

## 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:

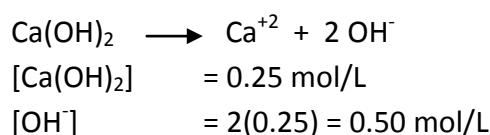
Since STRONG ACIDS totally dissociate  $[H_3O^+] = [acid]$

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



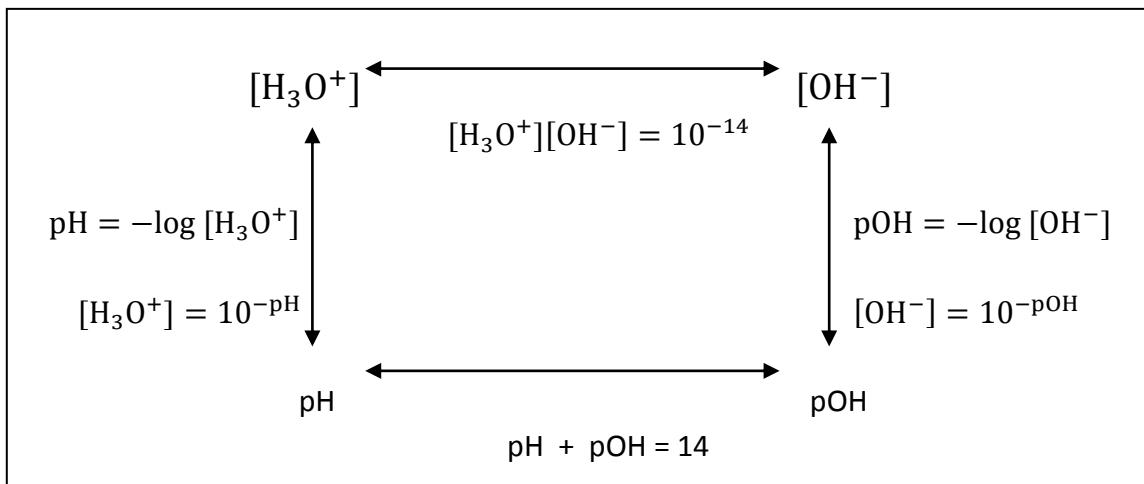
Therefore,  $[H_3O^+] = 0.25 \text{ mol/L}$

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  $4.00 \times 10^{-2}$  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 =  $2.51 \times 10^{-3}$  mol/L]

EG. Phenol,  $\text{C}_6\text{H}_5\text{OH}$ , 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  $[\text{H}_3\text{O}^+]$ ,  $[\text{OH}^-]$ , and pOH of the solution.

## WEAK ACIDS – form an equilibrium

Let HA represent a weak monoprotic acid.



$$K_c = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}][\text{H}_2\text{O}]}$$

ACID DISSOCIATION CONSTANT:

$$K_a = K_c [\text{H}_2\text{O}] = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

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  $8.0 \times 10^{-5}$ ? 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  $7.5 \times 10^{-3}$  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?