Non-Logarithmic Equation for Acid Base State in the Blood

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Abstract: Acid-base balance is maintained and regulated by the renal and respiratory systems *via* changing the bicarbonate pair (HCO_3^- and PCO_2) which is considered the most important buffering system in our body. Henderson-Hasselbalch equation was used for a long time to describe acid base state but we found a new simpler and easier way by reversing their negative logarithmic equation. This method enables us to know type of acidosis (whether respiratory or renal), calculate acid excess instead of negative base excess and compensate this excess by a simple way.

Keywords: New equation, acid, base.

INTRODUCTION

According to Ad Hoc Committee of the New York Academy of Sciences [1], the law of mass action can express any chemical reaction reaching equilibrium and in case of a weak acid:

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

At equilibrium the product of the concentrations of H^+ and A- is a constant fraction of the concentration of HA or:

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

where K is constant.

This equation can be rearranged as:

$$[H+] = \frac{K \times [HA]}{[A-]}$$

The negative log of this equation is Henderson-Hasselbalch equation [2]:

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

Bicarbonate-carbonic acid buffer is the most important buffer in human body that can be presented [3] as:

$$[H^+] = K \times \frac{H_2CO_3}{[HCO_3^-]}$$

By substitution of $H_2CO_3 = Pa\ CO_2\ X\ k_H\ CO_2$

$$[H^+] = K \times \frac{Pa\ CO_2 X\ k_H\ CO_2}{[HCO_3\text{-}]}$$

Where:

$[H^+]$ is the arterial blood H^+ ion concentration in nmol/L.

K=24 which the numerical value of the solubility coefficient of carbon dioxide and the dissociation constant of carbonic acid.

$PaCO_2$ is the partial pressure of carbon dioxide in the arterial blood.

$k_H\ CO_2$ is the Henry's law constant for the solubility of carbon dioxide in blood which is approximately 0.03 mmol/(mL-mmHg).

$[HCO_3^-]$ is the arterial blood bicarbonate concentration in mmol/L.

Henderson-Hasselbalch equation is the negative logarithm of previous equation which reverses the ratio between CO_2 and HCO_3:

$$PH = PK + \log \frac{[HCO_3^-]}{PaCO_2 \times 0.03}$$

Where:

pH is the negative logarithm of the H^+ ion activity (Ha^+).

pK is the negative logarithm of the dissociation constant of carbonic acid = 6.1.

Our suggestion is to return the negative logarithm of the equation to normal value:

$$[H^+] = K \times \frac{Pa\ CO_2 X\ k_H\ CO_2}{[HCO_3^-]}$$

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First, reverse the ratio to normal one:

The left side of the equation: $H^+ = \frac{1}{10^{7.4}} = 10^{-7.4}\, g/L$

The right side of the equation: $K = \frac{1}{10^{6.1}} = 10^{-6.1}$

Second, get rid of the logarithm by multiplying both sides of the equation by 10^9 to change gram to nanogram.

The left side of the equation:

$H^+ = 10^{-7.4} \times 10^9 = 10^{1.6} = 40$ nanogram,

The right side of the equation:

$$K = 10^{-6.1} \times 10^9 = 10^{2.9} = 800$$

By substitution of these Figures in the normal equation:

H^+ = 40 nanogram, K = 800, $PaCO_2$ is 40 mmHg, $k_H\, CO_2$ is approximately 0.03 mmol/(mL-mmHg), and HCO_3^- is 24 mEq/L.

$$40 = 800 \times \frac{40\ X\ 0.03}{24}$$

That is the final equation of normal individual.

OUR OBSERVATION

- $\mathbf{40} = 800 \times \frac{\mathbf{40}\ \mathrm{X}\ 0.03}{24}$

The Figure of H^+ = the Figure of Pa CO_2.

- ○ When **pH = 7.4**, so H^+ concentration is = $10^{-7.4} \times 10^9 = 10^{1.6}$ = ***40*** nanogram.
- ○ When **pH = 7.3,** so H^+ concentration is = $10^{-7.3} \times 10^9 = 10^{1.7}$ = ***50*** nanogram.
- ○ When **pH = 7.2**, so H^+ concentration is = $10^{-7.2} \times 10^9 = 10^{1.8} \approx$ ***60*** nanogram
- ○ When **pH = *7.1,*** so H^+ concentration is = $10^{-7.1} \times 10^9 = 10^{1.9}$ = ***80*** nanogram.

- When both values (pH and Pa CO_2) increase by the same manner, we found that HCO_3^- = 24 mEq/L (a constant value).

That is because the incremental increase in both sides has the same ratio i.e.

H^+ concentration can increase $\frac{10\ \textit{times}}{800\ (K)} = \frac{1}{80}$

$PaCO_2$ when increases $10 \text{ times} = 10 \times \frac{3}{100} \div 24 \text{mEq}.HCO_3^- = \frac{1}{80}$

We can deduce that:

- When there is acidosis (i.e. H^+ >40), and both H^+ (e.g. 50) and CO_2 (e.g. 50) are equally increasing, the acidosis is *totally respiratory*.
- When there is acidosis (e.g. H^+ ion 50) and $PaCO_2$ is normal (e.g. 40), the acidosis is *totally metabolic*.
- When there is acidosis and $PaCO_2$ is increased (e.g. 50) but not equal to H^+ ion (e.g. 60) so it is *mixed respiratory and metabolic acidosis*.
- When there is acidosis (e.g. H^+ ion 50) and $PaCO_2$ is decreased (e.g. 30), so there is *metabolic acidosis partially corrected by respiratory alkalosis*.
- We can calculate the difference between H^+ and $PaCO_2$ as an acid excess instead of base excess in minus value in regular manner.
- Each 10 of acid access can be compensated by:

$$10 \quad \times \quad \frac{3}{10} \quad \times \quad \frac{1}{3} = \boxed{1}\ \text{mEq HCO}_3\text{/Kg}$$

↓ ↓ ↓

(Increase in H^+) $\left(10[CO_2] \times \frac{3}{100} = \frac{3}{10}\right)$ (Third body weight)

So we can give 1 mEq HCO_3/Kg for each 10 difference between H^+ ion and $PaCO_2$ to correct the acid excess which is calculated in simpler and easier way than the base excess.

CONCLUSION

We found a new simpler and easier way to know type of acidosis (whether respiratory or metabolic), calculate acid excess instead of negative base excess and compensate this excess by a simple way.

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