2.3.3

Buffers

Because it is absolutely essential that blood pH is maintained within the narrow range, the body has several mechanisms to regulate the H^+ concentration of the blood. One important defense employed by the body is the various **buffer systems**. Buffers are chemicals that tend to resist changes in pH. Note that buffers do not prevent changes; they resist changes. Let's see if we can figure out how this works.

A typical buffer system is composed of a weak acid and the conjugate base of that acid. Remember, weak acids are those that do not dissociate completely but reach an equilibrium between the reactants and the products of the reaction. An important buffer system in our blood is the bicarbonate buffer system. The components of this system are shown below.

$$H_2CO_3 \longleftrightarrow H^+ + HCO_3^-$$

Carbonic Acid   Hydrogen Ion   Bicarbonate Ion

In this case, the carbonic acid is the weak acid, and the bicarbonate ion is its conjugate base. The entire reaction is in equilibrium. If the equilibrium is disrupted by the addition of more hydrogen ions, the reaction will proceed to the left until equilibrium is restored. When it proceeds to the left, some of the excess hydrogen ions will combine with bicarbonate forming carbonic acid, hence removing some of the excess hydrogen ions from the solution. Essentially, the buffer has "soaked up" some of the extra hydrogen ions, thus preventing a large change in pH.

Another way of thinking of this system is to assume it behaves like a teeter-totter. If we have equal weights on each side, the teeter-totter is balanced (in equilibrium). If we add excess weight to one side (excess hydrogen ions), it will be out of balance. The only way to restore balance (equilibrium) is to move some of the excess weight to the opposite side until the teeter-totter is balanced again (equilibrium restored). Obviously, in this simple example, we realize that we cannot move all of the added weight to the opposite side because it would again be out of balance, but if some of the excess weight is moved to the other side, balance can be restored. Like the teeter-totter, when extra hydrogen ions are added, not all can be combined with bicarbonate, so there will still be a few more hydrogen ions than at the beginning (this is why buffers *resist* pH changes instead of *prevent* changes in pH). The pH will decrease, but not nearly as much as it would have if all added hydrogen ions were allowed to remain without being buffered. We could use the same analogy to see what happens when hydrogen ions are removed from the solution by the addition of a base. Since the equation is again out of equilibrium, the reaction will proceed to the right until some of the hydrogen ions have been replaced. Again, there will be a slight increase in pH, but not nearly as great as would happen in the absence of the buffer.

We will be discussing the bicarbonate buffer system in greater depth in upcoming modules. Here is an introduction to the full equation that shifts to maintain balance in our bodies:

$$CO_2 \quad + \quad H_2O \quad \leftrightarrow \quad H_2CO_3 \quad \leftrightarrow \quad HCO_3^- \quad + \quad H^+$$

Carbon Dioxide	Water	Carbonic acid (weak acid)	Bicarbonate ion (conjugate base)	Hydrogen ion