Chem12 Acid-Base Reactions-40

Pure water is often considered a non-electrolyte. It is actually a weak electrolyte. A few water molecules do break apart to form ions. This is called self-ionization. The equations below represent this process.

$$H_2O(l) <-> H^+(aq) + OH^-(aq)$$

$$2H_2O(l) <-> H_3O^+(aq) + OH^-(aq) \qquad \Delta H = 57 \text{ kJ}$$

The second equation is considered more accurate as water is now known to be a proton acceptor as well as a proton donor. This is an equilibrium endothermic reaction. This reaction occurs in all aqueous solutions. In pure water $[H_3O^+] = [OH^-]$. If an acid is added to the water the equilibrium shifts left and $[OH^-]$ decreases. If a base is added the equilibrium shifts left and $[H_3O^+]$ decreases. The equilibrium expression for the reaction is given below.

$$K_{eq} = K_w = [H_3O^+][OH^-] = 1.00\text{x}10^{-14} \text{ at } 25°C$$

This relationship is true for all aqueous solutions.

Exercise 1: Find $[H_3O^+]$ and $[OH^-]$ in a 1.00 L beaker of pure water.

Exercise 2 : What is the value of $[H_3O^+]$ and $[OH^-]$ in

a) 0.010M HBr

b) 5.0M NaOH

Exercise 3 : Calculate the $[H_3O^+]$ and $[OH^-]$ in a $2.0\text{x}10^{-3}$ M solution of $HClO_4$.

Exercise 4 : If 1.00g of KOH is dissolved in 2.00L of water, find the

$[H_3O^+]$ and $[OH^-]$ for this solution.

Exercise 5 : If 0.100 mol of HCl are dissolved in 1.5L of water, find $[H_3O^+]$ and $[OH^-]$.

Answers : 1) $1.00x10^{-7}$, $1.00x10^{-7}$, 2)a) 0.010, $1.00x10^{-12}$, b) $2.00x10^{-15}$, 5.0, 3) $2.0x10^{-3}$, $5.0x10^{-12}$, 4) $1.12x10^{-12}$, $8.91x10^{-3}$, 5) 0.0667 M, $1.5x10^{-13}$ M.