

* Reaction

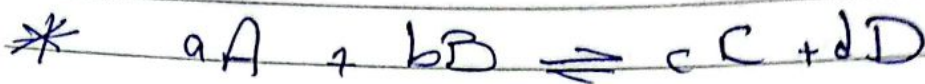
irreversible



Reversible

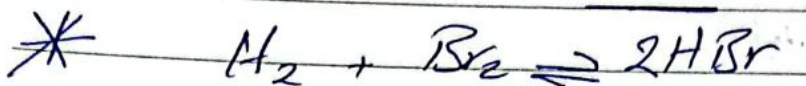


at some point the
rate of forward = rate of backward
 $\therefore$ equilibrium Ex: weak acids & weak bases



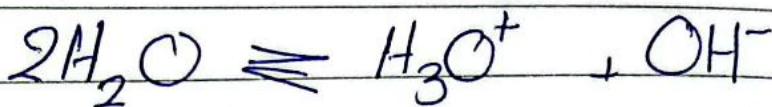
$$K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

eq^m Constant



$$K = \frac{[HBr]^2}{[H_2][Br_2]}$$

For H_2O



$$K \times \frac{[OH^-][H_3O^+]}{[H_2O]^2}$$

$$K [H_2O]^2 = K_w = [OH^-][H^+]$$

ion Product Constant for water

$$1 \times 10^{-14} K_w = \frac{[OH^-][H^+]}{1 \times 10^{-7} \times 1 \times 10^{-7}}$$

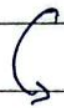
taking $-\log$ of Both sides:

$$-\log(K_w) = -\log([OH^-]) - \log([H^+])$$

$$pK_w = pOH + pH$$

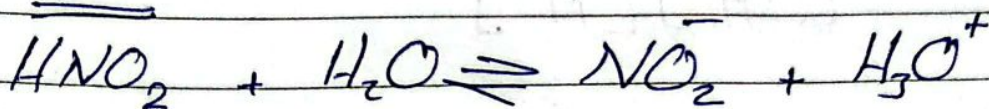
$$14 = pOH + pH$$

for water, $pH = pOH = \underline{\underline{7}}$



Some molecules of water dissociate into H_3O^+ (for every H_3O^+ , OH^- is formed)

For weak Acids



$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]}$$

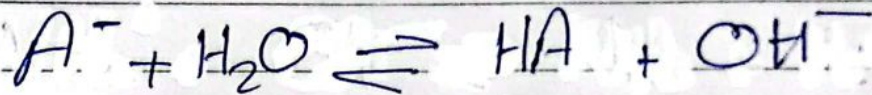
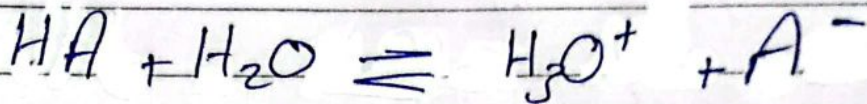
Constant
for Acid

التاريخ:
 * For weak bases:-



$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

Relation betⁿ K_a & K_b



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

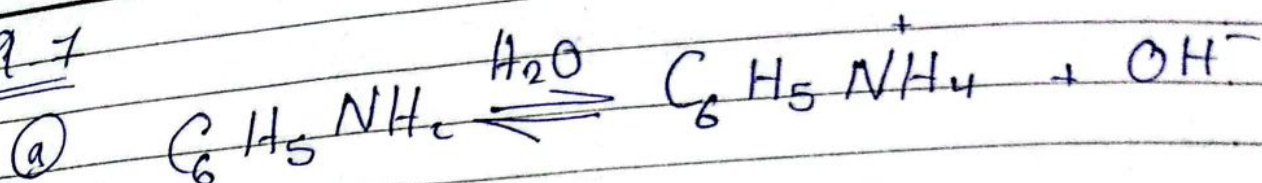
$$K_b = \frac{[\text{OH}^-][\text{HA}]}{[\text{A}^-]}$$

$$\cancel{K_w} = \cancel{K_a} \cdot \cancel{K_b}$$

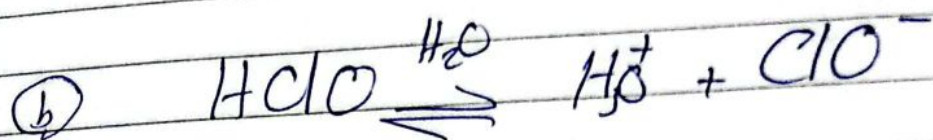
$$K_a \cdot K_b = [\text{H}_3\text{O}^+][\text{OH}^-] = K_w = 1 \times 10^{-14}$$

#

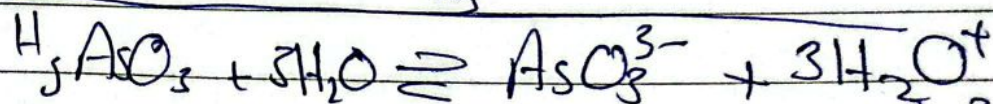
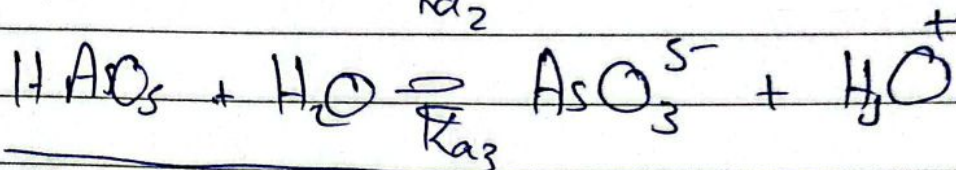
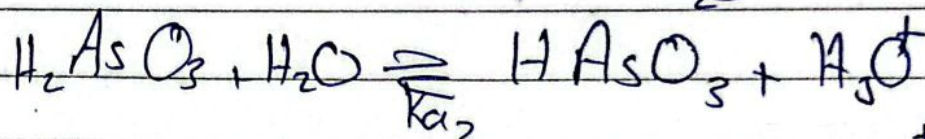
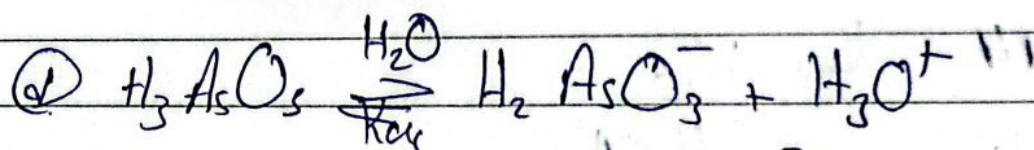
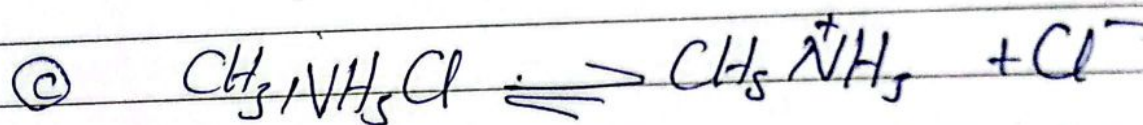
97



$$K_b = \frac{[\text{C}_6\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_6\text{H}_5\text{NH}_2]} = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{K_a}$$



$$K_a = \frac{[\text{ClO}^-][\text{H}_3\text{O}^+]}{[\text{HClO}]} = 3 \times 10^{-8}$$



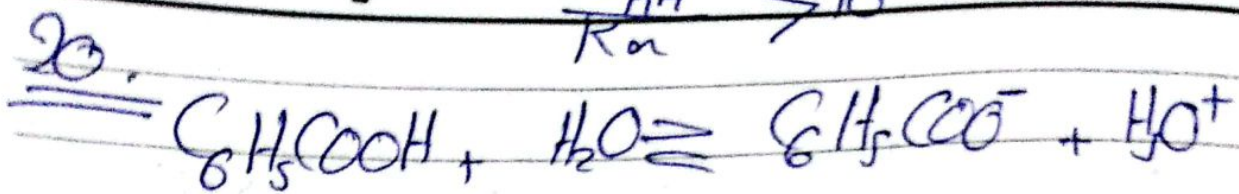
$$K = K_{a1} \cdot K_{a2} \cdot K_{a3} = \frac{[\text{AsO}_3^{3-}][\text{H}_3\text{O}^+]^3}{[\text{H}_3\text{AsO}_3]} = \text{---}$$

التاريخ:

$$\frac{C_{HA}}{K_a} > 10^3$$

error < 1.6%

20.



$$0.0300M$$

two

two

$$0.0300 - x$$

x

x

$$K_a = \frac{[H_3O^+][C_6H_5COO^-]}{[C_6H_5COOH]}$$

$$K_a = \frac{x^2}{[0.03 - x]}$$

$$\text{assume } [C_6H_5COOH] = [HA]$$

$$K_a = \frac{x^2}{0.03}$$

$$0.03 K_a = x^2$$

$$x = \sqrt{0.03 K_a}$$

$$K_a = \frac{x^2}{0.03}$$

$$x = \sqrt{6.5 \times 10^{-5} \times 0.03}$$

$$[H_3O^+] = x = 1.39 \times 10^{-3}$$

$$[OH^-] = \frac{1 \times 10^{-14}}{1.39 \times 10^{-3}} \approx \frac{K_w}{[H_3O^+]}$$

$$[OH^-] = 7.1 \times 10^{-12}$$

التاريخ:

No assumptions made

$$K_a = \frac{x^2}{0.03 - x}$$

$$K_a \times 0.03 - x K_a = x^2$$

$$1.95 \times 10^{-6} - x K_a = x^2$$

$$x^2 + 6.5 \times 10^{-5} x - 1.95 \times 10^{-6} = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - (4 \times a \times c)}}{2a}$$

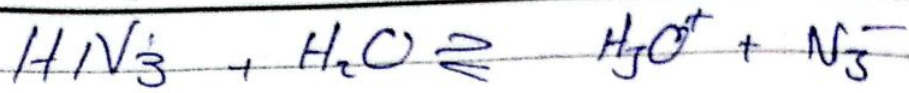
$$[H_3O^+] = x = 1.3 \times 10^{-3}$$

$$[OH^-] = 7.6 \times 10^{-12}$$

$$\text{Error} = \frac{(1.39 - 1.3) \times 10^{-3}}{1.39} \times 100$$

$$= \frac{6.5}{1.39} \%$$

Hydrazoic Acid



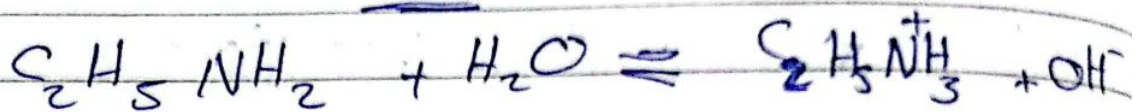
$$0.06$$



$$0.06 - x$$



$$K_a \neq \frac{[x][x]}{0.06 - x}$$



$$0.1$$



$$0.1 - x$$

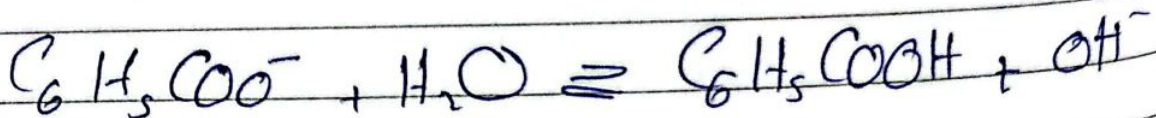
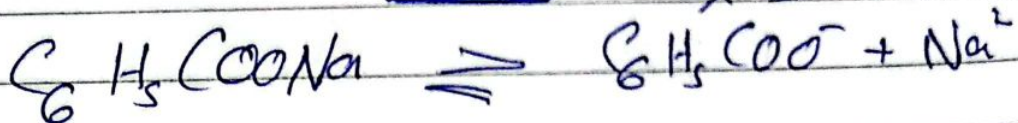


$$K_b \neq \frac{[x]^2}{0.1 - x}$$

$$0.1 K_b = [x]^2$$

$$x = \sqrt{0.1 K_b}$$

$$x =$$



$$0.2$$



$$0.2 - x$$



$$K_b \neq \frac{[x]^2}{0.2 - x}$$

$$\frac{1 \times 10^{-14}}{6.28 \times 10^{-5}} = \frac{[X^+]}{0.2} \quad \text{ignore}$$

$$[OH^-] = \sqrt{\frac{1 \times 10^{-14}}{6.28 \times 10^{-5}}} \times 0.2 = 5.64 \times 10^{-7}$$

$$[H^+] = \frac{1 \times 10^{-14}}{5.64 \times 10^{-7}} = 1.77 \times 10^{-8}$$

* Buffer Solutions

Are solutions that resist small changes in pH due to the addition of strong acid or strong base.



$$K_a = \frac{[H^+][B^-]}{[HB]}$$

$$[H^+] = \frac{K_a [HB]}{[B^-]}$$

$$\log [H^+] = \log [K_a] + \log [HB] - \log [B^-]$$

$$\log [H^+] = \log K_a + \log \left(\frac{HB}{B^-} \right) \quad \times -1$$

$$pH = pK_a + \log \left(\frac{B^-}{HB} \right)$$

Example 9

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

$$= -\log(1.4 \times 10^{-4}) + \log\left(\frac{1}{0.2}\right)$$

Lactic Acid $pK_a = 3.85$ #

$$[HA] = \frac{1}{550 \times 10^{-3}}$$

$$[A^-] = \frac{1}{550 \times 10^{-3}}$$

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

$$pH = -\log(1.4 \times 10^{-4}) + \log\left(\frac{1.81}{1.81}\right)$$

$$pH = 3.85$$

$$[H^+] = 1.41 \times 10^{-4} M$$

$$[A^-] = \frac{34.6}{(25 + (3 \times 12) + 5 + (3 \times 16)) \times 550 \times 10^{-3}} = 0.56 M$$

$$[HA] = 1.20 M$$

$$pH = 3.523$$

$$[H^+] = 2.99 \times 10^{-4} M$$

التاريخ:

$$(a) \quad pH = pK_a + \log \left(\frac{A^-}{HA} \right)$$

$$pH = -\log(1.8 \times 10^{-5}) + \log \left(\frac{8 \times 10^{-3}}{200 \times 10^{-3}} \right)$$

$$pH = 4.34$$

~~$$pH = -\log(1.8 \times 10^{-5}) + \log \left(\frac{8 \times 10^{-3}}{200 \times 10^{-3}} \right)$$~~



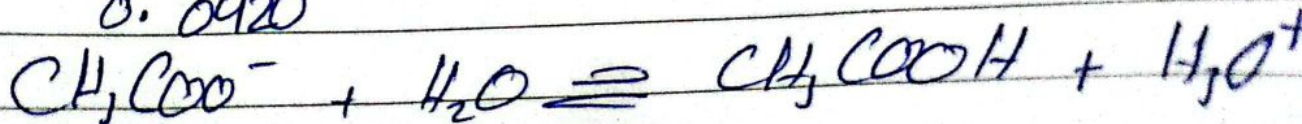
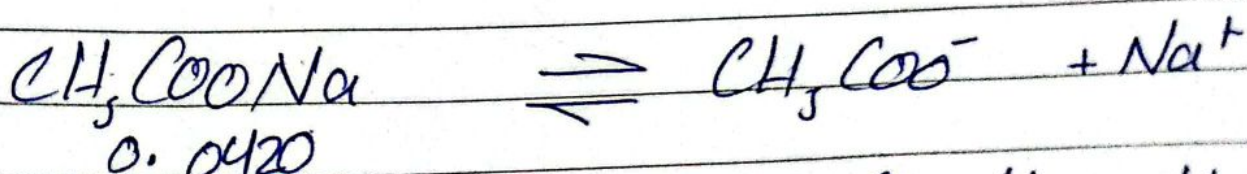
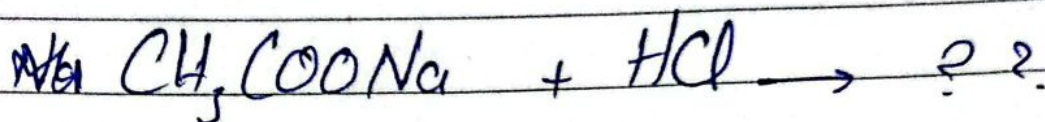
$$10^3 \times 0.175 \times 100 \quad 0.05 \times 100 \times 10^3 \quad \text{Zero} \quad \text{Zero}$$

$$0.0175 \text{ mol} \quad 5 \times 10^{-3} \text{ mol} \quad \text{Zero} \quad \text{Zero}$$

$$0.0125 \text{ mol} \quad \text{Zero} \quad \frac{5 \times 10^{-3} \text{ mol}}{0.0125} \quad 5 \times 10^{-3}$$

$$pH = -\log(1.8 \times 10^{-5}) + \log \left(\frac{5 \times 10^{-3}}{200} \right)$$

$$pH = 4.546$$



التاريخ:

$$(a) \quad pH = pK_a + \log \left(\frac{A^-}{HA} \right)$$

$$pH = -\log(1.8 \times 10^{-5}) + \log \left(\frac{8 \times 10^{-5}}{200 \times 10^{-3}} \right)$$

$$pH = 4.34$$

~~$$pH = -\log(1.8 \times 10^{-5}) + \log \left(\frac{8 \times 10^{-5}}{200 \times 10^{-3}} \right)$$~~



$$10^3 \times 0.175 \times 100 \quad 0.05 \times 100 \times 10^{-3} \quad \text{Zero} \quad \text{Zero}$$

$$0.0175 \text{ mol} \quad 5 \times 10^{-3} \text{ mol} \quad \text{Zero} \quad \text{Zero}$$

$$0.0125 \text{ mol} \quad \text{Zero} \quad 5 \times 10^{-3} \text{ mol} \quad 5 \times 10^{-3}$$

$$pH = -\log(1.8 \times 10^{-5}) + \log \left(\frac{5 \times 10^{-3}}{200} \right)$$

$$pH = 4.546$$

