Chem12 Buffers : Notes-200

Buffer solutions are equilibrium solutions that maintain a relatively constant pH, when small amounts of an acid or a base are added.

A **Buffer** can be prepared from a mixture of a weak acid and significant amounts of its conjugate base, or a weak base and significant amounts of its conjugate acid.

We can prepare a buffer by adding formic acid, HCOOH, to a solution of the salt, sodium formate, HCOONa. Usually, equal amounts of the acid and its conjugate base are used. The equilibrium is:

$$HCOOH + H_2O \quad H_3O^+ + HCOO^-$$

The equilibrium expression is : $[H_3O^+][HCOO^-]/[HCOOH] = 1.78 \times 10^{-4}$.

If $[HCOO^-] = [HCOOH]$, then $[H_3O^+] = 1.78 \times 10^{-4}$, and the pH = 3.75. This buffer is an example of an **acidic buffer** .

A **basic buffer** can be prepared from an NH_3/NH_4^+ solution.

Problems :

1) Find the change in pH when 10.0 mL of 1.00 M HCl is added to 1.00 L of pure water.

2) Suppose we have 1.00 L of the formic acid/formate ion buffer with 0.500 M concentrations of each species. The pH is 3.75. Find the change in pH by adding 10.0 mL Of 1.00 M HCl.

Solutions :

1) Pure water has a pH of 7.00. If 10.0 mL of 1.00 M HCl is added to the water, the $[H_3O^+]$ will become 0.0100 M. The new pH will then be 2.00. The change in pH will be -5.00.

2) The equilibrium is:

$$HCOOH + H_2O \quad H_3O^+ + HCOO^-$$

Since 0.0100 moles of H_3O^+ are added, the equilibrium shifts left and the 0.0100 moles of H_3O^+ are "absorbed" by the HCOO- ions, and 0.0100 moles of HCOOH are "created".

The new concentrations of formic acid and formate ion are; [0.510] and [0.490] respectively. Therefore, since:

$$K_a = \frac{[H_3O^+][HCOO^-]}{[HCOOH]} = 1.78 \times 10^{-4}$$

we have,

$$[H_3O^+] = \frac{[HCOOH]}{[HCOO^-]} \cdot 1.78 \times 10^{-4}$$

and,

$$[H_3O^+] = \frac{[0.510]}{[0.490]} \cdot 1.78 \times 10^{-4}$$

therefore, $[H_3O^+] = 1.85 \times 10^{-4}$. The new pH = 3.73.

So the pH changed from 3.75 to 3.73. Compare with problem 1.