Neutralization of Strong Acids and Bases

➢ Example

- 1- How many ml of 0.025 M H_2SO_4 are required to neutralize exactly 525 ml of 0.06 M KOH?
- 2- What is the pH of the neutralized solution?
- No. of moles (equivalents) of H^+ required = no. of moles (equivalents) of OH^- present

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

- For H_2SO_4 = 0.025 M, N = M * n = 0.025 * 2 = 0.05 N
- For KOH since n =1 then, N = M = 0.06

$L_{acid} \times 0.05 = 0.525 \times 0.06$

$L_{acid} = (0.525 \times 0.06) / 0.05$

$L_{acid} = 0.63$

- The volume of the acid required = 630 ml.
- The neutralized solution contains only K_2SO_4 “a salt” of a strong acid and strong base has no effect on pH thus pH =7

➢ Example

- How many ml of 0.05 N HCl are required to neutralize exactly 8g of NaOH?

At the equivalent point:

The no. of moles H^+ added = no. of moles OH^- present

$$L_{acid} \times N_{acid} = \text{no. of (moles) equivalents of } H^+ \text{ added}$$

$$wt_{NaOH} / Mwt_{NaOH} = \text{no. of moles of NaOH } (OH^-) \text{ present}$$

$$L_{acid} \times N_{acid} = wt_{NaOH} / Mwt_{NaOH}$$

$$L_{acid} \times 0.05 = 8 / 40$$

$$L_{acid} = 0.2/ 0.05 = 4 \text{ L or } 4000 \text{ ml.}$$

Henderson Hasselbalch Equation

Henderson Hasselbalch Equation

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

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

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

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

- Multiply by log

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

- Multiply by -1

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

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

Henderson Hasselbalch Equation

$$MOH \rightleftharpoons M^+ + OH^-$$

$$K_b = \frac{[M^+][OH^-]}{[MOH]}$$

$$[M^+][OH^-] = [MOH]\,K_b$$

$$[OH^-] = \frac{K_b[MOH]}{[M^+]}$$

- Multiply by log

$$\log[OH^-] = \log K_b + \log\frac{[MOH]}{[M^+]}$$

- Multiply by -1

$$-\log[OH^-] = -\log K_b - \log\frac{[MOH]}{[M^+]}$$

$$pOH = pK_b + \log\frac{[M^+]}{[MOH]}$$