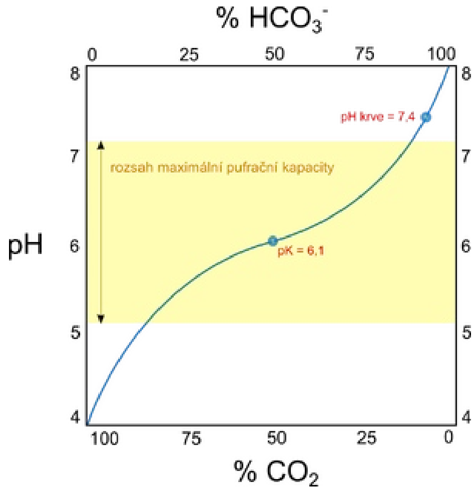
Bicarbonate buffer

The "bicarbonate buffer system" (also "bicarbonate") is the most important and effective buffering system in the body. Especially in the blood, where it accounts for up to 53% buffering capacity.[1] Its importance lies in its good ability to maintain a stable pH mainly due to the fact that the concentration of both components can change independently - **CO_2</ sub>** by breathing, **HCO_3^-** by kidney and liver activity. Therefore, the bicarbonate buffer in the body is referred to as an **open buffering system**.

The reaction proceeds as follows: $CO_2 + H_2O \rightleftharpoons HCO_3^- + H$< sup>+.



Bicarbonate buffer titration curve

Buffers composed of weak acids and their salts (or weak bases and their salts) with the same substance concentration have the greatest buffering capacity, i.e., more precisely, for which **$pH = pK_A$'**. The optimal blood pH value is **7.4'** ± 0.04. The pK_A value for bicarbonate buffer is **6.1**. Thus, it appears that this buffer will not buffer pH fluctuations very well. But the opposite is true. For a better idea, here is an example:

To Henderson-Hasselbalch equation

$$pH = pK_A + \log \frac{[HCO_3^-]}{[CO_2]}$$

we substitute physiological concentrations

HCO_3^- = 24 mmol/l and CO_2 1.2 mmol/l.
(Ratio of base to acid is therefore 20:1.)

$$pH = 6.1 + \log \frac{24 mmol/l}{1.2 mmol/l}$$ Resulting pH = **7.4** .

In the case of a closed system, after the addition of H^+, conjugated acid CO_2 is formed, which cannot escape from the system, and thus its concentration rises. An increase in the concentration of CO_2 by 2 mmol/l is reciprocally balanced by a decrease in the concentration of HCO_3^-.

$$pH = 6.1 + \log \frac{22 mmol/l}{3.2 mmol/l}$$ Resulting pH = **6 ,93**. (In this case, the buffering capacity of the buffer is very small, because the value of 6.93 is quite far from 7.4.)

However, if the resulting CO_2 is removed (exhaled) from the system, as is the case with the bicarbonate **open system**, only the concentration changes with the addition of H^+ HCO_3^-. The ratio of HCO_3^- to CO_2, and thus the pH value, will shift much less.

$$pH = 6.1 + \log \frac{22 mmol/l}{1.2 mmol/l}$$ Resulting pH = **7 ,36**.

Summary: An increase in H^+ in the blood leads to the production of CO_2, which is soon exhaled in the lungs, allowing a constant pCO_2< /sub>, i.e. a concentration of 1.2 mmol/l.

Links

Related Articles

- Buffers
- Henderson-Hasselbalch equation
- pH
- Protein Buffer System
- Buffer systems

Reference

1. FONTÁNA, Josef. *Acid-base balance* [lecture for subject Biochemistry, specialization General medicine, 3.LF Charles University]. Prague. 30.3.2011.

References

- LEDVINA, M. *Biochemie pro studující medicíny II.* 2. edition. Nakladatelství Karolinum, 2009. 281 pp. ISBN 978-80-246-1415-1.