16.3 The Autoionization of Water

$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + H_3O^+_{(aq)}$$

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

Ion Product of Water

Equilibrium expression

$$K_c = [H_3O^+][OH^-] = K_w = 1.0 \times 10^{-14} \text{ (@ 25°C)}$$

because H^+ is the same as H_3O^+

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

when $[H^+] = [OH^-]$ in a solution
↳ the solution is neutral

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

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$
$$X \cdot X = 1.0 \times 10^{-14}$$
$$X^2 = 1.0 \times 10^{-14}$$
$$X = 1.0 \times 10^{-7}$$

If $[H^+] > [OH^-]$ the solution is acidic
If $[H^+] < [OH^-]$ the solution is basic (alkaline)
If $[H^+] = [OH^-]$ the solution is neutral

What is the $[H^+]$ in a solution in which $[OH^-] = 0.010$ M.

$$[H^+][OH^-] = 1.0\times10^{-14}$$

$$[H^+] = \frac{1.0\times10^{-14}}{[OH^-]} = \frac{1.0\times10^{-14}}{.010} = 1.0\times10^{-12}\text{ M}$$

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

16.4 pH Scale

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

What is the pH of a neutral solution

$$pH = -\log 1.0\times10^{-7} = 7.00$$

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

$$pH + pOH = 14.00$$

pH scale

- < 7 acidic
- 7 neutral
- > 7 basic

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

$$pH = 14.00 - 10.24 = 3.76$$

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

$$[H^+] = 10^{-pH} = 10^{-3.76} = 1.7\times10^{-4}\text{ M}$$