SCH 4U

EQUILIBRIUM of ACIDS

The "Reaction" of Pure water: $2H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + OH^-{}_{(aq)}$

- 2 out of every billion water molecules will dissociate into ions -- 2 ppb

At 25°C, $[\boldsymbol{H_3O^+}] = \mathbf{1.0 \times 10^{-7}}\, \boldsymbol{mol/L} = [\boldsymbol{OH^-}]$

- The concentration changes with changes in temperature.

$$K_c = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

$$\begin{aligned} K_c[H_2O]^2 &= [H_3O^+][OH^-] \\ &= (1.0 \times 10^{-7}\, mol/L)(1.0 \times 10^{-7}\, mol/L) \end{aligned}$$

$\boldsymbol{K_w} = \mathbf{1.0 \times 10^{-14}}$ **Ion-Product Constant for water**

Since $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$, as $[H_3O^+]$ increases, the $[OH^-]$ decreases, AND the equilibrium of pure water shifts to the left.

RECALL:

Since STRONG ACIDS totally dissociate $[H_3O^+] = [acid]$

EG. Given [HCl] = 0.25 mol/L

Since HCl is a strong acid. $HCl_{(aq)} + H_2O \rightarrow H_3O^+{}_{(aq)} + Cl^-{}_{(aq)}$

Therefore, $[H_3O^+] = 0.25\, mol/L$

EG. What is the $[OH^-]$ of a 0.25 mol/L solution of calcium hydroxide?

$Ca(OH)_2 \longrightarrow Ca^{+2} + 2\,OH^-$

$[Ca(OH)_2]$ = 0.25 mol/L

$[OH^-]$ = 2(0.25) = 0.50 mol/L

STRONG ACIDS & BASES: Conversions

- **Equilibrium does NOT occur with strong acids/bases**

$$\begin{tikzcd}[column sep=huge, row sep=huge]
{[\mathrm{H_3O^+}]} \arrow[r, leftrightarrow, "{[\mathrm{H_3O^+}][\mathrm{OH^-}] = 10^{-14}}"'] \arrow[d, leftrightarrow, "{\mathrm{pH} = -\log[\mathrm{H_3O^+}] \quad [\mathrm{H_3O^+}] = 10^{-\mathrm{pH}}}"'] & {[\mathrm{OH^-}]} \arrow[d, leftrightarrow, "{\mathrm{pOH} = -\log[\mathrm{OH^-}] \quad [\mathrm{OH^-}] = 10^{-\mathrm{pOH}}}"] \\
\mathrm{pH} \arrow[r, leftrightarrow, "{\mathrm{pH} + \mathrm{pOH} = 14}"'] & \mathrm{pOH}
\end{tikzcd}$$

EG. What is the pH of a 4.00×10^{-2} mol/L barium hydroxide solution?
[ANS = 12.9]

EG. A sodium hydroxide solution has a pH of 11.4. What is the concentration of NaOH?
[ANS = 2.51×10^{-3} mol/L]

EG. Phenol, C_6H_5OH, is used as a disinfectant. An aqueous solution of phenol was found to have a pH of 4.72. Is phenol acidic, neutral, or basic? Calculate $[H_3O^+]$, $[OH^-]$, and pOH of the solution.

WEAK ACIDS – form an equilibrium

Let HA represent a weak monoprotic acid.

Dissociation Equilibrium Equation $HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$

$$K_c = \frac{[H_3O^+][A^-]}{[HA][H_2O]}$$

ACID DISSOCIATION CONSTANT: $$K_a = K_c[H_2O] = \frac{[H_3O^+][A^-]}{[HA]}$$

① GIVEN pH & [WEAK ACID], DETERMINE K_a & % DISSOCIATION

1. Prepare an ICE table. Use the dissociation equilibrium equation for acid in water.

2. Determine $[H_3O^+]$ at equilibrium using pH equation.

3. Set up K_a expression using ICE table.

4. Substitute for x and solve for K_a.

5. Calculate $\boxed{\%\ \text{dissociation} = \frac{[\boldsymbol{H_3O^+}]_{aq}}{[\boldsymbol{acid}]_{initial}} \times \mathbf{100}\%}$

EXAMPLES:

A) Aspirin is a weak monoprotic acid (H-Asp). A 0.100 mol/L solution has a pH of 2.24. Determine the K_a and % dissociation.

B) Propanoic acid, CH_3CH_2COOH, is a weak monoprotic acid that is used to inhibit mould formation in bread. A student prepared a 0.10 mol/L solution of propanoic acid and found that the pH was 2.96. What is the acid dissociation constant for propanoic acid? What percent of its molecules were dissociated in the solution?

② GIVEN K_a and [WEAK ACID], DETERMINE pH and % DISSOCIATION

1. Set up an ICE table and substitute.
2. Set up K_a expression using ICE table.
3. Substitute for K_a and check for approximation method.
4. Solve for x.
5. Solve for pH using $x = [H_3O^+]$
6. Calculate % dissociation = $\frac{[H_3O^+]_{aq}}{[acid]_{initial}} \times 100\%$

EXAMPLES:

A) What is the pH of a 0.045 mol/L solution of the weak chlorous acid? Use table E.9

B) What is the pH of a 0.100 mol/L vitamin C solution having an acid ionization constant of 8.0×10^{-5}? What is the % dissociation?

C) Hydrosulfuric acid, $H_2S(aq)$, is a weak diprotic acid that is often used to precipitate metal sulfides, which tend to be very insoluble.
Calculate the pH and $[HS^-(aq)]$ of a 7.5×10^{-3} mol/L solution.

D) Oxalic acid, HOOCCOOH, is a weak diprotic acid that occurs naturally in some foods, including rhubarb. Calculate the pH of a solution of oxalic acid that is prepared by dissolving 2.5 g in 1.0 L of water. What is the concentration of hydrogen oxalate, $HOOCCOO^-$, in the solution?