Ionic Equilibria II

Paper-III, Bsc II, General Chemistry

Dr. Md Asif Amin, *Balurghat Mahila Mahavidyalaya*

pH Scale

Introduction and Definitions:

The pH scale measures how acidic or basic a substance is. The pH scale ranges from 0 to 14 at 298 K. A pH of 7 is neutral. A pH less than 7 is acidic. A pH greater than 7 is basic.

The pH scale is logarithmic and as a result, each whole pH value below 7 is ten times more acidic than the next higher value. For example, pH 4 is ten times more acidic than pH 5 and 100 times (10 times 10) more acidic than pH 6. The same holds true for pH values above 7, each of which is ten times more alkaline (another way to say basic) than the next lower whole value. For example, pH 10 is ten times more alkaline than pH 9 and 100 times (10 times 10) more alkaline than pH 8.

Pure water is neutral. But when chemicals are mixed with water, the mixture can become either acidic or basic. Examples of acidic substances are vinegar and lemon juice. Lye, milk of magnesia, and ammonia are examples of basic substances.

Definition of pH, pOH, and pKw:

The concentrations of hydrogen ions and indirectly hydroxide ions are given by a pH number. **pH is defined as the negative logarithm of the hydrogen ion concentration.** The equation is:

$pH=-\log[H^+]$
similarly,$pOH=-\log[OH^-]$
and $pKw = -\log[Kw]$.

Following equation 4 concentration of H^+ is αC, then $pH= \frac{1}{2} pK_a - \frac{1}{2} \log C$ 8
As K depends on temperature pH is also depends on temperature.

Buffer Solution On addition of small amount acid or base the pH of which solution remain unchanged.

Henderson equation

The **Henderson** Hasselbalch **equation** is an approximate **equation** that shows the relationship between the pH or pOH of a solution and the pK_a or pK_b and the ratio of the concentrations of the dissociated **chemical** species. In order to use the **equation**, the acid dissociation constant must be known. Taking negative logarithm of equation 6 we get

$$pK_a = pH - \log\frac{[A^-]}{[HA]} \text{ or } pH = pK_a + \log\frac{[A^-]}{[HA]} \qquad 9$$

$$pH = pK_a + \log\frac{[Salt]}{[Acid/Base]}$$

Buffer Capacity Gram equivalent of strong acid or base needed to change the pH of a buffer solution by one unit.