

**Strong Electrolytes**

✓ SA  
✓  $\text{H}_2\text{SO}_4$ ,  $\text{HClO}_4$ ,  $\text{HClO}_3$   
✓  $\text{HNO}_3$ ,  $\text{HCl}$ ,  $\text{HBr}$ ,  $\text{HI}$   
All type of salts

✓ SB

All alkali metal hydroxides and  $\text{Ba}(\text{OH})_2$



## Weak Electrolytes

WA  
H<sub>2</sub>SO<sub>3</sub>, HClO<sub>2</sub>, HClO  
HNO<sub>2</sub>, All oxy acids of P, All  
organic carboxylic acid, HCN  
H<sub>3</sub>BO<sub>3</sub>, H<sub>2</sub>CO<sub>3</sub>, HF etc

WB

All alkaline earth metal hydroxides  
Except Ba(OH)<sub>2</sub>, Al(OH)<sub>3</sub>, All organic  
bases, NH<sub>3</sub>, d block metal hydroxides

## Arrhenius Acid Base theory

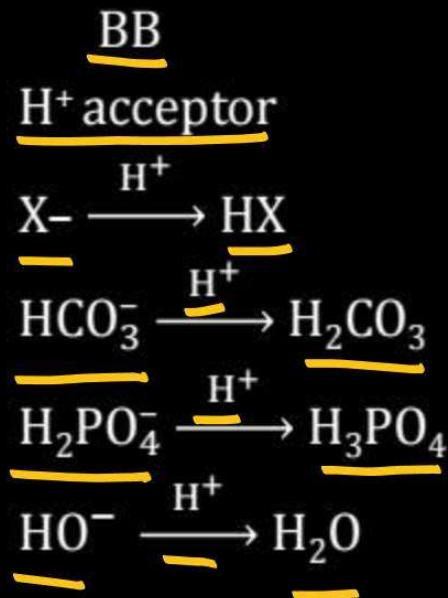
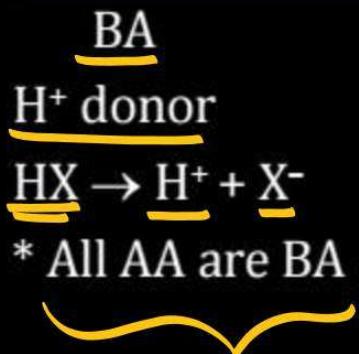
AA ✓

Donate H<sup>+</sup> ion in aq Sol<sup>n</sup>  
HX, H<sub>2</sub>SO<sub>4</sub>, H<sub>3</sub>PO<sub>4</sub>  
HClO<sub>4</sub>, HCN etc.

✓ AB

OH donor in aq Sol<sup>n</sup>  
MOH, M(OH)<sub>2</sub>, M(OH)<sub>3</sub>

## Bronsted Acid Base theory



## Lewis Acid Base theory

L. A

L. P acceptor

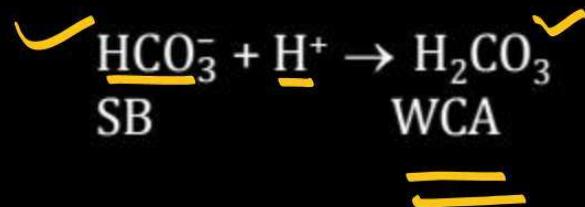
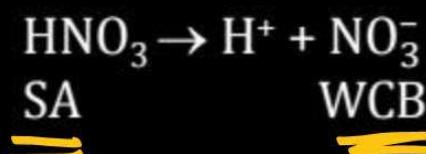
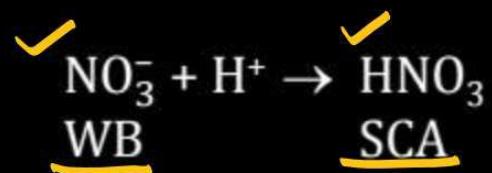
H<sup>+</sup>, CO<sub>2</sub>, SF<sub>4</sub>, PF<sub>5</sub>, BCl<sub>3</sub>, AlCl<sub>3</sub>  
FeCl<sub>3</sub> etc.

L. B

L. P donor

X<sup>-</sup>, -OH, NH<sub>3</sub>, H<sub>2</sub>O etc.

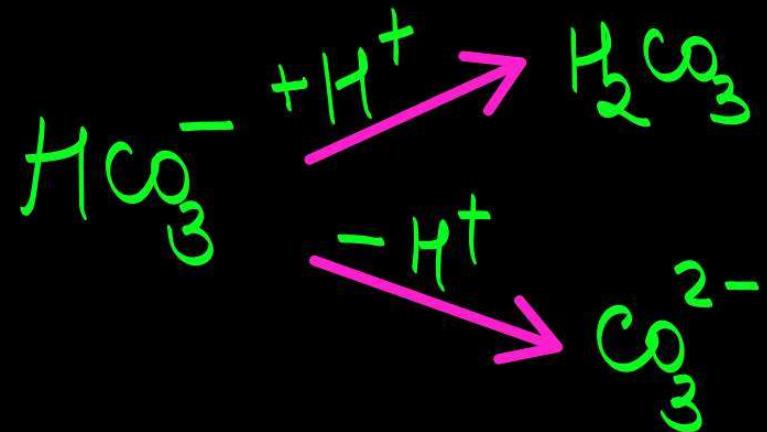
## **Conjugate Acid Base theory**



## Amphiprotic species

H<sup>+</sup> donor as well as acceptor

Bronsted acid as well as bronsted base



## pH Calculation

$$pH = -\log[H^+]$$

$$pOH = -\log[OH^-]$$

$$pH + pOH = 14 \text{ at } 25^\circ C$$

$$pH + pOH = 12 \text{ at } 90^\circ C$$

$T \uparrow K_w \uparrow$

**Case 1** Strong acid / Base

① pH of  $10^{-8} M HCl$   $[H^+] < 10^{-6} M$

$$[H^+]_{\text{net}} = [H^+]_{\text{from acid}} + [H^+]_{\text{from } H_2O}$$

$$[H^+]_{\text{net}} = 10^{-8} + 10^{-7}$$

$$pH = -\log [H^+]_{\text{new}} = 6.9$$



$$\textcircled{2} \quad [H^+] \geq 10^{-6} \text{ M}$$

find  $pH$  of  $10^{-4} \text{ M HCl}$

$$[H^+] = 10^{-4}$$

$$pH = -\log 10^{-4}$$

$$pH = 4$$

## Case 2

pH of mix of two SA & SB  
1)  $pH + pOH = 14$   
2) for conju acid base pair  $pK_a + pK_b = 14$



① Two SA

$$[H^+]_{\text{mix}} = \frac{M_1 V_1 n_f_1 + M_2 V_2 n_f_2}{V_1 + V_2}$$

$$pH = -\log [H^+]$$

② Two strong Bases

$$[\bar{OH}]_{\text{mix}} = \frac{M_1 V_1 n_f_1 + M_2 V_2 n_f_2}{V_1 + V_2}$$

$$pOH = -\log [\bar{OH}]$$

$$pH = 14 - pOH$$

③ one SA + one SB

$$[\text{H}^+] / [\text{OH}^-] = \frac{|M_1 V_1 n f_1 - M_2 V_2 n f_2|}{V_1 + V_2}$$

Q. Find pH



$$\begin{aligned}
 [\text{H}^+]_{\text{mix}} &= \frac{\frac{1}{5} \times 50 \times 1 + \frac{1}{10} \times 50 \times 2}{100} \\
 &= \frac{10 + 10}{100} \\
 &= \frac{20}{100} \\
 &= \frac{1}{5}
 \end{aligned}$$

$$pH = -\log \frac{1}{5}$$

$$pH = +\log 5$$

$$pH = 0.7$$

Q. Find pH  
 75 mL of  $\frac{M}{2}$  Ba(OH)<sub>2</sub> + 25 mL of  $\frac{M}{1}$  HCl

$$\frac{[H^+]/[\bar{OH}]}{100} = \frac{\left| \frac{1}{2} \times 75 - 1 \times 25 \right|}{100}$$

$$[\bar{OH}] = \frac{|75 - 25|}{100}$$

$$[\bar{OH}] = \frac{1}{2}$$

$$P^{OH} = -\log \frac{1}{2}$$

$$P^{OH} = +\log 2$$

$$P^{OH} = 0.3$$

$$P^H = 14 - 0.3$$

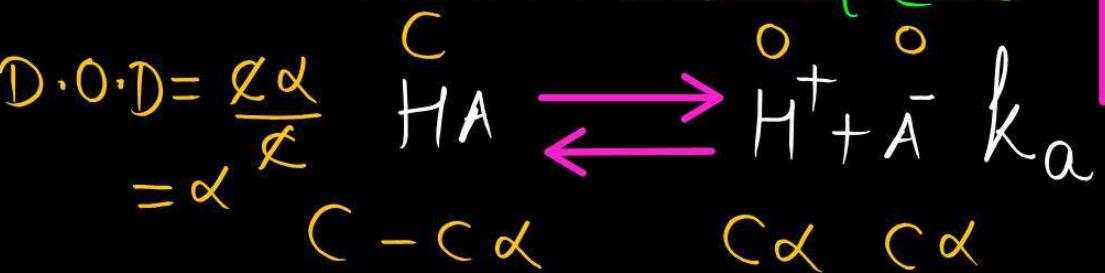
$$= 13.7$$

(P  
W)

**Case 3** pH of weak mono basic acid or weak mono acidic base



Ostwald dilution Law



BOH

$$\alpha \propto \frac{1}{\sqrt{C}}$$

$$1 - \alpha \approx 1$$

$1 \gg \alpha$

$$k_a = C\alpha^2 \quad [\text{H}^+] = C\alpha$$

$$\alpha = \sqrt{\frac{k_a}{C}} \quad [\text{H}^+] = \sqrt{k_a C}$$

$$pH = \frac{1}{2} [p k_a - \log C]$$

Q. Calculate pH of  $\underline{0.1 \text{ M CH}_3\text{COOH}}$  ( $K_a = \underline{10^{-5}}$ )

$$pH = \frac{1}{2} [5 - \log 10^1]$$

$$pH = \frac{1}{2} [5 + 1]$$

$$pH = 3$$

$$\begin{aligned} pK_a &= -\log K_a \\ &= -\log 10^{-5} \\ &= 5 \end{aligned}$$

## Case 4 pH of weak poly basic acid or Weak poly acidic base

$H_2S$ ,  $H_3PO_4$ ,  $H_3PO_3$  etc

$$\begin{array}{ccc} K_{a_1} & K_{a_1} & K_{a_1} \\ K_{a_2} & K_{a_2} & K_{a_2} \\ & K_{a_3} & \end{array}$$

$Mg(OH)_2$  etc  
 $Al(OH)_3$



$$pH = \frac{1}{2} [pK_{a_1} - \log c]$$

Q. Find pH of 0.1 M  $\text{H}_3\text{PO}_4$   $K_{a_1} = 10^{-3}$  ✓  $PK_{a_1} = -\log 10^{-3} = 3$   
 $K_{a_2} = 10^{-8}$  ✗  
 $K_{a_3} = 10^{-13}$  ✗

$$pH = \frac{1}{2} [PK_{a_1} - \log c]$$

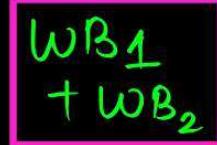
$$= \frac{1}{2} [3 - \log 10^1]$$

$$= \frac{1}{2} [3+1]$$

$$pH = 2$$

**Case 5** pH of mix of two weak acids or two weak bases

OR



$$\boxed{[H^+] = \sqrt{K_a_1 c_1 + K_a_2 c_2}}$$

 $c_1$  &  $c_2$  conc<sup>n</sup> after mixing

Q. Find pH of sol<sup>n</sup> obtained by mixing equal volumes of 0.02 HOCl

$(K_a = 2 \times 10^{-4})$  and 0.2 CH<sub>3</sub>COOH ( $K_a = 2 \times 10^{-5}$ )

$$[H^+] = A \times 10^{-B}$$

$$[H^+] = \sqrt{K_a c_1 + K_a c_2}$$

$$c_1 = \frac{0.02}{2} = 0.01$$

$$[H^+] = \sqrt{2 \times 10^{-4} \times 10^{-2} + 2 \times 10^{-5} \times 10^{-1}}$$

$$c_2 = \frac{0.2}{2} = 0.1$$

$$= \sqrt{2 \times 10^{-6} + 2 \times 10^{-6}}$$

$$= \sqrt{4 \times 10^{-6}}$$

$$= 2 \times 10^{-3}$$

$$pH = 3 - \log 2 = 3 - 0.3 = 2.7$$

**Case 6** pH of mix of WA and SA or WB and SB

$$[H^+] = \frac{-C_2 \pm \sqrt{C_2^2 + 4K_{a1}C_1}}{2}$$

 $C_2$  = conc<sup>n</sup> of SA $C_1$  = conc<sup>n</sup> of WA

Q. Find pH of mix of  $10^{-4}$  M HCl +  $10^{-2}$  M CH<sub>3</sub>COOH ( $K_a = 2 \times 10^{-5}$ )

$$\begin{array}{l} \text{SA} \\ C_2 = 10^{-4} \end{array} \quad \begin{array}{l} \text{WA} \\ C_1 = 10^{-2} \end{array} \quad 1$$

$$[\text{H}^+] = -10^{-4} \pm \sqrt{(10^{-4})^2 + 4 \times 2 \times 10^{-5} \times 10^{-2}}$$

$$= \frac{-10^{-4} + \sqrt{10^{-8} + 8 \times 10^{-7}}}{2}$$

$$= \frac{-10^{-4} + \sqrt{10^{-8} + 80 \times 10^{-8}}}{2} = \frac{-10^{-4} + \sqrt{81 \times 10^{-8}}}{2}$$

$$[\text{H}^+] = \frac{-10^{-4} + 9 \times 10^{-4}}{2}$$

$$= \frac{8 \times 10^{-4}}{2}$$

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

$$\text{pH} = 4 - \log 4$$

$$= 4 - 0.6$$

$$\text{pH} = 3.4$$

**Case 7** pH of Amphiprotic Species

$\text{H}^+$  donor & acceptor

$$K_{a_1} \vee K_{a_2}$$

$$\text{pH} = \frac{\text{p}K_{a_1} + \text{p}K_{a_2}}{2}$$

Q Find pH of NaHA



given  $K_{a1} = 10^{-3}$   $pK_{a_1} = 3$   
 $K_{a2} = 10^{-8}$   $pK_{a_2} = 8$

$$pH = \frac{3+8}{2} = \frac{11}{2} = 5.5$$



## Some Important points

### 1. Isohydric Sol<sup>n</sup>

Sol<sup>n</sup> having same [H<sup>+</sup>]

### 2. Relative strength of acids

$$RS = \frac{[H^+]_1}{[H^+]_2} = \frac{\sqrt{K_a_1 C_1}}{\sqrt{K_a_2 C_2}}$$

## Salt hydrolysis



$$1) K_H = \frac{K_w}{K_a}$$

$$2) \alpha = \sqrt{\frac{K_H}{C}}$$

$$3) P^H = 7 + \frac{1}{2} [PK_a + \log C]$$

4)  $P^H > 7$ , Nature of sol<sup>n</sup>  
Basic



$$1) K_H = \frac{K_w}{K_b}$$

$$2) \alpha = \sqrt{\frac{K_H}{C}}$$

$$3) P^H = 7 - \frac{1}{2} [PK_b + \log C]$$

4)  $P^H < 7$ , Nature of sol<sup>n</sup>  
acidic



$$1) K_H = \frac{K_w}{K_a K_b}$$

$$2) \alpha = \sqrt{K_H}$$

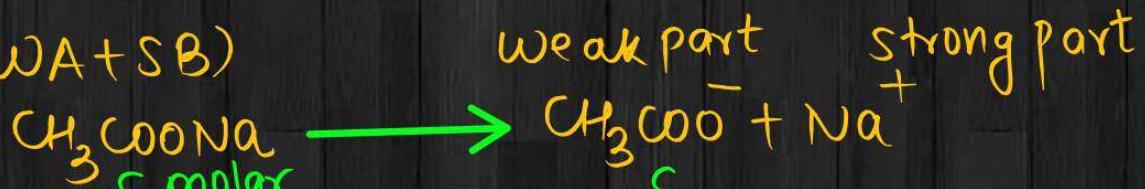
$$3) P^H = 7 + \frac{1}{2} [PK_a - PK_b]$$

a)  $K_a > K_b$ ;  $PK_a < PK_b$ ;  $P^H < 7$

b)  $K_a < K_b$ ;  $PK_a > PK_b$ ;  $P^H > 7$

c)  $K_a = K_b$ ;  $PK_a = PK_b$ ;  $P^H = 7$

(WA + SB)



Hydrolysis const

$$k_H$$

$$k_H = c\alpha^2$$

$$\alpha = \sqrt{\frac{k_H}{c}}$$

$$K_H = \frac{k_w}{k_a}$$

Q. Which salts will have highest pH?

- (A) KCl  $pH = 7$
- (B)  $\text{Na}_2\text{CO}_3$   $pH > 7$
- (C)  $\text{CuSO}_4$
- (D) NOT

$SA + SB$

$pH < 7$

Q. PK<sub>b</sub> for CN<sup>-</sup> is 4.7 The pH of 0.5 M aq NaCN Sol<sup>n</sup> is  
WA+SB

$$pH = 7 + \frac{1}{2} [pK_a + \log c]$$

$$pH = 7 + \frac{1}{2} [9.3 + \log(5 \times 10^1)]$$

$$= 7 + \frac{1}{2} [9.3 + \log 5 + \log 10^1]$$

$$= 7 + \frac{1}{2} [9.3 + 0.7 - 1]$$

$$= 7 + \frac{1}{2} \times 9 = 11.5$$

$$pK_b = 4.7$$

$$\begin{aligned} pK_a &= 14 - 4.7 \\ &= 9.3 \end{aligned}$$

Q. Kb for NH<sub>4</sub>OH is  $\underbrace{2 \times 10^{-5}}$  Find k<sub>H</sub> of NH<sub>4</sub>Cl

$$K_H = \frac{K_w}{K_b} = \frac{10^{-14}}{2 \times 10^{-5}} = \frac{10^{-9}}{2}$$

Q. [Incorrect match] pH < 7

SA + WB pH < 7

(A) FeCl<sub>3</sub> in H<sub>2</sub>O - Basic

(B) NH<sub>4</sub>Cl in H<sub>2</sub>O - Acidic

(C) CH<sub>3</sub>COONH<sub>4</sub> in H<sub>2</sub>O - Neutral

(D) Na<sub>2</sub>CO<sub>3</sub> in H<sub>2</sub>O - Basic

WA + WB

WA + SB

CH<sub>3</sub>COOH  $K_a = 2 \times 10^{-5}$

pH > 7

NH<sub>4</sub>OH  $K_b = 2 \times 10^{-5}$

## Buffer Solution

Acidic Buffer

$WA + \text{Salt of } WA + SB$

Basic Buffer

$WB + \text{Salt of } WB + SA$

### Examples of B.S

#### Acidic

1. HCN + NaCN
2. H<sub>3</sub>BO<sub>3</sub> + Na<sub>2</sub>B<sub>4</sub>O<sub>7</sub>
3. H<sub>2</sub>CO<sub>3</sub> + NaHCO<sub>3</sub>
4. H<sub>3</sub>PO<sub>4</sub> + NaH<sub>2</sub>PO<sub>4</sub>
5. CH<sub>3</sub>COOH + CH<sub>3</sub>COONa
6. HNO<sub>2</sub> + NaNO<sub>2</sub>

### Examples of B.S

#### Basic

1. CH<sub>3</sub>NH<sub>2</sub> + CHNH<sub>3</sub><sup>+</sup>Cl<sup>-</sup>
2. NH<sub>4</sub>OH + NH<sub>4</sub>Cl
3. PhNH<sub>2</sub> + PhNH<sub>3</sub><sup>+</sup>Cl<sup>-</sup>
2. C<sub>5</sub>H<sub>5</sub>N + C<sub>5</sub>H<sub>5</sub>NH<sup>+</sup> Cl<sup>-</sup>



Salt Buffer -Salt of WA+WB

## Acidic Buffer can be prepared by

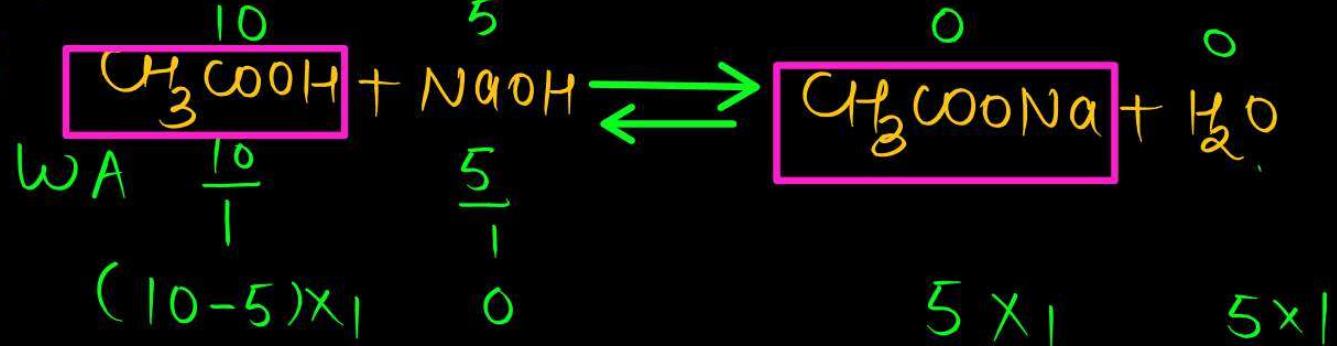
1. WA + Salt of WA+SB



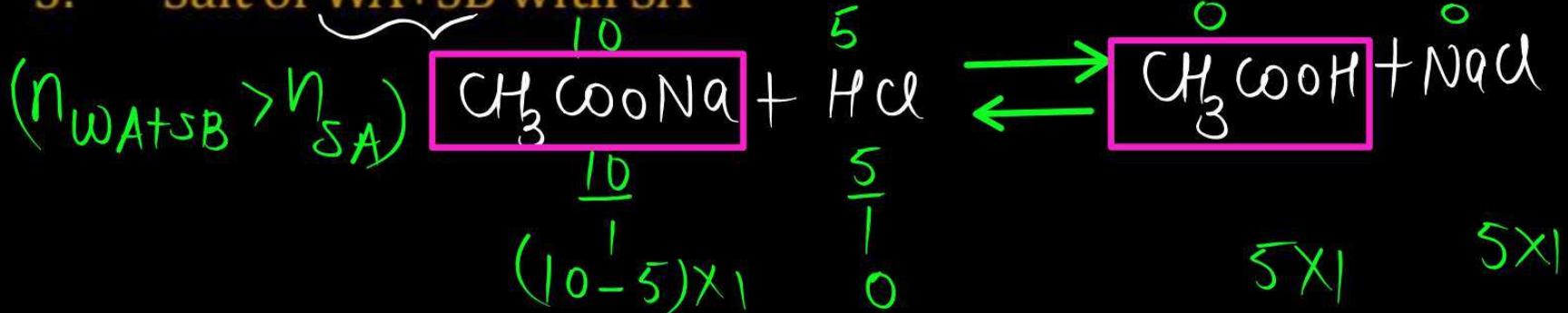
$n_{WA} > n_{SB}$

Salt of WA+SB

2. WA + SB



3. ✓ Salt of WA+SB with SA



Acidic Buffer

$$pH = pK_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

Basic Buffer

$$pOH = pK_b + \log \frac{[\text{salt}]}{[\text{base}]}$$

$$pH = 14 - pOH$$

# Buffer Range

$$\frac{1}{10} \leq \frac{[\text{conju base}]}{[\text{Acid}]} \leq 10$$
$$[P^k a - 1 \leq pH \leq P^k a + 1]$$

## Buffer Capacity ✓

Buffer Capacity = No of moles of Strong acid or Strong base added per L of sol<sup>n</sup>

Change in pH of buffer

Q. when 2 moles of HCl is added to 1 L of an acidic buffer sol<sup>n</sup> its pH changes from 3.9 to 3.4 find its buffer capacity

$$\text{B.C} = \frac{(2/1)}{3.9 - 3.4} = \frac{2}{0.5} = 4$$

Calculation of pH of buffer Sol<sub>n</sub>

2 M CH<sub>3</sub>COOH mixed with 1 M CH<sub>3</sub>COONa will have an approx pH of? (K<sub>a</sub> = 10<sup>-5</sup>)

$$\begin{aligned} \text{pH} &= 5 + \log \frac{1}{2} & \text{pK}_a &= 5 \\ &= 5 - \log 2 \\ &= 5 - 0.3 \\ &= 4.7 \end{aligned}$$

Q.  $\underbrace{0.05 \text{ M NH}_4\text{OH Soln}}$  is dissolved in  $\underbrace{0.001 \text{ M NH}_4\text{Cl Soln}}$  find pH ( $K_b = 10^{-5}$ )

PW

$$pK_b = 5$$

$$\begin{aligned} pOH &= 5 + \log\left(\frac{10^{-3}}{5 \times 10^{-2}}\right) & pOH &= 3.3 \\ &= 5 + \log\left(\frac{1}{50}\right) & pH &= 14 - 3.3 \\ &= 5 - \log 50 & &= 10.7 \\ &= 5 - 1.7 \\ &= 3.3 \end{aligned}$$

Q.  $\text{CH}_3\text{COOH}$  was titrated with  $\text{NaOH}$  when  $75\%$  of titration is over then find pH? ( $K_a = 10^{-5}$ )  $pK_a = 5$



25 l.

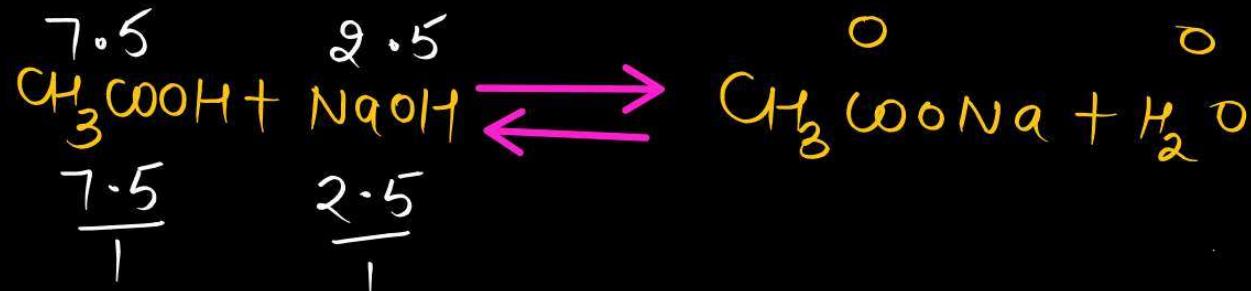
75 l.

$$\begin{aligned} p^H &= 5 + \log\left(\frac{75}{25}\right) \\ &= 5 + \log 3 \end{aligned}$$

# Caculation of pH when WA + SB mixing

$$k_a(\text{CH}_3\text{COOH}) = 10^{-5} \quad pK_a = 5$$

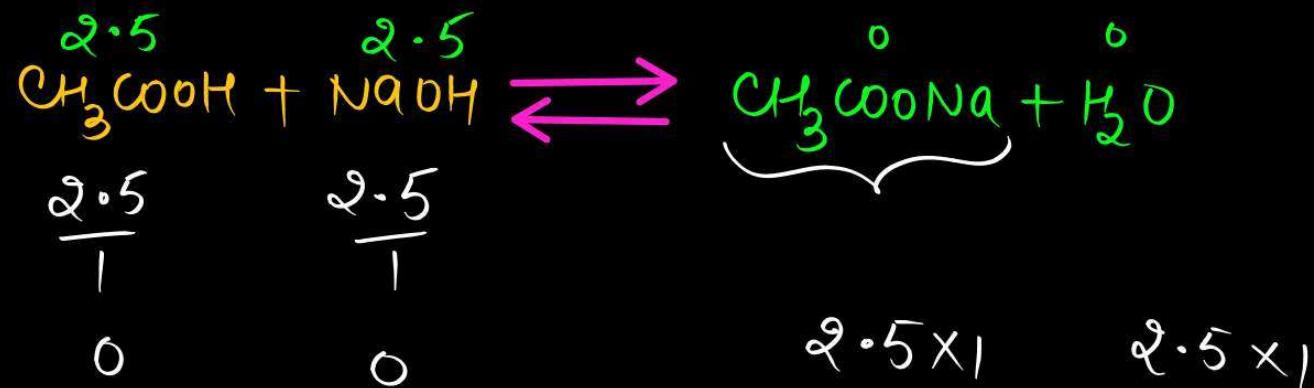
Case (1) 75 mL of 0.1 M CH<sub>3</sub>COOH + 25 mL of 0.1 M NaOH



$$1 \times (7.5 - 2.5) \quad 0 \quad 2.5 \times 1 \quad 2.5 \times 1$$

$$pH = pK_a + \log \frac{[S]}{[A]} = 5 + \log \left( \frac{2.5}{5} \right) = 5 + \log \frac{1}{2} = 4.7$$

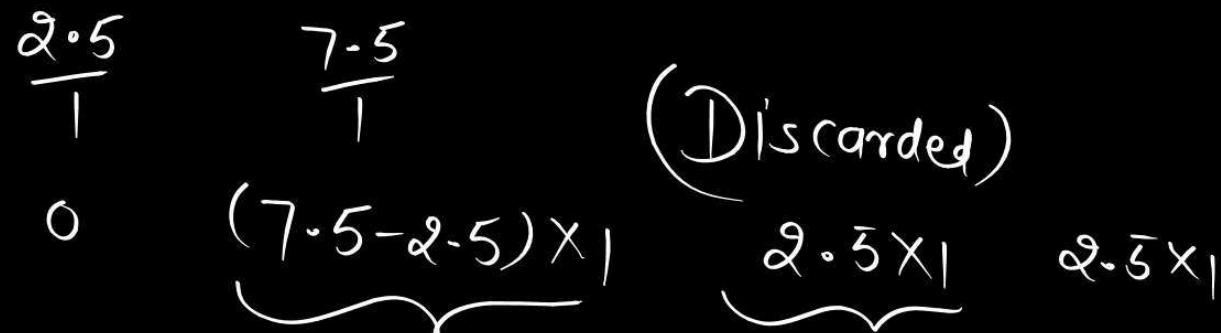
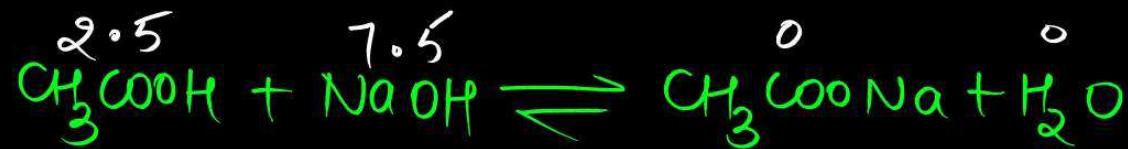
Case (2) 25 mL of 0.1 M CH<sub>3</sub>COOH + 25 mL of 0.1 M NaOH



$$\begin{aligned}
 \text{pH} &= 7 + \frac{1}{2} [\text{pK}_a + \log c] \\
 &= 7 + \frac{1}{2} [5 + \log \frac{1}{20}] \\
 &= 7 + \frac{1}{2} [5 - 1.3]
 \end{aligned}$$

$$\begin{aligned}
 c &= \frac{2.5}{50} \\
 &= \frac{25}{500} \\
 &= \frac{1}{20}
 \end{aligned}$$

Case (3) 25 mL of 0.1 M CH<sub>3</sub>COOH + 75 mL of 0.1 M NaOH



$$n_{\bar{\text{O}}\text{H}} = 5$$

$$[\bar{\text{O}}\text{H}] = \frac{5}{100} = \frac{1}{20}$$

$$\rho^{\text{H}} = 14 - 1.3$$

$$\rho^{\text{O}\text{H}} = -\log \frac{1}{20} = \log 20 = 1.3$$

Q.  $\text{pH} - \text{P}K_a = 1$  will be applicable to an acidic buffer when

- (A)  $[\text{Acid}] = [\text{Conju base}]$
- (B)  $[\text{Acid}] \times 10 = [\text{Conju base}]$
- (C)  $[\text{Acid}] = [\text{Conju base}] \times 10$
- (D) N.O.T

$$\log \frac{[S]}{[A]} = 1$$

$$\frac{[S]}{[A]} = 10$$

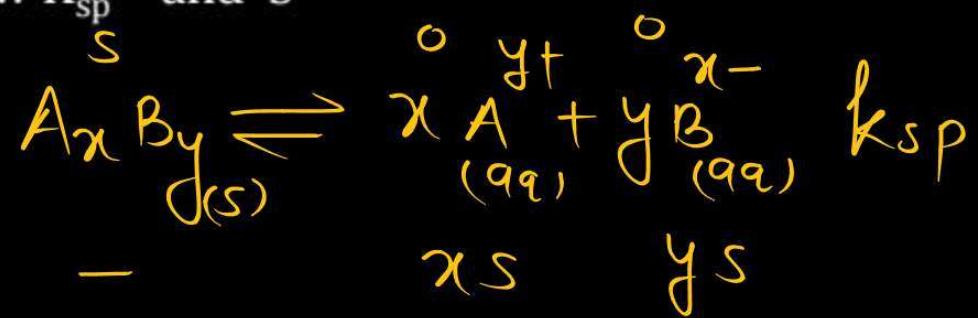
$(\frac{\text{mol}}{\text{L}})$ 

## Solubility and solubility product

Solubility : max no of moles of solute dissolved in a solvent to obtain 1L soln

$$\left( \overline{s} \left( \frac{\text{mol}}{\text{L}} \right) = \frac{\overline{s} (\text{gm/L})}{\text{M.W}} \right)$$

Relation b/w  $K_{sp}$  and S



$$K_{sp} = (x_S)^x (y_S)^y$$

$$K_{sp} = x^x y^y S^{x+y}$$

Compound	X	Y	Relation b/W S and $k_{sp}$
$\text{Ag}_2$	1	1	$1^1 \cdot 1^1 \cdot S^{1+1} = S^2$
$\text{MgCl}_2$	1	2	$1^1 \cdot 2^2 \cdot S^{1+2} = 4S^3$
$\text{AlCl}_3$	1	3	$1^1 \cdot 3^3 \cdot S^{1+3} = 27S^4$
* $\text{Hg}_2\text{Cl}_2$	<del>1</del> 1	2	$1^1 \cdot 2^2 \cdot S^{1+2} = 4S^3$
$\text{Hg}^+ - \text{Hg}^+$			

Q. if solubility of  $\text{Ag}_2\text{SO}_4$  is  $10^{-2} \frac{\text{mol}}{\text{L}}$  then find its  $k_{sp}$  ?

$$\overbrace{x=2}^{\text{Ag}}, \overbrace{y=1}^{\text{SO}_4^{2-}}$$

$$k_{sp} = 2^2 \cdot 1^1 \cdot 5^{2+1}$$

$$= 4 \cdot 5^3$$

$$= 4 \times (10^{-2})^3$$

$$= 4 \times 10^{-6}$$

Q. if solubility of AgCl is  $10^{-10}$  then find its solubility in gm/L ?  
product      s<sup>1+1</sup>

$$| \cdot | s^{1+1} = k_{sp}$$

$$k_{sp} = s^2$$

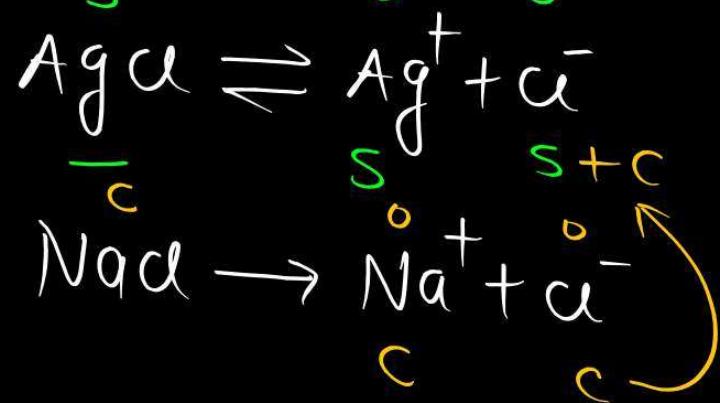
$$s = 10^{-5} \frac{\text{mol}}{\text{L}}$$

$$s (\text{gm/L}) = 10^{-5} \times 143.5$$

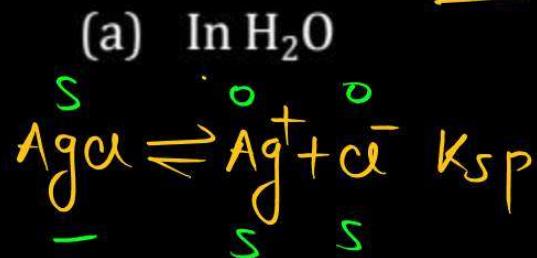
## Solubility in common ion effect

Solubility decreases by adding common ion

$\text{Ag}_\alpha$  is mixed with  $\text{NaCl}$



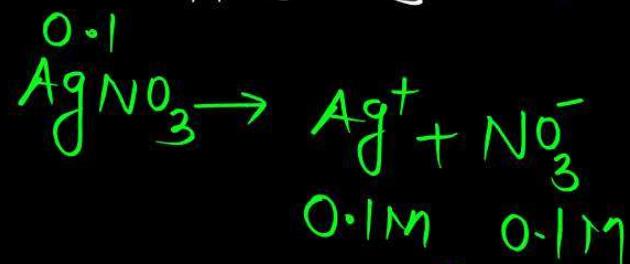
Q. Find solubility of AgCl ( $K_{sp} = 10^{-10}$ )



$$K_{sp} = S \times S$$

$$S = 10^{-5}$$

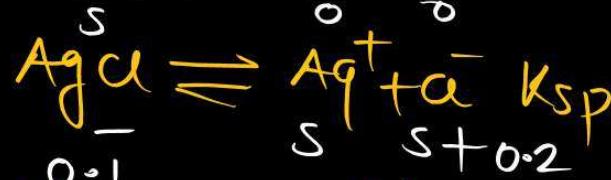
(b) In 0.1 M  $AgNO_3$



$$K_{sp} = (0.1 + S) \times S$$

$$S = \frac{10^{-10}}{0.1} = 10^{-9}$$

(c) In 0.1 M  $CaCl_2$



$$K_{sp} = S \times (S + 0.2)$$

$$S = \frac{10^{-10}}{0.2}$$

## Ionic Product of salt



$$K_{IP} = [A^{y+}]^x [B^{x-}]^y$$

Case ①  $K_{SP} = K_{IP}$  Saturated soln

Case ②  $K_{SP} > K_{IP}$  Unsaturated soln

Case ③  $K_{IP} > K_{SP}$  ppt occurs.

Q. The ppt of  $\text{CaF}_2$  ( $K_{\text{sp}} = 1.7 \times 10^{-10}$ ) is obtained when following are mixed in equal volume

- ~~X~~ (A)  $10^{-3} \text{ M}$      $10^{-5} \text{ M}$
- (B)  $10^{-5} \text{ M}$      $10^{-3} \text{ M}$
- ~~C~~ (C)  $10^{-2} \text{ M}$      $10^{-3} \text{ M}$
- (D)  $10^{-4} \text{ M}$      $10^{-4} \text{ M}$

$$\textcircled{c} \quad K_{\text{IP}} = 10^{-2} \times \left(\frac{10^{-3}}{2}\right)^2 > K_{\text{sp}}$$



$$K_{\text{IP}} = [\text{Ca}^{2+}] [\text{F}^-]^2$$

$$\textcircled{A} \quad K_{\text{IP}} = \frac{10^{-3}}{2} \left(\frac{10^{-5}}{2}\right)^2 < K_{\text{sp}}$$

$$\textcircled{B} \quad K_{\text{IP}} = \frac{10^{-5}}{2} \times \left(\frac{10^{-3}}{2}\right)^2$$

$$\frac{10^{-11}}{8}$$

## Some Important points

### 1. Types of salts

#### (a) Normal Salt

NaCl, Na<sub>2</sub>SO<sub>4</sub>, Na<sub>3</sub>PO<sub>4</sub> etc

#### (b) Acid salts

NaHCO<sub>3</sub>, NaHSO<sub>4</sub>

#### c) Basic salts

Zn(OH)Cl, Mg(OH)Cl & etc

#### (d) Double salts

K<sub>2</sub>SO<sub>4</sub>. Al<sub>2</sub>(SO<sub>4</sub>)<sub>3</sub>. 24H<sub>2</sub>O potash alum

#### (e) Complex salts

[Ag(NH<sub>3</sub>)<sub>2</sub>] Cl, [Cu(NH<sub>3</sub>)<sub>4</sub>] SO<sub>4</sub> etc

#### (f) mixed salts

NaKS, Na k Rb PO<sub>4</sub> etc.

—

pyQs



Q.

The molar solubility of Cd(OH)<sub>2</sub> is  $1.84 \times 10^{-5} \text{ M}$  in water.

The expected solubility of Cd(OH)<sub>2</sub> in a buffer solution of pH = 12 is  $\text{P}_{\text{OH}} = 2$   $[\text{OH}] = 10^{-2}$  (2019 Main, 12 April II)

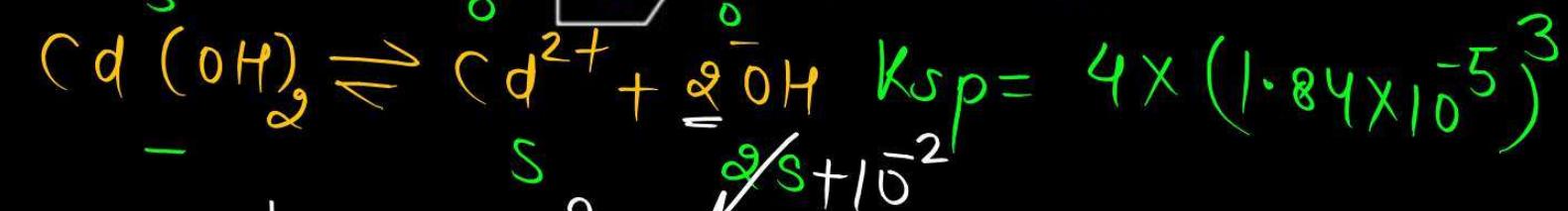
a.  $1.84 \times 10^{-9} \text{ M}$

b.  $\frac{2.49}{1.84} \times 10^{-9} \text{ M}$

c.  $6.23 \times 10^{-11} \text{ M}$

d.  $2.49 \times 10^{-10} \text{ M}$

Solution:



$$K_{\text{sp}} = s^1 \times (10^{-2})^2$$

$$s = \frac{K_{\text{sp}}}{10^{-4}}$$



Q.

What is the molar solubility of  $\text{Al(OH)}_3$  in 0.2 M NaOH solution? Given that, solubility product of

$$\text{Al(OH)}_3 = 2.4 \times 10^{-24}$$

(2019 Main, 12 April II)

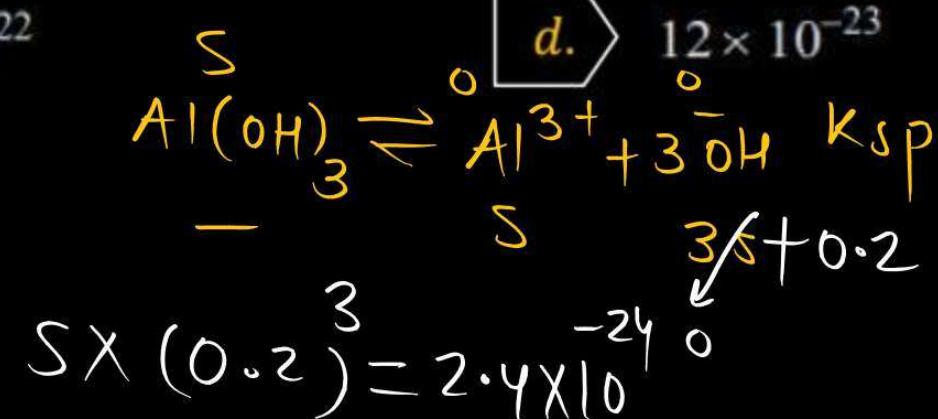
a.  $3 \times 10^{-19}$

b.  $12 \times 10^{-21}$

c.  $3 \times 10^{-22}$

d.  $12 \times 10^{-23}$

Solution:





Q.

The pH of a 0.02 M NH<sub>4</sub>Cl solution will be [Given  
 $K_b(\text{NH}_4\text{OH}) = 10^{-5}$  and  $\log 2 = 0.301$ ] (2019 Main, 10 April II)

- (a) 4.65      (b) 2.65      (c) 5.35      (d) 4.35

Solution:

$$\begin{aligned} \text{pH} &= 7 - \frac{1}{2} [\text{p}K_b + \log c] \\ &= 7 - \frac{1}{2} [5 + \log(2 \times 10^{-2})] \end{aligned}$$

**Q.**

Consider the following statements.

- I. The pH of a mixture containing 400 mL of 0.1 M  $\text{H}_2\text{SO}_4$  and 400 mL of 0.1 M  $\text{NaOH}$  will be approximately 1.3.
- II. Ionic product of water is temperature dependent. =
- III. A monobasic acid with  $K_a = 10^{-5}$  has a pH = 5. The degree of dissociation of this acid is 50%.
- IV. The Le-Chatelier's principle is not applicable to common-ion effect.

The correct statements are

(2019 Main, 10 April I)

- (a) I, II and IV
- (b) II and III
- (c) I and II
- (d) I, II and III

Solution:

$$[\text{H}^+] / [\text{OH}^-] = \frac{|0.1 \times 400 \times 2 - 0.1 \times 400 \times 1|}{800}$$

$$[\text{H}^+] = \frac{|80 - 40|}{800}$$

$$= \frac{40}{800}$$

$$= \frac{1}{20}$$

$$\text{pH} = -\log \frac{1}{20}$$

$$= 1.3$$

$$k_a = \frac{c\alpha^2}{1-\alpha}$$

$$k_a = \frac{\cancel{c}\alpha \times \alpha}{1-\alpha}$$

$$10^5 = \frac{10^5 \times \alpha}{1-\alpha}$$

$$\frac{\alpha}{1-\alpha} = 1 \quad \alpha = \frac{1}{2} \quad 50\%$$
$$\alpha = 1 - \alpha$$



Q.

$$3^3 \times 4^4 \times S^{3+4} = K_{sp}$$

If solubility product of  $\text{Zr}_3(\text{PO}_4)_4$  is denoted by  $K_{sp}$  and its molar solubility is denoted by  $S$ , then which of the following relation between  $S$  and  $K_{sp}$  is correct? (2019 Main, 8 April I)

(a)  $S = \left( \frac{K_{sp}}{144} \right)^{1/6}$

✓ (b)  $S = \left( \frac{K_{sp}}{6912} \right)^{1/7}$

(c)  $S = \left( \frac{K_{sp}}{929} \right)^{1/9}$

(d)  $S = \left( \frac{K_{sp}}{216} \right)^{1/7}$

Solution:

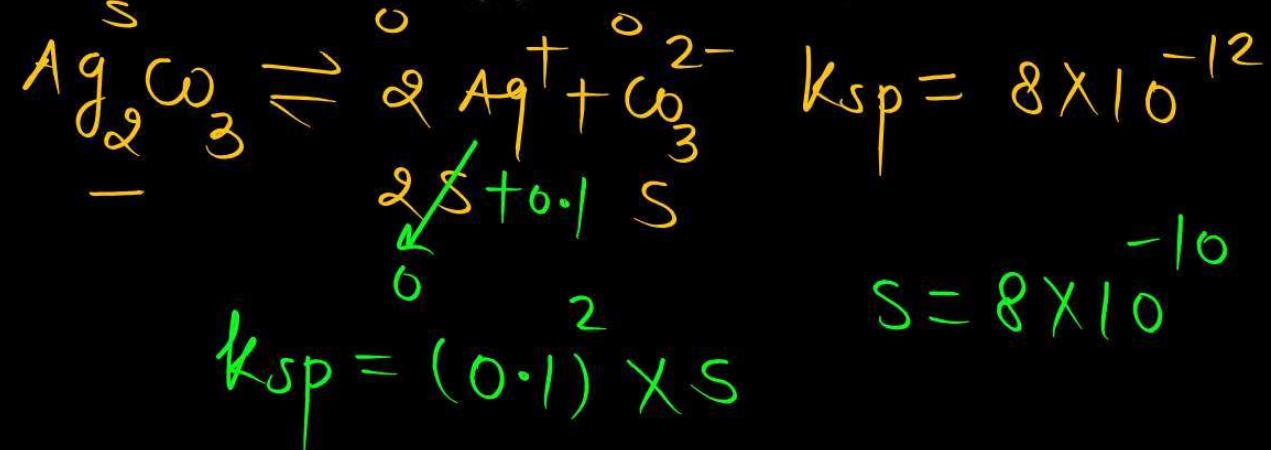


Q.

If  $K_{sp}$  of  $\text{Ag}_2\text{CO}_3$  is  $8 \times 10^{-12}$ , the molar solubility of  $\text{Ag}_2\text{CO}_3$  in  $\underbrace{0.1 \text{ M AgNO}_3}_{\text{in}}$  is

(2019 Main, 12 Jan II)

- (a)  $8 \times 10^{-12} \text{ M}$       (b)  $8 \times 10^{-13} \text{ M}$   
~~(c)  $8 \times 10^{-10} \text{ M}$~~       (d)  $8 \times 10^{-11} \text{ M}$

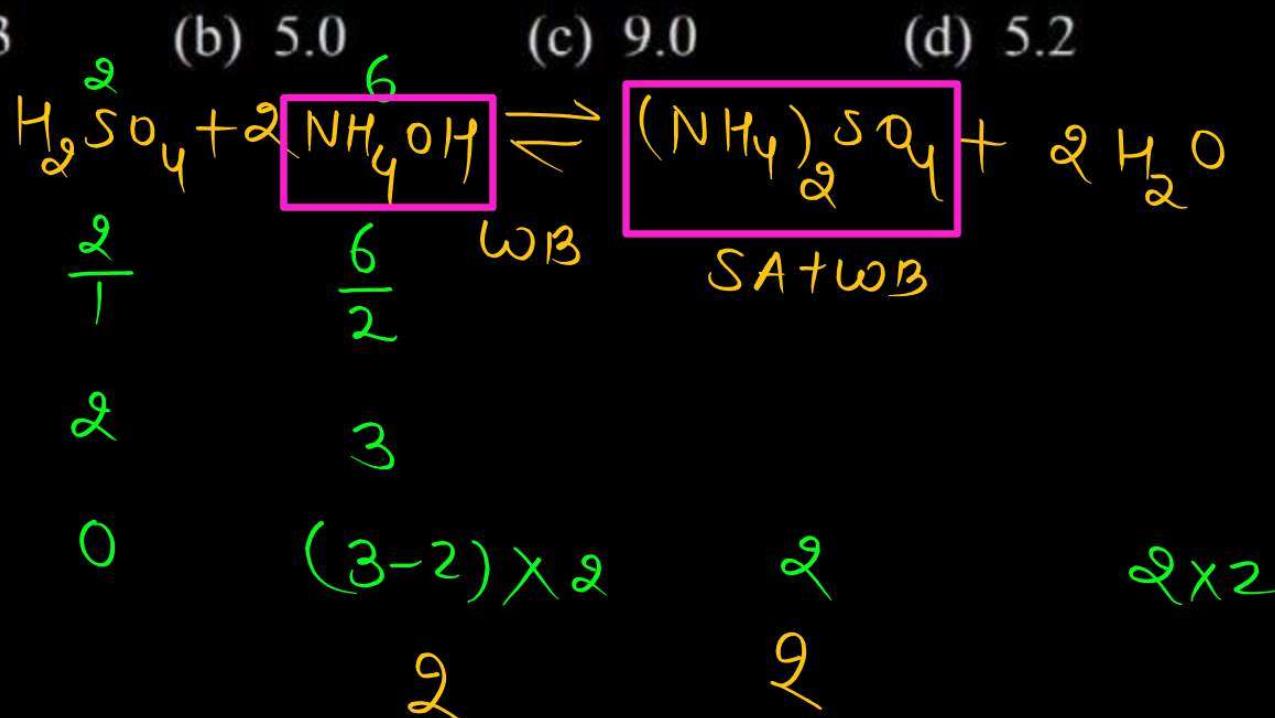
Solution:



Q. <sup>SA</sup> 20 mL of 0.1 M  $\text{H}_2\text{SO}_4$  solution is added to 30 mL of 0.2 M  $\text{NH}_4\text{OH}$  solution. The pH of the resultant mixture is [ $\text{p}K_b$  of  $\text{NH}_4\text{OH} = 4.7$ ] (2019 Main, 9 Jan I)

- ~~(a) 9.3~~ (b) 5.0 (c) 9.0 (d) 5.2

Solution:



$$p^{OH} = pK_b + \log \frac{[S]}{[B]}$$

$$= 4.7 + \log 1$$

$$= 4.7$$

$$p^H = 14 - 4.7$$

$$= 9.3$$



**Q.** Which of the following are Lewis acids?

(2018 Main)

- (a)  $\text{PH}_3$  and  $\text{BCl}_3$
- (b)  $\text{AlCl}_3$  and  $\text{SiCl}_4$
- (c)  $\text{PH}_3$  and  $\text{SiCl}_4$
- (d)  $\text{BCl}_3$  and  $\text{AlCl}_3$

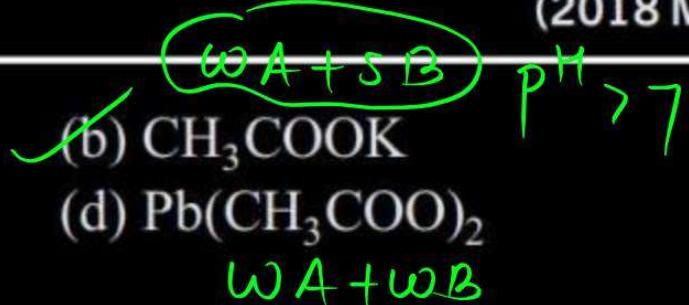
Solution:



Q.

Which of the following salts is the most basic in aqueous solution?

(2018 Main)



Solution:



Q.  $\text{p}K_a$  of a weak acid ( $\text{HA}$ ) and  $\text{p}K_b$  of a weak base ( $\text{BOH}$ ) are 3.2 and 3.4, respectively. The pH of their salt ( $\text{AB}$ ) solution is \_\_\_\_\_ (2017 Main)

- (a) 7.2      ✓(b) 6.9      (c) 7.0      (d) 1.0

Solution:

$$\begin{aligned}\text{pH} &= 7 + \frac{1}{2} [\text{p}K_a - \text{p}K_b] \\ &= 7 + \frac{1}{2} [3.2 - 3.4] \\ &= 7 + \frac{1}{2} [-0.2] \\ &= 6.9\end{aligned}$$



Q.

How many litres of water must be added to 1 L of an aqueous solution of HCl with a pH of 1 to create an aqueous solution with pH of 2? (2013 Main)

- (a) 0.1 L    (b) 0.9 L    (c) 2.0 L    (d) 9.0 L

Solution:  $\text{[H}^+ \text{]}_1 = 10^{-1}$      $V_1 = 1 \text{ L}$      $M_1 V_1 = M_2 V_2$   
 $\text{[H}^+ \text{]}_2 = 10^{-2}$      $V_2 = ?$      $10^{-1} \times 1 = 10^{-2} \times V_2$

$$1 \text{ L} + (9 \text{ L}) = 10 \text{ L}$$



Q.

Solubility product constant ( $K_{sp}$ ) of salts of types  $\underline{MX}$ ,  $\underline{MX}_2$  and  $\underline{M}_3X$  at temperature 'T' are  $4.0 \times 10^{-8}$ ,  $3.2 \times 10^{-14}$  and  $2.7 \times 10^{-15}$ , respectively. Solubilities (mol dm<sup>-3</sup>) of the salts at temperature 'T' are in the order (2008, 3M)

(a)  $MX > MX_2 > M_3X$       (b)  $M_3X > MX_2 > MX$

(c)  $MX_2 > M_3X > MX$       (d)  $MX > M_3X > MX_2$

Solution:

$$S_1^2 = K_{sp},$$

$$S_1 = 2 \times 10^{-4}$$

$$4S_2^3 = K_{sp_2}$$

$$4S_2^3 = 3.2 \times 10^{-14}$$

$$S_2 = 2 \times 10^{-5}$$

$$27S_3^4 = K_{sp_3}$$

$$27S_3^4 = 2.7 \times 10^{-15}$$

$$S_3 = 10^{-4}$$



Q.

$\text{CH}_3\text{NH}_2$  (0.1 mole,  $K_b = 5 \times 10^{-4}$ ) is added to 0.08 mole of  $\text{HCl}$  and the solution is diluted to one litre, resulting hydrogen ion concentration is (2005, 1M)

(a)  $1.6 \times 10^{-11}$

(c)  $5 \times 10^{-5}$

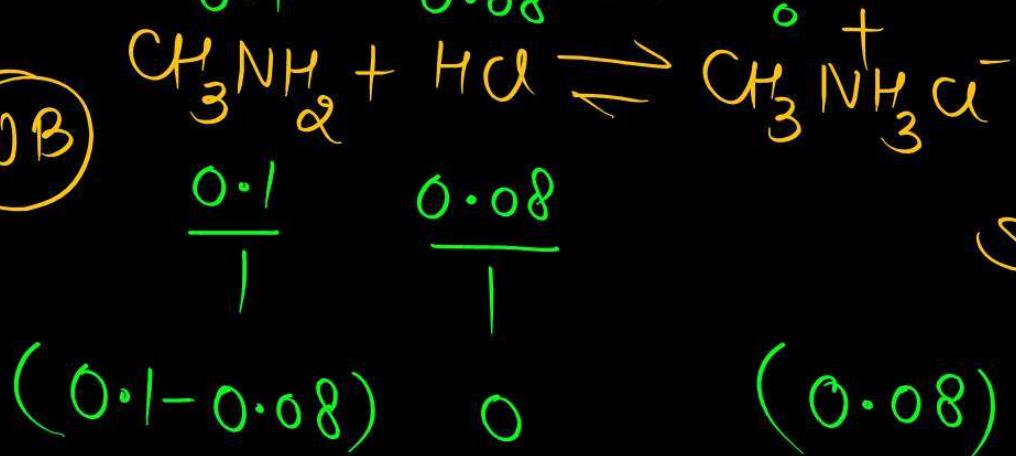
(b)  $8 \times 10^{-11}$

(d)  $8 \times 10^{-2}$

$$\begin{aligned} pK_b &= 4 - \log 5 \\ &= 3.3 \end{aligned}$$

Solution:

(WB)



Salt of SA + WB

$$p^{OH} = p^{K_b} + \log \frac{[S]}{[B]}$$
$$= 3.3 + \log \left( \frac{0.08}{0.02} \right)$$

$$= 3.3 + 0.6$$

$$p^{OH} = 3.9$$

$$p^H = 14 - 3.9$$

$$p^H = 10.1$$



Q.

HX is a weak acid ( $K_a = 10^{-5}$ ). It forms a salt NaX(0.1M) on reacting with caustic soda. The degree of hydrolysis of NaX is (2004, 1M)

- (a) 0.01%    (b) 0.0001%  
(c) 0.1%    (d) 0.5%

Solution:

$$\alpha = \sqrt{\frac{K_H}{C}} = \sqrt{\frac{10^{-4}}{10^{-5} \times 0.1}} = 10^4 = 0.01\%$$



Q.

A solution which is  $10^{-3}$  M each in  $\text{Mn}^{2+}$ ,  $\text{Fe}^{2+}$ ,  $\text{Zn}^{2+}$  and  $\text{Hg}^{2+}$  is treated with  $10^{-16}$  M sulphide ion. If  $K_{\text{sp}}$  of  $\text{MnS}$ ,  $\text{FeS}$ ,  $\text{ZnS}$  and  $\text{HgS}$  are  $10^{-15}$ ,  $10^{-23}$ ,  $10^{-20}$  and  $10^{-54}$  respectively, which one will precipitate first? (2003, 1M)

- (a) FeS      (b) MgS      (c) HgS      (d) ZnS

Solution:



**Q.** For a sparingly soluble salt  $A_pB_q$ , the relationship of its solubility product ( $L_s$ ) with its solubility ( $S$ ) is (2001, 1M)

- (a)  $L_s = S^{p+q} \cdot p^p \cdot q^q$       (b)  $L_s = S^{p+q} \cdot p^q \cdot q^p$   
(c)  $L_s = S^{pq} \cdot p^p \cdot q^q$       (d)  $L_s = S^{pq} \cdot (p \cdot q)^{(p+q)}$

Solution:

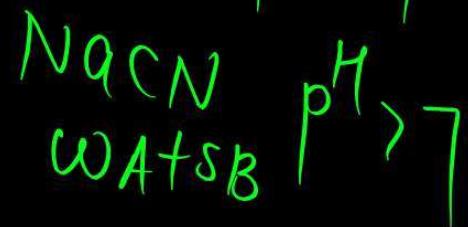


Q.

The pH of 0.1 M solution of the following salts increases in  
the order           

- (a)  $\text{NaCl} < \text{NH}_4\text{Cl} < \text{NaCN} < \text{HCl}$   
 (b)  $\text{HCl} < \text{NH}_4\text{Cl} < \text{NaCl} < \text{NaCN}$   
(c)  $\text{NaCN} < \text{NH}_4\text{Cl} < \text{NaCl} < \text{HCl}$   
(d)  $\text{HCl} < \text{NaCl} < \text{NaCN} < \text{NH}_4\text{Cl}$

Solution:





Q.

Which of the following solutions will have pH close to 1.0?

(1992, 1M)

$$P_{H^+} = 1$$

(a) ~~100 mL of (M/10) 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~~

$$\frac{1}{100}$$

Solution: (A)  $[H^+] / [OH^-] = \frac{|M_1 V_1 n_f_1 - M_2 V_2 n_f_2|}{V_1 + V_2}$

$$[H^+] = \frac{\left| \frac{1}{5} \times 75 \times 1 - \frac{1}{5} \times 25 \times 1 \right|}{100} = 10^{-1}$$

**Q.**

Which of the following is the strongest acid?

(1989, 1M)



- (a)  $\text{ClO}_3(\text{OH})$
- (b)  $\text{ClO}_2(\text{OH})$
- (c)  $\text{SO}(\text{OH})_2$
- (d)  $\text{SO}_2(\text{OH})_2$

Solution:



Q.

When equal volumes of the following solutions are mixed, precipitation of  $\text{AgCl}$  ( $K_{\text{sp}} = 1.8 \times 10^{-10}$ ) will occur only with

(1988, 1M)

- (a)  $10^{-4} \text{ M } (\text{Ag}^+)$  and  $10^{-4} \text{ M } (\text{Cl}^-)$
- (b)  $10^{-5} \text{ M } (\text{Ag}^+)$  and  $10^{-5} \text{ M } (\text{Cl}^-)$
- (c)  $10^{-6} \text{ M } (\text{Ag}^+)$  and  $10^{-6} \text{ M } (\text{Cl}^-)$
- (d)  $10^{-10} \text{ M } (\text{Ag}^+)$  and  $10^{-10} \text{ M } (\text{Cl}^-)$

$$\textcircled{A} \quad K_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-] \\ = \frac{10^{-4}}{2} \times \frac{10^{-4}}{2}$$

Solution:

Q.

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



(1985, 1M)



(a)  $\text{NH}_3$

(b)  $\text{NH}_2\text{OH}$

(c)  $\text{NH}_4^+$

(d)  $\text{N}_2\text{H}_4$

Solution:



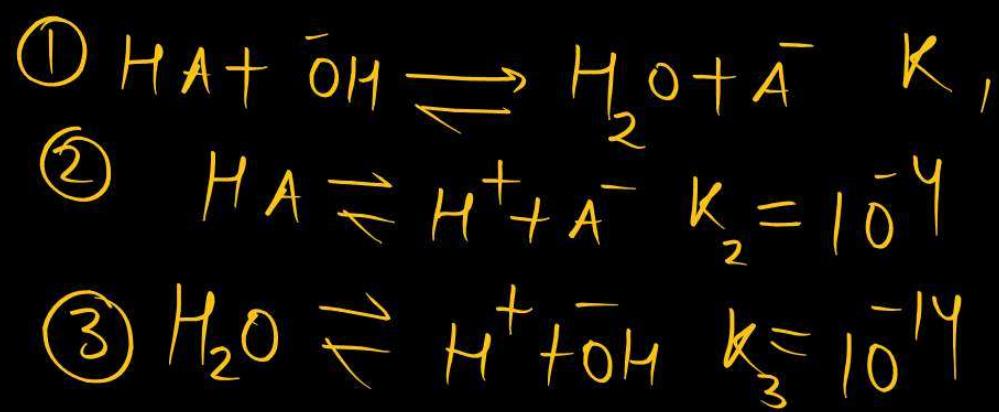
Q.

A certain weak acid has a dissociation constant of  $1.0 \times 10^{-4}$ .  
The equilibrium constant for its reaction with a strong base is  
(1984, 1M)

- (a)  $1.0 \times 10^{-4}$   
~~(c)  $1.0 \times 10^{10}$~~

- (b)  $1.0 \times 10^{-10}$   
(d)  $1.0 \times 10^{14}$

$$K_a = 10^{-4}$$
  
Solution:  $K_1 = \frac{10^{-4}}{10^{14}}$





Q.

A certain buffer solution contains equal concentration of  $X^-$  and  $HX$ . The  $K_b$  for  $X^-$  is  $10^{-10}$ . The pH of the buffer is

$$K_a \times K_b = 10^{-14}$$

(1984, 1M)

- Acidic buffer

- (a) 4  
(c) 10

- (b) 7  
(d) 14

Solution:

$$\begin{aligned}pH &= pK_a + \log \frac{[S]}{[A]} \\&= 4 + 0\end{aligned}$$

$$\begin{aligned}K_a(HX) &= \frac{10^{-14}}{10^{-10}} \\&= 10^{-4}\end{aligned}$$

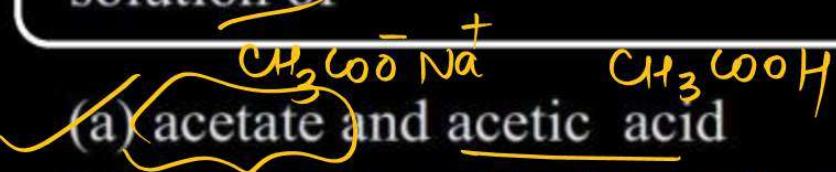
$$pK_a = 4$$



Q.

An acidic buffer solution can be prepared by mixing the solution of

(1981, 1M)



- (a) acetate and acetic acid
- (b) ammonium chloride and ammonium hydroxide
- (c) sulphuric acid and sodium sulphate
- (d) sodium chloride and sodium hydroxide

Basic Buffer  
 $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$

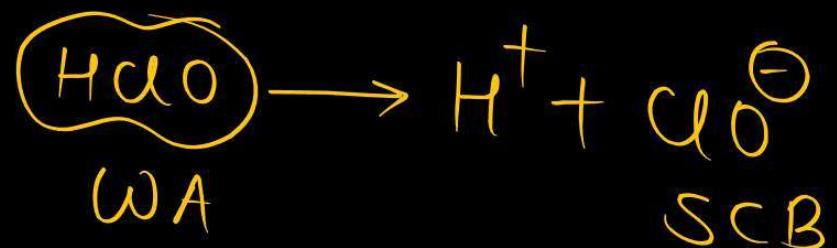
Solution:

**Q.** Of the given anions, the strongest base is

(1981, 1M)

- (a)  $\text{ClO}^-$
- (b)  $\text{ClO}_2^-$
- (c)  $\text{ClO}_3^-$
- (d)  $\text{ClO}_4^-$

Solution:





Q.

At 90°C, pure water has  $\text{[H}_3\text{O}^+]$  as  $10^{-6}$  mol L<sup>-1</sup>. What is the value of  $K_w$  at 90°C ? (1981, 1M)

- (a)  $10^{-6}$       (b)  $10^{-12}$       (c)  $10^{-14}$       (d)  $10^{-8}$

Solution:

$$\begin{aligned}K_w &= [\text{H}^+] [\text{OH}^-] = 10^{-6} \times 10^{-6} \\&= 10^{-12}\end{aligned}$$



**Q.** The pH of  $10^{-8}$  M solution of HCl in water is (1981, 1M)

- (a) 8
- (b) -8
- (c) between 7 and 8
- (d) between 6 and 7

Solution: