Q1: Calculate:

a) $[H^+]$, b) $[OH^-]$, c) pH, d) pOH of the final solution obtained after 100ml of 0.2M NaOH are added to 150ml of 0.4M H_2SO_4 .

-Note: Any question where an acid and a base are added together, calculate the following:

1- no. of moles of H^+ in the acid (putting in mind, if the acid is **NOT** monoprotic, then multiply the M with (n); here n = 2).

2- no. of moles of OH^- in the base (putting in mind the (n, no.of hydroxyl ions).

3- no. of moles of remaining ions (either H^+ or OH^- , according to the highest mole number)

Given values:

NaOH: V= 100ml, C= 0.2M + **H_2SO_4:** V= 150ml, C= 0.4M

Answer:

-No. of moles of H^+ = M (n) x V in L = 0.4 (2) x 0.15 = 0.12 moles

-No. of moles of OH^- = M (n) x V in L = 0.2 (1) x 0.1 = 0.02 moles

Both are strong → fully ionized

- No. of moles of remaining (H^+) = 0.12 -0.02 = 0.1 moles

a) $[H^+] = \frac{no.of\ moles\ remaining\ (H+)}{Tota\ volume\ of\ solution} = \frac{0.1}{0.1+0.15} = 0.4$ M

b) $k_w = [H^+][OH^-] \rightarrow [OH^-] = 1 \times 10^{-14} / 0.4 = 2.5 \times 10^{-14}$ M

c) pH= -log$[H^+]$ = -log (0.4) = 0.39

pH of strong acid is calculated by: **pH= -log $[H^+]$**

d) pk_w= pH + pOH → pOH = 14- 0.39 = 13.6

Q2: How many grams of solid KOH are required to neutralize 2L of an HCl solution of pH 2?

Given values:

KOH: wt=? g, + **HCl:** V= 2L , pH = 2

Answer:

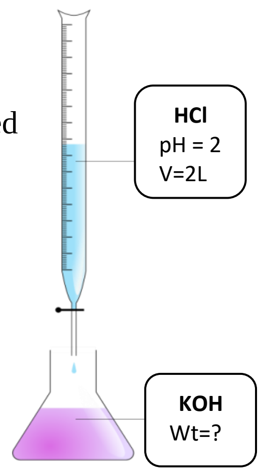
-HCl is a strong acid, it is fully ionized. From pH, conc. of acid (H^+) can be calculated

- $[H^+] = 10^{-pH} = 10^{-2} = 0.01$ M

- No. of moles of H^+ = no. of moles of OH^-

- M x V in L = wt / mwt

- 0.01 x 2 = wt / 56 → wt = 0.01 x 2 x 56 = 1.12g



Q3: The pH of a 0.27M solution of a weak acid, HA, is 4.3.

a) What is the $[H^+]$ in the solution?

- $[H^+] = 10^{-pH} = 10^{-4.3} = 5.011 \times 10^{-5}$ M

b) What is the degree of ionization of the acid?

-Degree of ionization = $\frac{[H+]}{[HA]}$ x 100 = $(5.011 \times 10^{-5} / 0.27) \times 100 = 0.018\%$

c) What is the k_a?

$HA \rightarrow H^+ + A^-$ (but since HA monoprotic acid, $[H^+] = [A^-]$)

-$k_a = [H^+]^2 / [HA]$

$= (5.011 \times 10^{-5})^2 \;/\; 0.27 = 9.3 \times 10^{-9}$

Q4: Calculate the pk_a, k_b and pk_b of the following weak acids:

a) CH_3COOH k_a=1.8 x 10^{-5}

- $pk_a = -\log k_a = -\log(1.8 \times 10^{-5}) = 4.744$

- $k_w = k_a \times k_b \rightarrow k_b = k_w / k_a$

$k_b = (1 \times 10^{-14}) / (1.8 \times 10^{-5}) = 5.55 \times 10^{-10}$

-$pk_b = -\log k_b = -\log(5.55 \times 10^{-10}) = 9.256$

OR *by ($pk_w = pk_a + pk_b$)*

b) Ammonium ion k_a=5.7 x 10^{-10}

- $pk_a = -\log k_a = -\log(5.7 \times 10^{-10}) = 9.24$

- $k_w = k_a \times k_b \rightarrow k_b = k_w / k_a$

$k_b = (1 \times 10^{-14}) / (5.7 \times 10^{-10}) = 1.754 \times 10^{-5}$

-$pk_b = -\log k_b = -\log(1.754 \times 10^{-5}) = 4.755$

OR *by ($pk_w = pk_a + pk_b$)*

Q5: The k_a of a weak acid HA, is 3 x 10^{-4}. Calculate

a) The hydroxyl ion concentration in the solution

b) The degree of dissociation of the acid in a 0.15M solution.

Given values:

HA→ k_a = 3 x 10^{-4} , **C=** 2 0.15M

Answer:

a) HA → H^+ + A^- (but since HA monoprotic acid, $[H^+] = [A^-]$)

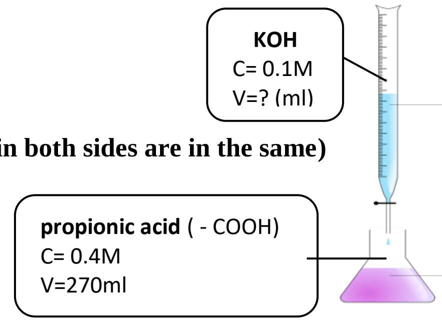
-$k_a = [H^+]^2 / [HA]$ → $[H^+] = \sqrt{ka[HA]}$

-$[H^+] = \sqrt{(3\,x10^{-4})x\,0.15}$ =6.7 x 10^{-3} M

-$k_w = [H^+][OH^-]$ → $[OH^-]$ = 1 x10^{-14} / 6.7 x 10^{-3} = 1.49 x 10^{-12} M

b) Degree of ionization = $\frac{[H+]}{[HA]}$ x 100 = (6.7 x 10^{-3} / 0.15) x 100 = 4.47%

Q6: How many ml of 0.1M KOH are required to titrate completely 270ml of 0.4M propionic acid?



- No. of moles of H^+ = no. of moles of OH^-
- M x V (ml) = M x V (ml) **(the important thing is the units in both sides are in the same)**
- 0.4 x 270 = 0.1 x V (ml)
- V(ml) of KOH = (0.4 x270) /0.1 =1080 ml

Q7: 200ml of 0.2M NaOH was mixed with 800ml of 0.1M HCOOH. Calculate the pH of the resulting solution. pKa = 3.75.

Given values:

NaOH: V= 200ml, C= 0.2M + **HCOOH:** V= 800ml, C= 0.1M

Answer:

-No. of moles of H^+ = M (n) x V in L = 0.1 (1) x 0.8 = 0.08 moles

-No. of moles of OH^- added = M (n) x V in L = 0.2 (1) x 0.2 = 0.04 moles

-No. of moles of remaining (H^+) = 0.08 -0.04 = 0.04 moles

pH= pka + log $\frac{[A^-]=(no.of\ moles\ of\ OH-added)}{[HA]=(no.of\ moles\ remainig\ H+)}$

pH of solution of strong base added to weak acid is calculated by:

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

= 3.75 + log $\frac{0.04}{0.04}$

= 3.75 → **pH=pKa (the solution act as buffer)**

Q8: In the following reaction:

$A + B \longrightarrow 2C + D$

If [A] = 2.1 x10^{-3} M, [B] = 3.4 x10^{-2} M , [D] = 1.8 x10^{-3} M , K_{eq}= 6.23 x10^{-6}

Calculate the concentration of [C].

$$k_{eq} = \frac{[C]^2\ [D]}{[A][B]}$$

$$6.23x10^{-6} = \frac{[C]^2\ (1.8x10^{-3})}{(2.1x10^{-3})(3.4x10^{-2})}$$

$$[C]^2 = 2.471x10^{-7}$$

$$[C] = \sqrt{2.471x10 - 7} = 4.969\ x10^{-4}\ M$$