

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, $[H_3O^+] = 1.0 \times 10^{-7} \text{ mol/L} = [OH^-]$

- The concentration changes with changes in temperature.

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

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

$$K_w = 1.0 \times 10^{-14} \quad \text{Ion-Product Constant for water}$$

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:

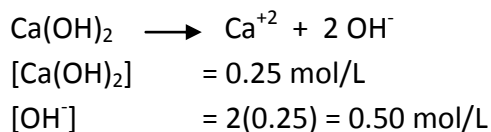
Since STRONG ACIDS totally dissociate $[H_3O^+] = [acid]$

EG. Given $[HCl] = 0.25 \text{ mol/L}$

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

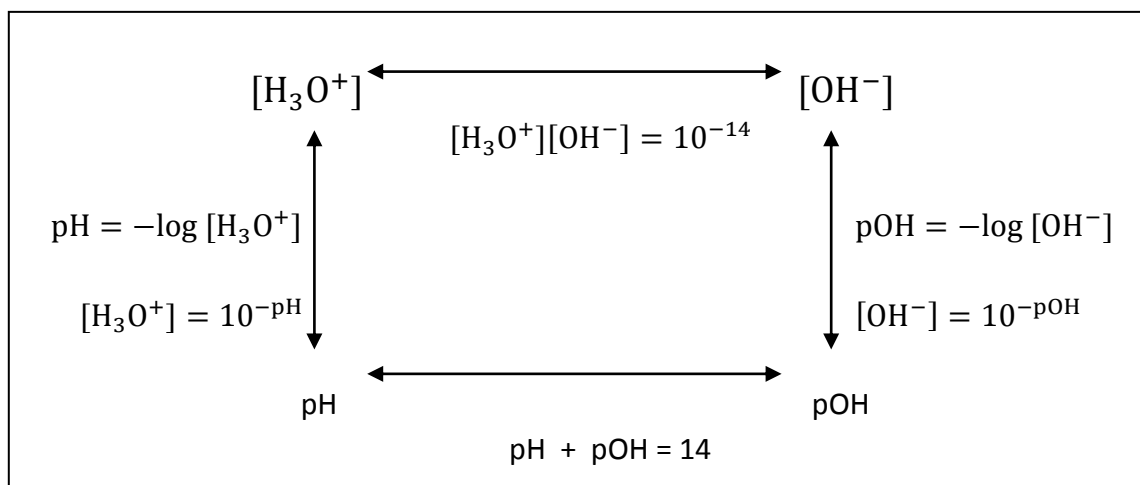
Therefore, $[H_3O^+] = 0.25 \text{ mol/L}$

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



STRONG ACIDS & BASES: Conversions

- Equilibrium does NOT occur with strong acids/bases



- 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.



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

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

$$\% \text{ dissociation} = \frac{[H_3O^+]_{aq}}{[acid]_{initial}} \times 100\%$$

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, $HOOC-COOH$, 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, $HOOC-COO^-$, in the solution?