

## Acids and Bases



## Acids and Bases

- **Acid:** is a substance that can donate protons (hydrogen ions).
- **Base:** is a substance that can accept protons.

## Ionization of Strong Acids and Bases

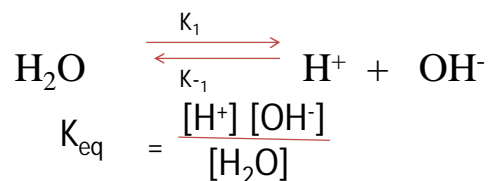
- A strong acid is a substance that ionizes 100% in aqueous solutions.



- A strong base is a substance that ionizes totally in solution to produce  $\text{OH}^-$  ions.



## Ionization of Water



- Water is amphoteric it can accept and donate protons.
- In pure water for every mole of  $[\text{H}^+]$ , a 1 mole of  $[\text{OH}^-]$  is produced, ie.  $[\text{H}^+] = [\text{OH}^-]$
- The pH of water = 7
- Then:  $[\text{H}^+] = [\text{OH}^-] = 10^{-7} \text{ M}$

## Ionization of Water Continue

- Thus the molarity of water:

$$M = \frac{\text{no. of moles}}{\text{volume of solution in L}}$$

- In 1 liter of water = 1000g of water
- Mw of H<sub>2</sub>O = (2\*1) + (1\*16) = 18g/mole.
- No. of moles = 1000 / 18 = 55.6 moles.
- M = 55.6 / 1 = 55.6 molar.
- Since part of water molecules is ionized, the actual conc. of the water is = 55.6 - 10<sup>-7</sup>.

## Ionization of Water continue

The 10<sup>-7</sup> is very small it can be neglected

$$K_{eq} = \frac{[H^+][OH^-]}{55.6}$$

- Since the concentration of the water is constant thus K<sub>eq</sub> of water can be written as follows:

$$K_{eq} = [H^+][OH^-]$$

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

$$K_w = 10^{-7} \times 10^{-7}$$

$$K_w = 10^{-14}$$

$$pK_w = -\log 10^{-14}$$

$$pK_w = 14$$

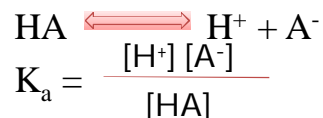
## Ionization of Weak Acids

- Weak acids have a weak affinity towards their proton



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]}$$

- The concentration of water is not considered since it is a constant. Thus



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

## Ionization of Weak Acids Continue

- Since weak acids ionize partially only thus, their  $K_a$  value will always be less than one because the concentration of  $[\text{HA}]$  is always higher than the concentration of both  $[\text{H}^+]$  and  $[\text{A}^-]$ .
- Between weak acids the higher the  $K_a$  the stronger the acid.

## Ionization of Weak Bases

- Weak bases have a weak affinity towards their proton.



$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_4\text{OH}]}$$

## pH of Solutions of Weak Acids

- The dissociation of a weak monoprotic acid, HA, yields,  $\text{H}^+$  and  $\text{A}^-$  in equal concentration.
- If  $K_a$  and the initial concentration of HA are known,  $\text{H}^+$  can be calculated easily:

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{[\text{H}^+]^2}{[\text{HA}]}$$

$$[\text{H}^+]^2 = K_a [\text{HA}]$$

$$[\text{H}^+] = \sqrt{K_a [\text{HA}]}$$

$$\text{Log}[\text{H}^+] = \frac{1}{2} \text{Log } K_a [\text{HA}]$$

## pH of Solutions of Weak Acids Continue

Multiply by -1

$$-\text{Log}[\text{H}^+] = \frac{1}{2} (-\text{Log } K_a - \text{Log } [\text{HA}])$$

$$\text{pH} = \frac{1}{2} (\text{p}K_a + \text{p} [\text{HA}])$$

- A similar relationship can be derived for weak bases:

$$[\text{OH}^-] = \sqrt{K_b [\text{A}^-]}$$

$$\text{pOH} = \frac{1}{2} (\text{p}K_b + \text{p} [\text{A}^-])$$

## Relationship Between $\text{p}K_a$ and $\text{p}K_b$ for Weak Acids and Bases

- $K_w = K_a * K_b$
- $\text{p}K_w = \text{p}K_a + \text{p}K_b$
- $\text{p}K_w = \text{pH} + \text{pOH}$

## ➤ Example

- A weak acid HA, is 0.1% ionized (dissociated) in a 0.2 M solution.
- a) What is the equilibrium constant of the acid  $K_a$ ?
- b) What is the pH of the solution?
- c) How many ml of 0.1 N KOH would be required to neutralize completely 500 ml of 0.2 M HA solution?

## Example Continue



The dissociation fraction=  $(0.1/100) \times 0.2$   
 $= 2 \times 10^{-4} \text{ M}$

Equilibrium:  $0.2 - 2 \times 10^{-4} \text{ M}$      $2 \times 10^{-4} \text{ M}$      $2 \times 10^{-4} \text{ M}$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$K_a = ((2 \times 10^{-4}) \times (2 \times 10^{-4})) / 0.2 - 2 \times 10^{-4}$$

When the amount of HA that has dissociated is small, 10% or less the  $K_a$  is simplified by ignoring the subtraction from [HA]

## Example Continue

$$K_a = ((2 \times 10^{-4}) \times (2 \times 10^{-4})) / 0.2$$

$$K_a = 4 \times 10^{-8} / 2 \times 10^{-1}$$

$$K_a = 2 \times 10^{-7}$$

B)  $\text{pH} = -\text{Log} [\text{H}^+]$

$$\text{pH} = -\text{Log } 2 \times 10^{-4}$$

$$\text{pH} = 3.7$$

C) No. of moles of  $\text{OH}^-$  required = no. of moles of  $\text{H}^+$  present

$$L_{\text{acid}} \times N_{\text{acid}} = L_{\text{base}} \times N_{\text{base}}$$

$$N = M$$

$$0.5 \times 0.2 = L_{\text{base}} \times 0.1$$

$$L_{\text{base}} = 0.1 / 0.1 = 1 \text{ liter}$$