Set 17: Buffers

1. Buffer solutions resist a change in pH even with the addition of substantial amounts of hydrogen or hydroxide ions. Many specific reactions that occur in biological systems occur only at specific pH values. Some reactions produce or use up hydrogen ions in these solutions. Buffers prevent large changes in the pH of solutions such as blood, cell contents and lymph system allowing vital reactions to continue.

2. When ammonia and ammonium chloride solutions are mixed the following equilibrium is established

$$H_2O(\ell) + NH_3(aq) \leftrightarrows NH_4^+(aq) + OH^-(aq)$$

(a) As more OH^- ions are added the $[OH^-]$ increases and causes the equilibrium to shift towards the reactants# (H_2O and NH_3). This results in the reduction in the $[OH^-]$ and uses up most, but not all of the added OH^- ions.

(b) As H^+ ions are added they react with OH^- ions according to the reaction represented by

$$H^+(aq) + OH^-(aq) \leftrightarrows H_2O(\ell)$$

This reduces the $[OH^-]$ which in turn causes the equilibrium to shift towards the products# ($NH_4^+ + OH^-$). This results in an increase in the $[OH^-]$ and replaces most, but not all of the OH^- ions that reacted with the H^+ ions.

The shift in equilibrium may be **predicted** by considering Le Chatelier's principle which states:

When a change is made to a chemical system in equilibrium a new equilibrium is established so that the change is partially counteracted.

OR

The shift in equilibrium may be **explained** in terms of changes in the relative reaction rates of the forward and reverse reactions.

For

(a) The increased $[OH^-]$ causes an increase in the rate of the reverse reaction with no immediate change in the rate of the forward reaction. This results in an increase in the concentrations of the reactants.

(b) The decreased $[OH^-]$ causes a decrease in the rate of the forward reaction with no immediate change in the rate of the reverse reaction. This results in an increase in the concentrations of the products.

3. Buffer solutions can be produced that have specific and known pH value. These values remain constant (at constant temperature) even with the inadvertent addition of small amounts of H^+ or OH^- ions. The fixed pH values are use as standards to calibrate the meters.

4. (a) Hypochlorite ions are added to pool water in the form of sodium hypochlorite solution or calcium hypochlorite solid to kill micro organisms. Some of the hypochlorite ions react with water to produce the weak acid hypochlorous acid hydroxide ions as illustrated in the equation:

$$OC\ell^-(aq) + H_2O(\ell) \leftrightarrows HOC\ell(aq) + OH^-(aq)$$

The buffer tends to use up some of these OH^- ions stopping the pool water becoming alkaline too quickly.

(b) $OH^-(aq) + HCO_3^-(aq) \leftrightarrows H_2O(\ell) + CO_3^{2-}(aq)$

And $HCO_3^-(aq) \leftrightarrows CO_3^{2-}(aq) + H^+(aq)$ where the H^+ reacts with the OH^-

Both processes use up OH^- ions.

5. (a) H_2O and HCO_3^- ion

(b) $CO_2(aq) + H_2O(\ell) \leftrightarrows HCO_3^-(aq) + H^+(aq)$

H^+ ion is used up by reacting with the HCO_3^- as indicated in the reverse reaction above and OH^- ions is used up by reacting with CO_2 directly

$$OH^-(aq) + CO_2^-(aq) \leftrightarrows HCO_3^-(aq)$$

or

with H^+ ions which results in more H^+ ions being produced.

$$H^+(aq) + OH^-(aq) \leftrightarrows H_2O(\ell)$$

6. Method 1 Add an equal volume of 1 mol L^{-1} sodium ethanoate solution

Method 2 Add an equal volume of a 0.5 mol L^{-1} solution of sodium hydroxide. This would react with half the ethanoic acid to produce ethanoate ion as in the equation

$$CH_3COOH(aq) + OH^-(aq) \leftrightarrows CH_3COO^-(aq) + H_2O(\ell)$$

7. Method 1 Make a solution in water of the sodium citrate mixed with an equal number of moles of citric acid.

Method 2 Make a solution in water of the sodium citrate then add 1½ times that number of moles of hydrochloric acid. This would react with half the ethanoate ion to produce ethanoic acid as in the equation

$$C_6H_5O_7^{3-}(aq) + 3H^+(aq) \leftrightarrows H_3C_6H_5O_7(aq)$$

8 A change in temperature changes the equilibrium concentration of the species in equilibrium and as one of the species is always either H^+ or OH^- ions, changing their concentrations changes the pH.

For example : in an ethanoic acid – ethanoate ion buffer the equilibrium that exists can be represented by the equation:

$$CH_3COOH(\ell) \leftrightarrows H^+(aq) + CH_3COO^-(aq) \qquad \Delta H \text{ is positive}$$

An increase in temperature causes an increase in the H^+ and so a decrease in pH.