Revised Arrhenius Theory

Acids → H^+

Bases → OH^- (NH_3, Na_2CO_3)

Bronsted-Lowry Theory

Acids → H^+ donors

Bases → H^+ acceptors

Water Equilibrium

Conductivity is due to the presence of ions. For water:

$$H_2O_{(l)} \Longleftrightarrow H^+_{(aq)} + OH^-_{(aq)}$$

- therefore $K = \dfrac{[H^+][OH^-]}{[H_2O]}$ is very small

$[H_2O] \leftarrow$ constant

- slight conductivity shows that equilibrium greatly favors water molecules (less than 2 H^+ per billion water)

- therefore the concentration of water in pure water and in dilute aqueous solutions is essentially constant and can be combined with the equilibrium constant to produce a new constant called the *Ion Product Constant*

Ionization Constant for water (ion product constant)

$\boxed{\mathbf{K_w = [H^+]\,[OH^-]}}$ **= 1.0 x 10^{-14}** at SATP

Since $[H^+]$ and $[OH^-]$ are found in 1:1 ratio
($H_2O_{(l)} \Longleftrightarrow H^+_{(aq)} + OH^-_{(aq)}$)

$[H^+_{(aq)}] = [OH^-_{(aq)}] = 1.0 \times 10^{-7}$ mol/L in **neutral** solutions.

+ Acid

Arrhenius's Theory - acid is a substance that ionizes water to produce H^+ ions.
- additional ions produced by the acid increases the H^+ concentration in the water. (more acid, more H^+)

Therefore acids always have a $[H^+] > 10^{-7}$ mol/L

Basic solutions produce a $[OH^-]$ greater than 10^{-7} mol/L

K_w can be used to calculate either $[H^+]$ or $[OH^-]$

since $\mathbf{K_w = [H^+]\,[OH^-]}$ then $[H^+] = K_w / [OH^-]$

and $[OH^-] = K_w / [H^+]$

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

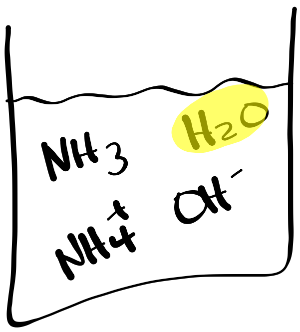
$$K = \frac{[H^+][OH^-]}{[H_2O_{(l)}]}$$

$$\underbrace{K[H_2O_{(l)}]} = [H^+][OH^-]$$

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

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_{4(aq)}^{+} + OH_{(aq)}^{-}$$

H^+



pH and pOH

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

$pH = -\log[H^+_{(aq)}]$ $pOH = -\log[OH^-_{(aq)}]$

$[H^+_{(aq)}] = 10^{-pH}$ $[OH^-_{(aq)}] = 10^{-pOH}$

$\boxed{pH + pOH = 14.00}$

Ex. Calculate the pH of a solution where $[H^+_{(aq)}] = 3.24 \times 10^{-4}$M.

$[H^+_{(aq)}] = 3.24 \times 10^{-4}$ M

$[OH^-] = ?$

$pH = ?$

$pOH = ?$

$pH = -\log[H^+_{(aq)}]$

$pH = -\log[3.24 \times 10^{-4}]$

EXP

EE

$\boxed{pH = 3.489}$

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

$[OH^-] = \frac{1.0 \times 10^{-14}}{3.24 \times 10^{-4}}$

$\boxed{[OH^-] = 3.09 \times 10^{-11} M}$

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

$pOH = -\log[3.09 \times 10^{-11}]$

$\boxed{pOH = 10.511}$

Strong Acids

Calculate the concentration of the hydroxide ions, pH and pOH of a 0.15 mol/L solution of hydrochloric acid at 25°C.

Strong acids will always completely ionize

$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$

0.15mol/L   0.15mol/L   0.15mol/L

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

$pH = -\log[0.15]$

$\boxed{pH = 0.82}$