A Buffer System

An Introduction to the
Bicarbonate Buffer System

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

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Equilibrium

When chemical reactions are in equilibrium, the rate of reaction is the same in both directions. Equilibrium does not mean that the concentrations are the same, only that the rate of reactions is proceeding at the same rate ($\rightleftharpoons$) in both directions.

$$CO_2 \quad + \quad H_2O \quad \rightleftharpoons \quad H_2CO_3$$

However, when a system is at equilibrium, and then the concentration on one side of the reaction changes, equilibrium will be disturbed. In this example as the CO_2 concentration increases, there will be more CO_2 to react with H_2O, and therefore the reaction will be proceed to the right and more H_2CO_3 will be produced. This will continue until equilibrium is reestablished.

$$\uparrow CO_2 \quad + \quad H_2O \quad \rightleftharpoons \quad H_2CO_3$$

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Likewise, if the concentration on the other side of the reaction increases, then the rate of reaction will shift in the other direction until equalibrium is reestablished.

Review of Essential Chemical Reactions

Note: reactions are in equilibrium: the rate of reaction is proceeding in the two directions at the same rate.

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

Below, equilibrium is disturbed. Increased CO_2 concentration allows for the production of more H_2CO_3. The increase in H_2CO_3 then causes an increase in $H^+$and HCO_3^-.

$$\uparrow CO_2 + H_2O \rightleftharpoons \uparrow H_2CO_3 \rightleftharpoons \uparrow H^+ + \uparrow HCO_3^-$$

As the reactions proceeds to the right, the concentration of CO_2 will go down, until equilibrium is reestablished.

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

Likewise, if the H^+ concentration goes up, the reaction will shift to the left, resulting in an increase in H_2CO_3, and then an increase is CO_2 and H_2O.

$$\uparrow CO_2 + \uparrow H_2O \rightleftharpoons \uparrow H_2CO_3 \rightleftharpoons \uparrow H^+ + HCO_3^-$$

As the reactions proceeds to the left, the concentration of H^+ will go down, until equilibrium is reestablished.

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

In the illustration below, reactions are depicted as being in perfect equilibrium. There is no metabolism, there is no movement of gases etc. This state does not exist in a living person.

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$$CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons H^+ + HCO_3^- \xrightleftharpoons[\text{Transport}]{\text{Blood}} H^+ + HCO_3^- \rightleftharpoons H_2CO_3 \rightleftharpoons CO_2 + H_2O$$

However, when alive, there is a "healthy disequilibrium", where body cells producing CO_2 will drive a chain reaction that will lead to the elimination of CO_2 by the lungs.

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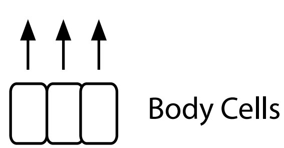
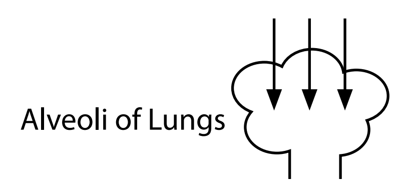
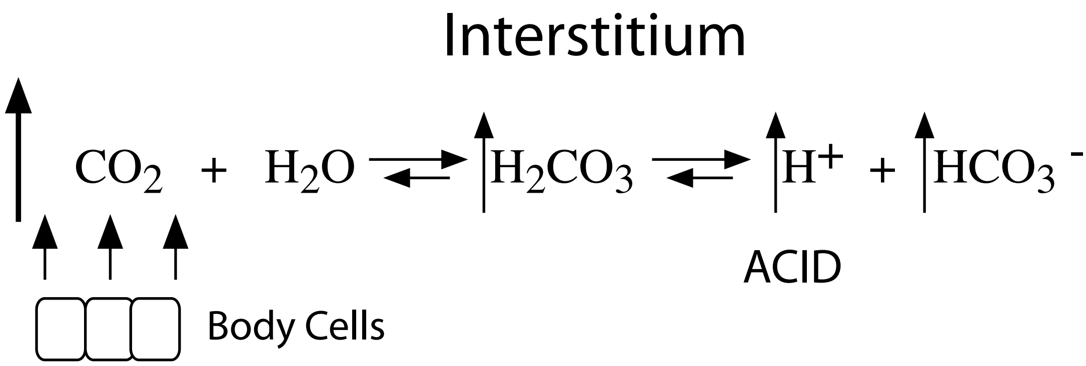
$$CO_2\uparrow + H_2O \rightleftharpoons \uparrow H_2CO_3 \rightleftharpoons \uparrow H^+ + \uparrow HCO_3^- \xrightleftharpoons[\text{Transport}]{\text{Blood}} \uparrow H^+ + \uparrow HCO_3^- \rightleftharpoons \uparrow H_2CO_3 \rightleftharpoons \uparrow CO_2 + H_2O$$



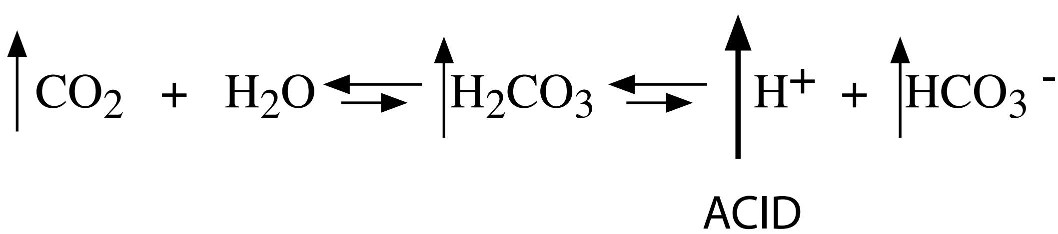


Carbon Dioxide and Acid Production

Note that if interstitial cells produce more CO_2, then there will be an increase in H^+. The pH is becoming more acid.



By extension, if there is an increase in H^+ for a reason other than CO_2 production, such as from the diet, the reaction can proceed to the left.

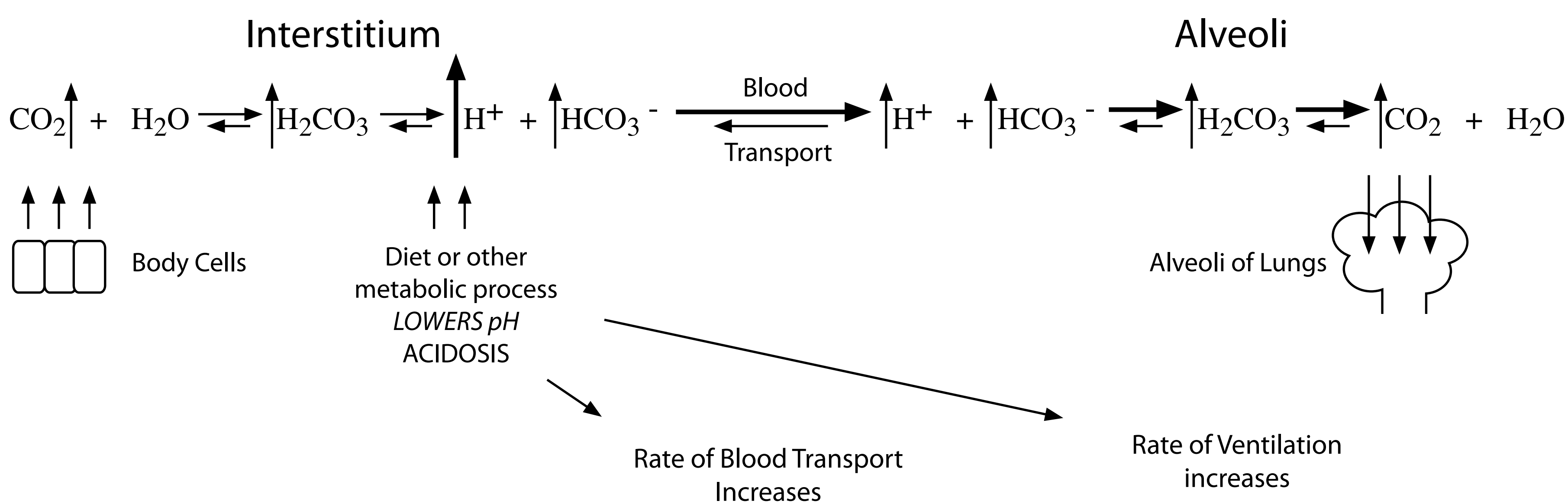


Buffer System

The pH of the blood needs to be maintained within a narrow range of the physiological set point of 7.35 - 7.45. Therefore, homeostatic mechanisms exist to "buffer against" changes in the acid / base balance. A buffer system is a collection of chemical reactions that can remove H^+ and OH^- thereby minimizing a disruption of pH. A buffer system can convert a strong acid into a weak acid and a strong base into a weak base. There are many buffer systems in the body. The carbonic acid / bicarbonate buffer system is one of the more important, and uses the chemical reactions and principles covered so far.

Acid Production

If there is an increase in acid concentration, such as from the diet or some other metabolic process, then equilibrium will be further disturbed and the reaction will proceed in a direction that favors equilibrium. Since the cells of the body will not absorb CO_2, the reaction will not go to the left. Since CO_2 can be eliminate by the lungs, the reaction will proceed to the right. Note that as the CO_2 is eliminated through the lungs, excess H+ is now in the form of water, and has been removed out of "harms way". We have restored proper pH.



Base Prodcuction

Recall that pure water H_2O is in perfect equilibrium with H^+ and OH^-.

$$H^+ + OH^- \rightleftharpoons H_2O$$

If there is an increase is OH , this will react with H+ and produce H_2O. H+ is still produced due to the production of CO_2, but there will now be less net H+ and therefore the reaction will proceed to the right at a slower pace, and ventilation rate will decrease. We will have restored pH.