

## **EXERCISE O-I**

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**Single correct**

Q.1 The conjugate acid of  $\text{NH}_2^-$  is

- (A)  $\text{NH}_3$       (B)  $\text{NH}_2\text{OH}$       (C)  $\text{NH}_4^+$       (D)  $\text{N}_2\text{H}_4$

To write conjugate acid add  $\text{H}^+$ .

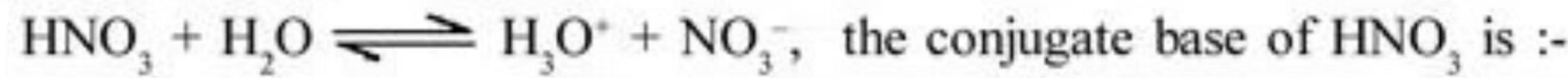
Q.2 Which of the following is not a Bronsted acid:-

- (A)  $\text{CH}_3\text{NH}_4^+$        (B)  $\text{CH}_3\text{COO}^-$       (C)  $\text{H}_2\text{O}$       (D)  $\text{HSO}_4^-$

Bronsted acid is proton donor.

$\text{CH}_3\text{COO}^-$  is a Bronsted base. It can accept a proton.

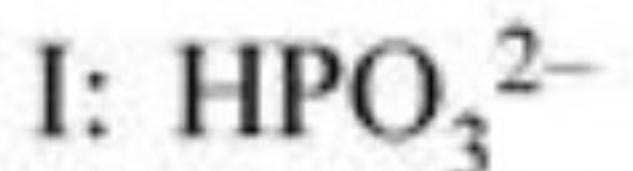
Q.3 In the reaction



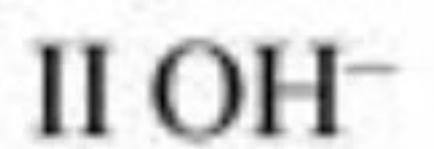
- (A)  $\text{H}_2\text{O}$       (B)  $\text{H}_3\text{O}^+$       (C)  $\text{NO}_3^-$       (D)  $\text{H}_3\text{O}^+$  and  $\text{NO}_3^-$

To write conjugate base remove  $\text{H}^+$ .

Q.4 Out of the following, amphiprotic species in aqueous medium are



(A) I, III, IV



(B) I and III



(C) III and IV



(D) All

→  $\text{H}_3\text{PO}_3$  is a dibasic acid, so  $\text{HPO}_3^{2-}$  is not an amphiprotic species.

→ Amphiprotic species give both  $\text{H}^+$  and  $\text{OH}^-$  in water.



⇒  $\text{H}_2\text{PO}_4^-$  is amphiprotic. Similarly  $\text{HCO}_3^-$  is also amphiprotic.

Q.5 When ammonia is added to water, it decreases the concentration of which of the following ion

- (A)  $\text{OH}^-$       (B)  $\text{H}_3\text{O}^+$       (C)  $\text{NH}_4^+$       (D)  $\text{NH}_4^+ & \text{ OH}^-$

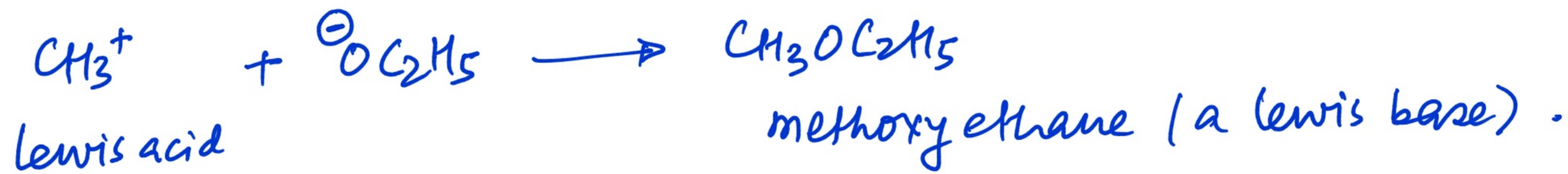
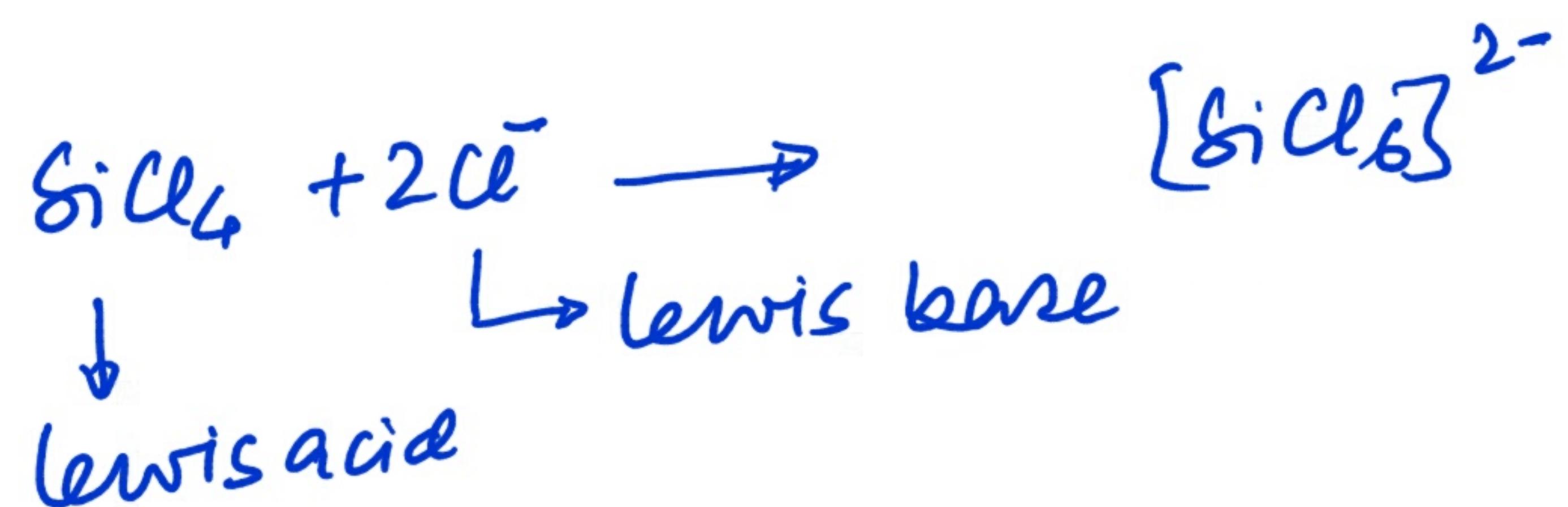
On adding  $\text{NH}_3$ , conc. of  $\text{OH}^-$  increases  $\Rightarrow [\text{H}_3\text{O}^+]$  decreases.

Q.6 Which of the following pair is Lewis acid & Lewis base & Product of these is also Lewis base

- (A)  $\text{BF}_3$ ,  $\text{NH}_3$       (B)  $\text{SiCl}_4$ ,  $2\text{CF}_3^+$       ✓ (C)  $\text{CH}_3^+$ ,  ${}^\ominus\text{OCH}_2\text{H}_5$       (D) All of these

Lewis acid  $\rightarrow$  Accepts lone pair of e<sup>-</sup>

Lewis base  $\rightarrow$  donates lone pair of e-.



Q.7 Ionic product of water will increase, if :-

- (A) Pressure is decreased
- (B)  $\text{H}^+$  is added
- (C)  $\text{OH}^-$  is increased
- (D) Temperature is increased

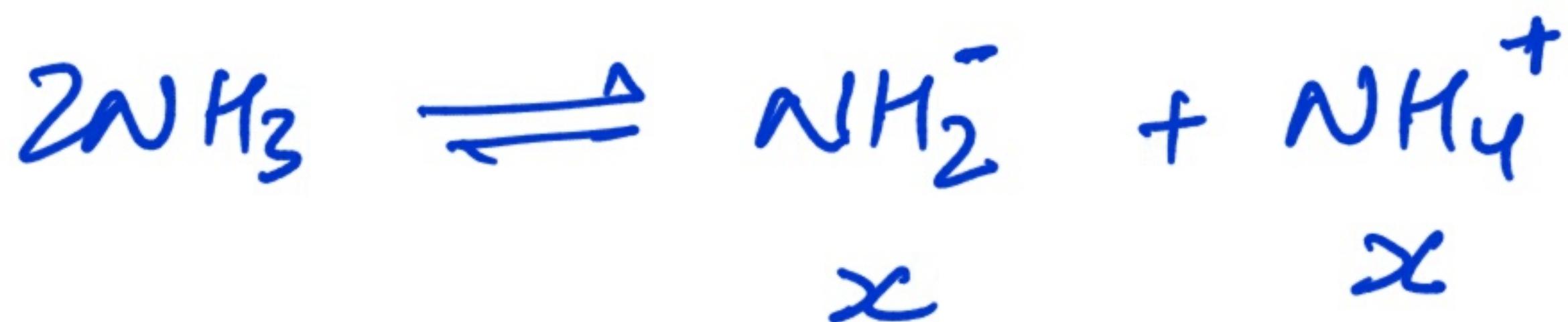
Ionic product ( $K_w$ ) depends only upon temp  
and it increases with increase in temp.

- Q.8 At 60°C, pure water has  $[H_3O^+] = 10^{-6.7}$  mol/lit. what is the value of  $K_w$  at 60°C :-
- (A)  $10^{-6}$       (B)  $10^{-12}$       (C)  $10^{-67}$       ✓(D)  $10^{-13.4}$

In pure water,  $[H_3O^+] = [OH^-] = 10^{-6.7}$   
 $\hookrightarrow$  Always -

$$K_w = [H_3O^+] \times [OH^-] = 10^{-6.7} \times 10^{-6.7} = 10^{-13.4}$$

- Q.9 Liquid  $\text{NH}_3$  ionises to a slight extent. At a certain temperature its self ionization constant  $K_{\text{SIC}(\text{NH}_3)} = 10^{-30}$ . The number of  $\text{NH}_4^+$  ions present per  $100 \text{ cm}^3$  of pure liquid are  
(A)  $10^{-15}$       (B)  $6.022 \times 10^8$       ✓ (C)  $6.022 \times 10^7$       (D) None



$$K_{\text{self ionization const}} = [\text{NH}_2^-][\text{NH}_4^+]$$

$$10^{-30} = x^2 \Rightarrow x = 10^{-15} \text{ mole/L}$$

$$\therefore [\text{NH}_4^+] = 10^{-15} \text{ mole/L}$$

$$= 10^{-15} \times 6.022 \times 10^{23} \text{ ions/1000 ml}$$

$$= 6.022 \times 10^8 \text{ ions/1000 ml.}$$

$$= \frac{6.022 \times 10^8 \times 100}{1000} \text{ ions/100 ml}$$

$$= \boxed{6.022 \times 10^7 \text{ ions/100 ml.}}$$

**Q. 10** The pH of solution is increased from 3 to 6. Its  $\text{H}^+$  ion conc. will be :-

- (A) Reduced to half
  - (B) Doubled
  - (C) Reduced by 1000 times
  - (D) Increased by 1000 times

$$PH_1 = 3 \Rightarrow [H^+]_1 = 10^{-3}$$

$$\text{pH}_2 = 6 \Rightarrow [\text{H}^+]_2 = 10^{-6}$$

$$\therefore \frac{[H^+]_1}{[H^+]_2} = \frac{10^{-3}}{10^{-6}} = 1000$$

∴  $[H^+]$  ion is reduced by 1000 times.

Q.11. pOH of [1/200] mol/m<sup>3</sup> H<sub>2</sub>SO<sub>4</sub> (aq.) solution at 25°C is-

(A) 2

(B) 5

✓(C) 9

(D) 12

$$\begin{aligned} [\text{H}_2\text{SO}_4] &= \frac{1}{200} \text{ mole/m}^3 \\ &= \frac{1}{200} \text{ mole/1000L} \quad \because 1\text{m}^3 = 1000\text{L} \\ &= \frac{1}{200 \times 1000} \text{ mole/L} \\ &= 0.5 \times 10^{-5} \text{ mole/L} = 5 \times 10^{-6} \text{ mole/L} \\ \therefore [\text{H}^+] &= 2 \times 5 \times 10^{-6} = 10^{-5} \text{ mole/L} \Rightarrow \text{pH} = 5 \end{aligned}$$

$$\therefore \text{pOH} = 14 - 5 = \boxed{9} \text{ Ans.}$$

Q.12 Degree of dissociation of 0.1 N CH<sub>3</sub>COOH is :- (Dissociation constant = 1 × 10<sup>-5</sup>)

(A) 10<sup>-5</sup>

(B) 10<sup>-4</sup>

(C) 10<sup>-3</sup>

(D) 10<sup>-2</sup>

For CH<sub>3</sub>COOH, Normality = molarity.

∴ C = 0.1 M.

$$\alpha = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{10^{-5}}{0.1}} = \boxed{10^{-2}}$$

Ans.

Q.13 The pH of a 0.02 M ammonia solution which is 5% ionised will be :-

(A) 2

(B) 11

(C) 5

(D) 7

$$[\text{NH}_3] = C = 0.02 \text{ M}, \alpha = \frac{5}{100} = 0.05 .$$

$$[\text{OH}^-] = C\alpha = 0.02 \times 0.05 = 0.0010 = 10^{-3} \text{ M}.$$

$$\therefore \text{pOH} = 3 \Rightarrow \boxed{\text{pH} = 11} \quad \underline{\text{Ans.}}$$

**Q. 14** The pH of an aqueous solution of 1.0 M solution of a weak monoprotic acid which is 1% ionised is

- (A) 1      (B) 2      (C) 3      (D) 11

$$C = 1M, \quad \mathcal{L} = 0.01.$$

$$[H^+] = C\alpha = 1 \times 0.01 = 10^{-2} M$$

$$\Rightarrow \text{PH} = 2$$

Q.15 The concentration of  $[H^+]$  and concentration of  $[OH^-]$  of a 0.1 M aqueous solution of 2% ionised weak acid is [ionic product of water =  $1 \times 10^{-14}$ ]

- (A)  $0.02 \times 10^{-3}$  M and  $5 \times 10^{-11}$  M      (B)  $1 \times 10^{-3}$  M and  $3 \times 10^{-11}$  M  
~~(C)  $2 \times 10^{-3}$  M and  $5 \times 10^{-12}$  M~~      (D)  $3 \times 10^{-2}$  M and  $4 \times 10^{-13}$  M

$$C = 0.1 \text{ M}, \alpha = \frac{2}{100} = 0.02$$

$$[H^+] = C\alpha = 0.1 \times 0.02 = \boxed{2 \times 10^{-3} \text{ M.}}$$

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{10^{-14}}{2 \times 10^{-3}} = 0.5 \times 10^{-11} = \boxed{5 \times 10^{-12} \text{ M.}}$$

Q.16 What is the quantity of NaOH present in 250 cc of the solution, so that it gives a pH = 13 :-

- (A)  $10^{-13}$  g      (B)  $10^{-1}$  g      ✓ (C) 1.0 g      (D) 4.0 g

$$V=250\text{ml}, \quad \text{pH}=13 \Rightarrow [\text{H}^+]=10^{-13} \Rightarrow [\text{OH}^-]=\frac{10^{-14}}{10^{-13}}=10^{-1}\text{M}.$$

$$\therefore [\text{NaOH}]=[\text{OH}^-]=10^{-1}\text{M}$$

$$\text{or, } \frac{n \times 1000}{250} = 0.1 \quad (\text{where } n = \text{moles of NaOH})$$

$$\text{or, } n = \frac{1}{40} \text{ mole}$$

$$\therefore \text{mass of NaOH} = n \times 40 = \frac{1}{40} \times 40 \text{ g} = \boxed{1 \text{g}}$$

Q.17 An aqueous solution of HCl is  $10^{-9}$  M HCl. The pH of the solution should be:-

- (A) 9                          (B) Between 6 and 7    (C) 7                          (D) Unpredictable

When [strong acid]  $\leq 10^{-9}$   $\Rightarrow$  pH  $\approx 7$ .

Exact pH will be very slightly less than 7 but very close to 7.

Q.18. The moles of  $H^+$  from  $H_2O$  in a  $1l$ ,  $\sqrt{5} \times 10^{-7}$  M HCl solution at  $25^\circ C$ , is

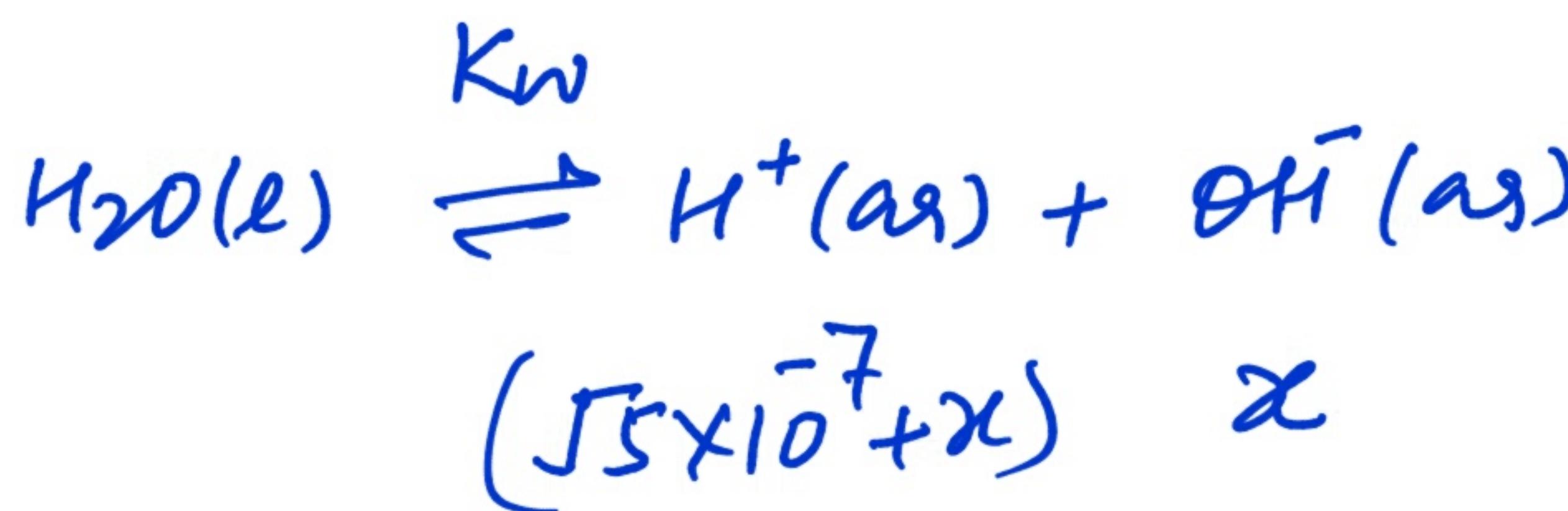
$$(\sqrt{5} = 2.23)$$

$$(A) 10^{-7}$$

$$(B) 6.85 \times 10^{-8}$$

$$\checkmark (C) 3.85 \times 10^{-8}$$

$$(D) 10^{-8}$$



$$10^{-14} = x(5 \times 10^{-7} + x)$$

$$\text{or, } x^2 + 5 \times 10^{-7}x - 10^{-14} = 0$$

$$\text{or, } x = \frac{-5 \times 10^{-7} \pm \sqrt{5 \times 10^{-14} + 4 \times 10^{-14}}}{2}$$

$$\therefore x > 0 \Rightarrow x = \frac{-5 \times 10^{-7} + 3 \times 10^{-7}}{2}$$

$$\begin{aligned} &\text{or, } x = \frac{10^7 (3-2.23)}{2} \\ &\text{or, } x = \frac{0.77}{2} \times 10^{-7} = 0.385 \times 10^{-7} \\ &\text{or, } x = 3.85 \times 10^{-8} \text{ M} \\ &\boxed{[H^+]_{H_2O} = x = 3.85 \times 10^{-8} \text{ M}} \end{aligned}$$

Ans.

Q.19 Which one of the following has highest pH:-

(A) Distilled water

$$PH = 7$$

(C) 1 M NaOH

$$PH = 14$$



(B) 1 M NH<sub>3</sub>

$7 < PH < 14$  ∵ weak base.

(D) Water saturated with chlorine



$$PH < 7$$



Q.20 8 gm NaOH and 4.9 gm H<sub>2</sub>SO<sub>4</sub> are present in one litre of the solution. What is its pH

(A) 1

(B) 13

(C) 12

(D) 2

$$\text{NaOH} = 8 \text{ g} = \frac{8}{40} \text{ mole} = 0.2 \text{ mole} \Rightarrow \text{OH}^- = 0.2 \text{ mole}$$

$$\text{H}_2\text{SO}_4 = 4.9 \text{ g} = \frac{4.9}{98} \text{ mole} = \frac{1}{20} \text{ mole} = 0.05 \text{ mole} \therefore \text{H}^+ = 2 \times 0.05 \\ = 0.1 \text{ mole}$$

$\therefore$  moles of OH<sup>-</sup> > moles of H<sup>+</sup>

$$\therefore [\text{OH}^-] = \frac{(0.2 - 0.1)}{1} = 0.1 \text{ mole/L} \Rightarrow \text{pOH} = 1$$

$$\Rightarrow \boxed{\text{pH} = 13}$$

Ans.

Q.21 10 ml of  $\frac{M}{200}$  H<sub>2</sub>SO<sub>4</sub> is mixed with 40 ml of  $\frac{M}{200}$  H<sub>2</sub>SO<sub>4</sub>. The pH of the resulting solution is

(A) 1

✓ (B) 2

(C) 2.3

(D) none of these

$$[H^+] = \frac{\frac{1}{200} \times 10 \times 2 + \frac{1}{200} \times 40 \times 2}{(10+40)} = \frac{0.1 + 0.4}{50} = \frac{0.5}{50} = 10^{-2} M$$

∴

$$\boxed{PH=2}$$

Ans.

Q.22 Which of the following solution will have pH close to 1.0?

- (A) 100 ml of M/100 HCl + 100 ml of M/10 NaOH
- (B) 55 ml of M/10 HCl + 45 ml of M/10 NaOH
- (C) 10 ml of M/10 HCl + 90 ml of M/10 NaOH
- (D) 75 ml of M/5 HCl + 25 ml of M/5 NaOH

A)  $H^+ = \frac{1}{100} \times 100 = 1 \text{ mmole}$      $OH^- = \frac{1}{10} \times 100 = 10 \text{ mmole}$

$\therefore OH^- > H^+ \Rightarrow pH > 7$

B)  $H^+ = \frac{1}{10} \times 55 = 5.5 \text{ mmole}$ ,     $OH^- = \frac{45}{10} = 4.5 \text{ mmole}$

$$[H^+] = \frac{5.5 - 4.5}{55 + 45} = \frac{1}{100} = 10^{-2} \Rightarrow pH = 2$$

C)  $H^+ = \frac{1}{10} \times 10 = 1 \text{ mmole}$ ,     $OH^- = \frac{1}{10} \times 90 = 9 \text{ mmole}$

$\therefore OH^- > H^+ \Rightarrow pH > 7.$

D)  $H^+ = 75 \times \frac{1}{5} = 15 \text{ mmole}$ ,  $OH^- = 25 \times \frac{1}{5} = 5 \text{ mmole}$ .

$$[H^+] = \frac{(15-5)}{(75+25)} = \frac{10}{100} = 0.1 = 10^{-1} M$$

$\therefore \boxed{PH=1}$  Ans.

Q.23 A solution with pH 2.0 is more acidic than the one with pH 6.0 by a factor of:

(A) 3

(B) 4

(C) 3000

(D) 10000

$$\text{pH}_1 = 2 \Rightarrow [\text{H}^+]_1 = 10^{-2} \text{ M}$$

$$\text{pH}_2 = 6 \Rightarrow [\text{H}^+]_2 = 10^{-6} \text{ M.}$$

$$\frac{[\text{H}^+]_1}{[\text{H}^+]_2} = \frac{10^{-2}}{10^{-6}} = \boxed{10^4} \text{ Ans.}$$

Q.24 The first and second dissociation constants of an acid  $\text{H}_2\text{A}$  are  $1.0 \times 10^{-5}$  and  $5.0 \times 10^{-10}$  respectively.

The overall dissociation constant of the acid will be :

(A)  $5.0 \times 10^{-5}$

(B)  $5.0 \times 10^{15}$

~~(C)  $5.0 \times 10^{-15}$~~

(D)  $0.2 \times 10^5$

$$\begin{aligned}K_{\text{overall}} &= K_{\text{a}_1} \times K_{\text{a}_2} \\&= 1 \times 10^{-5} \times 5 \times 10^{-10} = 5 \times 10^{-15}\end{aligned}$$

Q.25 If  $pK_b$  for fluoride ion at  $25^\circ C$  is 10.4, the ionisation constant of hydrofluoric acid in water at this temperature is : ( $\log 2 = 0.3$ )

(A)  $4 \times 10^{-11}$

(B)  $3 \times 10^{-3}$

$\checkmark$  (C)  $2.5 \times 10^{-4}$

(D)  $2 \times 10^{-2}$

$$PK_b F^- = 10.4$$

$$\therefore PK_a_{HF} + PK_b F^- = 14 \quad (\because HF \text{ and } F^- \text{ are conjugate acid-base pair})$$

$$\therefore PK_a_{HF} = 14 - 10.4 = 3.6$$

$$\text{or, } K_a_{HF} = 10^{-3.6} = 10^{-3-0.6} = \frac{10^{-3}}{10^{0.6}} = \frac{10^{-3}}{10^{2 \times 0.3}} = \frac{10^{-3}}{10^{2 \times \log 2}}$$

$$= \frac{10^{-3}}{10^{\log 4}} = \frac{10^{-3}}{4} = \boxed{\underline{2.5 \times 10^{-4}}} \\ \text{Ans.}$$

Q.26 pH of an aqueous solution of NaCl at 85°C should be

(A) 7

(B) > 7

~~(C) < 7~~

(D) 0

As  $T \uparrow K_w \uparrow \Rightarrow$  pH of pure water will be less than 7 at temps greater than  $25^\circ C$ .

Also pH of NaCl soln = pH of pure water  $\Rightarrow < 7$ .

Q.27 1 cc of 0.1 N HCl is added to 99 cc solution of NaCl. The pH of the resulting solution will be

(A) 7

(B) 3

(C) 4

(D) 1

For HCl, Normality = Molarity.

$$[\text{H}^+] = \frac{0.1 \times 1}{(1 + 99)} = \frac{0.1}{100} = 10^{-3} \text{ M} \Rightarrow \boxed{\text{pH}=3}$$

Q.28 The degree of hydrolysis of a salt of weak acid and weak base in it's 0.1 M solution is found to be 0.5. If the molarity of the solution is 0.2 M, the percentage hydrolysis of the salt should be

- (A) 100%      ~~(B) 50%~~      (C) 25%      (D) none of these

for a salt of weak acid and weak base,

$$\frac{h}{(1-h)} = \sqrt{K_h} = \text{independent of concentration.}$$

$\therefore$  'h' will remain unchanged.  $\therefore h=0.5$

$$\% \text{ hydrolysis} = h \times 100 = 50\%.$$

Q.29 What is the percentage hydrolysis of NaCN in N/80 solution when the dissociation constant for HCN is  $2 \times 10^{-9}$  and  $K_w = 1.0 \times 10^{-14}$

(A) 2

(B) 5.26

(C) 8.2

(D) 9.6

for NaCN 80<sup>1</sup>, Normality = molarity.

$$C = \frac{1}{80} M.$$

$$h = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{K_w}{K_a \times C}} = \sqrt{\frac{10^{-14}}{2 \times 10^{-9} \times \frac{1}{80}}} = \sqrt{4 \times 10^{-4}} = 0.02.$$

$$\% \text{ Hydrolysis} = h \times 100 = 0.02 \times 100 = 2\%$$

Q.30 The compound whose 0.1 M solution is basic is

- (A) Ammonium acetate
- (C) Ammonium sulphate

- (B) Ammonium chloride
- (D) Sodium acetate

A)  $\text{NH}_4\text{CH}_3\text{COO}$   $\text{pH} = 7 + \frac{1}{2}(\text{pK}_a - \text{pK}_b)$ , salt of weak acid + weak base.

B)  $\text{NH}_4\text{Cl}$  salt of strong acid + weak base  $\Rightarrow \text{pH} < 7$

C)  $(\text{NH}_4)_2\text{SO}_4$  salt of strong acid + weak base  $\Rightarrow \text{pH} < 7$ .

D)  $\text{CH}_3\text{COONa}$  salt of weak acid + strong base  $\Rightarrow \text{pH} > 7$

Q.31 If equilibrium constant of



is  $1.8 \times 10^{-5}$ , equilibrium constant for



(A)  $1.8 \times 10^{-9}$

(B)  $1.8 \times 10^9$

(C)  $5.55 \times 10^{-9}$

(D)  $5.55 \times 10^{10}$

$$K_a = 1.8 \times 10^{-5}$$

Required equilibrium const =  $\frac{1}{K_b} = \frac{K_a}{K_w} = \frac{1.8 \times 10^{-5}}{10^{-14}}$

$$= [1.8 \times 10^9] \text{ Ans.}$$

Q.32 The  $pK_a$  of a weak acid, HA, is 4.80. The  $pK_b$  of a weak base, BOH, is 4.78. The pH of an aqueous solution of the corresponding salt, BA, will be :

(A) 8.58

(B) 4.79

(C) 7.01

(D) 9.22

salt of  $(w_A + w_B)$

$$pH = 7 + \frac{1}{2} (pK_a - pK_b)$$

$$= 7 + \frac{1}{2} (4.8 - 4.78)$$

$$= 7 + \frac{1}{2} (0.02)$$

$$= \boxed{7.01}$$

Q.33 The highest pH value is of :-

- (A) 0.1 M NaCl  $\text{pH} = 7$
- ✓ (C) 0.1 M CH<sub>3</sub>COONa  $\text{pH} > 7$

↓

$\because$  salt of  
strong base + weak acid.

- (B) 0.1 M NH<sub>4</sub>Cl  $\text{pH} < 7$  (WB+SA)
- (D) 0.1 M CH<sub>3</sub>COONH<sub>4</sub>

$\text{pH} \approx 7$

( $\because K_a_{\text{CH}_3\text{COOH}} = K_b_{\text{NH}_4\text{OH}}$ )

**Q.34 pH of  $K_2S$  solution is:-**

- (A) 7      (B) Less than 7      ✓ (C) More than 7      (D) 0

$K_2S$  = salt of weak acid  $H_2S$  and strong base  $KOH$  -

$\therefore \text{PH} > 7.$

Q.35 Degree of Hydrolysis of  $\frac{N}{100}$  solution of KCN is (Given  $K_a = 1.6 \times 10^{-9}$ )

- (A)  $2.5 \times 10^{-3}$       (B)  $2.5 \times 10^{-2}$       (C)  $2.5 \times 10^{-4}$       (D)  $2.5 \times 10^{-5}$

For KCN, Normality = Molarity.

$$[KCN] = [CN^-] = C = \frac{1}{100} M.$$

$$K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{1.6 \times 10^{-9}} = \frac{10^{-5}}{1.6} = \frac{10^{-4}}{1.6}$$

$$h = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{10^{-4}}{1.6 \times \frac{1}{100}}} = \sqrt{\frac{10^{-2}}{16}} = \frac{0.1}{4} = 0.025$$

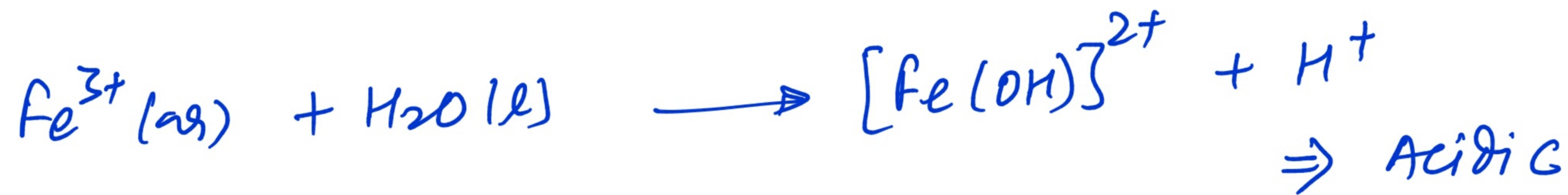
$= \boxed{2.5 \times 10^{-2}}$

Q.36 A solution of  $\text{FeCl}_3$  in water acts as acidic due to :-

- (A) Acidic impurities    (B) Ionisation    ✓ (C) Hydrolysis of  $\text{Fe}^{3+}$  (D) Dissociation

$\text{FeCl}_3$  = Salt of weak base  $\text{Fe}(\text{OH})_3$  and strong acid HCl.

Here hydrolysis of  $\text{Fe}^{3+}$  ion will occur.



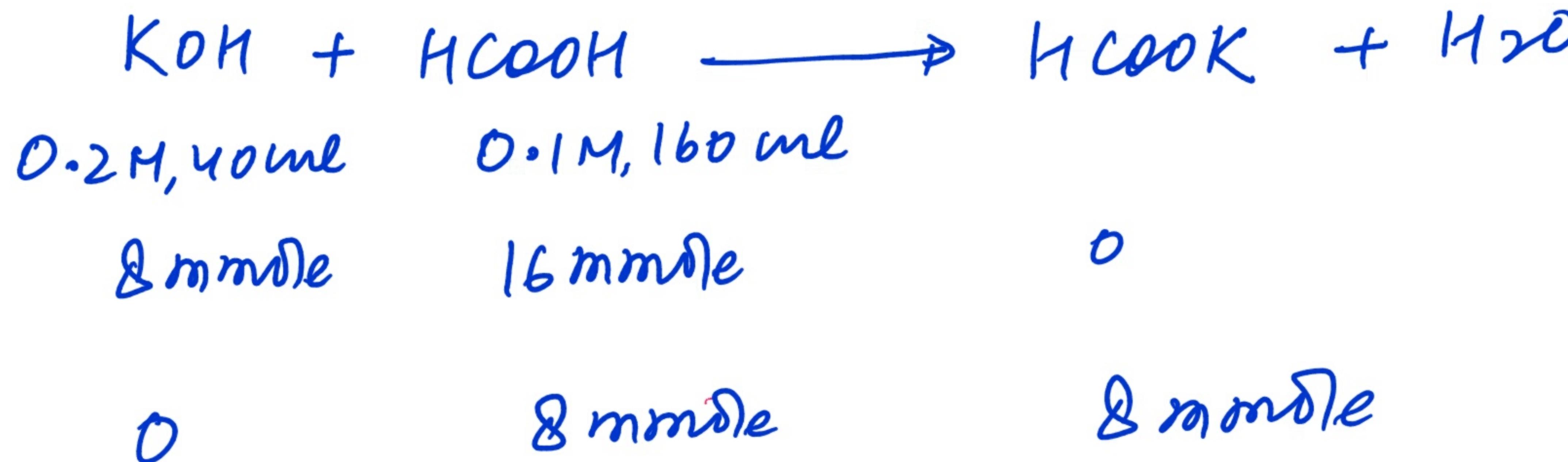
Q.37 If 40 ml of 0.2 M KOH is added to 160 ml of 0.1 M HCOOH [ $K_a = 2 \times 10^{-4}$ ], the pOH of the resulting solution is [log 2 = 0.3]

(A) 3.4

(B) 3.7

(C) 7

$$\xrightarrow{\text{PK}_a = 4 - \log 2 = 3.7} \quad \checkmark \quad \text{(D) } 10.3$$



mixture of weak acid HCOOH and salt HCOOK

$\Rightarrow$  Acidic buffer sol.

$$\text{pH} = \text{PK}_a + \log \frac{[\text{HCOOK}]}{[\text{HCOOH}]} = 3.7 + \log \frac{8}{8} = 3.7$$

$$\text{pOH} = 14 - 3.7 = \boxed{10.3} \quad \text{Ans}$$

**Q.38** 1 M NaCl and 1M HCl are present in an aqueous solution. The solution is

- (A) not a buffer solution and with pH < 7      (B) not a buffer solution with pH > 7  
 (C) a buffer solution with pH < 7      (D) a buffer solution with pH > 7

∴ Buffer sol<sup>n</sup> is a mixture of weak acid and its own salt. Since HCl is a strong acid, so above sol<sup>n</sup> is not a buffer sol<sup>n</sup>.

Q.39 The  $pK_a$  of a weak acid (HA) is 4.5. The  $pOH$  of an aqueous buffered solution of HA in which 50% of the acid is ionized is :

(A) 4.5

(B) 2.5

~~(C) 9.5~~

(D) 7.0

$$PH = PK_a + \log \frac{[A^-]}{[HA]}$$

$$\text{or, } PH = 4.5 + \log \frac{0.5}{0.5}$$

$$= 4.5$$

$$\therefore POH = 14 - 4.5 = \boxed{9.5} \text{ Ans ,}$$

Q.40 To a 50 ml. of 0.05M formic acid, how much volume of 0.10M sodium formate must be added to get a buffer solution of pH = 4.0 ?

(pK<sub>a</sub> of the acid is 3.7) (log2 = 0.3)

(A) 40 ml.

(B) 4 ml.

~~(C) 50 ml.~~

(D) 100 ml.

let volume of sodium formate required = V ml.

$$pH = pK_a + \log \frac{[HCOO^-]}{[HCOON]}$$

$$\text{or, } 4 = 3.7 + \log \frac{(0.1 \times V)}{(0.05 \times 50)}$$

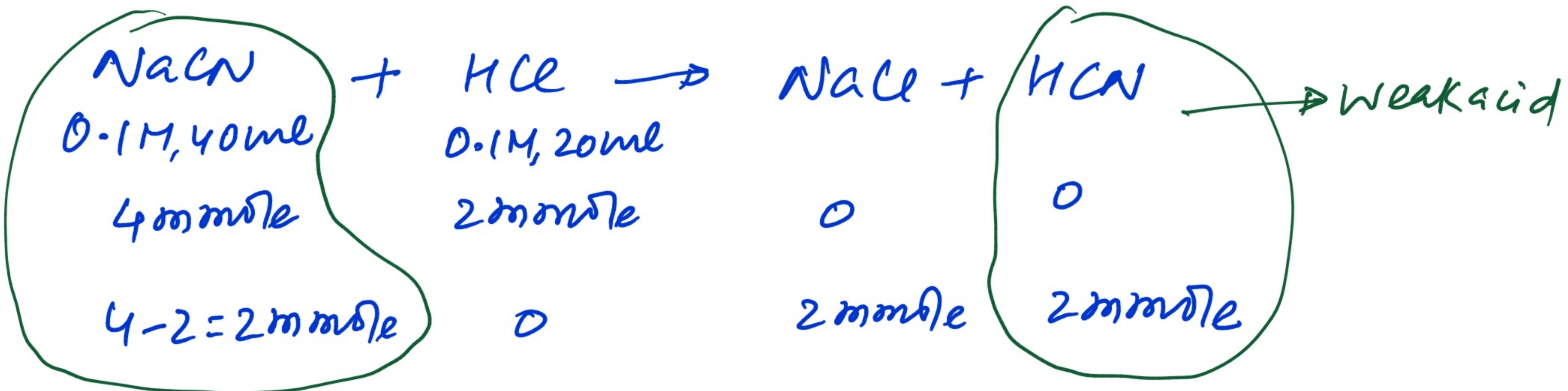
$$\text{or, } 0.3 = \log \frac{0.1V}{2.5}$$

$$\text{or, } \log 2 = \log \frac{0.1V}{2.5} \Rightarrow \frac{0.1V}{2.5} = 2 \Rightarrow V = \frac{5}{0.1} = \boxed{50 \text{ ml.}}$$

Ans.

Q.41 Which can act as buffer :-

- (A)  $\text{NH}_4\text{OH} + \text{NaOH} \rightarrow$  weak base + strong base
- (B)  $\text{HCOOH} + \text{HCl} \rightarrow$  weak acid + strong acid.
- (C) 40 ml. of 0.1 M NaCN + 20 ml. of 0.1 M HCl
- (D) All of them



→ salt of weak acid.

⇒ Acidic buffer soln.

Q.42 If equal volume of 0.05 M ammonium hydroxide solution is dissolved in 0.001 M ammonium chloride solution. What will be the  $\text{OH}^-$  ion concentration of this solution :  $\sqrt{32.8} = 5.73$

$$K_b(\text{NH}_4\text{OH}) = 1.8 \times 10^{-5}$$

(A)  $3.0 \times 10^{-3}$

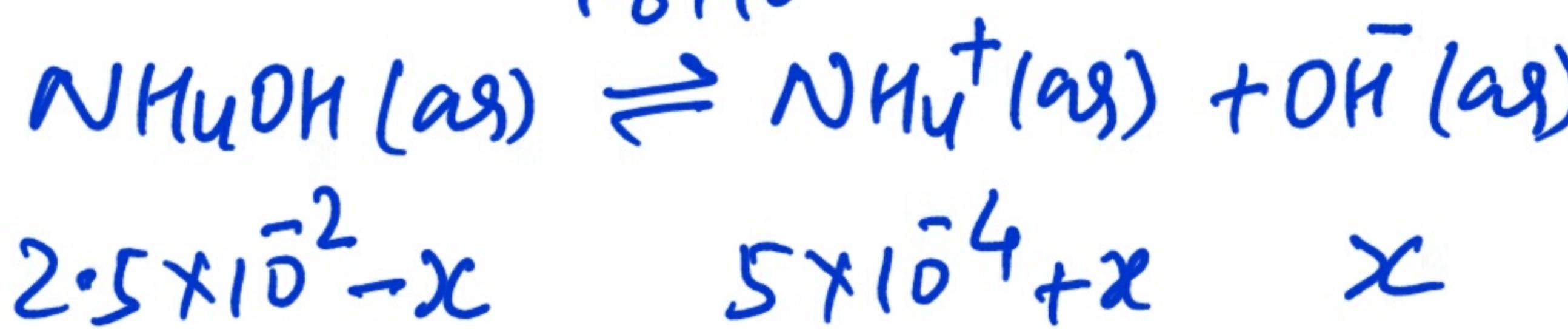
$\checkmark$  (B)  $4.66 \times 10^{-4}$

(C)  $9.0 \times 10^{-3}$

(D)  $9.0 \times 10^{-4}$

$$[\text{NH}_4\text{OH}] = \frac{0.05}{2} = 0.025 \text{ M} = 2.5 \times 10^{-2} \text{ M}$$

$$[\text{NH}_4\text{Cl}] = \frac{0.001}{2} = 0.0005 \text{ M} = 5 \times 10^{-4} \text{ M}$$



$$1.8 \times 10^{-5} = \frac{(5 \times 10^{-4} + x)x}{(2.5 \times 10^{-2} - x)}$$

$$\text{or}, 4.5 \times 10^{-7} - 1.8 \times 10^{-5}x = x^2 + 5 \times 10^{-4}x$$

$$\text{or}, x^2 + 5 \times 10^{-4}x + 1.8 \times 10^{-5}x - 4.5 \times 10^{-7} = 0$$

$$\left. \begin{aligned} & \text{or}, x^2 + 5.18 \times 10^{-4}x - 4.5 \times 10^{-7} = 0 \\ & x = \frac{-5.18 \times 10^{-4} + \sqrt{(5.18 \times 10^{-4})^2 + 18 \times 10^{-7}}}{2} \end{aligned} \right\} \therefore x > 0$$

$$\text{or}, x = \frac{-5.18 \times 10^{-4} + 14.38 \times 10^{-4}}{2}$$

$$\text{or}, x = 4.6 \times 10^{-4} \text{ M}$$

$$\boxed{\therefore [\text{OH}^-] = x = 4.6 \times 10^{-4} \text{ M}}$$

Q.43 Calculate the pH of a buffer prepared by mixing 600 cc of 0.6 M NH<sub>3</sub> and 400 cc of 0.5 M NH<sub>4</sub>Cl.

$$K_b \text{ for } \text{NH}_3 = 1.8 \times 10^{-5}, (\log 1.8 = 0.26)$$

(A) 11.3

(B) 9.0

✓ (C) 9.52

(D) 5

$$\text{NH}_3 = 0.6 \times 600 = 360 \text{ mmole}$$

$$\text{NH}_4\text{Cl} = 0.5 \times 400 = 200 \text{ mmole.}$$

$$\text{pOH} = \text{PK}_b + \log \frac{[\text{NH}_4^+]}{[\text{NH}_3]}$$

$$K_b = 1.8 \times 10^{-5}$$

$$\therefore \text{PK}_b = 5 - \log 1.8 \\ = 5 - 0.26 = 4.74$$

$$\text{or, } \text{pOH} = 4.74 + \log \frac{20\phi}{36\phi}$$

$$\text{or, } \text{pOH} = 4.74 - \log 1.8 = 4.74 - 0.26 = 4.48$$

$$\therefore \text{pH} = 14 - 4.48 = \boxed{9.52}$$

Aus-

Q.44  $pK_b$  for  $\text{NH}_4\text{OH}$  at certain temperature is 4.74. The pH of basic buffer containing equimolar concentration of  $\text{NH}_4\text{OH}$  and  $\text{NH}_4\text{Cl}$  will be:-

- (A) 7.74      (B) 4.74      (C) 2.37      ✓ (D) 9.26

$$PDH = PK_b + \log \frac{[NH_4^+]}{[NH_4OH]}$$

$$= 4.74 + \log \frac{1}{I}$$

$$= 4.74$$

$$PH = 14 - 4.74 = 9.26$$

Q.45 On addition of NaOH to  $\text{CH}_3\text{COOH}$  solution, 60% of the acid is neutralised. If  $pK_a$  of  $\text{CH}_3\text{COOH}$  is 4.7 then the pH of the resulting solution is :-

- (A) More than 4.7 but less than 5.0
  - (B) Less than 4.7 but more than 4.0
  - (C) More than 5.0
  - (D) Remains unchanged



$$\begin{aligned}
 \text{pH} &= \text{p}K_a + \log \frac{[\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} \\
 &= 4.7 + \log \frac{0.6}{0.4} = 4.7 + \log \frac{3}{2} = 4.7 + 0.18 \\
 &= 4.88
 \end{aligned}$$

Q.46 Henderson equation  $\text{pH} - \text{pK}_a = 5$  will be applicable to an acidic buffer when :-

- (A) [Acid] = [Conjugate base]      ✓ (B) [Acid]  $\times 10^5$  = [Conjugate base]  
(C) [Acid] = [Conjugate base]  $\times 10^5$       (D) [acid] = 2 [conjugate base]

Given,

$$\text{pH} - \text{pK}_a = 5$$

$$\Rightarrow \text{pH} = \text{pK}_a + 5 \quad \textcircled{1}$$

$$\therefore \text{pH} = \text{pK}_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]}$$

$$\text{or, } \text{pH} + 5 = \text{pK}_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]}$$

$$\Rightarrow \frac{[\text{conjugate base}]}{[\text{acid}]} = 10^5 \Rightarrow [\text{acid}] \times 10^5 = [\text{conjugate base}]$$

Q.47 What amount of sodium propanoate should be added to one litre of an aqueous solution containing 0.02 mole of propanoic acid ( $K_a = 3 \times 10^{-5}$  at  $25^\circ\text{C}$ ) to obtain a buffer solution of pH 4.7  
[ $\log 2 = 0.3$ ]

- (A)  $4.52 \times 10^{-2}$  mol (B)  $3.52 \times 10^{-2}$  mol (C)  $2.52 \times 10^{-2}$  mol (D)  $3 \times 10^{-2}$  mol

Let moles of sodium propanoate required =  $x$  in 1 L soln.



$$\text{PH} = 4.7$$

$$K_a = \frac{[\text{C}_2\text{H}_5\text{COO}^-] \times [\text{H}^+]}{[\text{C}_2\text{H}_5\text{COOH}]}$$

$$[\text{H}^+] = 2 \times 10^{-5}$$

$$\text{or, } 3 \times 10^{-5} = \frac{x \times 2 \times 10^{-5}}{0.02}$$

$$\text{or, } x = 3 \times 0.01 = 3 \times 10^{-2} \text{ mol}$$

Q.48 In a buffer solution the ratio of concentration of  $\text{NH}_4\text{Cl}$  and  $\text{NH}_4\text{OH}$  is 1 : 1. When it changes in 2 : 1, what will be the value of pH of buffer ?

- (A) Increase       (B) Decrease      (C) No effect      (D) None

$$POH_1 = PK_b + \log \frac{1}{1} = PK_b .$$

$$POH_2 = PK_b + \log \frac{2}{1} = PB + \log 2$$

$\Rightarrow POH_2 > POH_1 \Rightarrow POH$  increases

$\Rightarrow PH$  decreases.

Q.49 The buffer solution play an important role in :-

- (A) Increasing the pH value
- (B) Decreasing the pH value
- (C) Keeping the pH constant
- (D) Solution will be neutral

Buffer solution resist change in pH.

Q.50 The total number of different kind of acidic buffers obtained during the titration of  $\text{H}_3\text{PO}_4$  with NaOH are :

(A) 3

(B) 1

(C) 2

(D) 0

following three reactions will occurs :

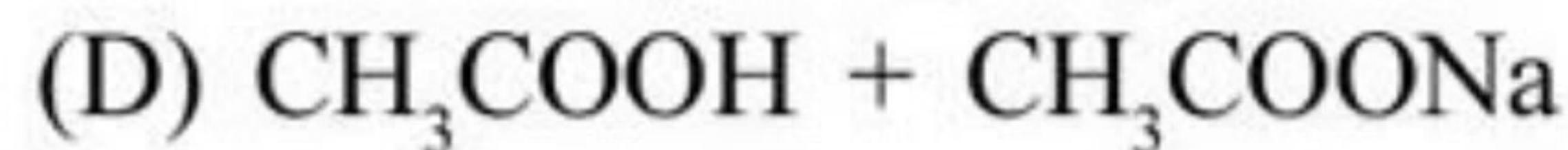
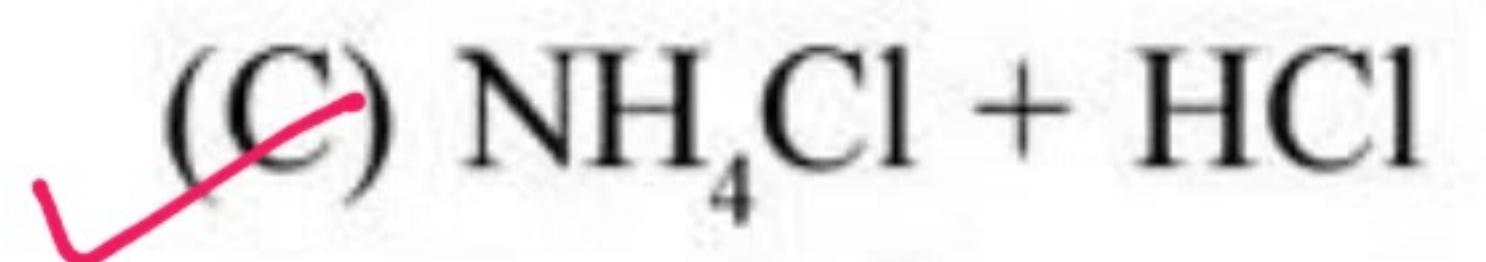
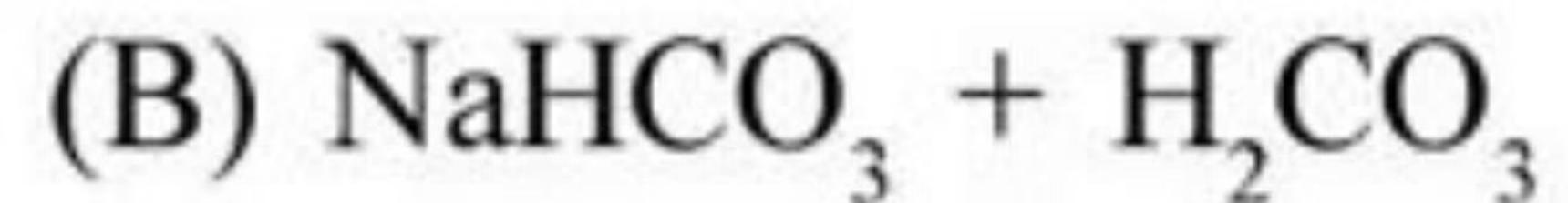
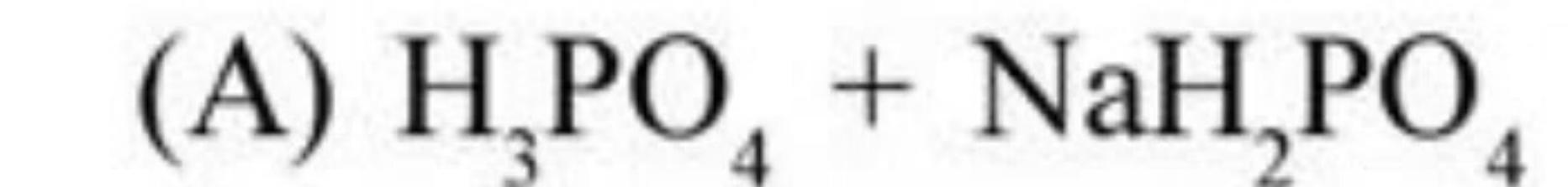


Therefore buffer solns will be:

- i)  $\text{H}_3\text{PO}_4 + \text{NaH}_2\text{PO}_4$
- ii)  $\text{NaH}_2\text{PO}_4 + \text{Na}_2\text{HPO}_4$
- iii)  $\text{Na}_2\text{HPO}_4 + \text{Na}_3\text{PO}_4 .$

no. of buffer solutions = 3.

Q.51 Which of the following solutions does not act as buffer :



$(\text{NH}_4\text{Cl} + \text{HCl})$  is a mixture of strong acid and its salt  
 $\Rightarrow$  not a buffer sol<sup>n</sup>.

Others are mixture of weak acid and its own salt  
with a strong base.

Q.52 Half of the formic acid solution is neutralised on addition of a KOH solution to it. If  $K_a(\text{HCOOH}) = 2 \times 10^{-4}$  then pH of the solution is : - ( $\log 2 = 0.3010$ )

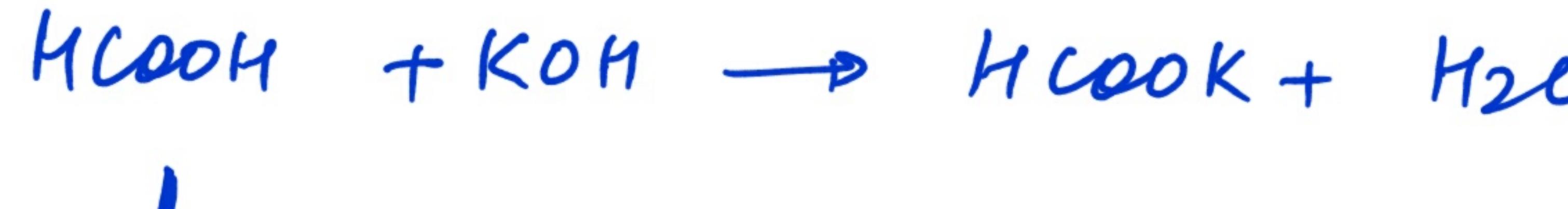
(A) 3.6990

(B) 10.3010

(C) 3.85

(D) 4.3010

$\rightarrow \text{PK}_a = 4 - \log 2 = 3.699$



$$1 - \frac{1}{2} = \frac{1}{2}$$

$\frac{1}{2}$   $\Rightarrow$  Acidic buffer soln.

$$\text{PH} = \text{PK}_a + \log \frac{[\text{HCOO}^-]}{[\text{HCOOH}]} = 3.699 + \log \frac{Y_2}{Y_1} = \boxed{3.699}.$$

Ans.

Q.53 When 0.02 moles of NaOH are added to a litre of buffer solution, its pH changes from 5.75 to 5.80.

What is its buffer capacity :-

✓ (A) 0.4

(B) 0.05

(C) - 0.05

(D) 2.5

$$BC = \frac{db}{d(\text{pH})} = \frac{x}{\Delta(\text{pH})} = \frac{0.02}{5.8 - 5.75} = \frac{0.02}{0.05} = \boxed{0.4}$$

Ans.

Q.54 Calculate pH when 100 ml of 0.2M NaOH is reacted with 100 ml of 0.2 M CH<sub>3</sub>COOH

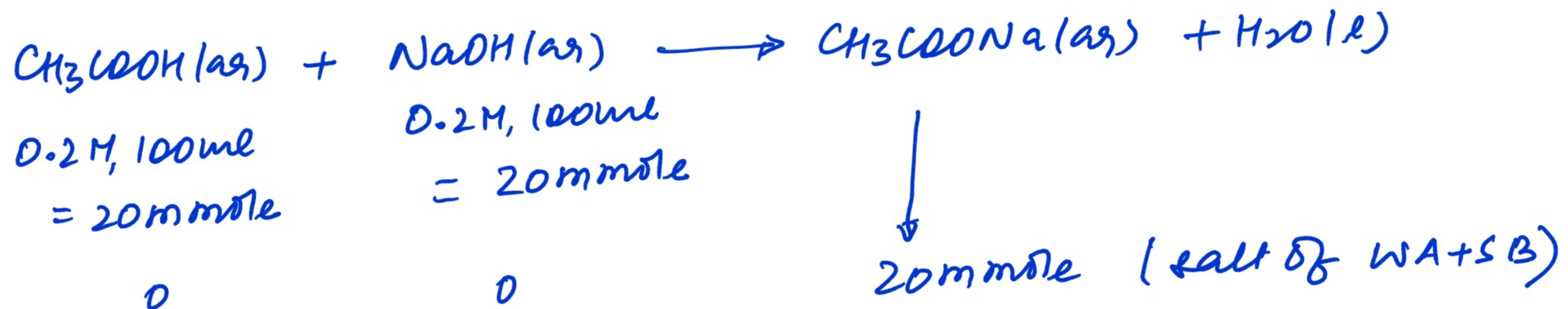
$$(K_a = 10^{-5}) \Rightarrow PK_a = 5$$

(A) 9

(B) 7

(C) 5

(D) 2



$$\therefore C = [\text{CH}_3\text{COO}^-] = \frac{20}{(100+100)} = \frac{20}{200} = 0.1\text{M}.$$

$$\frac{K_h}{C} < 10^{-2}, \quad \text{pH} = 7 + \frac{1}{2} (\text{PK}_a + \log C) = 7 + \frac{1}{2} (5 + \log 0.1) = 7 + \frac{1}{2} (5-1) = \boxed{9}$$

Ans.

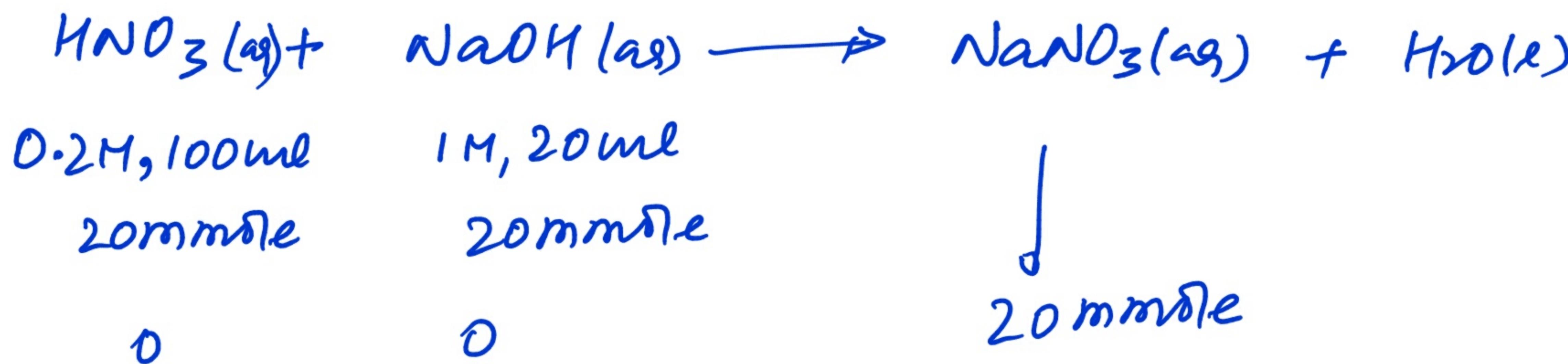
Q.55 At  $90^{\circ}\text{C}$ , pure water has  $[\text{H}^+] = 10^{-6} \text{ M}$ , if 100 ml of 0.2 M  $\text{HNO}_3$  is added to 20 ml of 1 M  $\text{NaOH}$  at  $90^{\circ}\text{C}$  then pH of the resulting solution will be

(A) 5

~~(B) 6~~

(C) 7

(D) None of these



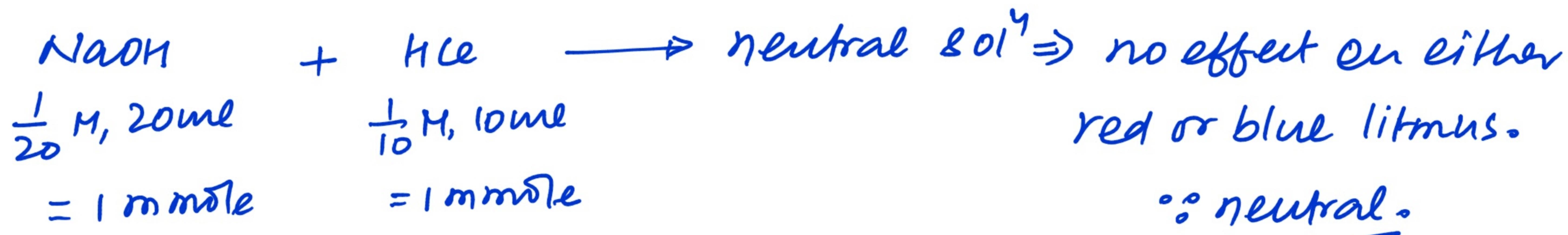
Salt of strong base and strong acid  $\Rightarrow$  neutral sol $^{\text{y}}$

$$\therefore \boxed{\text{PH}=6}$$

$\therefore$  At  $90^{\circ}\text{C}$ , neutral sol $^{\text{y}}$  has  $\text{PH}=6$ .

Q.56 When 20 ml of  $\frac{M}{20}$  NaOH are added to 10 ml of  $\frac{M}{10}$  HCl, the resulting solution will:-

- (A) Turn blue litmus red  
(B) Turn phenolphthalein solution pink colour  
(C) Turn methyl orange red  
(D) Will have no effect on either red or blue litmus



	Colour in acidic region	Colour in basic region.
litmus	red	blue
Phenolphthalein	colourless	pink
methyl orange	pink (red)	yellow

Q.57 The rapid change of pH near the stoichiometric point of an acid–base titration is the basis of indicator detection. pH of the solution is related to ratio of the concentrations of the conjugate acid ( $\text{HIn}$ ) and base ( $\text{In}^-$ ) forms of the indicator by the expression :-

(A)  $\log \frac{[\text{HIn}]}{[\text{In}^-]} = \text{pK}_{\text{In}} - \text{pH}$

(B)  $\log \frac{[\text{HIn}]}{[\text{In}^-]} = \text{pH} - \text{pK}_{\text{In}}$

(C)  $\log \frac{[\text{In}^-]}{[\text{HIn}]} = \text{pH} + \text{pK}_{\text{In}}$

(D)  $\log \frac{[\text{In}^-]}{[\text{HIn}]} = \text{pK}_{\text{In}} - \text{pH}$

$$\text{pH} = \text{pK}_{\text{In}} + \log \frac{[\text{In}^-]}{[\text{HIn}]}$$

∴  $\log \frac{[\text{In}^-]}{[\text{HIn}]} = \text{pH} - \text{pK}_{\text{In}}$

or,  $\log \frac{[\text{HIn}]}{[\text{In}^-]} = \text{pK}_{\text{In}} - \text{pH}.$

Q.58 Calculate the pH range in which an acid indicator with  $K_{acid}$  (indicator) =  $1.0 \times 10^{-5}$  changes colour when the concentration of the indicator is  $1 \times 10^{-3} M$ .

$$\xrightarrow{\text{PK}_H = 5}$$

- (A)  $5 \pm 1$       (B)  $11 \pm 1$       (C)  $3 \pm 1$       (D)  $8 \pm 1$

pH range of indicator =  $\text{PK}_H - 1$  to  $\text{PK}_H + 1$   
=  $5 - 1$  to  $5 + 1$

$$= 4 \text{ to } 6$$

i.e.,  $5 \pm 1$

Q.59 In what pH range will a  $1 \times 10^{-4}$  M solution of an indicator with  $K_b$  (indicator) =  $1 \times 10^{-11}$  change colour?

(A)  $7.0 \pm 1$

(B)  $3.0 \pm 1$

(C)  $5.5 \pm 1$

(D)  $11.0 \pm 1$

$$\xrightarrow{\text{PK}_b} \text{PK}_b = 11$$

$$\begin{aligned}\text{pOH range of the indicator} &= \text{PK}_b - 1 \text{ to } \text{PK}_b + 1 \\ &\approx 11 - 1 \text{ to } 11 + 1 \\ &= 10 \text{ to } 12\end{aligned}$$

$$\therefore \text{pH range} = 2 \text{ to } 4 \\ = 3 \pm 1$$

Q.60 Indicator which is used in the titration of  $\text{CH}_3\text{COOH}$  &  $\text{NaOH}$  :-

- (A) Methyl orange      (B) Methyl red      (C) Phenolphthalein      (D) Litmus

At equivalence point  $\text{pH} > 7$  ( $\because \text{WA} + \text{SB}$ )

pH range of phenolphthalein = 8.2 to 10.

Q.61 Phenolphthalein is a :-

- (A) Strong acid      (B) Strong base      (C) Weak base      (D) Weak acid



Q.62 pH-range of Methyl red indicator is :-

(A) 4.4–6.5

(B) 6.8–10.8

(C) 8 – 9.6

(D) 6.8 – 8.2

Q.63 In the volumetric estimation of HCl, if we make use of phenolphthalein as an indicator, which base is unsuitable for the titration :-

- (A) NaOH                    (B) RbOH                    (C) KOH                     (D) NH<sub>4</sub>OH

PH range of Phenolphthalein = 8.2 to 10

∴ we must take a strong base to neutralise HCl.

**Q.64** Phenolphthalein does not act as an indicator for the titration between :-

- (A) KOH and  $\text{H}_2\text{SO}_4$       (B) NaOH and  $\text{CH}_3\text{COOH}$   
~~(C) Oxalic acid and  $\text{KMnO}_4$~~       (D)  $\text{Ba}(\text{OH})_2$  and HCl

Phenolphthalein is used in titration of

- i) strong acid + strong base
  - ii) weak acid + strong base.

get can't be used in a Redox titration.

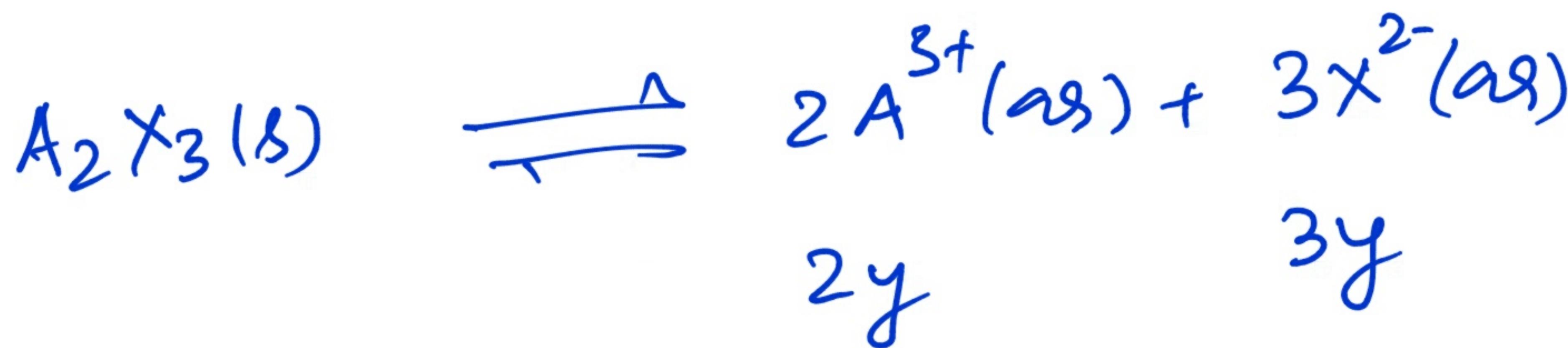
Q.65 For weak acid and strong base titration, the indicator used is :-

- (A) Potassium di-chromate  
(C) Litmus

- (B) Methyl orange  
✓ (D) Phenolphthalein

Q.66 The solubility of  $A_2X_3$  is  $y \text{ mol dm}^{-3}$ . Its solubility product is

- (A)  $6y^2$       (B)  $64y^4$       (C)  $36y^5$       (D)  $108y^5$



$$\begin{aligned} K_{sp} &= [A^{3+}]^2 [X^{2-}]^3 = (2y)^2 \times (3y)^3 \\ &= 4 \times 27 \times y^5 = \boxed{108y^5} \end{aligned}$$

Ans.

- Q.67 If  $K_{sp}$  for  $\text{HgSO}_4$  is  $6.4 \times 10^{-5}$ , then solubility of this substance in mole per  $\text{m}^3$  is

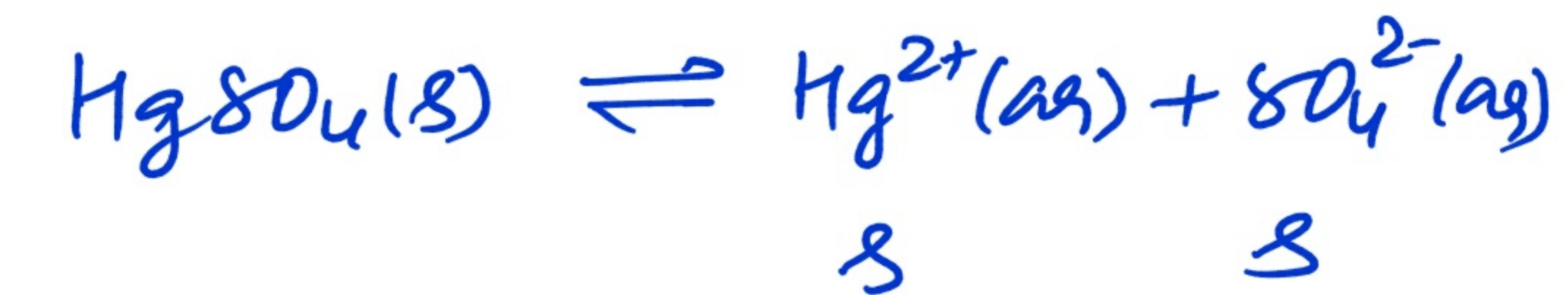
(A)  $8 \times 10^{-3}$

(B)  $6.4 \times 10^{-5}$

(C)  $8 \times 10^{-6}$

$\checkmark$  (D) 8

$$K_{sp} = [\text{Hg}^{2+}][\text{SO}_4^{2-}]^3$$



$$6.4 \times 10^{-5} = s \times s = s^2$$

$$\text{or, } s = \sqrt{6.4 \times 10^{-5}} = 8 \times 10^{-3} \text{ mole/L}$$

$$= 8 \times 10^{-3} \text{ mole} / 10^3 \text{ meter}^3.$$

$$= \boxed{8 \text{ mole/meter}^3}.$$

Ans.

Q.68 If the solubility of AgCl (formula mass=143) in water at 25°C is  $1.43 \times 10^{-4}$  gm/100 ml of solution then the value of  $K_{sp}$  will be :-

(A)  $1 \times 10^{-5}$

(B)  $2 \times 10^{-5}$

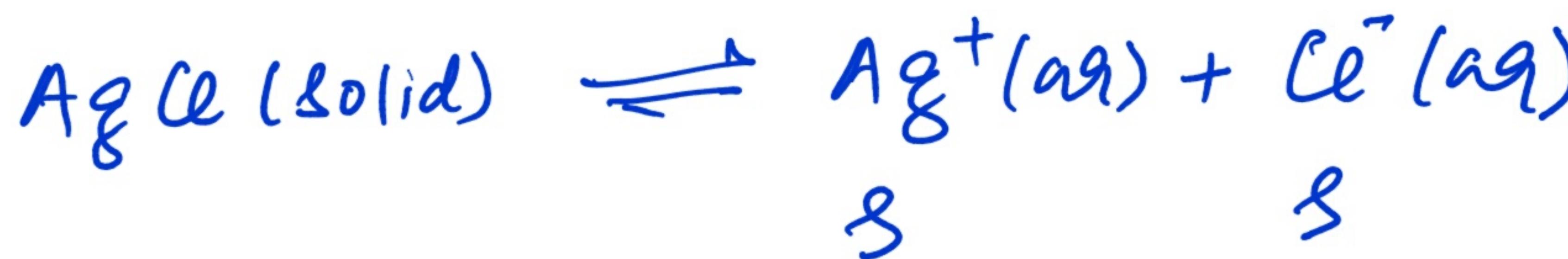
~~(C)  $1 \times 10^{-10}$~~

(D)  $2 \times 10^{-10}$

$$s = 1.43 \times 10^{-4} \text{ g/100ml} \quad M_{AgCl} = 143$$

$$\text{or, } s = \frac{1.43 \times 10^{-4}}{143} \times \frac{1000}{100} \text{ mole/L}$$

$$= 10^{-5} \text{ mole/L}$$



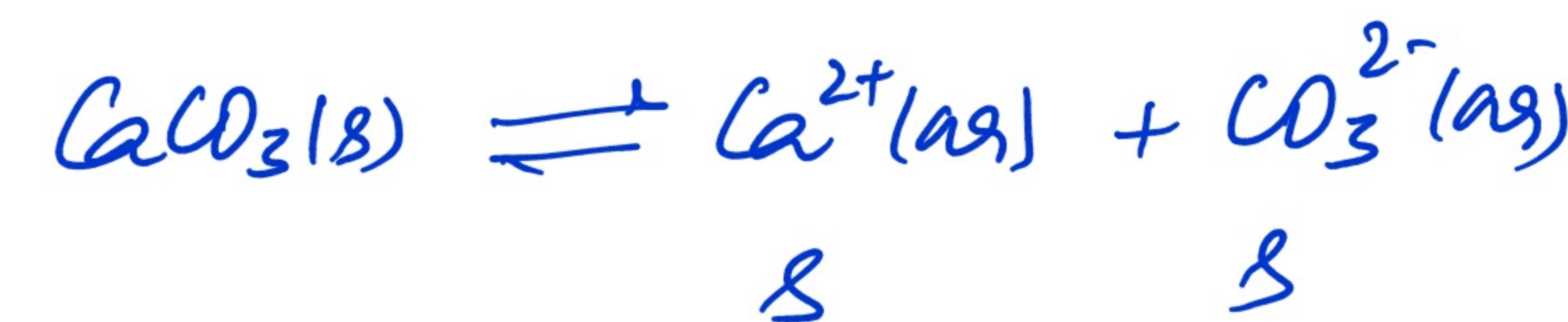
$$K_{sp} = [Ag^+][Cl^-] = s \times s = s^2 = (10^{-5})^2 = \boxed{10^{-10}} \text{ Ans.}$$

Q.69 One litre of saturated solution of  $\text{CaCO}_3$  is evaporated to dryness, 7.0 g of residue is left. The solubility product for  $\text{CaCO}_3$  is:-

- (A)  $4.9 \times 10^{-3}$       (B)  $4.9 \times 10^{-5}$       (C)  $4.9 \times 10^{-9}$       (D)  $4.9 \times 10^{-7}$

$$M_{\text{CaCO}_3} = 100 \text{ g.}$$

$$s = \frac{7}{100} \text{ mole/L}$$



$$K_{sp} = s^2 = \left(\frac{7}{100}\right)^2 = 49 \times 10^{-4}$$
$$= \boxed{4.9 \times 10^{-3}}$$

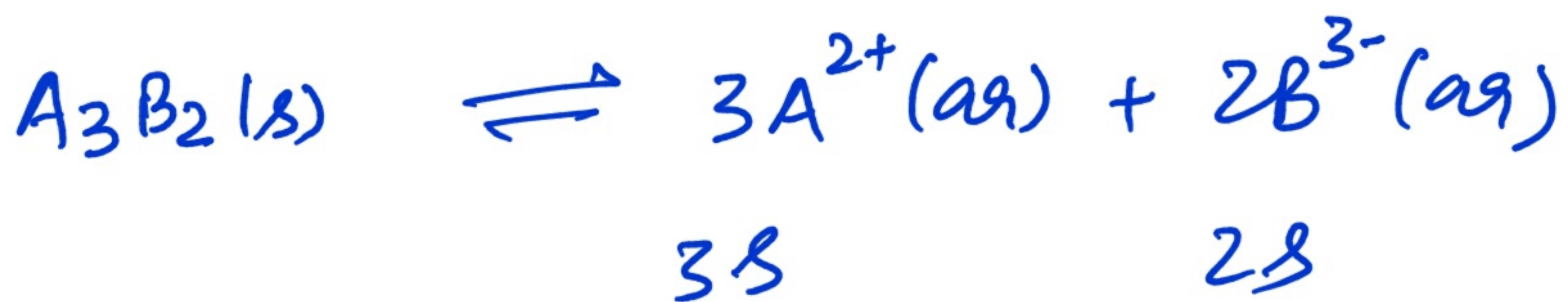
Q.70  $A_3B_2$  is a sparingly soluble salt of molar mass M ( $\text{g mol}^{-1}$ ) and solubility x g  $\text{lit}^{-1}$ . The ratio of the molar concentration of  $B^{3-}$  to the solubility product of the salt is

(A)  $108 \frac{x^5}{M^5}$

(B)  $\frac{1}{108} \frac{M^4}{x^4}$

*(C)*  $\frac{1}{54} \frac{M^4}{x^4}$

(D) None



Molar solubility 's' =  $\frac{x}{M}$ .

$$K_{sp} = (3s)^3 \times [B^{3-}]^2$$

$\therefore \frac{1}{27s^3} = \frac{[B^{3-}]^2}{K_{sp}}$

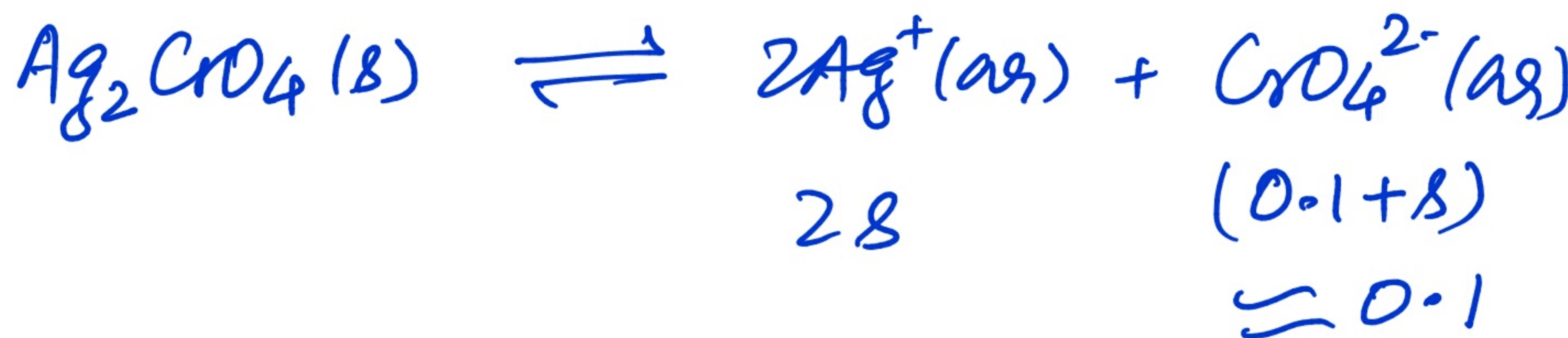
$\therefore \frac{1}{27s^3 \times [B^{3-}]} = \frac{[B^{3-}]}{K_{sp}}$

$\therefore \frac{[B^{3-}]}{K_{sp}} = \frac{1}{27s^3 \times 2s} = \frac{1}{54s^4} =$

$\frac{1}{54} \times \frac{1}{\left(\frac{x}{M}\right)^4} = \frac{\frac{M^4}{54 \times x^4}}{\text{Ans.}}$

Q.71. Solubility of  $\text{Ag}_2\text{CrO}_4$  ( $K_{sp} = 4 \times 10^{-13}$ ) in 0.1 M  $\text{K}_2\text{CrO}_4$  solution will be :-

- (A)  $10^{-3}$  M      (B)  $10^{-6}$  M      (C)  $4 \times 10^{-6}$  M      (D)  $5 \times 10^{-7}$  M



$$K_{sp} = (2s)^2 \times 0.1$$

$$\text{or, } 4 \times 10^{-13} = 4s^2 \times 0.1$$

$$\text{or, } s^2 = \frac{10^{-13}}{0.1} = 10^{-12}$$

$$\text{or, } s = 10^{-6} \text{ mole/L}$$

Aus

Q.72. How many times solubility of  $\text{CaF}_2$  is decreased in  $4 \times 10^{-3} \text{ M}$  KF (aq.) solution as compare to pure water at  $25^\circ\text{C}$ . Given  $K_{\text{sp}}(\text{CaF}_2) = 3.2 \times 10^{-11}$

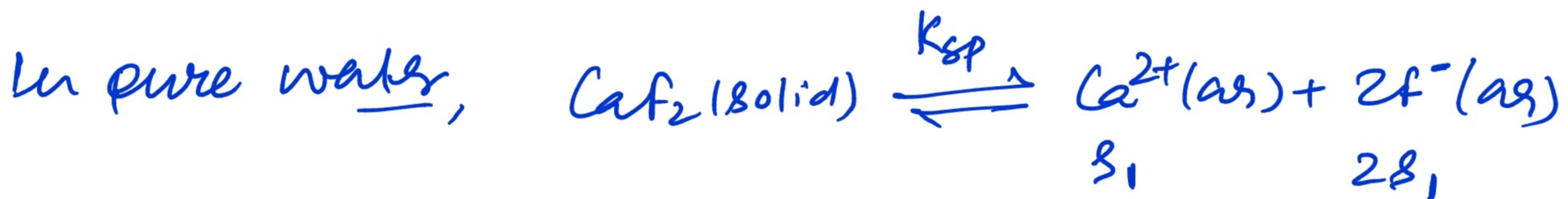
(A) 50

(B) 100

(C) 500

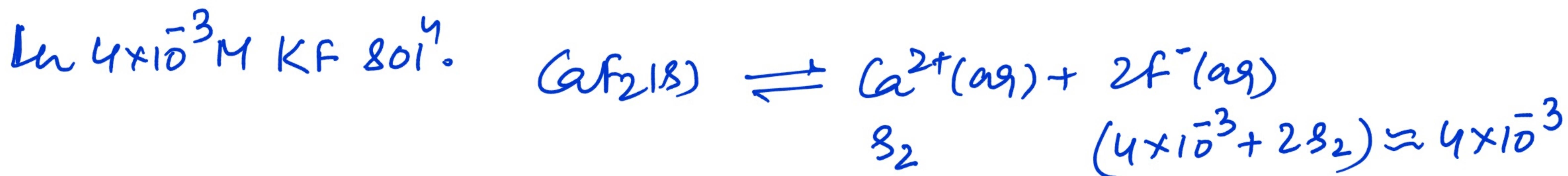
(D) 1000

let solubility in pure water =  $s_1$ , and in KF solution is  $s_2$ .



$$K_{\text{sp}} = s_1 \times (2s_1)^2 \propto, 3.2 \times 10^{-11} = 4 \times s_1^3 \Rightarrow 8 \times 10^{-12} = s_1^3$$

$$\Rightarrow s_1 = 2 \times 10^{-4} \text{ mole/L}.$$



$$K_{\text{sp}} = s_2 \times (4 \times 10^{-3})^2$$

$$\propto, 3.2 \times 10^{-11} = s_2 \times 16 \times 10^{-6} \Rightarrow s_2 = 2 \times 10^{-6} \text{ mole/L} \therefore \frac{s_1}{s_2} = \frac{2 \times 10^{-4}}{2 \times 10^{-6}}$$

$$= \boxed{100} \text{ Ans}$$

Q.73 At  $30^{\circ}\text{C}$ , In which of the one litre solution, the solubility of  $\text{Ag}_2\text{CO}_3$  (solubility product =  $8 \times 10^{-12}$ ) will be maximum :-

- (A) 0.05 M  $\text{Na}_2\text{CO}_3$     (B) Pure water    (C) 0.05 M  $\text{AgNO}_3$     (D) 0.05 M  $\text{NH}_3$

↓  
Common ion  $\text{CO}_3^{2-}$ , so solubility decreases.

↓  
Common ion  $\text{Ag}^+$ , so  
solubility decreases.

In  $\text{NH}_3(\text{aq})$  soln,  $\text{Ag}^+$  form complex  $[\text{Ag}(\text{NH}_3)_2]^+$ . So  
solubility will be maximum in  $\text{NH}_3(\text{aq})$ .

Q.74 What will happen if the pH of the solution of 0.001 M  $\text{Mg}(\text{NO}_3)_2$  solution is adjusted to pH = 9 ( $K_{\text{sp}}$  of  $\text{Mg(OH)}_2$  =  $8.9 \times 10^{-12}$ )

- (A) ppt will take place  
(C) Solution will be saturated

- (B) ppt will not take place  
(D) None of these

$$\text{pH} = 9 \Rightarrow [\text{H}^+] = 10^{-9} \Rightarrow [\text{OH}^-] = \frac{K_{\text{sp}}}{[\text{H}^+]} = \frac{10^{-14}}{10^{-9}} = 10^{-5} \text{ mole/L}$$

$$[\text{Mg}^{2+}] = 0.001 \text{ M} = 10^{-3} \text{ M}.$$



$$\text{ionic product} = [\text{Mg}^{2+}][\text{OH}^-]^2 = 10^{-3} \times (10^{-5})^2 = 10^{-13} < K_{\text{sp}} \text{ of } \text{Mg(OH)}_2$$

$\Rightarrow \underline{\text{no ppt.}}$

- Q.75  $\text{Na}_3\text{PO}_4$  which should be added in 10 L of  $1.0 \times 10^{-5} \text{ M}$  -  $\text{BaCl}_2$  solution without any precipitation of  $\text{Ba}_3(\text{PO}_4)_2$  is [Ksp of  $\text{Ba}_3(\text{PO}_4)_2$ ] =  $4 \times 10^{-23}$
- (A)  $2 \times 10^{-4} \text{ gm}$       (B)  $0.328 \text{ gm}$       (C)  $0.164 \text{ gm}$       (D)  $0.82 \text{ gm}$

Let  $x$  mole  $\text{Na}_3\text{PO}_4$  can be added in 10L  $10^{-5} \text{ M}$  -  $\text{BaCl}_2$  soln without any precipitation.

$$\therefore [\text{Na}_3\text{PO}_4] = [\text{PO}_4^{3-}] = \frac{x}{10} \text{ M.}, [\text{Ba}^{2+}] = 10^{-5} \text{ M}, M_{\text{Na}_3\text{PO}_4} = 164$$



$$K_{\text{sp}} = [\text{Ba}^{2+}]^3 [\text{PO}_4^{3-}]^2$$

$$\text{or, } 4 \times 10^{-23} = (10^{-5})^3 \times \left(\frac{x}{10}\right)^2$$

$$\text{or, } x^2 = \frac{4 \times 10^{-23}}{10^{-17}} = 4 \times 10^{-6}$$

$$\therefore x = 2 \times 10^{-3} \text{ mole}$$

$$= 2 \times 10^{-3} \times 164 = \boxed{0.328 \text{ gram.}}$$

Ans.