Role of the Respiratory System in Acid-Base Balance of Blood

As you have already learned, pulmonary ventilation is necessary for continuous oxygenation of the blood and removal of carbon dioxide (a waste product of cellular respiration) from the blood. Blood pH must be relatively constant for the cells of the body to function optimally. The carbonic acid-bicarbonate buffer system of the blood is extremely important because it helps stabilize arterial blood pH at 7.4 ± 0.02.

When carbon dioxide diffuses into the blood from the tissue cells, much of it enters the red blood cells, where it combines with water to form carbonic acid (Figure 39.15):

\[ H_2O + CO_2 \xrightarrow{\text{Carbonic anhydrase (RBC)}} H_2CO_3 \]

Some carbonic acid is also formed in the plasma, but the reaction is very slow because of the lack of the carbonic anhydrase enzyme. Shortly after it forms, carbonic acid dissociates to release bicarbonate (HCO_3^-) and hydrogen (H^+) ions. The hydrogen ions that remain in the cells are neutralized, or buffered, when they combine with hemoglobin molecules. If they were not neutralized, the intracellular pH would become very acidic as H^+ ions accumulated. The bicarbonate ions diffuse out of the red blood cells into the plasma, where they become part of the carbonic acid-bicarbonate buffer system. As HCO_3^- follows its concentration gradient into the plasma, an electrical imbalance develops in the RBCs that draws Cl^- into them from the plasma. This exchange phenomenon is called the chloride shift.

Acids (more precisely, H^+) released into the blood by the body cells tend to lower the pH of the blood by the body cells tend to lower the pH of the blood and to cause it to become acidic. On the other hand, basic substances that enter the blood tend to cause the blood to become more alkaline and the pH to rise. Both of these tendencies are resisted in large part by the carbonic acid-bicarbonate buffer system. If the H^+ concentration in the blood begins to increase, the H^+ ions combine with bicarbonate ions to form carbonic acid (a weak acid that does not tend to dissociate at physiological or acid pH) and are thus removed.

\[ H^+ + HCO_3^- \rightarrow H_2CO_3 \]

Likewise, as blood H^+ concentration drops below what is desirable and blood pH rises, H_2CO_3 dissociates to release bicarbonate ions and H^+ ions to the blood.

\[ H_2CO_3 \rightarrow H^+ + HCO_3^- \]

The released H^+ lowers the pH again. The bicarbonate ions, being weak bases, are poorly functional under alkaline conditions and have little effect on blood pH unless and until blood pH drops toward acid levels.

In the case of excessively slow or shallow breathing (hypoventilation) or fast deep breathing (hyperventilation), the amount of carbonic acid in the blood can be greatly modified – increasing dramatically during hypoventilation and decreasing substantially during hyperventilation. In either situation, if the buffering ability of the blood is inadequate, respiratory acidosis or alkalosis can result. Therefore, maintaining the normal rate and depth of breathing is important for proper control of blood pH.
Activity 6 - (Setup 1)

**Demonstrating the Reaction Between Carbon Dioxide (in Exhaled Air) and Water**

1. Fill a beaker with 100 ml of distilled water.
2. Add 0.5 ml of 0.05 M NaOH and two drops of phenol red. Swirl. Phenol red is a pH indicator that turns yellow in acidic solutions.
3. Blow through a straw into the solution.

What chemical reaction is taking place in the beaker?

Activity 7 - (Setup 2)

**Observing the Operation of Standard Buffers**

1. To observe the ability of a buffer system to stabilize the pH of a solution, obtain five 100-ml beakers and a wash bottle containing distilled water. Set up the following experimental samples:

   - **Beaker 1:** pH: _____
     (50 ml distilled water)
   - **Beaker 2:** pH: _____
     (50 ml distilled water and 1 drop concentrated HCl)
   - **Beaker 3:** pH: _____
     (50 ml distilled water and 1 drop concentrated NaOH)
   - **Beaker 4:** pH: _____
     (50 ml standard buffer solution [pH 7] and 1 drop concentrated HCl)
   - **Beaker 5:** pH: _____
     (50 ml standard buffer solution [pH 7] and 1 drop concentrated NaOH)

2. Using a pH meter standardized with a buffer solution of pH 7, determine the pH of the contents of each beaker and record above. After each and every pH recording, the pH probe should be thoroughly rinsed with a stream of distilled water from the wash bottle.
3. Add 3 more drops of concentrated HCl to beaker 4, stir, and record the pH: _____
4. Add 3 more drops of concentrated NaOH to beaker 5, stir, and record the pH: _____

How successful was the buffer solution in resisting pH changes when a strong acid (HCl) or a strong base (NaOH) was added?

Activity 8 - (Setup 3)

**Exploring the Operation of the Carbonic Acid-Bicarbonate Buffer System**

To observe the ability of the carbonic acid-bicarbonate buffer system of blood to resist pH changes, perform the following simple experiment.

1. Obtain two small beakers (50 ml), animal plasma, graduated cylinder, glass stirring rod, and a dropper bottle of 0.01 M HCl. Using the pH meter standardized with the buffer solution of pH 7.0, measure the pH of the animal plasma. Use only enough plasma to allow immersion of the electrodes and the measure the volume carefully.

   pH of the animal plasma: _______

2. Add 2 drops of the 0.01 M HCl solution to the plasma; stir and measure the pH again.

   pH of plasma plus 2 drops of HCl: _______

3. Rinse the electrode and then immerse them in a quantity of distilled water (pH 7) exactly equal to the amount of animal plasma used. Measure the pH of the distilled water.

   pH of distilled water: _______

4. Add 2 drops of 0.01 M HCl, swirl, and measure the pH again.

   pH of distilled water plus the two drops of HCl: _______

Is the plasma a good buffer? ______

What component of the plasma carbonic acid-bicarbonate buffer system was acting to counteract a change in pH when HCl was added?