Warm Up

Predict the products for the following reaction, and identify each reactant as an acid or a base.

$$\underset{A}{C_6H_5COOH_{(aq)}} + \underset{B}{OH^-_{(aq)}} \rightleftharpoons \underset{B}{C_6H_5COO^-_{(aq)}} + \underset{A}{H_2O_{(l)}}$$

$$\underset{B}{HCO_3^-{}_{(aq)}} + \underset{A}{H_3O^+_{(aq)}} \rightleftharpoons \underset{A}{H_2CO_{3(aq)}} + \underset{B}{H_2O_{(l)}}$$

Conjugate acid-base pairs

two substances that differ by a proton

Ex. H_2O and H_3O^+

NH_4^+ and NH_3

Revised Arrhenius Theory

ACID - H^+ ions

BASE - OH^- ions

NH_3

Na_2CO_3

Bronsted-Lowry Theory

ACID - H^+ (proton) donors

BASE H^+ (proton) acceptors

Water Equilibrium

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

$$H_2O_{(l)} <==> H^+_{(aq)} + OH^-_{(aq)}$$

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

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

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

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

Arrhenius's Theory - acid is a substance that ionizes water t 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 $[OH^-] = K_w / [H^+]$

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

$$K = \frac{[H^+][OH^-]}{[H_2O]}$$

$$K[H_2O] = [H^+][OH^-]$$

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

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

$$[H^+]^2 = 1.0 \times 10^{-14}$$

$$[H^+] = 1.0 \times 10^{-7}$$

pH and pOH

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

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

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

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

$pH + pOH = 14.00$

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

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

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

$pH = 3.489$

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

$[OH^-_{(aq)}] = \dfrac{1.0 \times 10^{-14}}{3.24 \times 10^{-4}}$

$[OH^-_{(aq)}] = 3.09 \times 10^{-11}$ M

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

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

$pOH = 10.510$

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

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

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

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

$pH = -\log[0.15]$

$pH = 0.82$

$pH + pOH = 14.00$

$pOH = 14.00 - 0.82$

$pOH = 13.18$

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

$[OH^-_{(aq)}] = 10^{-13.18}$

$[OH^-_{(aq)}] = 6.6\times10^{-14} M$