Acids and Bases

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- 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.

$$HCl + H_2O \Rightarrow H_3O^+ + Cl^-$$

- A strong base is a substance that ionizes totally in solution to produce OH^- ions.

$$KOH \Rightarrow K^+ + OH^-$$

Ionization of Water

$$H_2O \underset{K_{-1}}{\overset{K_1}{\rightleftharpoons}} H^+ + OH^-$$

$$K_{eq} = \frac{[H^+]\,[OH^-]}{[H_2O]}$$

- Water is amphoteric it can accept and donate protons.
- In pure water for every mole of $[H^+]$, a 1 mole of $[OH^-]$ is produced, ie. $[H^+] = [OH^-]$
- The pH of water = 7
- Then: $[H^+] = [OH^-] = 10^{-7}$ 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_2O = (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 = $55.6 - 10^{-7}$.

Ionization of Water continue

The 10^{-7} is very small it can be neglected

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

- Since the concentration of the water is constant thus K_{eq} 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

$$CH_3COOH + H_2O \Longleftrightarrow CH_3COO^- + H_3O^+$$

$$K_a = \frac{[H_3O^+]\,[CH_3COO^-]}{[CH_3COOH]}$$

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

$$HA \Longleftrightarrow H^+ + A^-$$

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

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 [HA] is always higher than the concentration of both $[H^+]$ and $[A^-]$.
- 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.

$$NH_4OH \Leftrightarrow NH_4^+ + OH^-$$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]}$$

pH of Solutions of Weak Acids

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

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

$$[H^+]^2 = K_a [HA]$$

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

$$\text{Log}[H^+] = ½ \text{ Log } K_a [HA]$$

pH of Solutions of Weak Acids Continue

Multiply by -1

$- \text{Log}[H^+] = ½ (-\text{Log } K_a - \text{Log } [HA])$

$pH = ½ (pK_a + p [HA])$

- A similar relationship can be derived for weak bases:

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

$pOH = ½ (pK_b + p [A^-])$

Relationship Between pK_a and pK_b for Weak Acids and Bases

- $K_w = K_a * K_b$
- $pK_w = pk_a + pK_b$
- $pK_w = pH + pOH$

➢ 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

• A)	HA	$\Leftrightarrow$	H^+ +	A^-
Start:	0.2 M		0	0

The dissociation fraction= (0.1/100) * 0.2
= 2×10^{-4} M

Equilibrium: $0.2-2\times10^{-4}$ M 2×10^{-4} M 2×10^{-4} M

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

$K_a = ((2\times10^{-4}) \times (2\times10^{-4})) / 0.2-2\times10^{-4}$

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

$K_a = ((2\times10^{-4}) \times (2\times10^{-4})) / 0.2$

$K_a = 4\times10^{-8} / 2\times10^{-1}$

$K_a = 2\times10^{-7}$

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

$pH = -\text{Log}\ 2\times10^{-4}$

$pH = 3.7$

C) No. of moles of OH^- required = no. of moles of H^+ present

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

$$N = M$$

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

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