base*	*Strength of base*
Very strong	–7	$HCl = H^+ + Cl^-$	Very weak
↑	–2	$H_3O^+ = H^+ + H_2O$	↓
	3.6	Acetoacetic acid = H^+ + Acetoacetate$^-$	
	3.86	Lactic acid = H^+ + lactate$^-$	
	4 6.1	$H_2CO_3 = H^+ + HCO_3^-$ (when closed system) $H_2CO_3 = H^+ + HCO_3^-$ (when the system is open)	
	4.7	β-hydroxybutyric acid = H^+ + butyrate$^-$	
	5	Creatinine = H^+ + Cr^-	
	5.35	Uric acid = H^+ + urate$^-$	
	5.6	Acetic acid = H^+ + acetate$^-$	
	5.5–8.5	Protein = H^+ + Prot$^-$	
	6–7.5	Organic protein = H^+ + Org Pro$^-$	
	7	$H_2PO_4 = H^+ + HPO_4^-$	
	7–8	H hemoglobin = H^+ + Hemoglobin$^-$	
	9.5	$NH_4^+ = H^+ + NH_3$	
Very weak	16	$H_2O = H^+ + OH^-$	Very strong

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A few examples in **Table 2**, illustrate that an acid may be either an uncharged molecule (electrically not dissociated), a cation, or an anion. When a strong acid such as hydrochloric acid is dissolved in water the uncharged acid is so much stronger than OH_3^+ that it is almost completely converted into H^+ and the weak conjugate base Cl^-.

Thus, $HCl + H_2O = OH_3^+ + Cl^-$ with equilibrium very far indeed to the right.

The bases encountered in the body are either anions or uncharged molecules. It is clear that cations cannot be called bases in the sense of the present definition. Indeed, on account of their positive charges, cations must always repel protons and could not be expected to unite with them. Neither a cation like Na^+ can give off a proton. The metallic cations are neither acids nor bases. They neutralize the electrostatic charges upon the anions which accompany them. In plasma at 37°C, the pK of the bicarbonate-pCO_2 buffer system is 6.1.

It also allows us to say that the acid–base disorders are attributable to the numerator of the ratio (metabolic process), to the denominator of the ratio (respiratory process), or to both (mixed or complex acid–base disorders of HHE. A broad and functional view of HHE can be seen in the picture **(Fig. 3)**.

Similar HHE can be stated for other buffer systems.

$$pH = pK + \log \frac{(HPO_4^{-2})}{H_2PO_4^-}$$

Where, pK = 6.8

$$pH = pK + \log \frac{(NH_3)}{NH_4^+}$$

Where, pK = 9

All the buffer systems are in relation with each other and in equilibrium while encountering a change in acid–base.

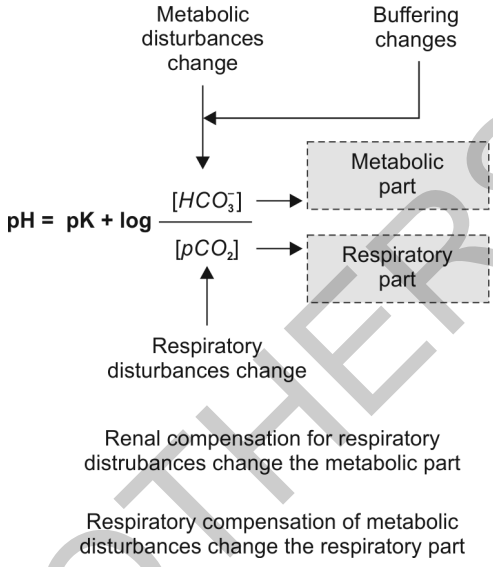

Fig. 3: The centrality of HHE.

$$(H^+) = K_1 \frac{0.03pCO_2}{HCO_3^-} = K_2 \frac{H_2PO_4^-}{HPO_4^{-2}} = K_3 \frac{(HA)}{A^-}$$

is the isohydric principle.

ELECTRICAL NEUTRALITY

Acid–base mixtures are mixtures of ions. In all physiological fluids electrical neutrality must be maintained. The number of positive charges on cations must be balanced (covered) by an equivalent number of negative charges on anions. It is important to consider the behavior of individual ions separately and not in terms of electrolytes from which they are derived. It is equally important to remember that a given solution cannot receive a supplement of additional cation or anion without that being accompanied by an equivalent coverage of anion or cation. The electrolyte composition of body fluids differs in body fluids, but the main compositions are given. The proteins are minimal in interstitial fluids compared to ECF and intracellular fluids **(Figs. 4 and 5)**.

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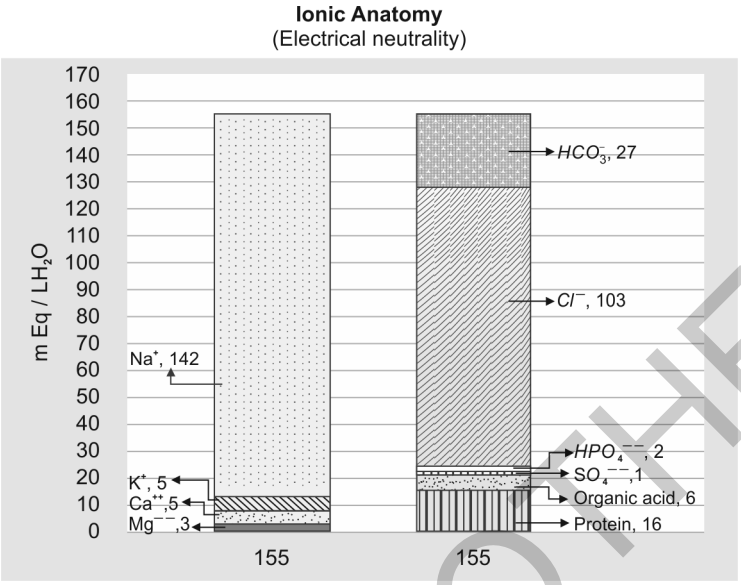

Fig. 4: The ionic anatomy of serum after James L Gamble.

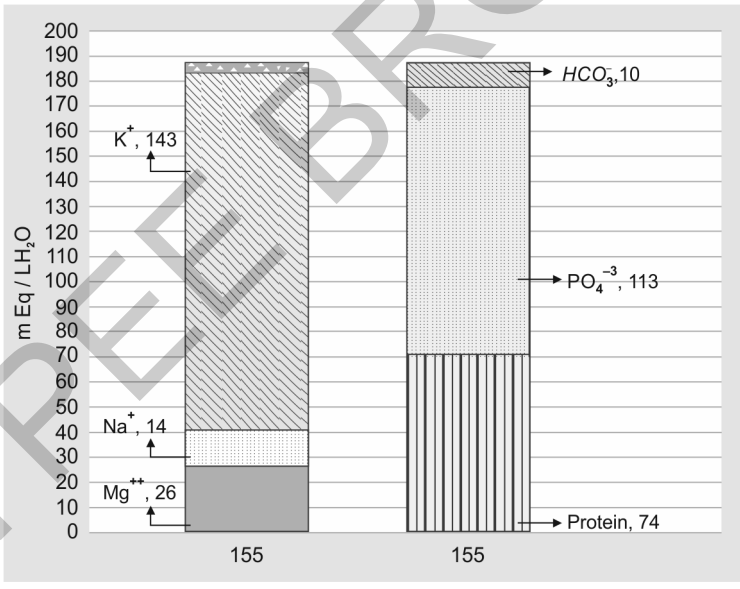

Fig. 5: Ionic composition of Intracellular fluid.

The electrolyte imbalances can have profound effects on buffering mechanisms specifically renal and hemoglobin buffers.

PRACTICAL POINTS

- pK represents the pH at which maximum buffering takes place. If the substance (acid) has a pK value greater/nearer to the pH of the urine, it is excreted in the urine in an undissociated form.
- Substances pK value less than the pH of urine are excreted as salts.
- The HHE is respected by every cell for solute transportation.

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SELF-ASSESSMENT REVIEW

One or more options are correct.

1. **What is true about buffers?**
   a. Withstands radical changes in pH
   b. Helps maintain the pH of the blood and other tissue fluids in a narrow range of 7.35–7.45.
   c. It is a mixture of either acid and its alkali salt or a weak base and its acid salt.
   d. All of the above
2. **Which is the most important ECF buffer?**
   a. Hb-HHb
   b. Carbonate–carbonic acid
   c. Phosphate buffers
   d. Plasma proteins
3. **Which is not an intracellular buffer?**
   a. ADP
   b. Creatinine
   c. Bicarbonate
   d. Hemoglobin
4. **What percentage of total chemical buffering is intracellular?**
   a. 60–70%
   b. 50–60%
   c. 30–50%
   d. None of the above
5. **Which statement is correct? pCO_2 is maintained by:**
   a. Metabolic oxidation
   b. Respiratory ventilation
   c. Kidney participation
   d. Both metabolism and respiratory ventilation.
6. **In a normal steady state pCO_2 value is close to:**
   a. 40 mm Hg
   b. >40 mm Hg
   c. <40 mm Hg
   d. None of the above
7. **Which statement is correct if K = 1?**
   a. Products are more than reactants
   b. Reactants are more than products
   c. Products and reactants are equal in number
   d. None of the above
8. **The value of K does not depend on:**
   a. Catalyst
   b. Nature of reactants
   c. Prevailing temperature
   d. Concentration of reactants
9. **Which of the following is a solid-state electrolyte?**
   a. HCI b. NaCl
   c. Proteins d. NH_4CI
10. **The pH of the urine of a normal individual is:**
    a. 6 b. 4
    c. 4.5–8 d. 9
11. **pK and pH of a chemical species are concerned when the solvent is water:**
    a. Determines the quantum of hydrogen ion concentration in solution, unionized quantum of acid in solution
    b. pK is more specific compared to pH
    c. pK is the pH value at which a chemical species will accept or donate a proton
    d. All of the above
12. **If the pK of the chemical species is more than the pH of urine it is mostly excreted as:**
    a. Ionized form
    b. Unionized form
    c. both equally ionized and nonionized
    d. All of the above
13. **If the pK of the chemical species is less than the pH of urine it is mostly excreted as:**
    a. Ionized form
    b. Unionized form

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c. Equally, both ionized and unionized form
d. All of the above

ANSWERS

1. (d); **2.** (b); **3.** (c); **4.** (a); **5.** (d); **6.** (d); **7.** (c); **8.** (a); **9.** (b); **10.** (c); **11.** (d); **12.** (b); **13.** (a)

SUGGESTED READING

1. Berend K, de Vries APJ, Gans ROB. Physiological approach to assessment of acid-base disturbances. N Eng J Med. 2014;371(15):1434-45.
2. Bidoni A, Tuazam DM, Henring TA. Regulation of whole acid-base balance in Dubose Hamm—acid base and Electrolyte Disorders. A companion to Brenner and Rectors Kidney. Philadelphia: Elsevier; 2002. pp. 1-21.
3. Frank B. Chemistry: A Modern Introduction, 2nd edition. WB Saunders Company; 1978.
4. Ghosh AK. Diagnosing Acid-Base Disorders. J Assoc Physicians India. 2006;54:720-4.
5. Khitan ZJ, Shapiro JI. Pathogenesis and Management of Metabolic Acidosis and Alkalosis in Renal and Electrolyte Disorders, 8th edition. Baltimore: Wolters Kluwer Publication; 2018.
6. Nuttal FQ. Serum Electrolytes and Their Relation to Acid-Base Balance. Arch Intern Med. 1965;116(5):670-80.
7. Regulation of hydrogen ion concentration of body fluids 4 = 103-112. In: Brobeck JR (Ed). Best and Taylor's Physiological Basis of Medical Practice, 9th edition. Baltimore: The Williams & Wilkins Company; 1973. p. 1390.
8. Seal US. The chemistry of buffers. Arch Intern Med. 1965;116(5):658-64.

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ACID–BASE REGULATION

This informative book about Acid–Base Regulation describes in detail with examples of various diseases encountered in various divisions of medicine in critical state. It is a well-meaning book meant for practicing physicians, pulmonologists, intensivists, house officers, medical students doing clinical clerkship, nurses in ICU, and lab technicians.

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SK Sahoo MBBS MD MBA (HCA) is associated with Apollo Hospital Noida, Uttar Pradesh, India as a Senior Consultant in Medicine and Diabetes. He has more than 30 years of experience in various fields of medicine and management of critically ill patients. This book is an affirmative testimony to his work and long close association with patients.

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