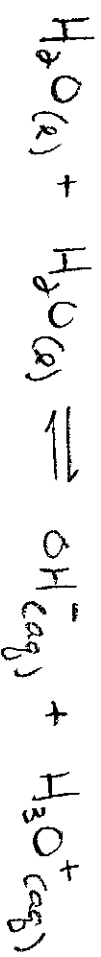


16.3 The Autoionization of Water



→ extremely rapid reaction

→ 2 out of 10^9 water molecules ionize at any given moment

Ion Product of Water

Equilibrium expression

$$K_c = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w = 1.0 \times 10^{-14} \text{ (@ } 25^\circ\text{C)}$$

because H^+ is the same as H_3O^+

$$\text{so } K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

When $[\text{H}^+] = [\text{OH}^-]$ in a solution

↳ the solution is neutral

What is $[\text{H}^+]$ and $[\text{OH}^-]$ in a neutral solution @ 25°C ?

$$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$x \cdot x = 1.0 \times 10^{-14}$$

$$x^2 = 1.0 \times 10^{-14}$$

$$x = 1.0 \times 10^{-7}$$

If $[\text{H}^+] > [\text{OH}^-]$ the solution is acidic

If $[\text{H}^+] < [\text{OH}^-]$ the solution is basic (alkaline)

If $[\text{H}^+] = [\text{OH}^-]$ the solution is neutral

What is the $[H^+]$ in a solution in which

$$[OH^-] = 0.010 \text{ M.}$$

$$[H^+][OH^-] = 1.0 \times 10^{-14}$$

$$[H^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{.010} = 1.0 \times 10^{-12} \text{ M}$$

$[H^+] < [OH^-]$ basic

16.4 pH Scale

$$pH = -\log [H^+]$$

What is the pH of a neutral solution

$$pH = -\log 1.0 \times 10^{-7} = 7.00$$

$$pOH = -\log [OH^-]$$

$$pH + pOH = 14.00$$

pH scale

< 7 acidic

7 neutral

> 7 basic

If pOH is 10.24, calculate $[H^+]$

$$pH = 14.00 - 10.24 = 3.76$$

$$pH = -\log [H^+]$$

$$[H^+] = 10^{-pH} = 10^{-3.76} = 1.7 \times 10^{-4} \text{ M}$$