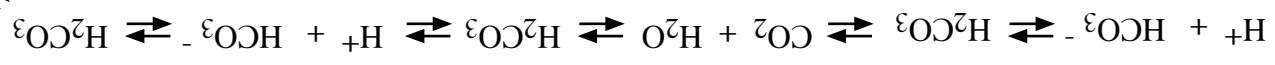


# A Buffer System

## An Introduction to the *Bicarbonate Buffer System*



*By Noel Ways*

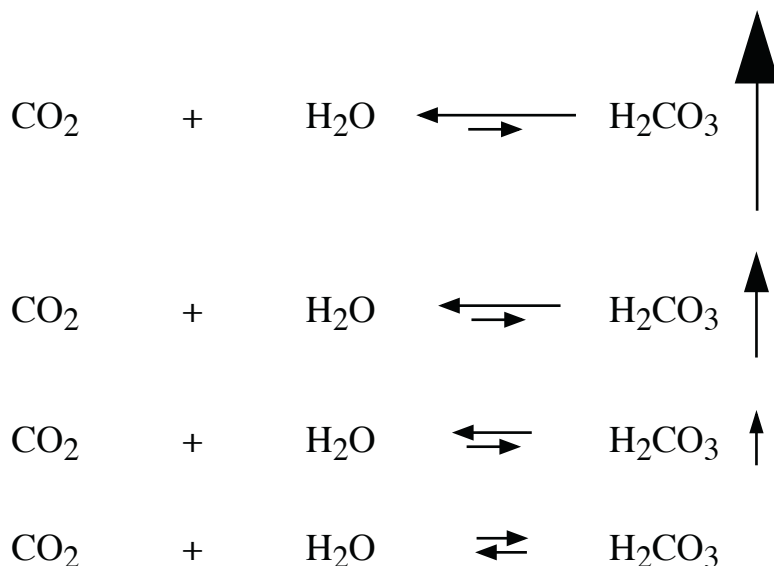
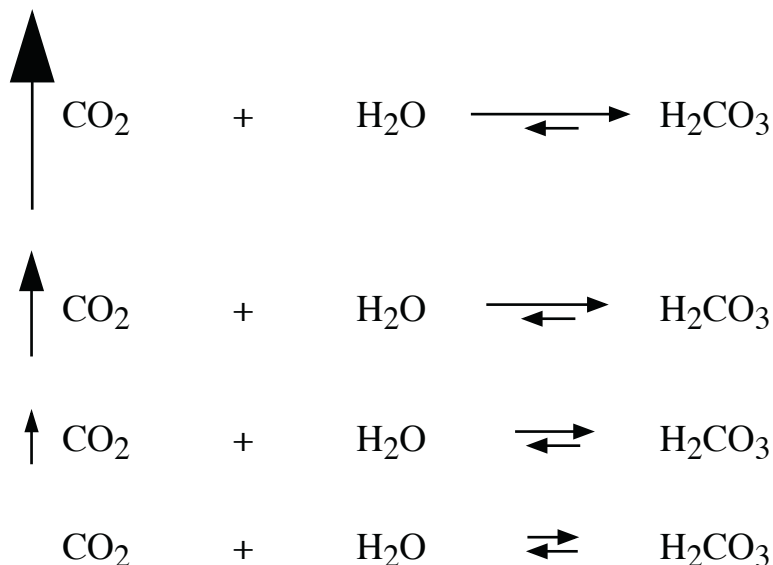


## Equilibrium

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.



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  $\text{CO}_2$  concentration increases, there will be more  $\text{CO}_2$  to react with  $\text{H}_2\text{O}$ , and therefore the reaction will be proceed to the right and more  $\text{H}_2\text{CO}_3$  will be produced. This will continue until equilibrium is reestablished.



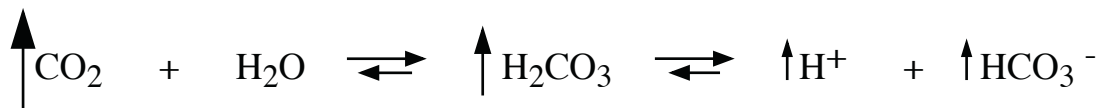
Likewise, if the concentration on the other side of the reaction increases, then the rate of reaction will shift in the other direction until equilibrium is reestablished.

## Review of Essential Chemical Reactions

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



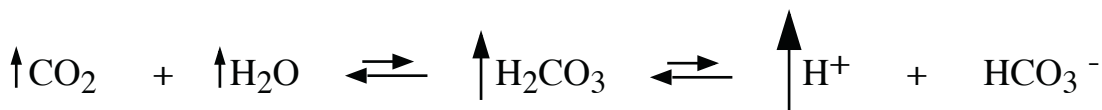
Below, equilibrium is disturbed. Increased  $\text{CO}_2$  concentration allows for the production of more  $\text{H}_2\text{CO}_3$ . The increase in  $\text{H}_2\text{CO}_3$  then causes an increase in  $\text{H}^+$  and  $\text{HCO}_3^-$ .



As the reaction proceeds to the right, the concentration of  $\text{CO}_2$  will go down, until equilibrium is reestablished.



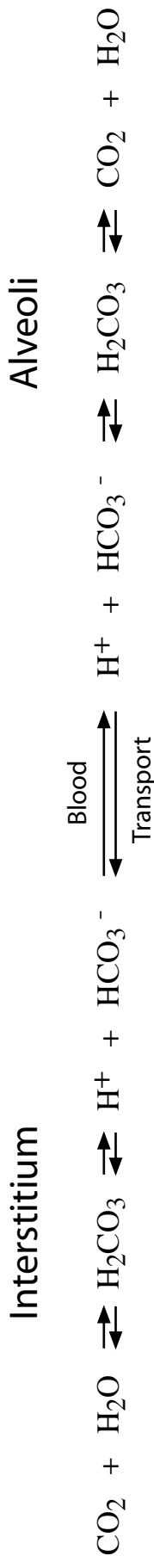
Likewise, if the  $\text{H}^+$  concentration goes up, the reaction will shift to the left, resulting in an increase in  $\text{H}_2\text{CO}_3$ , and then an increase in  $\text{CO}_2$  and  $\text{H}_2\text{O}$ .



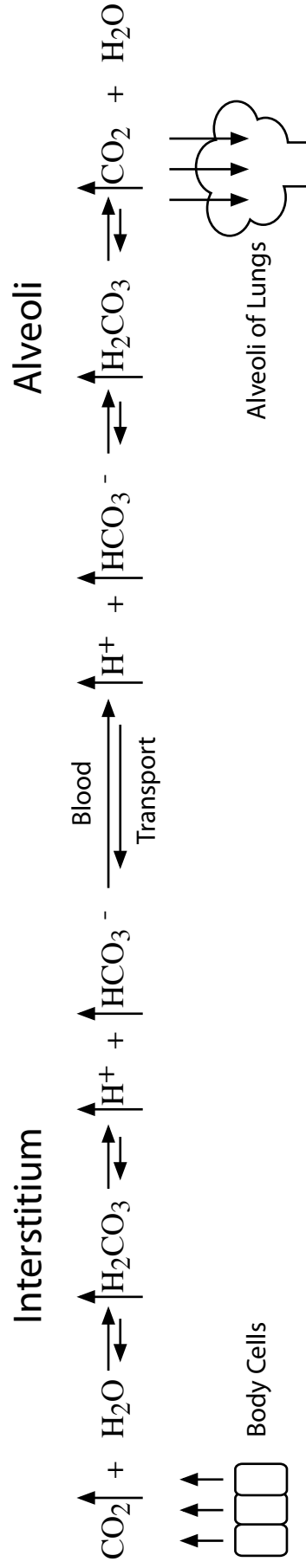
As the reaction proceeds to the left, the concentration of  $\text{H}^+$  will go down, until equilibrium is reestablished.



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.

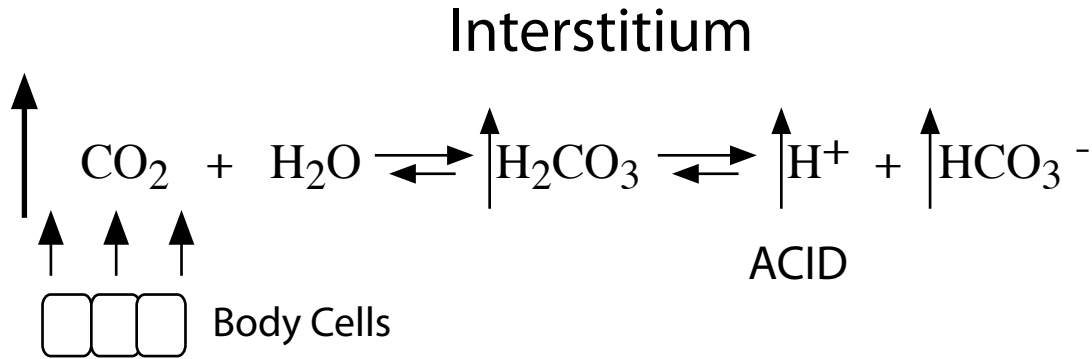


However, when alive, there is a “healthy disequilibrium”, where body cells producing CO<sub>2</sub> will drive a chain reaction that will lead to the elimination of CO<sub>2</sub> by the lungs.

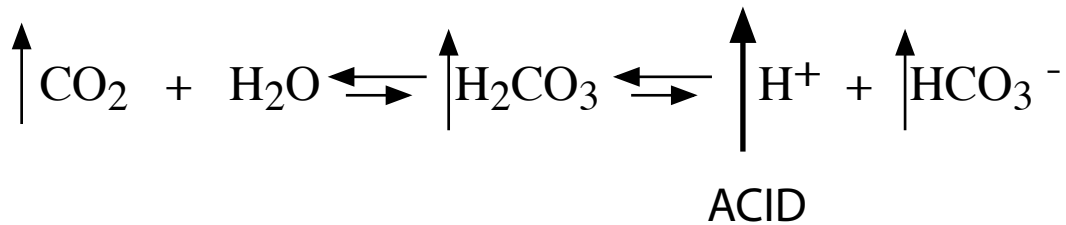


## Carbon Dioxide and Acid Production

Note that if interstitial cells produce more  $\text{CO}_2$ , then there will be an increase in  $\text{H}^+$ . The pH is becoming more acid.



By extension, if there is an increase in  $\text{H}^+$  for a reason other than  $\text{CO}_2$  production, such as from the diet, the reaction can proceed to the left.



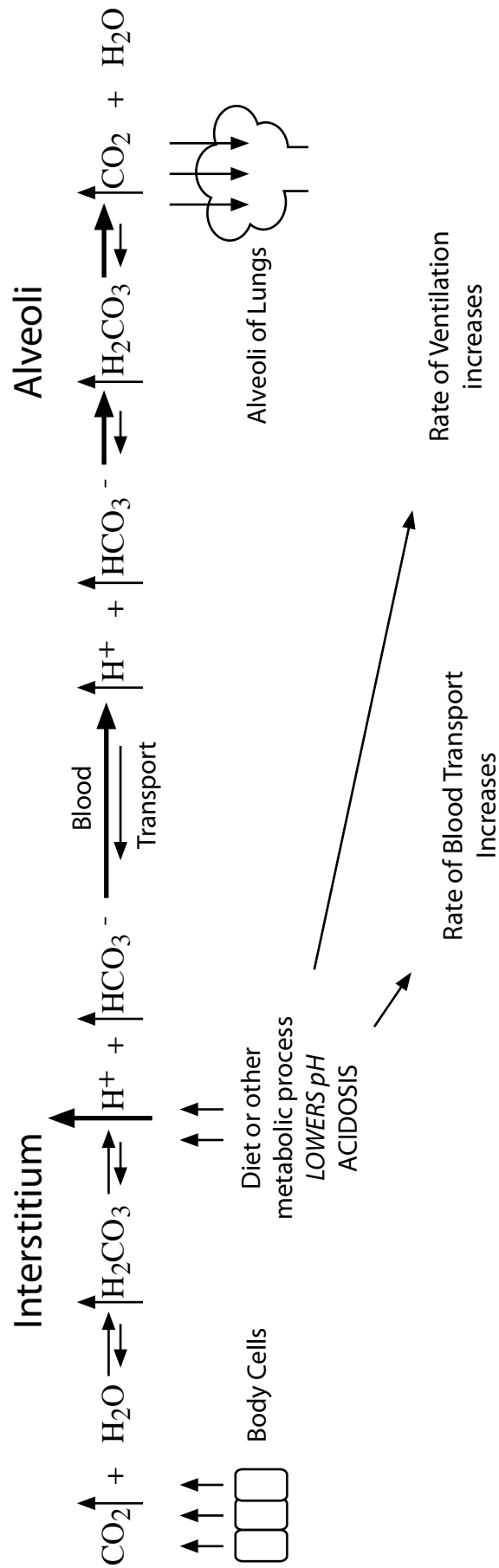
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## Buffer System

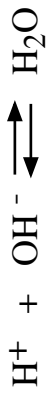
The pH of the blood needs to be maintained within a narrow range of the physiological set point of 7.35 - 7.45. Therefore, homeostatic mechanisms exist to “buffer against” changes in the acid / base balance. A buffer system is a collection of chemical reactions that can remove  $\text{H}^+$  and  $\text{OH}^-$  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  $\text{CO}_2$ , the reaction will not go to the left. Since  $\text{CO}_2$  can be eliminated by the lungs, the reaction will proceed to the right. Note that as the  $\text{CO}_2$  is eliminated through the lungs, excess  $\text{H}^+$  is now in the form of water, and has been removed out of “harm’s way”. We have restored proper pH.



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



### Base Production

If there is an increase in  $OH^-$ , 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.

