

# Ionic Equilibrium

## Question1

Identify conjugate acid and conjugate base for  $\text{HCO}_3^-$  ion respectively

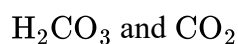
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Options:

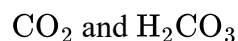
A.



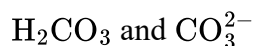
B.



C.



D.



**Answer: D**

**Solution:**

To find conjugate acid add  $\text{H}^+$  to given  $\text{HCO}_3^-$ .

$\therefore$  Conjugated acid is formed when a BronstedLowry base accepts a proton.

Hence, conjugate acid of  $\text{HCO}_3^-$  is  $\text{H}_2\text{CO}_3$ .

Conjugate base is formed when a BronstedLowry acid donates a proton.

$\therefore$  The conjugate base of  $\text{HCO}_3^-$  is  $\text{CO}_3^{2-}$ .

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## Question2

Find the concentration of sodium acetate when added to 0.1 M solution of acetic acid to form a buffer solution of pH = 5.5 ?

(  $pK_a$  of  $CH_3COOH = 4.5$  )

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**Options:**

A.

0.1 M

B.

0.01M

C.

1.0 M

D.

10.0 M

**Answer: C**

**Solution:**

$$[CH_3COOH] = [Acid] = 0.1M,$$

$$[CH_3COONa] = [Salt] = xM,$$

$$pK_a = 4.5, pH = 5.5$$

$$\text{For acidic buffer, } pH = pK_a + \log_{10} \frac{[Salt]}{[Acid]}$$

$$5.5 = 4.5 + \log_{10} \frac{x}{0.1}$$

$$\log_{10} \frac{x}{0.1} = 1$$

$$\frac{x}{0.1} = \text{Antilog}(1) = 10$$

$$x = 10 \times 0.1 = 1M$$

∴ The concentration of sodium acetate

$$[\text{CH}_3\text{COONa}] = 1.0\text{M}$$

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## Question3

Calculate the ionisation constant of  $0.08 \text{ mol dm}^{-3}$  of a monobasic acid having  $\text{pH} = 2$ .

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Options:

A.

$$3531 \times 10^{-7}$$

B.

$$2.081 \times 10^{-6}$$

C.

$$3.456 \times 10^{-8}$$

D.

$$1.25 \times 10^{-3}$$

**Answer: D**

**Solution:**

Given:

- Concentration of acid,  $C = 0.08 \text{ mol dm}^{-3}$
- $\text{pH} = 2$  (so solution is acidic)

**Step 1: Calculate  $[H^+]$**

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

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

**Step 2: Degree of dissociation ( $\alpha$ )**

The acid dissociates:  $HA \rightleftharpoons H^+ + A^-$

If total acid concentration is  $C = 0.08$  M, then:

$$[H^+] = C\alpha \implies \alpha = \frac{[H^+]}{C} = \frac{0.01}{0.08} = 0.125$$

### Step 3: Expression for $K_a$

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

Since it is monobasic:

$$[H^+] = [A^-] = 0.01 \text{ M}, \quad [HA] = C - \alpha C = 0.08 - 0.01 = 0.07$$

So,

$$K_a = \frac{(0.01)(0.01)}{0.07}$$

### Step 4: Simplify

$$K_a = \frac{1 \times 10^{-4}}{0.07} \approx 1.43 \times 10^{-3}$$

### Step 5: Match with options

Closest is:

**Option D:**  $1.25 \times 10^{-3}$

**Final Answer:**

$1.25 \times 10^{-3}$

## Question4

The solubility product of salt  $B_2 A$  is  $3.2 \times 10^{-11}$  at 298 K . What is solubility of the salt at same temperature?

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**Options:**

A.

$$5.52 \times 10^{-5} \text{ moldm}^{-3}$$

B.

$$4.92 \times 10^{-4} \text{ moldm}^{-3}$$

C.

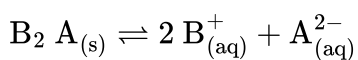
$$2.00 \times 10^{-4} \text{ moldm}^{-3}$$

D.

$$3.52 \times 10^{-5} \text{ moldm}^{-3}$$

**Answer: C**

**Solution:**



Here,  $x = 2, y = 1$

$$K_{sp} = x^x y^y S^{x+y} = (1)^1 (2)^2 S^{1+2} = 4 S^3$$

$$\therefore 3.2 \times 10^{-11} = 4 \times (S)^3$$

$$S = \sqrt[3]{8 \times 10^{-12}}$$

$$= 2.00 \times 10^{-4} \text{ moldm}^{-3}$$

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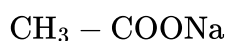
## Question5

**Which among the following salts turns red litmus blue in its aqueous solution?**

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**Options:**

A.



B.



C.



D.

KCl

**Answer: A**

**Solution:**

Sodium acetate ( $\text{CH}_3\text{COONa}$ ) is a salt of strong base ( $\text{NaOH}$ ) and weak acid ( $\text{CH}_3\text{COOH}$ ). Hence, it is a basic in nature and it turns red litmus blue.

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## Question6

**What is pH of weak dibasic acid, that is 2% dissociated in its M/100 solution at 298 K ?**

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**Options:**

A.

1.6990

B.

2.3979

C.

3.3970

D.

4.6990

**Answer: C**

**Solution:**

**Step 1: Data provided**

- Acid is **dibasic** → each molecule gives **2 H<sup>+</sup> ions** when fully dissociated.

- Concentration of acid solution =  $\frac{M}{100} = 0.01 M$ .
- Degree of dissociation =  $\alpha = 2\% = 0.02$ .

### Step 2: Calculate effective $[H^+]$

For a weak dibasic acid  $H_2A$ , when fraction  $\alpha$  dissociates:

$$[H^+] = 2 \times C \times \alpha$$

because each molecule gives 2 protons upon dissociation.

$$[H^+] = 2 \times 0.01 \times 0.02 = 0.0004 M$$

### Step 3: Calculate pH

$$pH = -\log_{10}[H^+] = -\log_{10}(4.0 \times 10^{-4})$$

$$pH = -(\log_{10} 4 + \log_{10} 10^{-4}) = -(0.60206 - 4)$$

$$pH = 3.39794 \approx 3.397$$

### Final Answer:

The pH = 3.397

Correct Option: C (3.3970)

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## Question 7

The solubility product of  $AgBr$  is  $4.9 \times 10^{-13}$  at a certain temperature. Calculate the solubility.

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Options:

A.  $4 \times 10^{-6} \text{ mol dm}^{-3}$

B.  $4 \times 10^{-7} \text{ mol dm}^{-3}$

C.  $7 \times 10^{-7} \text{ mol dm}^{-3}$

D.  $3 \times 10^{-8} \text{ mol dm}^{-3}$

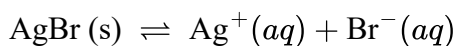
Answer: C

## Solution:

We are given:

$$K_{sp}(\text{AgBr}) = 4.9 \times 10^{-13}$$

### Step 1: Dissociation Equation



### Step 2: Expression for $K_{sp}$

If solubility =  $s$  mol  $\text{dm}^{-3}$ , then:

$$[\text{Ag}^+] = s, \quad [\text{Br}^-] = s$$

$$K_{sp} = s \times s = s^2$$

### Step 3: Solve for $s$

$$s = \sqrt{K_{sp}} = \sqrt{4.9 \times 10^{-13}}$$

$$s = \sqrt{4.9} \times 10^{-6.5}$$

$$\text{Now, } \sqrt{4.9} \approx 2.21.$$

$$\text{Also, } 10^{-6.5} = 10^{-7} \cdot \sqrt{10} \approx 3.16 \times 10^{-7}.$$

So,

$$s \approx 2.21 \times 3.16 \times 10^{-7} \approx 6.98 \times 10^{-7}$$

### Final Answer:

$$\text{Solubility} \approx 7 \times 10^{-7} \text{ mol dm}^{-3}$$

Correct option: C

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## Question8

Dissociation constant of 0.01 M weak acid is  $10^{-4}$ . What is percent dissociation of acid?

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Options:

- A. 2%
- B. 6%
- C. 10%
- D. 1.5%

**Answer: C**

## Solution:

**Acid dissociation constant problem:**

- Concentration of acid,  $c = 0.01 \text{ M} = 1 \times 10^{-2} \text{ M}$
- $K_a = 1 \times 10^{-4}$

We need **percent dissociation**,  $\alpha \times 100\%$ .

### Step 1: Formula

For a weak acid  $\text{HA} \rightleftharpoons \text{H}^+ + \text{A}^-$ ,

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

If  $\alpha$  is small, approximate denominator as 1:

$$K_a \approx c\alpha^2$$

### Step 2: Calculate $\alpha$

$$\alpha \approx \sqrt{\frac{K_a}{c}} = \sqrt{\frac{10^{-4}}{10^{-2}}} = \sqrt{10^{-2}} = 0.1$$

### Step 3: Convert to percentage

$$\% \text{ dissociation} = \alpha \times 100\% = 0.1 \times 100 = 10\%$$

 **Final Answer:**

**Option C — 10%**

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## Question9

**Which among the following salts forms basic solution when dissolved in water?**

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## Options:

- A.  $\text{NaNO}_3$
- B.  $\text{CH}_3\text{COONH}_4$
- C. KCN
- D.  $\text{NH}_4\text{F}$

**Answer: C**

## Solution:

### Step 1: Recall the rule

- A salt solution can be **acidic, basic, or neutral** depending on the strength of the acid and base from which it is formed.
- Strong acid + Strong base  $\rightarrow$  Neutral
- Strong acid + Weak base  $\rightarrow$  Acidic
- Weak acid + Strong base  $\rightarrow$  Basic
- Weak acid + Weak base  $\rightarrow$  Depends on the relative strengths of conjugate acid and conjugate base.

### Step 2: Analyze each option

#### (A) $\text{NaNO}_3$

$\text{NaOH}$  (strong base) +  $\text{HNO}_3$  (strong acid).

$\rightarrow$  Neutral solution. Not basic.

#### (B) $\text{CH}_3\text{COONH}_4$

$\text{CH}_3\text{COO}^-$  (from weak acid  $\text{CH}_3\text{COOH}$ ) +  $\text{NH}_4^+$  (from weak base  $\text{NH}_3$ ).

Both weak  $\rightarrow$  The pH depends on relative  $K_a$  and  $K_b$ .

- $\text{CH}_3\text{COOH}$  has  $K_a \sim 1.8 \times 10^{-5}$
- $\text{NH}_4^+$  ( $K_a \sim 5.6 \times 10^{-10}$ )

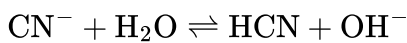
Here, acetate is a stronger conjugate base than ammonium is a conjugate acid.

So the solution is slightly **basic or close to neutral** but not strongly so. Usually, it's considered **slightly acidic or near neutral**, not distinctly basic. (**Not a strong candidate**)

**(C) KCN**

$K^+$  (from KOH, strong base)  $\rightarrow$  neutral.

$CN^-$  (from weak acid HCN)  $\rightarrow$  conjugate base, undergoes hydrolysis:



$\rightarrow$  Solution is **basic**.

**(D)  $NH_4F$**

$NH_4^+$  (from weak base  $NH_3$ ) = conjugate acid  $\rightarrow$  acidic.

$F^-$  (from weak acid HF) = conjugate base  $\rightarrow$  basic.

Compare  $K_a$  values:  $K_a(NH_4^+) \approx 5.6 \times 10^{-10}$ ;  $K_a(HF) \approx 7.2 \times 10^{-4} \rightarrow$  So  $F^-$  is a much weaker base compared to the acidity of  $NH_4^+$ .

Thus solution will be **acidic**, not basic.

**Final Answer:**

Option C: KCN

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## Question10

**Calculate the pH of centimolar solution of monoacidic weak base. Which is 10% dissociated in its aqueous solution?**

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**Options:**

A. 9

B. 10

C. 11

D. 12

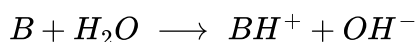
**Answer: C**

**Solution:**

### Step 1: Interpret the problem

- "Centimolar solution" means  $0.01\text{ M}$  (since centi =  $10^{-2}$ ).
- Base is monoacidic (accepts only one proton).
- Dissociation = 10%.

Reaction:



### Step 2: Concentration of $OH^-$

Initial conc of base =  $0.01\text{ M}$ .

Dissociation = 10% = 0.10.

So,

$$[OH^-] = 0.01 \times 0.10 = 0.001\text{ M} = 1.0 \times 10^{-3}\text{ M}$$

### Step 3: pOH

$$pOH = -\log[OH^-] = -\log(1.0 \times 10^{-3}) = 3$$

### Step 4: pH

$$pH = 14 - pOH = 14 - 3 = 11$$

**Final Answer:**

11

So the correct option is C. 11.

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## Question11

**A monobasic weak acid dissociates 2% in its 0.002 M solution. Calculate the dissociation constant of weak acid.**

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**Options:**

A.  $2 \times 10^{-9}$

B.  $8 \times 10^{-7}$

C.  $6 \times 10^{-7}$

D.  $4 \times 10^{-6}$

**Answer: B**

## **Solution:**

### **Step 1: Write given values**

- Initial concentration of acid,  $C = 0.002 \text{ M}$
- Degree of dissociation,  $\alpha = \frac{2}{100} = 0.02$

### **Step 2: Formula for dissociation constant**

For a weak acid  $HA \rightleftharpoons H^+ + A^-$ ,

$$K_a = C\alpha^2$$

*(Approximate because  $\alpha$  is small, so equilibrium acid concentration  $\approx C$ ).*

### **Step 3: Calculation**

$$K_a = 0.002 \times (0.02)^2$$

$$K_a = 0.002 \times 0.0004$$

$$K_a = 8 \times 10^{-7}$$

 **Final Answer:**

The dissociation constant is

$8 \times 10^{-7}$
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**Correct Option: B**

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## **Question12**

**Which among the following salts turns blue litmus red in its aqueous solution?**

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## Options:

A. KCN

B.  $\text{Na}_2\text{CO}_3$

C.  $\text{NaNO}_3$

D.  $\text{CuCl}_2$

**Answer: D**

## Solution:

**Step 1: Consider each option.**

**(A) KCN**

KCN consists of  $\text{K}^+$  (neutral, from a strong base KOH) and  $\text{CN}^-$  (conjugate base of HCN, a weak acid).

→  $\text{CN}^-$  undergoes hydrolysis to form  $\text{OH}^-$ .

Result: Solution is basic. So it turns red litmus blue (not the asked case).

**(B)  $\text{Na}_2\text{CO}_3$**

$\text{Na}^+$  (neutral, from strong base) and  $\text{CO}_3^{2-}$  (conjugate base of  $\text{H}_2\text{CO}_3$ , weak acid).

→  $\text{CO}_3^{2-}$  hydrolyzes to form  $\text{OH}^-$ .

Result: Basic solution. Not acidic.

**(C)  $\text{NaNO}_3$**

$\text{Na}^+$  (neutral) and  $\text{NO}_3^-$  (conjugate base of  $\text{HNO}_3$ , a strong acid).

→ No hydrolysis, solution is neutral.

Does not affect litmus.

**(D)  $\text{CuCl}_2$**

$\text{Cu}^{2+}$  (cation of weak base,  $\text{Cu}(\text{OH})_2$ ) tends to hydrolyze in water:

- Produces  $\text{H}^+$  ions.

→ An acidic solution is formed.

Thus  $\text{CuCl}_2$  solution turns **blue litmus red**.

**✓ Correct answer: Option D —  $\text{CuCl}_2$**

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## Question13

The solubility product of the sparingly soluble salt  $AB_2$  is  $2.56 \times 10^{-10}$  at 298 K . Calculate its solubility in  $\text{mol dm}^{-3}$  at the same temperature?

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Options:

A.  $1 \times 10^{-4}$

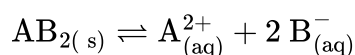
B.  $2 \times 10^{-4}$

C.  $4 \times 10^{-4}$

D.  $3 \times 10^{-2}$

Answer: C

Solution:



$$x = 1, y = 2$$

$$K_{sp} = x^x y^y S^{x+y} = (1)^1 (2)^2 S^{1+2} = 4 S^3$$

The molar solubility S of  $AB_2$  is given by

$$\begin{aligned} S &= \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{2.56 \times 10^{-10}}{4}} \\ &= \sqrt[3]{\frac{256 \times 10^{-12}}{4}} = \sqrt[3]{64 \times 10^{-12}} \\ &= 4 \times 10^{-4} \text{ mol dm}^{-3} \end{aligned}$$

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## Question14

Calculate the value of dissociation constant of weak acid, which dissociates to 0.01% in its 0.1 M solution?

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**Options:**

A.  $10^{-3}$

B.  $10^{-4}$

C.  $10^{-5}$

D.  $10^{-9}$

**Answer: D**

**Solution:**

- Degree of dissociation =  $0.01\% = 0.01/100 = 1 \times 10^{-4}$
- Concentration  $C = 0.1 \text{ M}$

**Step 1: Relation between  $K_a$ ,  $\alpha$ , and  $C$**

$$K_a = C\alpha^2$$

(for weak acids where  $\alpha \ll 1$ ).

**Step 2: Substitute the values**

$$\alpha = 1 \times 10^{-4}, \quad C = 0.1$$

$$K_a = 0.1 \times (1 \times 10^{-4})^2$$

$$K_a = 0.1 \times 10^{-8}$$

$$K_a = 1 \times 10^{-9}$$

 **Final Answer:**

$$\boxed{10^{-9}}$$

**Correct Option: D**

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## Question15

**The pH of monoacidic base is 10 . Calculate its percentage dissociation in 0.01 M solution at 298 K ?**

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**Options:**

A. 10%

B. 5%

C. 2%

D. 1%

**Answer: D**

**Solution:**

**Step 1: Understand the problem**

- A "monoacidic base" = base that accepts only one proton → contributes one  $\text{OH}^-$  ion per molecule when fully dissociated.
- Given  $\text{pH} = 10 \rightarrow$  then  $\text{pOH} = 14 - 10 = 4$ .
- So,  $[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-4} \text{ M}$ .

**Step 2: Relating to dissociation**

The base concentration is  $C = 0.01 \text{ M}$ .

If it dissociates to produce  $[\text{OH}^-]$ , then:

$$\alpha = \frac{[\text{OH}^-]}{C}$$

where  $\alpha$  = degree of dissociation.

$$\alpha = \frac{1.0 \times 10^{-4}}{0.01} = \frac{1.0 \times 10^{-4}}{1.0 \times 10^{-2}} = 0.01$$

**Step 3: Convert to percentage**

% dissociation =  $\alpha \times 100 = 0.01 \times 100 = 1\%$ .

**✔ Final Answer:**

**1%**

So the correct option is **D. 1%**.

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## Question16

**Which among the following salts forms basic solution when dissolved in water?**

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**Options:**

- A.  $\text{NH}_4\text{NO}_3$
- B.  $\text{Na}_2\text{CO}_3$
- C.  $\text{NaNO}_3$
- D.  $\text{CuSO}_4$

**Answer: B**

**Solution:**

- (A)  $\text{NH}_4\text{NO}_3$  : Salt of strong acid and weak base: Solution is acidic.
  - (B)  $\text{Na}_2\text{CO}_3$  : Salt of weak acid and strong base: Solution is basic.
  - (C)  $\text{NaNO}_3$  : Salt of strong acid and strong base: Solution is neutral.
  - (D)  $\text{CuSO}_4$  : Salt of strong acid and weak base: Solution is acidic.
- 

## Question17

**Calculate the equilibrium concentration of  $\text{Pb}^{++}$  ions in a solution of  $\text{PbS}$  containing  $1 \times 10^{-11} \text{ mol dm}^{-3}$  of sulphide ions.**

**(Given  $K_{\text{sp}}$  for  $\text{PbS} = 8.0 \times 10^{-28}$  )**

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**Options:**

A.  $4 \times 10^{-14}$

B.  $4 \times 10^{-17}$

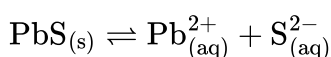
C.  $8 \times 10^{-17}$

D.  $8 \times 10^{-11}$

**Answer: C**

**Solution:**

**Step 1: Write the dissociation equation for PbS**



**Step 2: Write the solubility product ( $K_{sp}$ ) expression**

$$K_{sp} = [\text{Pb}^{2+}][\text{S}^{2-}]$$

**Step 3: Substitute the given values**

We are told that  $[\text{S}^{2-}] = 1 \times 10^{-11} \text{ mol dm}^{-3}$  and  $K_{sp} = 8.0 \times 10^{-28}$ .

Put these values into the  $K_{sp}$  formula:

$$8.0 \times 10^{-28} = [\text{Pb}^{2+}] \times 1 \times 10^{-11}$$

**Step 4: Solve for  $[\text{Pb}^{2+}]$**

Divide both sides by  $1 \times 10^{-11}$ :

$$[\text{Pb}^{2+}] = \frac{8.0 \times 10^{-28}}{1 \times 10^{-11}} = 8 \times 10^{-17}$$

**Final Answer:**

The equilibrium concentration of  $\text{Pb}^{2+}$  ions in the solution is  $8 \times 10^{-17} \text{ mol dm}^{-3}$ .

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## Question 18

**Which among the following salts forms basic solution in water?**

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### Options:

A.  $\text{NH}_4\text{NO}_3$

B.  $\text{NH}_4\text{Cl}$

C.  $\text{Na}_2\text{CO}_3$

D.  $\text{Na}_2\text{SO}_4$

**Answer: C**

### Solution:

#### Step 1: Analyse each option

**Option A:**  $\text{NH}_4\text{NO}_3$

- This salt is formed from a **weak base** ( $\text{NH}_4\text{OH}$ ) and a **strong acid** ( $\text{HNO}_3$ ).
- Solution is **acidic** due to  $\text{NH}_4^+$  hydrolysis.

Acidic, not basic.

**Option B:**  $\text{NH}_4\text{Cl}$

- This salt is formed from a **weak base** ( $\text{NH}_4\text{OH}$ ) and a **strong acid** ( $\text{HCl}$ ).
- Solution is **acidic** due to  $\text{NH}_4^+$ .

Acidic, not basic.

**Option C:**  $\text{Na}_2\text{CO}_3$

- This salt is formed from a **strong base** ( $\text{NaOH}$ ) and a **weak acid** ( $\text{H}_2\text{CO}_3$ ).
- In water,  $\text{CO}_3^{2-}$  hydrolyzes:  
$$\text{CO}_3^{2-} + \text{H}_2\text{O} \leftrightarrow \text{HCO}_3^- + \text{OH}^-$$
- Produces  **$\text{OH}^-$  ions**, making the solution **basic**.

This is **basic**.

**Option D:**  $\text{Na}_2\text{SO}_4$

- This salt is formed from a **strong base** ( $\text{NaOH}$ ) and a **strong acid** ( $\text{H}_2\text{SO}_4$ ).
- Hydrolysis is negligible; solution is **neutral**.

Neutral.

Correct Answer:

Option C:  $\text{Na}_2\text{CO}_3$

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## Question 19

A weak base is 5% dissociated in its 0.01 M solution. Calculate the dissociation constant.

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Options:

A.  $3.5 \times 10^{-6}$

B.  $2.0 \times 10^{-5}$

C.  $2.3 \times 10^{-4}$

D.  $2.5 \times 10^{-5}$

**Answer: D**

**Solution:**

**Step 1. Recall the formula:**

For a weak base,

$$K_b = C \cdot \alpha^2$$

Where:

- $C$  is the initial concentration,
- $\alpha$  is the degree of dissociation.

**Step 2. Values given:**

- $C = 0.01 \text{ M}$
- $\alpha = \frac{5}{100} = 0.05$

**Step 3. Substitute:**

$$K_b = (0.01)(0.05^2)$$

$$K_b = 0.01 \times 0.0025$$

$$K_b = 2.5 \times 10^{-5}$$

**Step 4. Match with option:**

The correct option is:

**Option D:**  $2.5 \times 10^{-5}$

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## Question20

**Solubility of binary sparingly soluble salt is  $1.12 \times 10^{-4} \text{ g/dm}^3$ .  
Calculate its solubility product (molar mass of salt =  $112 \text{ g mol}^{-1}$ )**

**MHT CET 2025 23rd April Morning Shift**

**Options:**

A.  $1 \times 10^{-3}$

B.  $1 \times 10^{-6}$

C.  $1 \times 10^{-9}$

D.  $1 \times 10^{-12}$

**Answer: D**

**Solution:**

We are given:

- Solubility of a sparingly soluble binary salt =  $1.12 \times 10^{-4} \text{ g/dm}^3$ .
- Molar mass of salt =  $112 \text{ g mol}^{-1}$ .

The salt is a **binary sparingly soluble salt**, i.e. of type  $AB \rightleftharpoons A^+ + B^-$ .

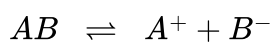
**Step 1: Convert solubility to molar solubility**

$$\text{Molar solubility} = \frac{\text{given solubility in g/dm}^3}{\text{molar mass}} = \frac{1.12 \times 10^{-4}}{112}$$

$$= 1.0 \times 10^{-6} \text{ mol/dm}^3$$

### Step 2: Write dissociation

For a binary salt  $AB$ :



If solubility =  $S$ , then at equilibrium:

$$[A^+] = S, [B^-] = S.$$

So,

$$K_{sp} = [A^+][B^-] = S^2$$

### Step 3: Calculate solubility product

$$K_{sp} = (1.0 \times 10^{-6})^2 = 1.0 \times 10^{-12}$$

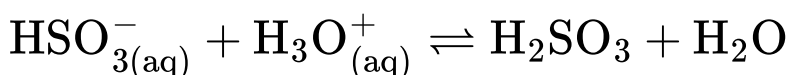
 **Final Answer:**

$$\boxed{1 \times 10^{-12}} \quad (\text{Option D})$$

---

## Question21

**Identify conjugate acid-base pair from following equilibrium reaction.**



### MHT CET 2025 22nd April Evening Shift

**Options:**

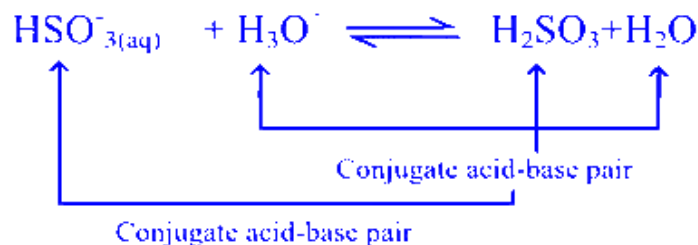
- A.  $\text{H}_2\text{SO}_3$  and  $\text{HSO}_3^-$
- B.  $\text{HSO}_3^-$  and  $\text{H}_3\text{O}^+$
- C.  $\text{H}_2\text{SO}_3$  and  $\text{H}_2\text{O}$
- D.  $\text{H}_3\text{O}^+$  and  $\text{H}_2\text{SO}_3$

**Answer: A**

## Solution:

Conjugate acid is formed when a Bronsted. Lowry base accepts a proton.

Conjugate base is formed when a Bronsted. Lowry acid donates a proton.



## Question22

**A weak monoacidic base dissociates to 1.5% in 0.001 M solution at 298 K . Calculate the dissociation constant of weak base.**

**MHT CET 2025 22nd April Evening Shift**

**Options:**

A.  $2.25 \times 10^{-7}$

B.  $3.05 \times 10^{-7}$

C.  $2.5 \times 10^{-5}$

D.  $3.725 \times 10^{-6}$

**Answer: A**

## Solution:

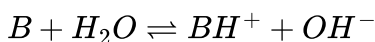
We are tasked to calculate the dissociation constant  $K_b$  of a weak **monoacidic base**.

**Step 1: Information given**

- Base concentration:  $C = 0.001 \text{ M}$
- Degree of dissociation:  $\alpha = 1.5\% = 0.015$

## Step 2: Expression for base dissociation constant

For a weak base:



The dissociation constant ( $K_b$ ) is:

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

With degree of dissociation:

- Initial concentration of base:  $C$
- At equilibrium:
- $[OH^-] = C\alpha$
- $[BH^+] = C\alpha$
- $[B] = C(1 - \alpha) \approx C$  (since  $\alpha \ll 1$ )

Thus:

$$K_b \approx \frac{C\alpha^2}{1} = C\alpha^2$$

## Step 3: Substitution

$$K_b = 0.001 \times (0.015)^2$$

$$K_b = 0.001 \times 0.000225$$

$$K_b = 2.25 \times 10^{-7}$$

 **Final Answer:**

$$\boxed{2.25 \times 10^{-7}} \quad (\text{Option A})$$

---

## Question23

The solubility product of NiS is  $4.9 \times 10^{-5}$  at 298 K . Calculate its solubility in  $\text{mol dm}^{-3}$  at the same temperature?

**MHT CET 2025 22nd April Evening Shift**

**Options:**

A.  $1.69 \times 10^{-3}$

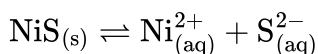
B.  $7.0 \times 10^{-3}$

C.  $2.45 \times 10^{-3}$

D.  $6.18 \times 10^{-3}$

**Answer: B**

**Solution:**



$$x = 1, y = 1$$

$$K_{sp} = x^x y^y S^{x+y} = (1)^1 (1)^1 S^{1+1} = S^2$$

The molar solubility S of NiS is given by

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

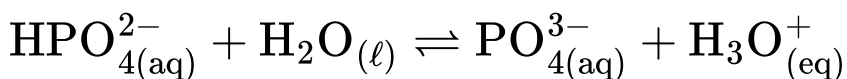
$$= \sqrt{4.9 \times 10^{-5}}$$

$$S = 7.0 \times 10^{-3} \text{ mol dm}^{-3}$$

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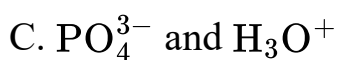
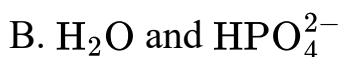
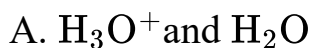
## Question24

**Identify the conjugate acid-base pair respectively from following equilibrium reaction.**



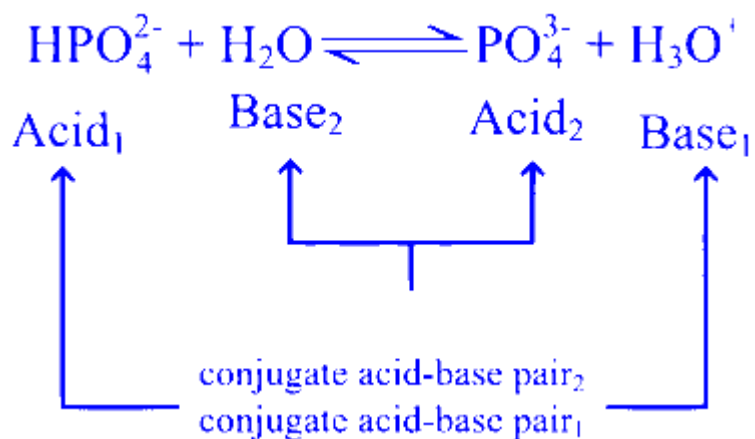
**MHT CET 2025 22nd April Morning Shift**

**Options:**



**Answer: A**

**Solution:**



## Question25

The solubility of sparingly soluble salt  $\text{AX}_2$  is  $1 \times 10^{-4} \text{ mol dm}^{-3}$  at 298 K . Calculate its solubility product.

**MHT CET 2025 22nd April Morning Shift**

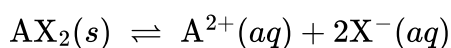
**Options:**

- A.  $2 \times 10^{-12}$
- B.  $4 \times 10^{-12}$
- C.  $2 \times 10^{-10}$
- D.  $4 \times 10^{-10}$

**Answer: B**

**Solution:**

**Step 1: Write the dissociation reaction**



**Step 2: Represent solubility**

If solubility =  $s = 1 \times 10^{-4} \text{ mol dm}^{-3}$ :

- Concentration of  $A^{2+} = s = 1 \times 10^{-4}$
- Concentration of  $X^{-} = 2s = 2 \times 10^{-4}$

### Step 3: Define solubility product

$$K_{sp} = [A^{2+}][X^{-}]^2$$

Substitute values:

$$K_{sp} = (1.0 \times 10^{-4}) \times (2.0 \times 10^{-4})^2$$

### Step 4: Calculation

$$K_{sp} = (1.0 \times 10^{-4}) \times (4.0 \times 10^{-8})$$

$$K_{sp} = 4.0 \times 10^{-12}$$

 **Final Answer:**

**Option B:**  $4 \times 10^{-12}$

---

## Question26

**Calculate the value of dissociation constant of weak monoacidic base if it dissociates to 2% in 0.1 M solution?**

**MHT CET 2025 22nd April Morning Shift**

**Options:**

A.  $6 \times 10^{-5}$

B.  $4 \times 10^{-5}$

C.  $2 \times 10^{-5}$

D.  $1 \times 10^{-5}$

**Answer: B**

**Solution:**

### Step 1: Recall formula for dissociation constant

For a weak electrolyte:

$$K_b = C\alpha^2$$

where

- $C$  = concentration of the base (M)
- $\alpha$  = degree of dissociation

### Step 2: Insert given values

- Concentration  $C = 0.1 \text{ M}$
- Degree of dissociation =  $2\% = 0.02$

So,

$$K_b = 0.1 \times (0.02)^2$$

### Step 3: Calculate

$$(0.02)^2 = 0.0004 = 4 \times 10^{-4}$$

$$K_b = 0.1 \times 4 \times 10^{-4} = 4 \times 10^{-5}$$

**Final Answer:**

$$K_b = 4 \times 10^{-5}$$

**Correct Option: B.**

---

## Question27

Which from following formulae is used to find the  $[\text{OH}^-]$  ion concentration of a weak monoacidic base?

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**Options:**

A.  $K_b \cdot c$

B.  $\sqrt{K_b \cdot c}$

C.  $\sqrt{\frac{K_b}{c}}$

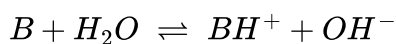
D.  $\sqrt{K_b}$

**Answer: B**

## Solution:

### Step 1: Recall equilibrium concept

For a weak base  $B$  (monoacidic, i.e., accepts one proton):



Let the initial concentration of the base be  $c$ . Suppose degree of ionisation =  $\alpha$ .

Then equilibrium  $[OH^-] = c\alpha$ .

### Step 2: Apply definition of $K_b$

$$K_b = \frac{[BH^+][OH^-]}{[B]}.$$

At equilibrium:

- $[OH^-] = c\alpha$ ,
- $[BH^+] = c\alpha$ ,
- $[B] = c(1 - \alpha) \approx c$  (since base is weak,  $\alpha$  is small).

So:

$$K_b \approx \frac{(c\alpha)(c\alpha)}{c} = c\alpha^2.$$

### Step 3: Solve for $\alpha$

$$\alpha = \sqrt{\frac{K_b}{c}}.$$

So,

$$[OH^-] = c\alpha = c \cdot \sqrt{\frac{K_b}{c}} = \sqrt{K_b \cdot c}.$$

### Step 4: Match with options

The correct formula is:

$$[OH^-] = \sqrt{K_b \cdot c}.$$

That corresponds to **Option B**.

**✓ Final Answer: Option B**

---

## Question28

A monobasic weak acid dissociates to 1.2% in its 0.01 M solution at 298 K . Calculate dissociation constant of it.

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Options:

- A.  $1.04 \times 10^{-8}$
- B.  $1.44 \times 10^{-6}$
- C.  $1.30 \times 10^{-6}$
- D.  $1.18 \times 10^{-5}$

Answer: B

**Solution:**

**Step 1: Write the data**

- Weak acid concentration:  $c = 0.01 \text{ M}$
- Degree of dissociation:  $\alpha = 1.2\% = 0.012$

**Step 2: Formula for dissociation constant**

For a weak monobasic acid  $HA \rightleftharpoons H^+ + A^-$ :

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

**Step 3: Approximation**

Since  $\alpha = 0.012$ , which is much less than 1,

$$1 - \alpha \approx 0.988 \approx 1$$

So,

$$K_a \approx c\alpha^2$$

**Step 4: Calculation**

$$K_a \approx (0.01)(0.012)^2$$

$$= 0.01 \times 0.000144 = 1.44 \times 10^{-6}$$

**Final Answer:**

$$\boxed{1.44 \times 10^{-6}}$$

**Correct Option: B**

---

## Question29

The solubility of calcium carbonate at 298 K is  $6.4 \times 10^{-5} \text{ mol dm}^{-3}$ . Calculate the value of solubility product at the same temperature?

**MHT CET 2025 21st April Evening Shift**

**Options:**

A.  $5.06 \times 10^{-10}$

B.  $4.096 \times 10^{-9}$

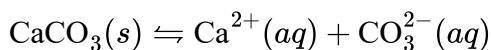
C.  $3.05 \times 10^{-10}$

D.  $2.8 \times 10^{-9}$

**Answer: B**

**Solution:**

**Step 1: Write the dissociation equilibrium.**



**Step 2: Define solubility.**

Let solubility of  $\text{CaCO}_3$  be  $s = 6.4 \times 10^{-5} \text{ mol dm}^{-3}$ .

Then at equilibrium:

$$[\text{Ca}^{2+}] = s = 6.4 \times 10^{-5} \text{ M}$$

$$[\text{CO}_3^{2-}] = s = 6.4 \times 10^{-5} \text{ M}$$

**Step 3: Expression for solubility product.**

$$K_{sp} = [\text{Ca}^{2+}] \cdot [\text{CO}_3^{2-}] = (s)(s) = s^2$$

**Step 4: Substitute value.**

$$K_{sp} = (6.4 \times 10^{-5})^2$$

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

$$= 40.96 \times 10^{-10}$$

$$= 4.096 \times 10^{-9}$$

**Final Answer:**

$$\boxed{4.096 \times 10^{-9}}$$

Correct option: **B**

---

## Question30

The solubility product of  $\text{PbI}_2$  is  $1.08 \times 10^{-7}$ .

Calculate its solubility in  $\text{mol dm}^{-3}$  at 298 K .

**MHT CET 2025 21st April Morning Shift**

**Options:**

A.  $2.018 \times 10^{-3}$

B.  $2.011 \times 10^{-9}$

C.  $1.259 \times 10^{-9}$

D.  $3.0 \times 10^{-3}$

**Answer: D**

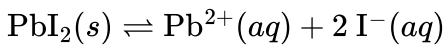
**Solution:**

Given:

Solubility product of  $\text{PbI}_2$ ,  $K_{sp} = 1.08 \times 10^{-7}$

Let the solubility of  $\text{PbI}_2$  be  $S$  mol/L.

**Step 1: Write the dissociation equation.**



**Step 2: Express ions' concentration in terms of  $S$ .**

- $[\text{Pb}^{2+}] = S$
- $[\text{I}^{-}] = 2S$

**Step 3: Write  $K_{sp}$  expression.**

$$K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2$$

$$K_{sp} = (S) \times (2S)^2 = S \times 4S^2 = 4S^3$$

**Step 4: Solve for  $S$ .**

$$K_{sp} = 4S^3$$

$$S^3 = \frac{K_{sp}}{4}$$

$$S^3 = \frac{1.08 \times 10^{-7}}{4}$$

$$S^3 = 2.7 \times 10^{-8}$$

$$S = \sqrt[3]{2.7 \times 10^{-8}}$$

**Step 5: Calculate the cube root.**

First, write  $2.7 \times 10^{-8}$  as:

$$S = (\sqrt[3]{2.7}) \times (\sqrt[3]{10^{-8}})$$

$$\sqrt[3]{2.7} \approx 1.4$$

$$\sqrt[3]{10^{-8}} = 10^{-8/3} = 10^{-2.666\dots} \approx 2.15 \times 10^{-3}$$

But the easiest is:

$$S \approx \sqrt[3]{2.7 \times 10^{-8}}$$

$$= 1.4 \times 10^{-8/3}$$

$$= 1.4 \times 10^{-2.667}$$

$$= 1.4 \times (10^{-2} \times 10^{-0.667})$$

$$10^{-0.667} \approx 0.215$$

$$S \approx 1.4 \times 0.215 \times 10^{-2}$$

$$= 0.301 \times 10^{-2}$$

$$= 3.01 \times 10^{-3}$$

**Step 6: Match with options.**

So, the solubility is

$$3.0 \times 10^{-3} \text{ mol dm}^{-3}$$

Option D is correct.

---

## Question31

**What is the value of pH of a NaOH solution that dissociates 2% in its 0.01 M solution?**

**MHT CET 2025 21st April Morning Shift**

**Options:**

A. 9.704

B. 10.301

C. 8.621

D. 8.750

**Answer: B**

**Solution:**

Given:

- Molarity of NaOH ( $C$ ) = 0.01 M
- Percentage dissociation = 2%

**Step 1: Calculate the concentration of  $\text{OH}^-$  ions formed.**

NaOH dissociates as:



The concentration of  $\text{OH}^-$  ions produced is:

$$[\text{OH}^-] = \text{Molarity} \times \frac{\text{Percentage dissociation}}{100}$$

So,

$$[\text{OH}^-] = 0.01 \times \frac{2}{100} = 0.01 \times 0.02 = 0.0002 \text{ M}$$

**Step 2: Calculate  $pOH$ .**

$$pOH = -\log_{10}[\text{OH}^-]$$

Substitute the value:

$$pOH = -\log_{10}(0.0002)$$

$$0.0002 = 2 \times 10^{-4}$$

So,

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

$$= -\log_{10} 2 - \log_{10} 10^{-4}$$

$$= -0.3010 + 4$$

$$= 3.699$$

**Step 3: Calculate pH.**

At  $25^\circ\text{C}$ ,

$$pH + pOH = 14$$

So,

$$pH = 14 - pOH = 14 - 3.699 = 10.301$$

**Final Answer:**

Option B is correct.

10.301

---

## Question32

**Which among the following salts turns red litmus blue in its aqueous solution?**

**MHT CET 2025 20th April Evening Shift**

**Options:**

A. KCN

B.  $\text{NaNO}_3$

C. NaCl

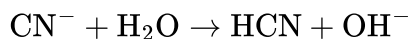
D. KCl

**Answer: A**

### **Solution:**

KCN (Potassium cyanide) is a salt of a strong base (KOH) and a weak acid (HCN). When KCN dissolves in water, it hydrolyses to form a basic solution.

Hydrolysis reaction:



The production of  $\text{OH}^-$  ions makes the solution basic, which turns red litmus blue.

**Other options:**

- $\text{NaNO}_3$ ,  $\text{NaCl}$ , and  $\text{KCl}$  are all salts of strong acids and strong bases. Their aqueous solutions are neutral and do not affect litmus.

**Correct Answer:**

**Option A: KCN**

---

## **Question33**

**4 gram of NaOH is added in water to form 500 mL solution at 298 K . What is pH of solution? (Molar mass of NaOH =  $40 \text{ g mol}^{-1}$  )**

**MHT CET 2025 20th April Evening Shift**

**Options:**

- A. 8.6990
- B. 10.3010
- C. 10.6990
- D. 13.3010

**Answer: D**

**Solution:**

Given:

- Mass of NaOH added = 4 g
- Volume of solution = 500 mL = 0.5 L
- Molar mass of NaOH = 40 g/mol

**Step 1: Calculate moles of NaOH**

$$\text{Moles of NaOH} = \frac{4 \text{ g}}{40 \text{ g/mol}} = 0.1 \text{ mol}$$

**Step 2: Calculate molarity (M) of NaOH solution**

$$\text{Molarity, } M = \frac{\text{moles of solute}}{\text{volume of solution in litres}} = \frac{0.1}{0.5} = 0.2 \text{ mol/L}$$

**Step 3: Calculate  $[OH^-]$**

NaOH is a strong base, so it dissociates completely:



Therefore,

$$[OH^-] = 0.2 \text{ mol/L}$$

**Step 4: Calculate  $pOH$**

$$pOH = -\log_{10}[OH^-]$$

$$pOH = -\log_{10}(0.2)$$

$$0.2 = 2 \times 10^{-1}$$

$$\log_{10} 0.2 = \log_{10} 2 + \log_{10} 10^{-1} = 0.3010 - 1 = -0.6990$$

So,

$$pOH = -(-0.6990) = 0.6990$$

**Step 5: Calculate pH using the relation**

At 298 K,

$$pH + pOH = 14$$

$$pH = 14 - pOH = 14 - 0.6990 = 13.3010$$

**Final Answer:**

$$\boxed{13.3010} \text{ (Option D)}$$

---

## Question34

The solubility of salt  $BA_2$  is  $4 \times 10^{-4} \text{ moldm}^{-3}$ .

**What is solubility product of the salt?**

**MHT CET 2025 20th April Evening Shift**

**Options:**

A.  $1.55 \times 10^{-8}$

B.  $2.56 \times 10^{-10}$

C.  $3.60 \times 10^{-8}$

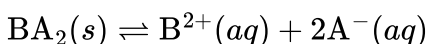
D.  $6.41 \times 10^{-10}$

**Answer: B**

**Solution:**

Let the salt be  $BA_2$ .

When  $BA_2$  dissolves in water, it dissociates as:



Let the solubility of  $BA_2$  be  $s = 4 \times 10^{-4} \text{ mol dm}^{-3}$ .

**Concentrations at equilibrium:**

- $[B^{2+}] = s = 4 \times 10^{-4} \text{ mol dm}^{-3}$
- $[A^{-}] = 2s = 2 \times 4 \times 10^{-4} = 8 \times 10^{-4} \text{ mol dm}^{-3}$

**Expression for solubility product  $K_{sp}$ :**

$$K_{sp} = [B^{2+}][A^{-}]^2$$

**Substitute the values:**

$$\begin{aligned} K_{sp} &= (4 \times 10^{-4}) \times (8 \times 10^{-4})^2 \\ &= (4 \times 10^{-4}) \times (8^2 \times (10^{-4})^2) \\ &= (4 \times 10^{-4}) \times (64 \times 10^{-8}) \\ &= 4 \times 64 \times 10^{-4} \times 10^{-8} \\ &= 256 \times 10^{-12} \\ &= 2.56 \times 10^{-10} \end{aligned}$$

**Correct answer:**

**Option B**

$$2.56 \times 10^{-10}$$

---

## Question35

Solubility of  $\text{Ca}_3(\text{PO}_4)_2$  is ' S '  $\text{mol dm}^{-3}$ . Find solubility product.

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Options:

A.  $S^5$

B.  $108 S^5$

C.  $54 S^5$

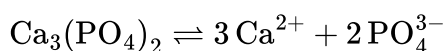
D.  $12 S^5$

Answer: B

Solution:

**Step 1: Write the Dissolution Equation**

When  $\text{Ca}_3(\text{PO}_4)_2$  dissolves in water, it breaks into ions like this:



**Step 2: Relate Solubility to Ion Concentrations**

Let  $S$  be the solubility in  $\text{mol dm}^{-3}$ . This means:

- For every 1 mole of salt that dissolves, 3 moles of  $\text{Ca}^{2+}$  form. So  $[\text{Ca}^{2+}] = 3S$ .
- For every 1 mole of salt, 2 moles of  $\text{PO}_4^{3-}$  form. So  $[\text{PO}_4^{3-}] = 2S$ .

**Step 3: Write the Expression for Solubility Product ( $K_{\text{sp}}$ )**

The solubility product is:  $K_{\text{sp}} = [\text{Ca}^{2+}]^3 \times [\text{PO}_4^{3-}]^2$  Substitute the values from Step 2:  $K_{\text{sp}} = (3S)^3 \times (2S)^2$

**Step 4: Calculate and Simplify**

$$(3S)^3 = 27S^3 \text{ and } (2S)^2 = 4S^2 \text{ so: } K_{\text{sp}} = 27S^3 \times 4S^2 = 108S^5$$

---

## Question36

The solubility of AgBr is  $7.1 \times 10^{-7} \text{ mol dm}^{-3}$ . Calculate its solubility product at the same temperature.

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Options:

A.  $7.08 \times 10^{-13}$

B.  $3.67 \times 10^{-13}$

C.  $5.89 \times 10^{-13}$

D.  $5.04 \times 10^{-13}$

Answer: D

Solution:



$$\therefore x = 1, y = 1$$

$$\therefore K_{sp} = x^x y^y S^{x+y} = (1)^1 (1)^1 S^{1+1} = S^2$$

$$\therefore K_{sp} = (7.1 \times 10^{-7})^2 = 5.04 \times 10^{-13}$$

---

## Question37

What is the pH of buffer solution formed by mixing 0.01 M acetic acid and 0.05 M sodium acetate? ( $pK_a = 4.7447$ )

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Options:

A. 2.80

B. 3.00

C. 5.44

D. 6.5

**Answer: C**

### **Solution:**

For a buffer solution made of a weak acid (acetic acid,  $\text{CH}_3\text{COOH}$ ) and its salt (sodium acetate,  $\text{CH}_3\text{COONa}$ ), we use the **Henderson-Hasselbalch equation**:

$$\text{pH} = \text{p}K_a + \log \left( \frac{[\text{Salt}]}{[\text{Acid}]} \right)$$

Given:

- $[\text{Salt}] = 0.05 \text{ M}$
- $[\text{Acid}] = 0.01 \text{ M}$
- $\text{p}K_a = 4.7447$

**Step 1: Write the equation with values:**

$$\text{pH} = 4.7447 + \log \left( \frac{0.05}{0.01} \right)$$

**Step 2: Calculate the ratio:**

$$\frac{0.05}{0.01} = 5$$

**Step 3: Find the logarithm:**

$$\log 5 \approx 0.699$$

**Step 4: Substitute the values:**

$$\text{pH} = 4.7447 + 0.699 = 5.4437 \approx 5.44$$

**Final Answer:**

5.44 (Option C)

---

## **Question38**

**What is the value of  $K_{sp}$  for saturated solution of  $\text{Ba}(\text{OH})_2$  having pH 12 ?**

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## Options:

A.  $4 \times 10^{-4}$

B.  $4 \times 10^{-6}$

C.  $5 \times 10^{-6}$

D.  $5 \times 10^{-7}$

**Answer: D**

## Solution:

Given:

pH of saturated solution of  $\text{Ba}(\text{OH})_2 = 12$

Step 1: Calculate pOH

$$\text{pH} + \text{pOH} = 14$$

$$12 + \text{pOH} = 14$$

$$\text{pOH} = 2$$

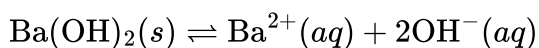
Step 2: Calculate  $[\text{OH}^-]$

$$\text{pOH} = -\log [\text{OH}^-]$$

$$2 = -\log [\text{OH}^-]$$

$$[\text{OH}^-] = 10^{-2} \text{ mol/L}$$

Step 3: Write the dissociation equation and  $K_{sp}$  expression



Let the solubility of  $\text{Ba}(\text{OH})_2 = s \text{ mol/L}$ .

Then,

$$[\text{Ba}^{2+}] = s$$

$$[\text{OH}^-] = 2s$$

But from above,  $[\text{OH}^-] = 10^{-2}$ .

So,

$$2s = 10^{-2} \implies s = 5 \times 10^{-3}$$

Step 4: Calculate  $K_{sp}$

$$K_{sp} = [\text{Ba}^{2+}][\text{OH}^{-}]^2$$

$$= (s) \times (2s)^2$$

$$= s \times 4s^2$$

$$= 4s^3$$

Plug in  $s = 5 \times 10^{-3}$ :

$$K_{sp} = 4 \times (5 \times 10^{-3})^3$$

$$= 4 \times 125 \times 10^{-9}$$

$$= 500 \times 10^{-9}$$

$$= 5 \times 10^{-7}$$

**Final Answer:**

**Option D:**  $5 \times 10^{-7}$

---

## Question39

**What is the pH of buffer solution prepared by mixing 0.01 M weak acid and 0.02 M salt of weak acid with strong base? ( $\text{pK}_a = 4.680$ )**

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**Options:**

A. 4.379

B. 2.379

C. 4.981

D. 2.981

**Answer: C**

**Solution:**

Given:

- Concentration of weak acid,  $[HA] = 0.01 \text{ M}$

- Concentration of salt (conjugate base),  $[A^-] = 0.02 \text{ M}$
- $pK_a = 4.680$

We use the Henderson–Hasselbalch equation for buffer solutions:

$$\text{pH} = pK_a + \log\left(\frac{[\text{Salt}]}{[\text{Acid}]}\right)$$

**Step 1:** Substitute the values:

- $[\text{Salt}] = [A^-] = 0.02 \text{ M}$
- $[\text{Acid}] = [HA] = 0.01 \text{ M}$

**Step 2:** Plug the values into the formula:

$$\text{pH} = 4.680 + \log\left(\frac{0.02}{0.01}\right)$$

**Step 3:** Simplify  $\frac{0.02}{0.01}$ :

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

**Step 4:** Calculate  $\log(2)$ :

- $\log(2) \approx 0.301$

**Step 5:** Put it all together:

$$\text{pH} = 4.680 + 0.301 = 4.981$$

**Final Answer:**

The pH of the buffer solution is

4.981

Correct option: **C (4.981)**

---

## Question40

**If pH of solution changes from 4 to 5 , then the  $\text{H}_3\text{O}^+$  ion concentration of solution**

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**Options:**

A. decreases by one times

B. increase by one times

C. increase by 10 times

D. decrease by 10 times

**Answer: D**

**Solution:**

Given: pH changes from 4 to 5.

Step 1: Recall the relationship between pH and  $\text{H}_3\text{O}^+$  ion concentration:

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$$

Step 2: Let the initial  $\text{H}_3\text{O}^+$  ion concentration (at pH 4) be  $[\text{H}_3\text{O}^+]_1$ :

$$4 = -\log_{10}[\text{H}_3\text{O}^+]_1 \implies [\text{H}_3\text{O}^+]_1 = 10^{-4} \text{ mol/L}$$

Step 3: The new concentration at pH 5,  $[\text{H}_3\text{O}^+]_2$ :

$$5 = -\log_{10}[\text{H}_3\text{O}^+]_2 \implies [\text{H}_3\text{O}^+]_2 = 10^{-5} \text{ mol/L}$$

Step 4: Compare the concentrations:

$$\frac{[\text{H}_3\text{O}^+]_2}{[\text{H}_3\text{O}^+]_1} = \frac{10^{-5}}{10^{-4}} = 10^{-1} = 0.1$$

This means the concentration **decreased by 10 times**.

**\*\*Correct Option: D**

decrease by 10 times\*\*

---

## Question41

**The pH of a sample of vinegar is 3.76. Calculate the concentration of hydrogen ion in it in  $\text{mol dm}^{-3}$  ?**

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**Options:**

A.  $1.97 \times 10^{-4}$

B.  $1.738 \times 10^{-4}$

C.  $1.84 \times 10^{-4}$

D.  $1.283 \times 10^{-4}$

**Answer: B**

### **Solution:**

Given:

pH of vinegar, = 3.76

We know,

$$\text{pH} = -\log_{10}[\text{H}^+]$$

Rearrange to find  $[\text{H}^+]$ :

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

Substitute the value of pH:

$$[\text{H}^+] = 10^{-3.76}$$

Now, write  $-3.76$  as  $-4 + 0.24$ :

$$10^{-3.76} = 10^{-4+0.24} = 10^{-4} \times 10^{0.24}$$

We know that  $10^{0.24} \approx 1.7378$  (using log table or calculator).

So,

$$[\text{H}^+] = 1.7378 \times 10^{-4} \text{ mol dm}^{-3}$$

**Final answer:**

$1.738 \times 10^{-4} \text{ mol dm}^{-3}$
--

So, the correct option is **Option B**.

---

## **Question42**

**The solubility product of a sparingly soluble salt AX is  $4.9 \times 10^{-13}$ .  
What is its solubility in  $\text{mol dm}^{-3}$  ?**

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**Options:**

A.  $2.4 \times 10^{-13}$

B.  $4.9 \times 10^{-7}$

C.  $7.0 \times 10^{-7}$

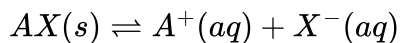
D.  $7.0 \times 10^{-13}$

**Answer: C**

### **Solution:**

Let the sparingly soluble salt be  $AX$ .

When  $AX$  dissolves in water, it dissociates as:



Let the solubility of  $AX$  be  $S \text{ mol.dm}^{-3}$ .

At equilibrium:

- $[A^+] = S$
- $[X^-] = S$

The solubility product expression is:

$$K_{sp} = [A^+][X^-]$$

Substitute values:

$$K_{sp} = S \times S = S^2$$

Given:

$$K_{sp} = 4.9 \times 10^{-13}$$

So,

$$S^2 = 4.9 \times 10^{-13}$$

Taking square root on both sides:

$$S = \sqrt{4.9 \times 10^{-13}}$$

$$S = \sqrt{4.9} \times \sqrt{10^{-13}}$$

$$S = 2.2136 \times 10^{-6.5}$$

Now,

$$10^{-6.5} = 10^{-6} \times 10^{-0.5}$$

And,

$$10^{-0.5} \approx 0.316$$

Thus,

$$S = 2.2136 \times 10^{-6} \times 0.316$$

$$S = 0.7 \times 10^{-6}$$

$$S = 7.0 \times 10^{-7} \text{ mol dm}^{-3}$$

**Correct Answer:**

**Option C:**  $7.0 \times 10^{-7} \text{ mol dm}^{-3}$

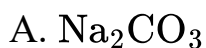
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## Question43

**Which among the following salts is NOT hydrolysed in water?**

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**Options:**



**Answer: C**

**Solution:**

Let us analyze each salt using NCERT concepts:

**Step 1: Nature of Ions**

- **Salt from strong acid and strong base:** No hydrolysis.
- **Salt from strong acid and weak base:** Hydrolysis (cation hydrolyses).
- **Salt from weak acid and strong base:** Hydrolysis (anion hydrolyses).
- **Salt from weak acid and weak base:** Both cation and anion hydrolyse.

**Step 2: Check Each Option**

**Option A:**  $\text{Na}_2\text{CO}_3$

- $\text{NaOH}$  (strong base) +  $\text{H}_2\text{CO}_3$  (weak acid)
- Salt from strong base and weak acid.
- **Hydrolysed** in water.

**Option B:  $\text{NH}_4\text{CN}$**

- $\text{NH}_4\text{OH}$  (weak base) +  $\text{HCN}$  (weak acid)
- Salt from weak base and weak acid.
- **Hydrolysed** in water.

**Option C:  $\text{KNO}_3$**

- $\text{KOH}$  (strong base) +  $\text{HNO}_3$  (strong acid)
- Salt from strong base and strong acid.
- **NOT hydrolysed** in water.

**Option D:  $\text{KCN}$**

- $\text{KOH}$  (strong base) +  $\text{HCN}$  (weak acid)
- Salt from strong base and weak acid.
- **Hydrolysed** in water.

**Final answer:**

**Option C:  $\text{KNO}_3$  is NOT hydrolysed in water.**

---

## Question44

**Which from following substances acts as a base when reacted with water?**

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**Options:**

A.  $\text{CH}_3\text{COOH}$

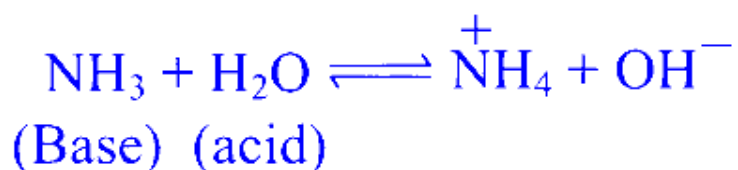
B.  $\text{H}_2\text{C}_2\text{O}_4$

C.  $\text{HCl}$

D.  $\text{NH}_3$

**Answer: D**

**Solution:**



## Question45

Calculate the pH of 0.02 M monobasic acid having 2% dissociation.

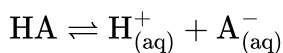
**MHT CET 2024 16th May Evening Shift**

**Options:**

- A. 3.4
- B. 4.5
- C. 5.1
- D. 5.8

**Answer: A**

**Solution:**



$$[\text{H}]^+ = \alpha C = \frac{2}{100} \times 0.02 \text{ m} = 0.0004 \text{ m}$$

$$\text{pH} = -\log [\text{H}^+] = -\log(0.0004) = 3.4$$

---

## Question46

Calculate the solubility product of sparingly soluble salt BA at 27°C if its solubility is  $1.8 \times 10^{-5} \text{ mol dm}^{-3}$  at same temperature.

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Options:

A.  $3.24 \times 10^{-10}$

B.  $2.44 \times 10^{-10}$

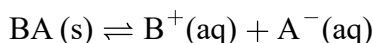
C.  $1.64 \times 10^{-10}$

D.  $4.00 \times 10^{-10}$

**Answer: A**

**Solution:**

The solubility product ( $K_{sp}$ ) is a constant for a given substance at a certain temperature, reflecting the level at which the substance dissolves in water to form a saturated solution. For a sparingly soluble salt like BA dissociating into its ions:



The equilibrium expression for the solubility product is given by:

$$K_{sp} = [\text{B}^+][\text{A}^-]$$

Given the solubility of BA is  $1.8 \times 10^{-5} \text{ mol dm}^{-3}$ , in a saturated solution at equilibrium, the concentration of  $\text{B}^+$  ions is  $1.8 \times 10^{-5} \text{ mol dm}^{-3}$ , and the concentration of  $\text{A}^-$  ions is also  $1.8 \times 10^{-5} \text{ mol dm}^{-3}$ , since the stoichiometry is 1:1.

Therefore, the solubility product is calculated as follows:

$$K_{sp} = (1.8 \times 10^{-5})(1.8 \times 10^{-5})$$

$$K_{sp} = 3.24 \times 10^{-10}$$

Therefore, the correct option is **Option A:  $3.24 \times 10^{-10}$** .

---

## Question47

**Calculate the pH of 0.01 M sulphuric acid.**

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**Options:**

A. 1.699

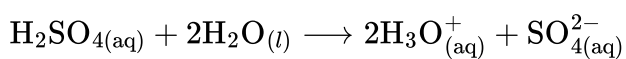
B. 2.00

C. 0.699

D. 3.398

**Answer: A**

**Solution:**



$$\begin{aligned}\text{Hence, } [\text{H}_3\text{O}^+] &= 2 \times c = 2 \times 0.01\text{M} \\ &= 2 \times 10^{-2}\text{M}\end{aligned}$$

$$\begin{aligned}\text{pH} &= -\log_{10} [\text{H}_3\text{O}^+] \\ &= -\log_{10} (2 \times 10^{-2}) \\ &= -\log_{10} 2 - \log_{10} 10^{-2} \\ &= -\log_{10} 2 + 2 \\ &= 2 - 0.3010\end{aligned}$$

$$\text{pH} = 1.699$$

---

## Question48

**Which from following mixtures in water acts as a buffer?**

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**Options:**

A. Acetic acid and sodium acetate

B. Acetic acid and ammonium chloride

C. Ammonium hydroxide and sodium chloride

D. Formic acid and acetic acid

**Answer: A**

## Solution:

Acetic acid is a weak acid and sodium acetate is a salt of weak acid and strong base. Thus, it acts as a buffer solution in water.

---

## Question49

Calculate the solubility in  $\text{mol dm}^{-3}$  of sparingly soluble salt  $BA$  if its solubility product  $4.9 \times 10^{-13}$  at same temperature.

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#### Options:

A.  $7.0 \times 10^{-7}$

B.  $7.5 \times 10^{-7}$

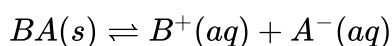
C.  $8.0 \times 10^{-7}$

D.  $4.9 \times 10^{-7}$

**Answer: A**

## Solution:

The solubility product ( $K_{\text{sp}}$ ) of a sparingly soluble salt  $BA$  is related to its solubility in solution. For a salt that dissolves according to:



The equilibrium expression of the solubility product is:

$$K_{\text{sp}} = [B^+][A^-]$$

If we let the solubility of  $BA$  in  $\text{mol dm}^{-3}$  be  $s$ , then at equilibrium:

$$[B^+] = s \quad \text{and} \quad [A^-] = s$$

Thus, the  $K_{\text{sp}}$  becomes:

$$K_{\text{sp}} = s^2$$

Given that  $K_{\text{sp}} = 4.9 \times 10^{-13}$ , we can solve for  $s$ :

$$s^2 = 4.9 \times 10^{-13}$$

Taking the square root of both sides:

$$s = \sqrt{4.9 \times 10^{-13}}$$

$$s = 7.0 \times 10^{-7} \text{ mol dm}^{-3}$$

Therefore, the solubility of the salt  $BA$  is  $7.0 \times 10^{-7} \text{ mol dm}^{-3}$ , which corresponds to Option A.

---

## Question50

**An aqueous solution of strong monoacidic base is of  $1 \times 10^{-4} \text{M}$ .  
What is the value of pH at  $25^\circ \text{C}$  ?**

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**Options:**

A. 7

B. 4

C. 3

D. 10

**Answer: D**

### **Solution:**

To find the pH of an aqueous solution of a strong monoacidic base, such as sodium hydroxide (NaOH), you start by determining the concentration of hydroxide ions ( $\text{OH}^-$ ). A strong base like NaOH dissociates completely in water:



Given the concentration of the base is  $1 \times 10^{-4} \text{M}$ , this is also the concentration of hydroxide ions [ $\text{OH}^-$ ] in the solution.

The relationship between pOH and the concentration of hydroxide ions is given by:

$$\text{pOH} = -\log_{10}[\text{OH}^-]$$

Substituting the concentration of hydroxide ions:

$$\text{pOH} = -\log_{10}(1 \times 10^{-4}) = 4$$

At room temperature ( $25^{\circ}\text{C}$ ), the relationship between pH and pOH is:

$$\text{pH} + \text{pOH} = 14$$

Thus, substituting the value of pOH:

$$\text{pH} = 14 - 4 = 10$$

Therefore, the pH of the solution is:

**Option D: 10.**

---

## Question51

**Acidic buffer solution is prepared by mixing proportionate quantity of**

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**Options:**

- A. strong acid and its salt with weak base.
- B. strong base and its salt with weak acid.
- C. weak acid and its salt with strong base.
- D. weak base and its salt with strong acid.

**Answer: C**

**Solution:**

An acidic buffer solution is prepared by mixing a weak acid and its salt with a strong base.

Thus, the correct choice is:

**Option C:** weak acid and its salt with strong base.

Buffers consist of a mixture of a weak acid and its conjugate base (often in the form of its salt) or a weak base and its conjugate acid (typically in its salt form). In the case of an acidic buffer, the weak acid will partially ionize in the solution, and the salt of the weak acid (which completely dissociates) will provide additional conjugate base. This combination helps to maintain the pH of the solution on the acidic side, resisting changes in pH upon the addition of small amounts of strong acids or bases.

For example, a common acidic buffer is composed of acetic acid ( $\text{CH}_3\text{COOH}$ ) and sodium acetate ( $\text{CH}_3\text{COONa}$ ). The equilibrium can be represented by:



The addition of acetate ions from sodium acetate will shift the equilibrium position to the left according to Le Chatelier's principle, ensuring that the pH of the solution remains relatively constant.

---

## Question52

The degree of dissociation of 0.01 M solution of  $\text{NH}_4\text{OH}$  is  $4.2 \times 10^{-2}$ . What is the percent dissociation of  $\text{NH}_4\text{OH}$  ?

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**Options:**

A. 2.4%

B. 4.2%

C. 0.2%

D. 0.4%

**Answer: B**

**Solution:**

The degree of dissociation, represented by  $\alpha$ , is given as  $4.2 \times 10^{-2}$ . To find the percent dissociation, convert the degree of dissociation from a fraction to a percentage by multiplying by 100.

The calculation is as follows:

$$\begin{aligned}\text{Percent dissociation} &= \alpha \times 100 = 4.2 \times 10^{-2} \times 100 \\ &= 4.2\%\end{aligned}$$

Therefore, the percent dissociation of  $\text{NH}_4\text{OH}$  is 4.2%.

The correct option is **Option B: 4.2%**.

---

## Question53

**Conjugate acid of  $\text{NH}_2^-$  and  $\text{NH}_3$  are respectively**

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**Options:**

A.  $\text{NH}_4\text{OH}$  and  $\text{NH}_2\text{OH}$

B.  $\text{NH}_3$  and  $\text{NH}_2^-$

C.  $\text{NH}_3$  and  $\text{NH}_4^+$

D.  $\text{NH}_4^+$  and  $\text{NH}_3$

**Answer: C**

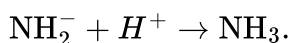
**Solution:**

**Definition of Conjugate Acid :**

The conjugate acid of a species is formed by adding one proton ( $\text{H}^+$ ) to it.

**Conjugate Acid of  $\text{NH}_2^-$  :**

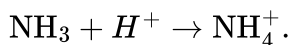
Starting with  $\text{NH}_2^-$  (the amide ion), add an  $\text{H}^+$  :



Thus, the conjugate acid of  $\text{NH}_2^-$  is  $\text{NH}_3$ .

**Conjugate Acid of  $\text{NH}_3$  :**

Starting with  $\text{NH}_3$  (ammonia), add an  $\text{H}^+$  :



Thus, the conjugate acid of  $\text{NH}_3$  is  $\text{NH}_4^+$ .

**Conclusion :**

Conjugate acid of  $\text{NH}_2^-$  is  $\text{NH}_3$ .

Conjugate acid of  $\text{NH}_3$  is  $\text{NH}_4^+$ .

**Correct Option :**

C)  $\text{NH}_3$  and  $\text{NH}_4^+$

---

## Question54

A buffer solution is prepared by mixing 0.2 M  $\text{NH}_4\text{OH}$  and 1 M  $\text{NH}_4\text{Cl}$ . What is the pH value of buffer solution? ( Give  $\text{pK}_b = 7.744$ )

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**Options:**

A. 5.56

B. 6.99

C. 8.44

D. 9.56

**Answer: A**

**Solution:**

For a basic buffer solution,

$$\begin{aligned}\text{pOH} &= \text{pK}_b + \log_{10} \frac{[\text{Salt}]}{[\text{Base}]} \\ &= 7.744 + \log_{10} \frac{1}{0.2} \\ &= 7.744 + 0.698\end{aligned}$$

$$\therefore \text{pOH} = 8.44$$

$$\therefore \text{Now, } \text{pH} + \text{pOH} = 14$$

$$\therefore \text{pH} = 14 - \text{pOH} = 14 - 8.44 = 5.56$$

---

## Question55

The dissociation constant of a weak monobasic acid is  $3.2 \times 10^{-4}$ . Calculate the degree of dissociation in its 0.04 M solution.

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**Options:**

A. 0.0128

B. 0.0151

C. 0.078

D. 0.089

**Answer: D**

**Solution:**

$$K_a = 3.2 \times 10^{-4}; c = 0.04M$$

For a monobasic acid,

$$\alpha = \sqrt{\frac{K_a}{c}} = \sqrt{\frac{3.2 \times 10^{-4}}{0.04}} = \sqrt{8 \times 10^{-3}} = 0.089$$

---

## Question56

The solubility of  $\text{CaCO}_3$  is  $7 \times 10^{-5} \text{ moldm}^{-3}$  at  $25^\circ\text{C}$ . What is its solubility product at same temperature?

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**Options:**

A.  $6.7 \times 10^{-9}$

B.  $9.0 \times 10^{-9}$

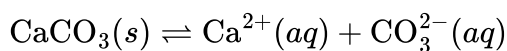
C.  $1.12 \times 10^{-9}$

D.  $4.9 \times 10^{-9}$

**Answer: D**

**Solution:**

The solubility product ( $K_{sp}$ ) of a sparingly soluble salt like  $\text{CaCO}_3$  can be determined from its solubility. Given that the solubility of  $\text{CaCO}_3$  is  $7 \times 10^{-5} \text{ mol dm}^{-3}$ , the dissolution of  $\text{CaCO}_3$  in water can be represented by the equation:



The solubility of  $\text{CaCO}_3$  provides the concentration of  $\text{Ca}^{2+}$  ions and  $\text{CO}_3^{2-}$  ions in a saturated solution:

$$[\text{Ca}^{2+}] = 7 \times 10^{-5} \text{ mol dm}^{-3}$$

$$[\text{CO}_3^{2-}] = 7 \times 10^{-5} \text{ mol dm}^{-3}$$

The solubility product expression for  $\text{CaCO}_3$  is:

$$K_{sp} = [\text{Ca}^{2+}][\text{CO}_3^{2-}]$$

Substitute the ion concentrations into the  $K_{sp}$  expression:

$$K_{sp} = (7 \times 10^{-5})(7 \times 10^{-5})$$

Calculate:

$$K_{sp} = 49 \times 10^{-10}$$

Simplifying gives:

$$K_{sp} = 4.9 \times 10^{-9}$$

Therefore, the solubility product of  $\text{CaCO}_3$  at  $25^\circ\text{C}$  is:

**Option D:**  $4.9 \times 10^{-9}$

---

## Question 57

Which among the following is the conjugate base of  $\text{HClO}_4$  ?

**MHT CET 2024 11th May Evening Shift**

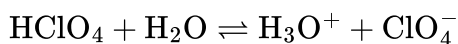
**Options:**



**Answer: D**

### **Solution:**

Conjugate base is formed when a Bronsted-Lowry acid donates a proton.



∴ Conjugate base of  $\text{HClO}_4$  is  $\text{ClO}_4^-$ .

---

## **Question58**

**What is pH of a centimolar solution of  $\text{H}_2\text{SO}_4$  ?**

**MHT CET 2024 11th May Evening Shift**

**Options:**

A. 1.7

B. 2.0

C. 3.2

D. 4.5

**Answer: A**

### **Solution:**

$$[\text{H}_2\text{SO}_4] = 2 \times C = 2 \times 0.01\text{M} = 2 \times 10^{-2}\text{M}$$

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

$$= -\log_{10} 2 - \log_{10} 10^{-2}$$

$$= -\log_{10} 2 + 2$$

$$\therefore \text{pH} = 2 - 0.3010 = 1.7$$

---

## **Question59**

**A buffer solution is prepared by mixing 0.01 M HCN and 0.02 M NaCN. If  $K_a$  for HCN is  $6.6 \times 10^{-10}$ , what is the concentration of  $H^+$  ions in solution?**

**MHT CET 2024 11th May Morning Shift**

**Options:**

A.  $3.3 \times 10^{-6} \text{M}$

B.  $3.3 \times 10^{-10} \text{M}$

C.  $1.32 \times 10^{-6} \text{M}$

D.  $1.32 \times 10^{-10} \text{M}$

**Answer: B**

**Solution:**

To determine the concentration of hydrogen ions ( $[H^+]$ ) in a buffer solution composed of a weak acid (HCN) and its conjugate base (NaCN), we can use the Henderson-Hasselbalch equation:

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

Where:

$$\text{p}K_a = -\log(K_a)$$

$[A^-]$  is the concentration of the conjugate base NaCN.

$[HA]$  is the concentration of the weak acid HCN.

**Step 1: Calculate  $\text{p}K_a$ .**

Given  $K_a = 6.6 \times 10^{-10}$ , the  $\text{p}K_a$  is:

$$\text{p}K_a = -\log(6.6 \times 10^{-10}) \approx 9.18$$

**Step 2: Use the Henderson-Hasselbalch equation.**

Plug in the given concentrations into the equation:

$$[A^-] = 0.02 \text{ M (NaCN)}$$

$$[HA] = 0.01 \text{ M (HCN)}$$

$$\text{pH} = 9.18 + \log \left( \frac{0.02}{0.01} \right)$$

The logarithmic term becomes:

$$\log\left(\frac{0.02}{0.01}\right) = \log(2) \approx 0.301$$

Thus, the pH of the solution is:

$$\text{pH} = 9.18 + 0.301 = 9.481$$

**Step 3: Convert pH to  $[\text{H}^+]$ .**

To find the hydrogen ion concentration:

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-9.481} \approx 3.3 \times 10^{-10} \text{ M}$$

Therefore, the concentration of  $\text{H}^+$  ions in the solution is:

**Option B:**  $3.3 \times 10^{-10} \text{ M}$

---

## Question60

**A monobasic acid is 5% dissociated in its 0.02 M solution. Calculate the dissociation constant of acid.**

### MHT CET 2024 11th May Morning Shift

**Options:**

A.  $2 \times 10^{-2}$

B.  $4 \times 10^{-4}$

C.  $5 \times 10^{-5}$

D.  $2.5 \times 10^{-4}$

**Answer: C**

### **Solution:**

To calculate the dissociation constant  $K_a$  of the acid, we start with the following information:

The acid dissociation can be represented as:



The initial concentration of the acid,  $[\text{HA}]_0$ , is 0.02 M.

The degree of dissociation,  $\alpha$ , is  $5\% = 0.05$ .

The concentration of dissociated hydrogen ions  $[H^+]$  is equal to  $\alpha \cdot [HA]_0$ . Therefore,

$$[H^+] = 0.05 \times 0.02 \text{ M} = 0.001 \text{ M}$$

This is also equal to the concentration of  $[A^-]$  due to the stoichiometry of the dissociation.

The concentration of undissociated  $[HA]$  is

$$[HA] = [HA]_0 \cdot (1 - \alpha) = 0.02 \cdot (1 - 0.05) = 0.019 \text{ M}$$

The acid dissociation constant,  $K_a$ , is calculated using the expression:

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

Substituting the concentrations, we have:

$$K_a = \frac{(0.001)(0.001)}{0.019}$$

$$K_a = \frac{0.000001}{0.019}$$

$$K_a = 5.26 \times 10^{-5}$$

By comparing with the given options, the closest value is option C:

$$5 \times 10^{-5}$$

Thus, the dissociation constant of the acid is  $5 \times 10^{-5}$ .

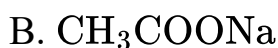
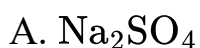
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## Question61

**Which among the following salts turns red litmus blue in its aqueous solution?**

**MHT CET 2024 11th May Morning Shift**

**Options:**



**Answer: B**

**Solution:**

The pH of a solution depends on hydrolysis of the salt. The salt  $\text{Na}_2\text{SO}_4$  is derived from strong acid and strong base, therefore it hydrolyses to a neutral solution. The other two salts are products of weak bases such as  $\text{NH}_4\text{OH}$  and  $\text{Cu}(\text{OH})_2$ . Only  $\text{CH}_3\text{COONa}$  comes from strong base  $\text{NaOH}$ , forming a basic solution and therefore colour changes from red to blue.

---

## Question62

**Calculate dissociation constant of a weak monobasic acid if it is 0.05% dissociated in 0.02 M solution.**

**MHT CET 2024 10th May Evening Shift**

**Options:**

A.  $2.0 \times 10^{-9}$

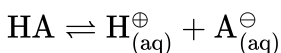
B.  $3.0 \times 10^{-9}$

C.  $4.0 \times 10^{-9}$

D.  $5.0 \times 10^{-9}$

**Answer: D**

**Solution:**



$$\alpha = 0.05\% = 0.05 \times 10^{-2}$$

$$\begin{aligned} K_a &= \alpha^2 C = (0.05 \times 10^{-2})^2 \times 0.02\text{M} \\ &= 5.0 \times 10^{-9} \end{aligned}$$

---

## Question63

**A buffer solution contains equal concentrations of weak acid and its salt with strong base. Calculate pH of buffer solution if dissociation**

**constant of weak acid is  $1.8 \times 10^{-5}$ .**

## **MHT CET 2024 10th May Evening Shift**

**Options:**

A. 4.7447

B. 5.1420

C. 5.8496

D. 4.0128

**Answer: A**

**Solution:**

The pH of a buffer solution containing equal concentrations of a weak acid and its salt with a strong base can be calculated using the Henderson-Hasselbalch equation:

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{Salt}]}{[\text{Acid}]}\right)$$

In this case, given that the concentrations of the weak acid and its conjugate base (salt) are equal, the equation simplifies to:

$$\text{pH} = \text{p}K_a$$

The dissociation constant ( $K_a$ ) is given as  $1.8 \times 10^{-5}$ . The pKa is calculated as follows:

$$\text{p}K_a = -\log(K_a)$$

Substitute the value of  $K_a$ :

$$\text{p}K_a = -\log(1.8 \times 10^{-5})$$

Using this calculation:

Calculate the logarithm:  $\log(1.8) \approx 0.2553$  and  $\log(10^{-5}) = -5$ .

Combine these logs:  $-\log(1.8 \times 10^{-5}) = 5 - 0.2553 = 4.7447$ .

Thus, the pH of the buffer solution is:

$$\boxed{4.7447}$$

This corresponds to Option A.

-----

## Question 64

Which from the following equations represents the relation between solubility (  $\text{mol L}^{-1}$  ) and solubility product for a salt  $\text{B}_3\text{A}_2$  ?

**MHT CET 2024 10th May Evening Shift**

**Options:**

A.  $S = \left( \frac{K_{\text{sp}}}{108} \right)^{\frac{1}{5}}$

B.  $S = (108 \times K_{\text{sp}})^{1/5}$

C.  $S = \left( \frac{K_{\text{sp}}}{27} \right)^{\frac{1}{5}}$

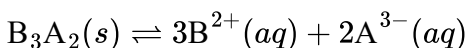
D.  $S = (27 \times K_{\text{sp}})^{1/5}$

**Answer: A**

**Solution:**

The equation that represents the relation between solubility ( $S$ ) and solubility product ( $K_{\text{sp}}$ ) for a salt  $\text{B}_3\text{A}_2$  can be determined by considering the dissolution equation and the corresponding expression for the solubility product:

When the salt  $\text{B}_3\text{A}_2$  dissolves, it dissociates as follows:



If  $S$  is the molar solubility, then at equilibrium:

The concentration of  $\text{B}^{2+}$  will be  $3S$ .

The concentration of  $\text{A}^{3-}$  will be  $2S$ .

The solubility product  $K_{\text{sp}}$  is given by:

$$K_{\text{sp}} = [\text{B}^{2+}]^3[\text{A}^{3-}]^2$$

Substituting the equilibrium concentrations:

$$K_{\text{sp}} = (3S)^3(2S)^2$$

$$K_{\text{sp}} = 27S^3 \cdot 4S^2$$

$$K_{sp} = 108S^5$$

Therefore, solving for  $S$ :

$$S = \left( \frac{K_{sp}}{108} \right)^{\frac{1}{5}}$$

Thus, the correct equation is:

**Option A**

$$S = \left( \frac{K_{sp}}{108} \right)^{\frac{1}{5}}$$

---

## Question 65

Calculate solubility ( $\text{mol dm}^{-3}$ ) of a sparingly soluble electrolyte AB at 298 K if its solubility product is  $1.6 \times 10^{-5}$  ?

**MHT CET 2024 10th May Morning Shift**

**Options:**

A.  $1.6 \times 10^{-3}$

B.  $2.5 \times 10^{-3}$

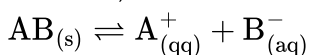
C.  $4.0 \times 10^{-3}$

D.  $8.0 \times 10^{-3}$

**Answer: C**

**Solution:**

For AB,



Here,  $x = 1, y = 1$

$$\therefore K_{sp} = x^x y^y S^{x+y} = (1)^1 (1)^1 S^{1+1} = S^2$$

$$\begin{aligned} \therefore S &= \sqrt{K_{sp}} \\ &= \sqrt{1.6 \times 10^{-5}} = \sqrt{16 \times 10^{-6}} \\ &= 4.0 \times 10^{-3} \text{ mol dm}^{-3} \end{aligned}$$

---

## Question66

What is the pH of  $10^{-8}$ M HCl solution?

**MHT CET 2024 10th May Morning Shift**

**Options:**

A. 8

B. 7

C. <7

D. >8

**Answer: C**

**Solution:**

$10^{-8}$ M indicates a very dilute solution.

Thus,  $[H^+]$  from water cannot be ignored.

But dissociation of water is suppressed due to common ion effect.

$[H^+] \neq 10^{-7}$ , but less than  $10^{-7}$ .

Hence, pH of  $10^{-8}$ M of HCl is less than 7.

---

## Question67

Which from following buffers is used to maintain the pH of human blood naturally?

**MHT CET 2024 9th May Evening Shift**

**Options:**

- A. Hydrogen cyanide and sodium cyanide
- B. Copper hydroxide and copper chloride
- C. Carbonic acid and salt of carbonic acid
- D. Ammonium hydroxide and ammonium chloride

**Answer: C**

**Solution:**

The pH of human blood is maintained naturally at 7.36 – 7.42 by  $(\text{HCO}_3^- + \text{H}_2\text{CO}_3)$  buffer.

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## Question68

**Calculate the solubility of sparingly soluble salt BA in  $\text{mol dm}^{-3}$  at 300 K if its solubility product is  $4.9 \times 10^{-9}$  at same temperature.**

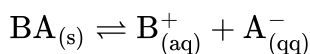
### MHT CET 2024 9th May Evening Shift

**Options:**

- A.  $5.72 \times 10^{-5}$
- B.  $6.40 \times 10^{-5}$
- C.  $7.00 \times 10^{-5}$
- D.  $7.81 \times 10^{-5}$

**Answer: C**

**Solution:**



$$x = 1, y = 1$$

$$K_{\text{sp}} = x^x y^y S^{x+y} = (1)^1 (1)^1 S^{1+1} = S^2$$

$$\therefore S = \sqrt{K_{\text{sp}}} = \sqrt{4.9 \times 10^{-9}}$$

$$= \sqrt{49 \times 10^{-10}}$$

$$= 7.00 \times 10^{-5} \text{M}$$

---

## Question69

Calculate the pH of buffer solution containing 0.04 M NaF and 0.02 M HF [ $\text{pK}_a = 3.142$ ].

**MHT CET 2024 9th May Evening Shift**

**Options:**

A. 4.5

B. 3.4

C. 2.6

D. 5.1

**Answer: B**

**Solution:**

0.04 M NaF and 0.02 M HF forms an acidic buffer.

$$\begin{aligned} \text{pH} &= \text{pK}_a + \log_{10} \frac{[\text{Salt}]}{[\text{Acid}]} \\ &= 3.142 + \log_{10} \frac{(0.04)}{(0.02)} \\ &= 3.142 + 0.3010 \\ &= 3.44 \approx 3.4 \end{aligned}$$

---

## Question70

**What is the pOH of millimolar solution of  $\text{Ca}(\text{OH})_2$  ?**

**MHT CET 2024 9th May Morning Shift**

**Options:**

A. 2.7

B. 10.3

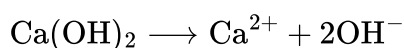
C. 12.3

D. 11.3

**Answer: A**

**Solution:**

$$1 \text{ mM} = 10^{-3} \text{ M}$$



$$10^{-3} \text{ M} \qquad 2 \times 10^{-3} \text{ M}$$

$$\therefore \text{pOH} = -\log_{10}[\text{OH}] = -\log_{10}(2 \times 10^{-3}) = 2.7$$

---

## **Question71**

**Molar conductivity of 0.02 M weak acid is  $7.92\Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$  and its molar conductivity at infinite dilution is  $232.7\Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$ . Calculate degree of dissociation of weak acid.**

**MHT CET 2024 9th May Morning Shift**

**Options:**

A. 0.0112

B. 0.0341

C. 0.0694

D. 0.292

**Answer: B**

### **Solution:**

The degree of dissociation ( $\alpha$ ) of a weak acid can be calculated using the formula for molar conductivity:

$$\alpha = \frac{\Lambda_c}{\Lambda_0}$$

where:

$\Lambda_c$  is the molar conductivity at a given concentration  $c$ , and

$\Lambda_0$  is the molar conductivity at infinite dilution.

Given values are:

$$\Lambda_c = 7.92 \Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$$

$$\Lambda_0 = 232.7 \Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$$

Calculate  $\alpha$ :

$$\alpha = \frac{7.92}{232.7}$$

Perform the division:

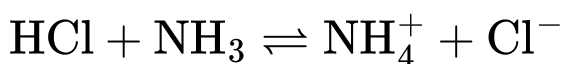
$$\alpha \approx 0.0341$$

Thus, the degree of dissociation of the weak acid is approximately 0.0341. Therefore, the correct option is **Option B: 0.0341**.

---

## **Question 72**

**Which among the following is correct conjugate acid base pair for the equation stated below?**



**MHT CET 2024 9th May Morning Shift**

**Options:**

A.  $\text{Cl}^-$  and  $\text{NH}_4^+$

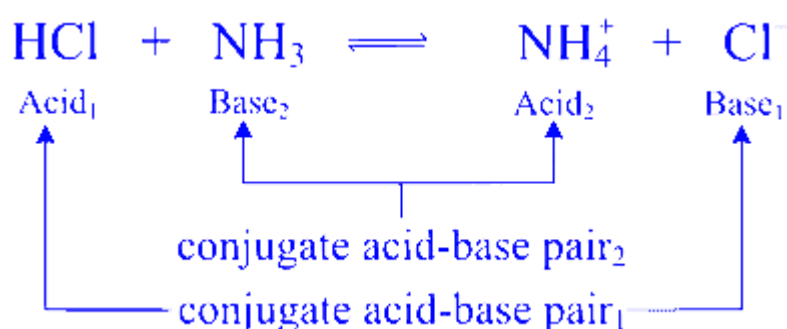
B. HCl and NH<sub>3</sub>

C. NH<sub>4</sub><sup>+</sup> and NH<sub>3</sub>

D. NH<sub>4</sub><sup>+</sup> and HCl

**Answer: C**

**Solution:**



## Question73

**Dissociation constant and degree of dissociation of weak acid are  $1.8 \times 10^{-5}$  and 0.03 respectively. What will be the concentration of solution of weak acid?**

**MHT CET 2024 9th May Morning Shift**

**Options:**

A. 0.2 M

B. 0.02 M

C. 0.5 M

D. 0.05 M

**Answer: B**

**Solution:**

The dissociation constant ( $K_a$ ) for a weak acid is related to its concentration ( $C$ ) and degree of dissociation ( $\alpha$ ) by the equation:

$$K_a = C\alpha^2$$

Given:

$$\text{Dissociation constant, } K_a = 1.8 \times 10^{-5}$$

$$\text{Degree of dissociation, } \alpha = 0.03$$

We need to find the concentration of the solution,  $C$ .

Rearrange the formula to solve for  $C$ :

$$C = \frac{K_a}{\alpha^2}$$

Substituting the given values:

$$C = \frac{1.8 \times 10^{-5}}{(0.03)^2}$$

Calculate  $(0.03)^2$ :

$$(0.03)^2 = 0.0009$$

Now calculate the concentration  $C$ :

$$C = \frac{1.8 \times 10^{-5}}{0.0009} = 2 \times 10^{-2} \text{ M} = 0.02 \text{ M}$$

Therefore, the concentration of the solution of the weak acid is:

**Option B: 0.02 M**

---

## Question 74

**Which from following salts is NOT derived from weak acid and weak base?**

**MHT CET 2024 4th May Evening Shift**

**Options:**

A. Ammonium fluoride

B. Ammonium cyanide

C. Ammonium acetate

D. Ammonium chloride

**Answer: D**

**Solution:**

Ammonium chloride ( $\text{NH}_4\text{Cl}$ ) is a salt of strong acid, HCl and weak base,  $\text{NH}_4\text{OH}$ .

---

## Question 75

Calculate the solubility product of sparingly soluble salt BA at  $25^\circ\text{C}$  if its solubility is  $7.2 \times 10^{-7} \text{ mol dm}^{-3}$  at same temperature.

**MHT CET 2024 4th May Evening Shift**

**Options:**

A.  $4.810 \times 10^{-13}$

B.  $5.184 \times 10^{-13}$

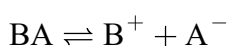
C.  $6.454 \times 10^{-13}$

D.  $5.925 \times 10^{-13}$

**Answer: B**

**Solution:**

The solubility product ( $K_{\text{sp}}$ ) of a sparingly soluble salt BA, which dissociates in water as follows:



is given by the expression:

$$K_{\text{sp}} = [\text{B}^+][\text{A}^-]$$

Since the solubility of the salt BA is  $7.2 \times 10^{-7} \text{ mol dm}^{-3}$ , the concentrations of the ions at equilibrium are:

$$[\text{B}^+] = 7.2 \times 10^{-7} \text{ mol dm}^{-3}$$

$$[A^-] = 7.2 \times 10^{-7} \text{ mol dm}^{-3}$$

Substituting these values into the expression for the solubility product:

$$K_{\text{sp}} = (7.2 \times 10^{-7})(7.2 \times 10^{-7})$$

Calculating the above expression:

$$K_{\text{sp}} = 7.2 \times 10^{-7} \times 7.2 \times 10^{-7} = 51.84 \times 10^{-14}$$

$$K_{\text{sp}} = 5.184 \times 10^{-13}$$

Therefore, the solubility product of the salt BA at 25°C is  $5.184 \times 10^{-13}$ .

Thus, the correct choice is **Option B**:  $5.184 \times 10^{-13}$ .

---

## Question 76

Calculate the  $[\text{OH}^-]$  if pOH of solution is 4.94

MHT CET 2024 4th May Evening Shift

Options:

A.  $2.356 \times 10^{-5} \text{ M}$

B.  $1.417 \times 10^{-5} \text{ M}$

C.  $1.881 \times 10^{-5} \text{ M}$

D.  $1.148 \times 10^{-5} \text{ M}$

**Answer: D**

**Solution:**

The concentration of hydroxide ions,  $[\text{OH}^-]$ , in a solution can be calculated from the pOH using the formula:

$$[\text{OH}^-] = 10^{-\text{pOH}}$$

Given that the pOH of the solution is 4.94, the calculation becomes:

$$[\text{OH}^-] = 10^{-4.94}$$

Using a calculator, we find:

$$[\text{OH}^-] \approx 1.148 \times 10^{-5} \text{M}$$

Therefore, the correct answer is Option D:  $1.148 \times 10^{-5} \text{M}$ .

---

## Question 77

**Which from following mixtures in water acts as a basic buffer?**

**MHT CET 2024 4th May Morning Shift**

**Options:**

- A.  $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$
- B.  $\text{C}_6\text{H}_5\text{COOH} + \text{C}_6\text{H}_5\text{COONa}$
- C.  $\text{HCOOH} + \text{HCOOK}$
- D.  $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$

**Answer: A**

**Solution:**

Basic buffer is the mixture of weak base and its salt with strong acid. In option A,  $\text{NH}_4\text{OH}$  is weak base and  $\text{NH}_4\text{Cl}$  is salt of strong acid and weak base. Hence,  $\text{NH}_4\text{OH} + \text{NH}_4\text{Cl}$  is basic buffer. Other options are examples of acidic buffer.

---

## Question 78

**Calculate the pH of a buffer solution containing 0.35 M weak acid and 0.70 M of its salt with strong base if  $\text{pK}_a$  is 4.56 .**

**MHT CET 2024 4th May Morning Shift**

**Options:**

- A. 6.11
- B. 3.72
- C. 4.86
- D. 5.65

**Answer: C**

### **Solution:**

For an acidic buffer,

$$\begin{aligned} \text{pH} &= \text{pK}_a + \log_{10} \frac{[\text{salt}]}{[\text{acid}]} \\ \therefore \text{pH} &= 4.56 + \log_{10} \frac{0.70}{0.35} \\ &= 4.56 + \log_{10} 2 = 4.56 + 0.3010 \\ &= 4.861 \end{aligned}$$

---

## **Question 79**

**Calculate the concentration of weak monobasic acid if its degree of dissociation and dissociation constant are  $5.0 \times 10^{-4}$  and  $5.0 \times 10^{-9}$  respectively.**

### **MHT CET 2024 4th May Morning Shift**

**Options:**

- A. 0.1 M
- B. 0.02 M
- C. 0.03 M
- D. 0.04 M

**Answer: B**

### **Solution:**

To calculate the concentration of a weak monobasic acid using its degree of dissociation ( $\alpha$ ) and dissociation constant ( $K_a$ ), the following relationship is used:

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

where:

$\alpha$  is the degree of dissociation,

$K_a$  is the dissociation constant,

$c$  is the concentration of the acid.

Given:

$$\alpha = 5.0 \times 10^{-4}$$

$$K_a = 5.0 \times 10^{-9}$$

Rearranging the formula to find  $c$  gives:

$$c = \frac{K_a}{\alpha^2}$$

Substitute the known values:

$$c = \frac{5.0 \times 10^{-9}}{(5.0 \times 10^{-4})^2}$$

Calculate  $\alpha^2$ :

$$(5.0 \times 10^{-4})^2 = 25 \times 10^{-8} = 2.5 \times 10^{-7}$$

Substitute back into the equation for  $c$ :

$$c = \frac{5.0 \times 10^{-9}}{2.5 \times 10^{-7}}$$

Simplify the expression:

$$c = 2.0 \times 10^{-2} = 0.02 \text{ M}$$

Therefore, the concentration of the weak monobasic acid is 0.02 M, which corresponds to option B.

---

## Question 80

**Calculate pH of 0.002 M KOH solution.**

**MHT CET 2024 3rd May Evening Shift**

**Options:**

A. 10.4

B. 11.3

C. 12.4

D. 13.2

**Answer: B**

### **Solution:**

To calculate the pH of a 0.002 M KOH solution, follow these steps:

#### **Solubility and Dissociation of KOH:**

Potassium hydroxide (KOH) is a strong base that dissociates completely in water:



This means that the concentration of hydroxide ions  $[\text{OH}^-]$  will be equal to the concentration of KOH, which is 0.002 M.

#### **Calculate pOH:**

The pOH is calculated using the formula:

$$\text{pOH} = -\log[\text{OH}^-]$$

Substitute the concentration of hydroxide ions:

$$\text{pOH} = -\log(0.002)$$

#### **Perform the logarithmic calculation:**

$$\text{pOH} = -\log(2 \times 10^{-3})$$

Using logarithmic properties:

$$\text{pOH} = -(\log(2) + \log(10^{-3}))$$

$$\text{pOH} = -(\log(2) - 3)$$

$$\text{pOH} = 3 - \log(2)$$

Approximate value of  $\log(2) \approx 0.301$

$$\text{pOH} \approx 3 - 0.301 = 2.699$$

#### **Calculate pH:**

Use the relationship between pH and pOH, given by:

$$\text{pH} + \text{pOH} = 14$$

Substitute the calculated pOH:

$$\text{pH} = 14 - \text{pOH}$$

$$\text{pH} = 14 - 2.699$$

$$\text{pH} \approx 11.301$$

Hence, the pH of a 0.002 M KOH solution is approximately 11.3.

**Answer:**

The correct option is **Option B: 11.3**.

---

## Question81

Calculate '  $\alpha$  ' for 0.1 M acetic acid ( $K_a = 1.0 \times 10^{-5}$ )

**MHT CET 2024 3rd May Evening Shift**

**Options:**

A.  $10^{-2}$

B.  $10^{-3}$

C.  $10^{-4}$

D.  $10^{-5}$

**Answer: A**

**Solution:**

For a weak acid HA,  $\alpha$  is very small, or  $(1 - \alpha) \cong 1$

$$\therefore \alpha = \sqrt{\frac{K_a}{c}} = \sqrt{\frac{1.0 \times 10^{-5}}{0.1}} = 10^{-2}$$

---

## Question82

Calculate  $[\text{H}_3\text{O}^+]$  in 0.02 M solution of monobasic acid if dissociation constant is  $1.8 \times 10^{-5}$ .

## MHT CET 2024 3rd May Morning Shift

### Options:

A.  $3.0 \times 10^{-4} \text{M}$

B.  $6.0 \times 10^{-4} \text{M}$

C.  $2.0 \times 10^{-4} \text{M}$

D.  $4.0 \times 10^{-4} \text{M}$

**Answer: B**

### Solution:

To calculate the concentration of hydronium ions  $[\text{H}_3\text{O}^+]$  in a 0.02 M solution of a monobasic acid, given that the dissociation constant ( $K_a$ ) is  $1.8 \times 10^{-5}$ , we can use the expression for the dissociation of a weak acid:

For a monobasic acid, HA, which dissociates as:



The expression for the dissociation constant  $K_a$  is:

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

Assuming  $x$  is the concentration of  $\text{H}^+$  ions that dissociate, we have:

$$[\text{H}^+] = x$$

$$[\text{A}^-] = x$$

The initial concentration of HA is 0.02 M, and the concentration after dissociation becomes  $0.02 - x \approx 0.02$

The assumption  $0.02 - x \approx 0.02$  is valid if  $x \ll 0.02$ , which is often true for weak acids.

Substitute these into the  $K_a$  expression:

$$1.8 \times 10^{-5} = \frac{x \cdot x}{0.02}$$

This simplifies to:

$$1.8 \times 10^{-5} = \frac{x^2}{0.02}$$

Solving for  $x$ , we get:

$$x^2 = 1.8 \times 10^{-5} \times 0.02$$

$$x^2 = 3.6 \times 10^{-7}$$

Taking the square root on both sides gives:

$$x = \sqrt{3.6 \times 10^{-7}}$$

$$x \approx 6.0 \times 10^{-4}$$

Thus, the concentration of hydronium ions  $[\text{H}_3\text{O}^+]$  in the solution is:

Option B:

$$6.0 \times 10^{-4} \text{ M}$$

---

## Question83

Which of the following equations represents the relation between solubility and solubility product for salt  $\text{BA}_3$ ?

**MHT CET 2024 3rd May Morning Shift**

**Options:**

A.  $S = \left(\frac{K_{sp}}{27}\right)^{1/4}$

B.  $S = (27 \times K_{sp})^{1/4}$

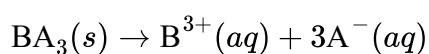
C.  $S = \left(\frac{K_{sp}}{4}\right)^{1/4}$

D.  $S = (4 \times K_{sp})^{1/4}$

**Answer: A**

**Solution:**

For a salt with the formula  $\text{BA}_3$ , it dissociates in water according to the following equation:



If the solubility of the salt  $\text{BA}_3$  is  $S$ , then the concentration of  $\text{B}^{3+}$  ions will be  $S$  and the concentration of  $\text{A}^-$  ions will be  $3S$ .

The solubility product  $K_{sp}$  for this dissociation can be expressed as:

$$K_{sp} = [\text{B}^{3+}][\text{A}^-]^3 = (S)(3S)^3$$

Simplifying the right-hand side:

$$K_{sp} = S(27S^3) = 27S^4$$

To find the relationship between the solubility  $S$  and the solubility product  $K_{sp}$ , we solve for  $S$ :

$$S^4 = \frac{K_{sp}}{27}$$

Taking the fourth root of both sides, we get:

$$S = \left(\frac{K_{sp}}{27}\right)^{1/4}$$

Thus, the correct equation that represents the relationship between solubility  $S$  and the solubility product  $K_{sp}$  for salt  $BA_3$  is:

$$S = \left(\frac{K_{sp}}{27}\right)^{1/4}$$

This corresponds to Option A.

---

## Question84

**Calculate the pH of buffer solution containing 0.027 M weak acid and 0.054 M of its salt with strong base if  $pK_a$  is 4.2.**

**MHT CET 2024 3rd May Morning Shift**

**Options:**

A. 4.5

B. 3.2

C. 5.6

D. 6.4

**Answer: A**

**Solution:**

To calculate the pH of the buffer solution, use the Henderson-Hasselbalch equation:

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

Where:

$pK_a$  is the acid dissociation constant, which is given as 4.2.

$[A^-]$  is the concentration of the salt of the weak acid with a strong base, given as 0.054 M.

$[HA]$  is the concentration of the weak acid, given as 0.027 M.

Substitute the given values into the Henderson-Hasselbalch equation:

$$pH = 4.2 + \log\left(\frac{0.054}{0.027}\right)$$

Calculate the ratio:

$$\frac{0.054}{0.027} = 2$$

Therefore, the equation becomes:

$$pH = 4.2 + \log(2)$$

The logarithm of 2 is approximately 0.301:

$$\log(2) \approx 0.301$$

Now substitute this value back into the equation:

$$pH = 4.2 + 0.301 = 4.501$$

Thus, the pH of the buffer solution is approximately 4.5, which corresponds to Option A.

-----

## Question85

**Calculate  $[H_3O^+]$  of a monobasic acid if it is 0.04% dissociated in 0.05 M solution.**

**MHT CET 2024 2nd May Evening Shift**

**Options:**

A.  $1 \times 10^{-5}$

B.  $1.5 \times 10^{-5}$

C.  $2.0 \times 10^{-5}$

D.  $3.0 \times 10^{-5}$

**Answer: C**

### **Solution:**

To find the concentration of hydronium ions,  $[\text{H}_3\text{O}^+]$ , for a monobasic acid in a solution where it is 0.04% dissociated, use the following approach:

For a monobasic acid, the dissociation can be represented as:



Given a 0.05 M solution with 0.04% dissociation, calculate the resulting concentration of  $[\text{H}_3\text{O}^+]$ :

Convert the percentage dissociation into a decimal:

$$0.04\% = \frac{0.04}{100} = 0.0004$$

Calculate the concentration of  $[\text{H}_3\text{O}^+]$ :

$$[\text{H}_3\text{O}^+] = 0.0004 \times 0.05 \text{ M}$$

$$= 2.0 \times 10^{-5} \text{ M}$$

Thus, the concentration of hydronium ions  $[\text{H}_3\text{O}^+]$  in the solution is  $2.0 \times 10^{-5} \text{ M}$ , corresponding to option C.

---

## **Question 86**

**Calculate the solubility product of sparingly soluble salt BA at 300 K if its solubility is  $9.1 \times 10^{-3} \text{ mol dm}^{-3}$  at same temperature.**

### **MHT CET 2024 2nd May Evening Shift**

**Options:**

A.  $9.635 \times 10^{-5}$

B.  $9.012 \times 10^{-5}$

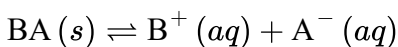
C.  $8.281 \times 10^{-5}$

D.  $7.816 \times 10^{-5}$

**Answer: C**

## Solution:

To calculate the solubility product ( $K_{sp}$ ) of a sparingly soluble salt, such as BA, we start from its dissociation in water, which can be represented as:



For each mole of BA that dissolves, it dissociates into one mole of  $\text{B}^+$  and one mole of  $\text{A}^-$  ions. If the solubility of BA is  $9.1 \times 10^{-3} \text{ mol dm}^{-3}$ , then the concentrations of both  $\text{B}^+$  and  $\text{A}^-$  ions in the solution at equilibrium will also be  $9.1 \times 10^{-3} \text{ mol dm}^{-3}$ .

The solubility product  $K_{sp}$  is given by the product of the concentrations of the ions, each raised to the power of their respective stoichiometric coefficients (which are 1 in this case):

$$K_{sp} = [\text{B}^+][\text{A}^-]$$

Substituting the known concentrations:

$$K_{sp} = (9.1 \times 10^{-3})^2$$

Calculating this gives:

$$K_{sp} = 9.1 \times 10^{-3} \times 9.1 \times 10^{-3}$$

$$K_{sp} = 82.81 \times 10^{-6}$$

$$K_{sp} = 8.281 \times 10^{-5}$$

Therefore, the solubility product of the sparingly soluble salt BA at 300 K is  $8.281 \times 10^{-5}$ , which corresponds to Option C.

---

## Question 87

Which from following species acts as base<sup>1</sup>, according to Bronsted-Lowry theory?



**MHT CET 2024 2nd May Evening Shift**

**Options:**

A.  $\text{Cl}^-$

B.  $\text{NH}_3$

C.  $\text{NH}_4^+$

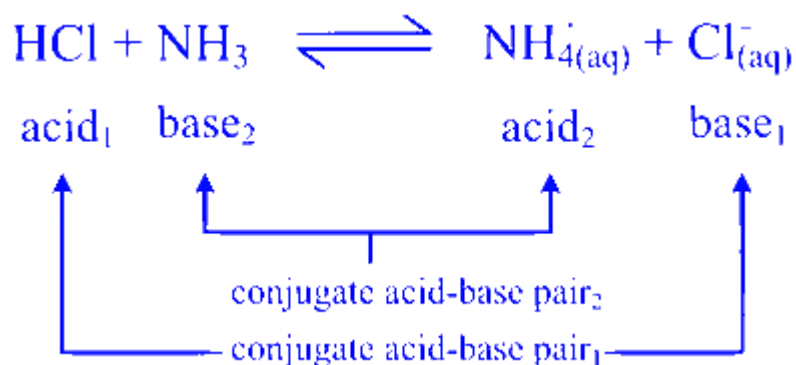
D. HCl

**Answer: A**

### Solution:

According to Bronsted - Lowry theory, a base is a proton acceptor.

For the given reaction,



## Question88

What is the ratio of concentration of salt to concentration of weak acid in buffer solution to maintain its pH value 7.2 ( $\text{pK}_a = 6.2$ ).

### MHT CET 2024 2nd May Morning Shift

Options:

A. 1.5

B. 10.0

C. 5.0

D. 8.5

**Answer: B**

### Solution:

The ratio of the concentration of salt to the concentration of the weak acid in a buffer solution can be calculated using the Henderson-Hasselbalch equation:

$$\text{pH} = \text{pK}_a + \log \left( \frac{[\text{Salt}]}{[\text{Acid}]} \right)$$

Given that:

$$\text{pH} = 7.2$$

$$\text{pK}_a = 6.2$$

Substituting these values into the Henderson-Hasselbalch equation:

$$7.2 = 6.2 + \log \left( \frac{[\text{Salt}]}{[\text{Acid}]} \right)$$

Simplify the equation:

$$7.2 - 6.2 = \log \left( \frac{[\text{Salt}]}{[\text{Acid}]} \right)$$

$$1.0 = \log \left( \frac{[\text{Salt}]}{[\text{Acid}]} \right)$$

To eliminate the logarithm, convert the expression to exponential form:

$$10^1 = \frac{[\text{Salt}]}{[\text{Acid}]}$$

Thus, the ratio of the concentration of salt to the concentration of weak acid is:

$$\frac{[\text{Salt}]}{[\text{Acid}]} = 10.0$$

Therefore, the correct option is **Option B: 10.0**.

---

## Question89

**What is the value of pOH if a buffer solution is prepared by mixing equal volumes of 0.4 M NH<sub>4</sub>OH and 0.5 M NH<sub>4</sub>Cl solutions. (pK<sub>b</sub> = 4.730)**

**MHT CET 2024 2nd May Morning Shift**

**Options:**

A. 6.0

B. 4.83

C. 10.42

D. 7.81

**Answer: B**

### **Solution:**

For a buffer solution consisting of a weak base and its conjugate acid, the pOH can be calculated using the Henderson-Hasselbalch equation for bases:

$$\text{pOH} = \text{pK}_b + \log \left( \frac{[\text{Salt}]}{[\text{Base}]} \right)$$

Given:

$$\text{pK}_b = 4.730$$

$$[\text{Base}] = [\text{NH}_4\text{OH}] = 0.4 \text{ M}$$

$$[\text{Salt}] = [\text{NH}_4\text{Cl}] = 0.5 \text{ M}$$

Substitute these values into the Henderson-Hasselbalch equation:

$$\text{pOH} = 4.730 + \log \left( \frac{0.5}{0.4} \right)$$

Calculate the logarithmic term:

$$\frac{0.5}{0.4} = 1.25$$

$$\log(1.25) \approx 0.0969$$

Now, substitute back into the equation:

$$\text{pOH} = 4.730 + 0.0969$$

$$\text{pOH} \approx 4.8269$$

Rounding to the appropriate significant figures, we get:

$$\text{pOH} \approx 4.83$$

Therefore, the value of pOH for this buffer solution is closest to Option B: **4.83**.

---

## **Question90**

**Which of the following salt solutions turns red litmus blue?**

**MHT CET 2024 2nd May Morning Shift**

## Options:

A.  $\text{NH}_4\text{CN}$

B.  $\text{NH}_4\text{Cl}$

C.  $\text{NH}_4\text{NO}_3$

D.  $\text{NaNO}_3$

**Answer: A**

## Solution:

To determine which salt solution turns red litmus paper blue, we need to consider whether the solution is basic (alkaline). A basic solution will turn red litmus paper blue.

The key concept for this is the nature of the ions produced when the salt is dissolved in water. A salt solution may turn basic if it contains a conjugate base of a weak acid.

Let's analyze the given options:

### Option A: $\text{NH}_4\text{CN}$

When dissolved in water,  $\text{NH}_4\text{CN}$  dissociates into  $\text{NH}_4^+$  (the ammonium ion) and  $\text{CN}^-$  (the cyanide ion). The ammonium ion  $\text{NH}_4^+$  is a weak acid, while  $\text{CN}^-$  is the conjugate base of the weak acid hydrocyanic acid (HCN). The  $\text{CN}^-$  ion will hydrolyze in water, producing  $\text{OH}^-$  ions and leading to a basic solution. Thus, this solution can turn red litmus blue.

### Option B: $\text{NH}_4\text{Cl}$

When  $\text{NH}_4\text{Cl}$  is dissolved in water, it dissociates into  $\text{NH}_4^+$  and  $\text{Cl}^-$ . Here,  $\text{NH}_4^+$  acts as a weak acid, but  $\text{Cl}^-$ , coming from a strong acid (HCl), does not significantly affect pH. Therefore, this solution is slightly acidic, not basic, and would not turn red litmus blue.

### Option C: $\text{NH}_4\text{NO}_3$

Upon dissolution,  $\text{NH}_4\text{NO}_3$  dissociates into  $\text{NH}_4^+$  and  $\text{NO}_3^-$ . Just like  $\text{Cl}^-$ , the  $\text{NO}_3^-$  ion, originating from a strong acid ( $\text{HNO}_3$ ), is neutral. Hence, the solution is slightly acidic due to the presence of  $\text{NH}_4^+$  and would not turn red litmus blue.

### Option D: $\text{NaNO}_3$

This salt dissociates into  $\text{Na}^+$  and  $\text{NO}_3^-$ . Both ions come from strong parents;  $\text{Na}^+$  from NaOH (strong base) and  $\text{NO}_3^-$  from  $\text{HNO}_3$  (strong acid). This solution is neutral and will not affect the pH significantly, thus it doesn't turn red litmus blue.

Therefore, the salt solution that turns red litmus blue is **Option A:  $\text{NH}_4\text{CN}$** .

---

## Question91

The solubility product of  $\text{PbCl}_2$  at 298 K is  $3.2 \times 10^{-5}$ . What is its solubility in  $\text{mol dm}^{-3}$  ?

### MHT CET 2023 14th May Evening Shift

Options:

A.  $8 \times 10^{-6}$

B.  $2 \times 10^{-2}$

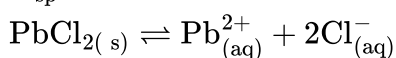
C.  $5.6 \times 10^{-3}$

D.  $5.0 \times 10^{-2}$

**Answer: B**

**Solution:**

$$K_{\text{sp}} = 3.2 \times 10^{-5}$$



$$\therefore K_{\text{sp}} = 4S^3$$

$$\therefore 4S^3 = 3.2 \times 10^{-5}$$

$$S = \sqrt[3]{\frac{3.2 \times 10^{-5}}{4}}$$

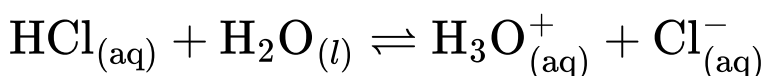
$$= \sqrt[3]{8 \times 10^{-6}}$$

$$= 2 \times 10^{-2} \text{ mol dm}^{-3}$$

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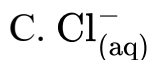
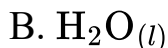
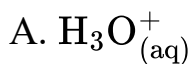
## Question92

Identify base<sub>2</sub> for following equation according to Bronsted-Lowry theory.



MHT CET 2023 14th May Evening Shift

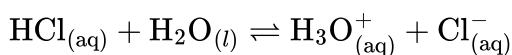
### Options:



**Answer: B**

### Solution:

The Bronsted-Lowry theory defines an acid as a substance that can donate a proton ( $\text{H}^+$  ion), and a base as a substance that can accept a proton. For the given chemical reaction:



In this reaction, hydrochloric acid (HCl) donates a proton to water ( $\text{H}_2\text{O}$ ), which acts as a base and accepts the proton, forming hydronium ion ( $\text{H}_3\text{O}^+$ ) and chloride ion ( $\text{Cl}^-$ ). Here is the role of each species according to the Bronsted-Lowry theory:

- HCl is a Bronsted-Lowry acid because it donates a proton to water.
- $\text{H}_2\text{O}$  is a Bronsted-Lowry base because it accepts a proton from HCl.
- $\text{H}_3\text{O}^+$  is the conjugate acid formed after  $\text{H}_2\text{O}$  accepts the proton.
- $\text{Cl}^-$  is the conjugate base formed after HCl donates the proton.

According to these roles, the correct option identifying the base that accepts a proton in this reaction is:

Option B:  $\text{H}_2\text{O}_{(l)}$

---

## Question93

**An organic monobasic acid has dissociation constant  $2.25 \times 10^{-6}$ . What is percent dissociation in its 0.01 M solution?**

### MHT CET 2023 14th May Evening Shift

#### Options:

A. 1.5%

B. 15%

C. 5%

D. 0.5%

**Answer: A**

### **Solution:**

The percent dissociation of an acid in solution can be determined by first calculating the concentration of the hydrogen ions ( $[H^+]$ ) that are produced when the acid dissociates. The dissociation constant ( $K_a$ ) gives us a measure of the extent to which the acid dissociates in solution. For a weak acid, which only partially dissociates, the equilibrium expression can be written as:



where  $HA$  is the weak acid,  $H^+$  is the hydrogen ion, and  $A^-$  is the conjugate base. The dissociation constant,  $K_a$ , is then given by:

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

The degree of dissociation ( $\alpha$ ) in weak acids is quite small. We can therefore assume that the change in concentration of the acid (the amount that dissociates) is relatively small and the equilibrium concentration of the acid  $[HA]$  can be approximated as the original concentration ( $C_0$ ) of the acid. Hence, taking  $\alpha$  to represent the degree of dissociation, we get:

$$[H^+] = [A^-] = \alpha C_0$$

$$[HA] = C_0 - \alpha C_0 \approx C_0$$

Now we plug in these into the expression for  $K_a$ :

$$K_a = \frac{\alpha C_0 \cdot \alpha C_0}{C_0}$$

$$K_a = \alpha^2 C_0$$

Solving for  $\alpha$ , we get:

$$\alpha = \sqrt{\frac{K_a}{C_0}}$$

Now let's plug in the given values:

$$K_a = 2.25 \times 10^{-6}$$

$$C_0 = 0.01 \text{ M}$$

$$\alpha = \sqrt{\frac{2.25 \times 10^{-6}}{0.01}}$$

$$\alpha = \sqrt{2225 \times 10^{-4}}$$

$$\alpha = 1.5 \times 10^{-2}$$

Now to find the percent dissociation, we convert the degree of dissociation into a percentage:

$$\text{Percent dissociation} = \alpha \times 100\%$$

$$\text{Percent dissociation} = 1.5 \times 10^{-2} \times 100\%$$

$$\text{Percent dissociation} = 1.5\%$$

Therefore, the percent dissociation of the acid in its 0.01 M solution is 1.5%, which corresponds to option A.

---

## Question94

**Calculate the pH of 0.01 M strong dibasic acid.**

**MHT CET 2023 14th May Morning Shift**

**Options:**

A. 5.5

B. 2.5

C. 2.0

D. 1.7

**Answer: D**

**Solution:**

$$[\text{H}_3\text{O}^+] = 2 \times c = 2 \times 0.01\text{M} = 2 \times 10^{-2}\text{M}$$

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+]$$

$$= -\log_{10} [2 \times 10^{-2}]$$

$$= -\log_{10} 2 - \log_{10} 10^{-2}$$

$$= -\log_{10} 2 + 2$$

$$= 2 - 0.3010$$

$$\text{pH} = 1.7$$

When  $\text{H}_3\text{O}^+$  concentration is  $2 \times 10^{-2}\text{M}$ , the pH should be less than 2. Only option (D) is valid.

---

## Question95

**Calculate dissociation constant of 0.001M weak monoacidic base undergoing 2% dissociation.**

## MHT CET 2023 14th May Morning Shift

**Options:**

A.  $4 \times 10^{-7}$

B.  $2 \times 10^{-6}$

C.  $2 \times 10^{-7}$

D.  $1 \times 10^{-7}$

**Answer: A**

**Solution:**

$$\alpha = \frac{\text{Percent dissociation}}{100} = \frac{2}{100} = 0.02$$

$$K_a = \alpha^2 c = (0.02)^2 \times 0.001 = 4 \times 10^{-7}$$

---

## Question96

**What is the pH of solution containing  $4.62 \times 10^{-4} \text{MH}^+$  ions?**

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**Options:**

A. 8.62

B. 4.64

C. 5.66

D. 3.34

**Answer: D**

**Solution:**

$$[\text{H}^+] = 4.62 \times 10^{-4}\text{M}$$

$$\text{pH} = ?$$

$$\text{Using, pH} = -\log [\text{H}^+] = -\log [4.62 \times 10^{-4}]$$

$$= -\log[4.62] - \log [10^{-4}] = -0.66 + 4$$

$$= 3.34$$

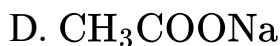
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## Question97

Which among the following salt solution in water is acidic in nature?

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Options:



**Answer: A**

**Solution:**

$\text{CuCl}_2$  is an acidic in water as  $\text{Cl}^-$  do not hydrolyse. Since, it is an anion of strong acid,  $\text{HCl}$ .  $\text{Cu}^{2+}$  will hydrolyse to form  $\text{Cu}(\text{OH})^+$ .

$\text{Cu}^{2+} + \text{H}_2\text{O}(\text{aq}) \rightleftharpoons \text{Cu}(\text{OH})^+(\text{aq}) + \text{H}^+(\text{aq})$  which is an acidic solution. So, the solution of  $\text{CuCl}_2$  in water is acidic.

---

## Question98

What is the molar concentration of acetic acid if value of it's, dissociation constant is  $1.8 \times 10^{-5}$  and degree of dissociation is 0.02 ?

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Options:

A.  $4.6 \times 10^{-3}$  M

B.  $4.5 \times 10^{-2}$  M

C.  $4.0 \times 10^{-4}$  M

D.  $3.6 \times 10^{-2}$  M

**Answer: B**

**Solution:**

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

$$\alpha = 0.02$$

$$C = ?$$

For acetic acid, dissociation is given as



Initial conc.	C	0	0
Conc. at equilibrium	$C - C\alpha$	$C\alpha$	$C\alpha$

$$\therefore k = \frac{\alpha^2 C^2}{(1 - \alpha)C} = \frac{\alpha^2 C}{(1 - \alpha)} \quad [\alpha \ll 1 \Rightarrow (1 - \alpha) \approx 1]$$

$$1.8 \times 10^{-5} = \frac{(0.02)^2 C}{1}$$

$$\text{or } C = 4.5 \times 10^{-2} \text{ M}$$

---

## Question99

A buffer solution is prepared by mixing equimolar acetic acid and sodium acetate. If ' $K_d$ ' of acetic acid is  $1.78 \times 10^{-5}$ , find the pH of buffer solution.

## MHT CET 2023 13th May Morning Shift

**Options:**

A. 4.75

B. 8.9

C. 9.4

D. 2.6

**Answer: A**

**Solution:**

For acidic buffer,

$$\begin{aligned} \text{pH} &= \text{pK}_a + \log_{10} \frac{[\text{Salt}]}{[\text{Acid}]} = \text{pK}_a + \log_{10} \frac{1}{1} = \text{pK}_a \\ \text{pK}_a &= -\log_{10} K_a = -\log_{10} (1.78 \times 10^{-5}) \\ &= -(\log_{10} 1.78 + \log_{10} 10^{-5}) \\ &= -(0.25 - 5) = 4.75 \end{aligned}$$

Therefore, pH of buffer solution = 4.75

---

## Question100

**The solubility product of  $\text{Mg}(\text{OH})_2$  is  $1.8 \times 10^{-11}$  at 298 K. What is its solubility in  $\text{mol dm}^{-3}$  ?**

## MHT CET 2023 13th May Morning Shift

**Options:**

A.  $1.650 \times 10^{-4}$

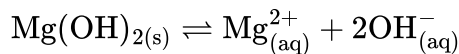
B.  $2.120 \times 10^{-4}$

C.  $3.184 \times 10^{-4}$

D.  $4.550 \times 10^{-4}$

**Answer: A**

**Solution:**



Here,  $x = 1, y = 2$

$$\therefore K_{sp} = x^x y^y S^{x+y} = (1)^1 (2)^2 S^{1+2} = 4S^3$$

$$\begin{aligned} \therefore S &= \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{1.8 \times 10^{-11}}{4}} \\ &= \sqrt[3]{4.5 \times 10^{-12}} = 1.650 \times 10^{-4} \end{aligned}$$

We know that,  $\sqrt[3]{1} = 1$  and  $\sqrt[3]{8} = 2$

$$\sqrt[3]{1} < \sqrt[3]{4.5} < \sqrt[3]{8}$$

Therefore,  $\sqrt[3]{8} < 2$ .

Only option (A) satisfies this condition.

---

## Question101

If  $K_{sp}$  is solubility product of  $\text{Al(OH)}_3$ , its solubility is expressed by formula,

**MHT CET 2023 13th May Morning Shift**

**Options:**

A.  $\sqrt[3]{\frac{4}{K_{sp}}}$

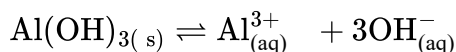
B.  $\sqrt[3]{\frac{K_{sp}}{4}}$

C.  $\sqrt[4]{\frac{K_{sp}}{27}}$

D.  $\sqrt[4]{K_{sp} \times 27}$

**Answer: C**

## Solution:



Here,  $x = 1, y = 3$

$$\therefore K_{\text{sp}} = x^x y^y S^{x+y} = (1)^1 (3)^3 S^{1+3} = 27 S^4$$

$$\therefore S = \sqrt[4]{\frac{K_{\text{sp}}}{27}}$$

---

## Question102

What is the solubility of  $\text{AgCl}_{(s)}$  if its solubility product is  $1.6 \times 10^{-10}$  ?

**MHT CET 2023 12th May Evening Shift**

**Options:**

A.  $1.26 \times 10^{-5} \text{M}$

B.  $1.00 \times 10^{-9} \text{M}$

C.  $2.6 \times 10^{-5} \text{M}$

D.  $1.56 \times 10^{-9} \text{M}$

**Answer: A**

## Solution:



$x = 1, y = 1$

$$K_{\text{sp}} = x^x y^y S^{x+y} = (1)^1 (1)^1 S^{1+1} = S^2$$

$$S = \sqrt{K_{\text{sp}}} = \sqrt{1.6 \times 10^{-10}}$$

$$= 1.26 \times 10^{-5} \text{M}$$

---

## Question103

**A buffer solution is prepared by mixing 0.01 M weak acid and 0.05 M solution of a salt of weak acid and strong base. What is the pH of buffer solution? ( $pK_a = 4.74$ )**

**MHT CET 2023 12th May Evening Shift**

**Options:**

A. 3.34

B. 4.80

C. 5.44

D. 6.93

**Answer: C**

**Solution:**

$$\begin{aligned} \text{pH} &= \text{pK}_a + \log_{10} \frac{[\text{salt}]}{[\text{acid}]} \\ &= 4.74 + \log_{10} \frac{0.05}{0.01} \\ &= 4.74 + 0.7 = 5.44 \end{aligned}$$

---

## **Question104**

**Which among the following is NOT an example of salt of strong acid and weak base?**

**MHT CET 2023 12th May Evening Shift**

**Options:**

A.  $\text{NH}_4\text{Cl}$

B.  $\text{NH}_4\text{NO}_3$

C.  $\text{CuSO}_4$

D.  $\text{Na}_2\text{SO}_4$

**Answer: D**

**Solution:**

$\text{Na}_2\text{SO}_4$  is a salt of strong base (NaOH) and strong acid ( $\text{H}_2\text{SO}_4$ ).

---

## Question105

**Find  $[\text{OH}^-]$  if a monoacidic base is 3% ionised in its 0.04 M solution.**

**MHT CET 2023 12th May Morning Shift**

**Options:**

A.  $3.1 \times 10^{-2} \text{ mol L}^{-1}$

B.  $4.5 \times 10^{-3} \text{ mol L}^{-1}$

C.  $9.0 \times 10^{-2} \text{ mol L}^{-1}$

D.  $1.2 \times 10^{-3} \text{ mol L}^{-1}$

**Answer: D**

**Solution:**

For a monoacidic base,

$$[\text{OH}^-] = c \times \alpha$$

$$[\text{OH}^-] = 0.04 \times \frac{3}{100}$$

$$[\text{OH}^-] = 1.2 \times 10^{-3} \text{ mol L}^{-1}$$

---

## Question106

**What is the expression for solubility product of silver chromate if its solubility is expressed as  $S \text{ mol L}^{-1}$  ?**

**MHT CET 2023 12th May Morning Shift**

**Options:**

A.  $2 S^2$

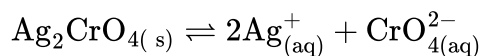
B.  $3 S^3$

C.  $4 S^3$

D.  $27 S^4$

**Answer: C**

**Solution:**



Here,  $x = 2, y = 1$

$$\therefore K_{sp} = x^x y^y S^{x+y} = (2)^2 (1)^1 S^{2+1} = 4 S^3$$

---

## Question107

**What is pH of solution containing 50 mL each of 0.1 M sodium acetate and 0.01 M acetic acid? ( $pK_a \text{CH}_3\text{COOH} = 4.50$ )**

**MHT CET 2023 12th May Morning Shift**

**Options:**

A. 2.5

B. 3.5

C. 4.5

D. 5.5

**Answer: D**

**Solution:**

For acidic buffer solution,

$$\text{pH} = \text{pK}_a + \log_{10} \frac{[\text{Salt}]}{[\text{Acid}]};$$

$$\text{pH} = 4.50 + \log_{10} \frac{0.1}{0.01};$$

$$\text{pH} = 4.50 + \log 10; \text{pH} = 5.50$$

For acidic buffer, if  $[\text{Salt}] > [\text{Acid}]$ , then  $\text{pH} > \text{pK}_a$  of acid. Hence, only option (D) is valid.

---

## Question108

**Identify the salt that undergoes hydrolysis and forms acidic solution from following.**

**MHT CET 2023 11th May Evening Shift**

**Options:**

A.  $\text{Na}_2\text{CO}_3$

B.  $\text{NH}_4\text{NO}_3$

C.  $\text{NH}_4\text{CN}$

D. KCN

**Answer: B**

**Solution:**

(A)  $\text{Na}_2\text{CO}_3$  : Salt of weak acid and strong base: Solution is basic.

(B)  $\text{NH}_4\text{NO}_3$  : Salt of strong acid and weak base: Solution is acidic.

(C)  $\text{NH}_4\text{CN}$  : Salt of weak acid and weak base for which  $K_a < K_b$  : Solution is basic.

(D) KCN : Salt of weak acid and strong base: Solution is basic.

---

## Question109

**A weak base is 1.42% dissociated in its 0.05 M solution. Calculate its dissociation constant.**

**MHT CET 2023 11th May Evening Shift**

**Options:**

A.  $5.5 \times 10^{-5}$

B.  $4.0 \times 10^{-5}$

C.  $1.8 \times 10^{-5}$

D.  $1.0 \times 10^{-5}$

**Answer: D**

**Solution:**

Percent dissociation = 1.42%

$$\therefore \alpha = 0.0142$$

For a weak monoacidic base,

$$\begin{aligned} K_b &= \alpha^2 C \\ &= (0.0142)^2 \times (0.05) \\ &= 1.0082 \times 10^{-5} \end{aligned}$$

---

## Question110

**Calculate the degree of dissociation of 0.01 M acetic acid at 25°C [ $\Lambda_c = 15.0\Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$  and  $\Lambda_0 = 300\Omega^{-1} \text{ cm}^2 \text{ mol}^{-1}$ ]**

**MHT CET 2023 11th May Evening Shift**

**Options:**

A. 0.042

B. 0.035

C. 0.025

D. 0.05

**Answer: D**

**Solution:**

$$\text{Degree of dissociation } (\alpha) = \frac{\Lambda_c}{\Lambda_0} = \frac{15.0}{300} = 0.05$$

---

## Question111

**What is the pH of 0.005 M NaOH solution?**

**MHT CET 2023 11th May Evening Shift**

**Options:**

A. 2.30

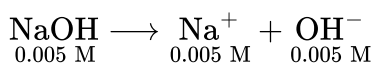
B. 12.6

C. 11.7

D. 3.2

**Answer: C**

**Solution:**



$$\begin{aligned} \text{pOH} &= -\log_{10}[\text{OH}] \\ &= -\log_{10}[0.005] \\ &= -\log_{10}(5 \times 10^{-3}) \\ &= -[\log 10^{-3} + \log 5] \\ &= -[-3 + 0.699] \\ &= -[-2.301] \\ &= 2.301 \end{aligned}$$

$$\text{pOH} + \text{pH} = 14$$

$$\begin{aligned} \therefore \text{pH} &= 14 - 2.301 \\ &= 11.699 \\ &\approx 11.7 \end{aligned}$$

---

## Question112

Which of the following salts turns red litmus blue in its aqueous solution?

**MHT CET 2023 11th May Morning Shift**

**Options:**



**Answer: C**

**Solution:**

$\text{NH}_4\text{CN}$  is the salt of weak acid  $\text{HCN}$  ( $K_a = 4.0 \times 10^{-10}$ ) and weak base  $\text{NH}_4\text{OH}$  ( $K_b = 1.8 \times 10^{-5}$ ) showing that  $K_a < K_b$ . If  $K_a < K_b$ , the aqueous solution of the salt will be basic. Thus, the solution of  $\text{NH}_4\text{CN}$  is basic and turns red litmus blue.

---

## Question113

**What is the concentration of  $\text{OH}^-$  ion in a solution containing  $0.05 \text{ M H}^+$  ions?**

**MHT CET 2023 11th May Morning Shift**

**Options:**

A.  $2.5 \times 10^{-13} \text{M}$

B.  $5.0 \times 10^{-2} \text{M}$

C.  $2.0 \times 10^{-13} \text{M}$

D.  $4.2 \times 10^{-12} \text{M}$

**Answer: C**

**Solution:**

$$K_w = [\text{H}_3\text{O}^+] [\text{OH}^-]$$

$$\therefore 10^{-14} = 0.05 \times [\text{OH}^-]$$

$$[\text{OH}] = \frac{10^{-14}}{0.05}$$

$$[\text{OH}^-] = 2.0 \times 10^{-13} \text{M}$$

---

## **Question114**

**Calculate the pH of  $1.36 \times 10^{-2} \text{M}$  solution of perchloric acid.**

**MHT CET 2023 11th May Morning Shift**

**Options:**

A. 1.43

B. 1.86

C. 2.43

D. 2.86

**Answer: B**

**Solution:**

Perchloric acid is a strong monobasic acid.

Hence,  $[\text{H}_3\text{O}^+] = 1.36 \times 10^{-2}\text{M}$

$$\begin{aligned}\therefore \text{pH} &= -\log_{10} [\text{H}_3\text{O}^+] \\ &= -\log_{10} [1.36 \times 10^{-2}] \\ &= -\log_{10} 1.36 - \log_{10} 10^{-2} \\ &= -\log_{10} 1.36 + 2 \\ &= 2 - 0.1335\end{aligned}$$

$$\therefore \text{pH} = 1.86$$

---

## Question115

**Find solubility in terms of mol L<sup>-1</sup> if solubility product of silver bromide is  $6.4 \times 10^{-13}$ .**

**MHT CET 2023 10th May Evening Shift**

**Options:**

A.  $4.0 \times 10^{-5} \text{ mol L}^{-1}$

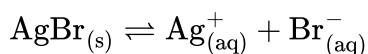
B.  $8.0 \times 10^{-7} \text{ mol L}^{-1}$

C.  $7.5 \times 10^{-5} \text{ mol L}^{-1}$

D.  $6.4 \times 10^{-4} \text{ mol L}^{-1}$

**Answer: B**

**Solution:**



$$\therefore x = 1, y = 1$$

$$K_{\text{sp}} = x^x y^y S^{x+y} = (1)^1 (1)^1 S^{1+1} = S^2$$

$$\begin{aligned}\therefore S &= \sqrt{K_{\text{sp}}} = \sqrt{6.4 \times 10^{-13}} = \sqrt{64 \times 10^{-14}} \\ &= 8 \times 10^{-7} \text{ mol L}^{-1}\end{aligned}$$

---

## Question116

Calculate the concentration of  $\text{H}^+$  ions in a solution if pOH is 11.

MHT CET 2023 10th May Morning Shift

Options:

A.  $10^{-11}$  M

B.  $10^{-8}$  M

C.  $10^{-6}$  M

D.  $10^{-3}$  M

**Answer: D**

**Solution:**

$$\text{pOH} = 11$$

$$\text{pH} + \text{pOH} = 14$$

$$\therefore \text{pH} = 14 - \text{pOH} = 14 - 11 = 3$$

$$\text{pH} = -\log_{10} [\text{H}^+]$$

$$\therefore [\text{H}^+] = 10^{-\text{pH}} = 10^{-3} \text{M}$$

---

## Question117

Acetic acid dissociated to 1.20% in its 0.01 M solution. What is the value of its dissociation constant?

## MHT CET 2023 10th May Morning Shift

**Options:**

A.  $2.20 \times 10^{-2}$

B.  $1.60 \times 10^{-4}$

C.  $1.44 \times 10^{-6}$

D.  $2.40 \times 10^{-4}$

**Answer: C**

**Solution:**

Percent dissociation = 1.20%

Degree of dissociation ( $\alpha$ ) = 0.012

For a weak monobasic acid,

$$K_a = \alpha^2 c = (0.012)^2 \times 0.01 = 1.44 \times 10^{-6}$$

---

## Question118

**Find solubility of  $PbI_2$  if its solubility product is  $7.0 \times 10^{-9}$ .**

## MHT CET 2023 9th May Evening Shift

**Options:**

A.  $1.21 \times 10^{-3} \text{ mol L}^{-1}$

B.  $3.228 \times 10^{-3} \text{ mol L}^{-1}$

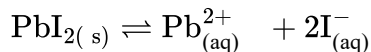
C.  $2.831 \times 10^{-3} \text{ mol L}^{-1}$

D.  $1.811 \times 10^{-3} \text{ mol L}^{-1}$

**Answer: A**

**Solution:**

For  $\text{PbI}_2$ ,



$$x = 1, y = 2$$

$$\therefore K_{\text{sp}} = x^x y^y S^{x+y} = (1)^1 (2)^2 S^{1+2} = 4 S^3$$

$$\therefore S = \sqrt[3]{\frac{K_{\text{sp}}}{4}} = \sqrt[3]{\frac{7.0 \times 10^{-9}}{4}} = 1.21 \times 10^{-3} \text{ mol L}^{-1}$$

---

## Question119

**What is the pH of a solution containing  $2.2 \times 10^{-6} \text{ M}$  hydrogen ions?**

**MHT CET 2023 9th May Evening Shift**

**Options:**

A. 6.34

B. 5.66

C. 4.34

D. 3.80

**Answer: B**

**Solution:**

$$\begin{aligned} \text{pH} &= -\log_{10} [\text{H}_3\text{O}^+] = -\log_{10} [2.2 \times 10^{-6}] \\ &= 6 - \log_{10}(2.2) = 6 - 0.34 \\ &= 5.66 \end{aligned}$$

---

## Question120

**What is the concentration of  $[\text{H}_3\text{O}^+]$  ion in  $\text{mol L}^{-1}$  of 0.001 M acetic acid ( $\alpha = 0.134$ ) ?**

## MHT CET 2023 9th May Evening Shift

**Options:**

A.  $1.34 \times 10^{-4}$

B.  $1.54 \times 10^{-4}$

C.  $1.80 \times 10^{-4}$

D.  $1.70 \times 10^{-4}$

**Answer: A**

**Solution:**

$$\begin{aligned}[\text{H}_3\text{O}^+] &= \alpha \times c = 0.134 \times 0.001 \\ &= 1.34 \times 10^{-4} \text{ mol dm}^{-3}\end{aligned}$$

---

## Question121

A buffer solution is prepared by mixing 0.2 M sodium acetate and 0.1 M acetic acid. If  $\text{pK}_a$  for acetic acid is 4.7, find the pH.

## MHT CET 2023 9th May Morning Shift

**Options:**

A. 3.0

B. 4.0

C. 5.0

D. 2.0

**Answer: C**

**Solution:**

Using Henderson's equation,

$$\text{pH} = \text{pK}_a + \log_{10} \frac{[\text{Salt}]}{[\text{Acid}]}$$

$$\text{pH} = 4.7 + \log_{10} \frac{0.2}{0.1}$$

$$\text{pH} = 4.7 + \log_{10} 2 = 4.7 + 0.3010 = 5.001$$

For acidic buffer, if  $[\text{Salt}] > [\text{Acid}]$ , then  $\text{pH} > \text{pK}_a$  of acid. Hence, only option (C) is valid.

---

## Question122

**Which salt from following forms aqueous solution having pH less than 7 ?**

**MHT CET 2023 9th May Morning Shift**

**Options:**

A.  $\text{CH}_3\text{COONa}$

B.  $\text{Na}_2\text{SO}_4$

C.  $\text{CuSO}_4$

D.  $\text{Na}_2\text{CO}_3$

**Answer: C**

**Solution:**

A salt of strong acid and weak base forms aqueous solution having pH less than 7 (that is, it forms an acidic solution).  $\text{CuSO}_4$  is salt of strong acid  $\text{H}_2\text{SO}_4$  and weak base  $\text{Cu}(\text{OH})_2$ .

---

## Question123

**Solubility of a salt  $\text{A}_2\text{B}_3$  is  $1 \times 10^{-3} \text{ mol dm}^{-3}$ . What is the value of its solubility product?**

## MHT CET 2023 9th May Morning Shift

**Options:**

A.  $1.08 \times 10^{-13}$

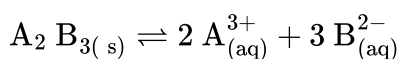
B.  $8.1 \times 10^{-15}$

C.  $2.7 \times 10^{-15}$

D.  $2.0 \times 10^{-13}$

**Answer: A**

**Solution:**



Here,  $x = 2, y = 3$

$$\begin{aligned} K_{sp} &= x^x y^y S^{x+y} \\ &= (2)^2 (3)^3 S^{2+3} \\ &= 108 S^5 \end{aligned}$$

$$\begin{aligned} K_{sp} &= 108 \times (1 \times 10^{-3})^5 \\ &= 108 \times 10^{-15} \\ &= 1.08 \times 10^{-13} \end{aligned}$$

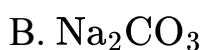
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## Question124

**Which of the following aqueous solutions of salts will have highest pH value?**

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**Options:**



C.  $\text{NH}_4\text{Cl}$

D.  $\text{NaCl}$

**Answer: B**

**Solution:**

$\text{Na}_2\text{CO}_3$  is the only alkaline solution having highest pH value. It is salt of weak acid ( $\text{H}_2\text{CO}_3$ ) and strong base ( $\text{NaOH}$ ).

---

## Question125

**What is the pH of  $2 \times 10^{-3}\text{M}$  solution of monoacidic weak base if it ionises to the extent of 5% ?**

**MHT CET 2022 11th August Evening Shift**

**Options:**

A. 14

B. 6

C. 4

D. 2

**Answer: C**

**Solution:**

$$[\text{H}^+] = c\alpha = 2 \times 10^{-3} \times 5 \times 10^{-2} = 10^{-4}\text{M}$$
$$\therefore \text{pH} = 4$$

---

## Question126

**The solubility of sparingly soluble salt  $\text{AB}_2$  is  $1.0 \times 10^{-4} \text{ mol dm}^{-3}$ . What is its solubility product?**

## MHT CET 2021 24th September Evening Shift

### Options:

A.  $2 \times 10^{-12}$

B.  $4 \times 10^{-8}$

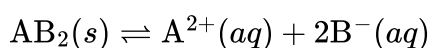
C.  $4 \times 10^{-12}$

D.  $2 \times 10^{-8}$

### Answer: C

### Solution:

To determine the solubility product ( $K_{sp}$ ) of a sparingly soluble salt  $AB_2$ , we start by considering its dissociation in water:



Given that the solubility of  $AB_2$  is  $1.0 \times 10^{-4} \text{ mol dm}^{-3}$ , let's denote the solubility by  $s$ .

Thus,  $s = 1.0 \times 10^{-4} \text{ mol dm}^{-3}$ . At equilibrium:

- The concentration of  $A^{2+}$  ions will be  $s = 1.0 \times 10^{-4} \text{ mol dm}^{-3}$ .

- The concentration of  $B^{-}$  ions will be  $2s = 2 \times 10^{-4} \text{ mol dm}^{-3}$  (because each formula unit of  $AB_2$  produces two  $B^{-}$  ions).

The solubility product  $K_{sp}$  is given by:

$$K_{sp} = [A^{2+}][B^{-}]^2$$

Substituting the equilibrium concentrations, we get:

$$K_{sp} = (1.0 \times 10^{-4}) \times (2 \times 10^{-4})^2$$

Calculating this value, we find:

$$K_{sp} = (1.0 \times 10^{-4}) \times (4 \times 10^{-8})$$

$$K_{sp} = 4 \times 10^{-12}$$

Thus, the solubility product of the sparingly soluble salt  $AB_2$  is  $4 \times 10^{-12}$ . Hence, the correct answer is:

Option C

$$4 \times 10^{-12}$$

---

## Question127

What is the pH of 0.005 M H<sub>2</sub>SO<sub>4</sub> solution?

MHT CET 2021 24th September Evening Shift

Options:

A. 5.0

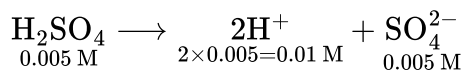
B. 2.3

C. 3.3

D. 2.0

**Answer: D**

**Solution:**



$$\begin{aligned} \text{pH} &= -\log_{10} [\text{H}^+] \\ &= -\log_{10} [10^{-2}] = 2.0 \end{aligned}$$

---

## Question128

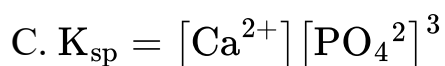
The solubility product expression for Ca<sub>3</sub>(PO<sub>4</sub>)<sub>2</sub> is represented as

MHT CET 2021 24th September Evening Shift

Options:

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

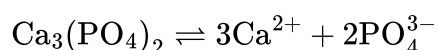
B.  $K_{\text{sp}} = [\text{Ca}^{2+}]^3 [\text{PO}_4^{3-}]^2$



**Answer: B**

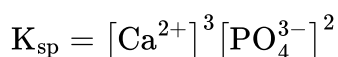
### **Solution:**

To determine the correct solubility product expression for  $\text{Ca}_3(\text{PO}_4)_2$ , we need to start by understanding its dissociation in water. When  $\text{Ca}_3(\text{PO}_4)_2$  dissolves, it dissociates into calcium ions ( $\text{Ca}^{2+}$ ) and phosphate ions ( $\text{PO}_4^{3-}$ ) according to the following equation:



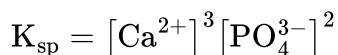
From this dissociation equation, for every 1 mole of  $\text{Ca}_3(\text{PO}_4)_2$  that dissolves, 3 moles of  $\text{Ca}^{2+}$  ions and 2 moles of  $\text{PO}_4^{3-}$  ions are produced.

The solubility product constant,  $K_{\text{sp}}$ , is given by the product of the concentrations of the ions, each raised to the power of its coefficient in the balanced equation. Therefore, the solubility product expression for  $\text{Ca}_3(\text{PO}_4)_2$  is:



Thus, the correct option is:

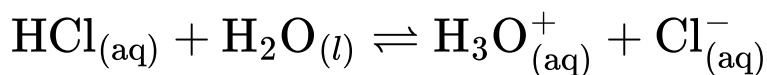
Option B



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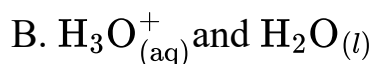
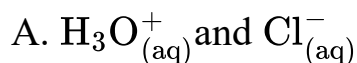
## **Question 129**

**Identify conjugate acid-base pair in the following reaction.**



**MHT CET 2021 24th September Morning Shift**

**Options:**

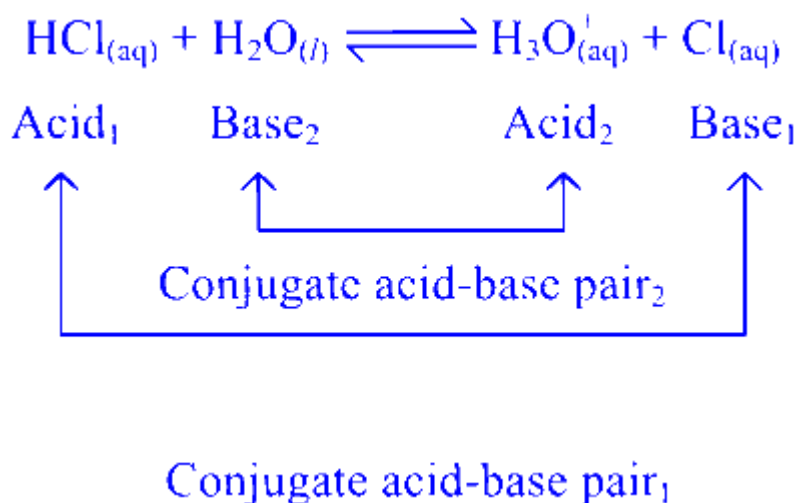


C.  $\text{HCl}_{(\text{aq})}$  and  $\text{H}_2\text{O}_{(l)}$

D.  $\text{Cl}^-_{(\text{aq})}$  and  $\text{H}_2\text{O}_{(l)}$

**Answer: B**

**Solution:**



---

## Question130

**Which among the following salts undergoes hydrolysis?**

**MHT CET 2021 23rd September Evening Shift**

**Options:**

A.  $\text{Na}_2\text{SO}_4$

B.  $\text{KCl}$

C.  $\text{NH}_4\text{Cl}$

D.  $\text{KNO}_3$

**Answer: C**

**Solution:**

$\text{NH}_4\text{Cl}$  is a salt of weak base,  $\text{NH}_4\text{OH}$  and strong acid,  $\text{HCl}$ . Hence, it undergoes hydrolysis. While  $\text{Na}_2\text{SO}_4$ ,  $\text{KCl}$  and  $\text{KNO}_3$  are salts of strong acids and strong bases and does not undergo hydrolysis.

---

## Question131

At 298 K, 0.1M solution of acetic acid is 1.34% ionized. What is the dissociation constant of acetic acid?

**MHT CET 2021 23rd September Evening Shift**

**Options:**

- A.  $1.4 \times 10^{-3}$
- B.  $1.8 \times 10^{-5}$
- C.  $1.6 \times 10^{-3}$
- D.  $1.34 \times 10^{-5}$

**Answer: B**

**Solution:**

$$c = 0.1 \text{ M}, \alpha = 1.34\% = \frac{1.34}{100} = 1.34 \times 10^{-2}$$

$$\begin{aligned} K_a &= \alpha^2 c \\ &= (1.34 \times 10^{-2})^2 \times 0.1 \\ &= 1.8 \times 10^{-5} \end{aligned}$$

---

## Question132

pH of soft drink is 3.6. Calculate the concentration of hydrogen ions in it.

**MHT CET 2021 23rd September Evening Shift**

**Options:**

A.  $2.51 \times 10^{-4}\text{M}$

B.  $2.3 \times 10^{-3}\text{M}$

C.  $2.0 \times 10^{-3}\text{M}$

D.  $2.81 \times 10^{-4}\text{M}$

**Answer: A**

**Solution:**

$$\text{pH} = 3.6, [\text{H}^+] = ?$$

$$\text{pH} = -\log_{10} [\text{H}^+]$$

$$\therefore \log_{10} [\text{H}^+] = -\text{pH}$$

$$= -3.6$$

$$= -3 - 0.6 - 1 + 1$$

$$= -4 + 0.4 = \bar{4}.4$$

$$\therefore [\text{H}^+] = \text{Antilog}(\bar{4}.4) = 2.512 \times 10^{-4}\text{M}$$

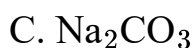
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## Question133

**Which among the following salt solution in water shows pH less than 7 ?**

**MHT CET 2021 23th September Morning Shift**

**Options:**



**Answer: A**

### **Solution:**

To determine which salt solution in water shows a pH less than 7, we need to examine the nature of the salts and their hydrolysis behavior:

1. **CuCl<sub>2</sub> (Copper(II) chloride):** Copper(II) chloride is the hydrolysis of a weak base (copper(II) hydroxide) and a strong acid (hydrochloric acid). The hydrolysis of CuCl<sub>2</sub> in water will produce H<sub>3</sub>O<sup>+</sup> ions, resulting in an acidic solution.

2. **CH<sub>3</sub>COONH<sub>4</sub> (Ammonium acetate):** This salt is derived from a weak acid (acetic acid) and a weak base (ammonium hydroxide). The pH of its solution is typically neutral because the weak acid and the weak base neutralize each other.

3. **Na<sub>2</sub>CO<sub>3</sub> (Sodium carbonate):** Sodium carbonate comes from a strong base (sodium hydroxide) and a weak acid (carbonic acid). The hydrolysis will produce OH<sup>-</sup> ions, resulting in a basic solution with pH greater than 7.

4. **KNO<sub>3</sub> (Potassium nitrate):** This salt is derived from a strong base (potassium hydroxide) and a strong acid (nitric acid). The solution of KNO<sub>3</sub> is neutral with pH around 7.

Given the above analysis, the salt solution in water that shows a pH less than 7 is **CuCl<sub>2</sub>**.

The correct answer is **Option A: CuCl<sub>2</sub>**.

---

## **Question134**

**The solubility of Ag<sub>2</sub>C<sub>2</sub>O<sub>4</sub> is  $2 \times 10^{-4}$  mol L<sup>-1</sup> at 298 K. What is it's solubility product?**

**MHT CET 2021 23th September Morning Shift**

**Options:**

A.  $1.6 \times 10^{-6}$

B.  $3.2 \times 10^{-11}$

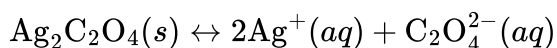
C.  $1.6 \times 10^{-11}$

D.  $3.2 \times 10^{-6}$

**Answer: B**

**Solution:**

To determine the solubility product ( $K_{sp}$ ) of  $\text{Ag}_2\text{C}_2\text{O}_4$ , we start by writing the dissociation equation:



Given that the solubility of  $\text{Ag}_2\text{C}_2\text{O}_4$  is  $2 \times 10^{-4} \text{ mol L}^{-1}$  at 298 K, we can say that the concentration of  $\text{Ag}_2\text{C}_2\text{O}_4$  that dissolves is  $2 \times 10^{-4} \text{ mol L}^{-1}$ .

From the dissociation equation, for each mole of  $\text{Ag}_2\text{C}_2\text{O}_4$  that dissolves, we get 2 moles of  $\text{Ag}^+$  and 1 mole of  $\text{C}_2\text{O}_4^{2-}$ .

Therefore, the concentrations at equilibrium are:

$$[\text{Ag}^+] = 2 \times (2 \times 10^{-4}) = 4 \times 10^{-4} \text{ mol L}^{-1}$$

$$[\text{C}_2\text{O}_4^{2-}] = 2 \times 10^{-4} \text{ mol L}^{-1}$$

The solubility product ( $K_{sp}$ ) is given by:

$$K_{sp} = [\text{Ag}^+]^2[\text{C}_2\text{O}_4^{2-}]$$

Substituting the equilibrium concentrations:

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

$$K_{sp} = 16 \times 10^{-8} \times 2 \times 10^{-4}$$

$$K_{sp} = 32 \times 10^{-12}$$

$$K_{sp} = 3.2 \times 10^{-11}$$

Therefore, the solubility product ( $K_{sp}$ ) of  $\text{Ag}_2\text{C}_2\text{O}_4$  is  $3.2 \times 10^{-11}$ . The correct option is:

**Option B**  $3.2 \times 10^{-11}$

---

## Question135

**A weak monobasic acid is 3.0% dissociated in it's 0.04 M solution. What is the dissociation constant of acid?**

**MHT CET 2021 23th September Morning Shift**

**Options:**

A.  $9 \times 10^{-4}$

B.  $3.6 \times 10^{-5}$

C.  $3 \times 10^{-2}$

D.  $4 \times 10^{-2}$

**Answer: B**

### **Solution:**

To determine the dissociation constant ( $K_a$ ) of the weak monobasic acid, we need to proceed with the following steps:

Given:

The degree of dissociation  $\alpha$  is 3.0% or 0.03 (since  $3.0\% = \frac{3}{100}$ ).

The concentration of the acid,  $C$ , is 0.04 M.

The formula for the dissociation constant  $K_a$  of a weak monobasic acid is:

$$K_a = C\alpha^2$$

Substitute the given values into the formula:

$$K_a = 0.04 \times (0.03)^2$$

First, calculate  $\alpha^2$ :

$$\alpha^2 = (0.03)^2 = 0.0009$$

Now, calculate  $K_a$ :

$$K_a = 0.04 \times 0.0009 = 3.6 \times 10^{-5}$$

Therefore, the dissociation constant of the acid is:

Option B:  $3.6 \times 10^{-5}$

---

## **Question136**

**Solubility product of AgBr is  $4.9 \times 10^{-13}$ . What is its solubility?**

**MHT CET 2021 22th September Evening Shift**

**Options:**

A.  $2.4 \times 10^{-7} \text{ mol dm}^{-3}$

B.  $3.2 \times 10^{-7} \text{ mol dm}^{-3}$

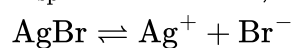
C.  $4.9 \times 10^{-7} \text{ mol dm}^{-3}$

D.  $7.0 \times 10^{-7} \text{ mol dm}^{-3}$

**Answer: D**

**Solution:**

$$K_{sp} = 4.9 \times 10^{-13}, S = ?$$



$$\therefore K_{sp} = S^2 \quad \therefore S = \sqrt{K_{sp}}$$

$$\therefore S = \sqrt{4.9 \times 10^{-13}} = 7.0 \times 10^{-7} \text{ mol dm}^{-3}$$

---

## Question137

Which of the following is NOT a correct mathematical equation for Ostwald dilution law?

**MHT CET 2021 22th September Evening Shift**

**Options:**

A.  $\alpha = \sqrt{\frac{K_a}{c}}$

B.  $K = \frac{\alpha^2}{V}$

C.  $K = \alpha^2 c$

D.  $\alpha = \sqrt{\frac{K_a}{V}}$

**Answer: D**

**Solution:**

$$\alpha = \sqrt{K_a \cdot V}$$

---

## Question138

A weak monobasic acid is 10% dissociated in 0.05 M solution. What is its percentage dissociation in 0.10 M solution?

MHT CET 2021 22th September Evening Shift

Options:

A. 5.27%

B. 7.17%

C. 10.3%

D. 4.5%

Answer: B

Solution:

$$\begin{aligned}\alpha_1 &= 10\% = \frac{10}{100} = 0.1, c_1 = 0.05\text{M} \\ c_2 &= 0.10\text{M}, \alpha_2 = ? \\ K_a &= \alpha_1^2 c_1 = \alpha_2^2 c_2 \\ \therefore (0.1)^2 \times 0.05 &= \alpha_2^2 \times 0.10 \\ \therefore \alpha_2^2 &= \frac{10^{-2} \times 5 \times 10^{-2}}{10^{-1}} = 50 \times 10^{-4} \\ \therefore \alpha_2 &= \sqrt{50 \times 10^{-4}} = 7.17 \times 10^{-2} = 7.17\%\end{aligned}$$

---

## Question139

The dissociation constant of weak monobasic acid is  $2.7 \times 10^{-5}$ . If degree of dissociation of acid is  $3 \times 10^{-2}$ , what is the concentration of acid?

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### Options:

- A. 0.24 M
- B. 0.03 M
- C. 0.3 M
- D. 0.11 M

**Answer: B**

### Solution:

To determine the concentration of the weak monobasic acid, we need to use the relationship between the dissociation constant ( $K_a$ ), the degree of dissociation ( $\alpha$ ), and the concentration of the acid ( $C$ ). The dissociation constant for a weak monobasic acid is given by the formula:

$$K_a = C \cdot \alpha^2$$

We are given the following values:

$$\text{Dissociation constant, } K_a = 2.7 \times 10^{-5}$$

$$\text{Degree of dissociation, } \alpha = 3 \times 10^{-2}$$

Now, let's substitute these values into the formula and solve for the concentration  $C$ :

$$2.7 \times 10^{-5} = C \cdot (3 \times 10^{-2})^2$$

Simplify the equation:

$$2.7 \times 10^{-5} = C \cdot 9 \times 10^{-4}$$

Dividing both sides by  $9 \times 10^{-4}$ :

$$C = \frac{2.7 \times 10^{-5}}{9 \times 10^{-4}}$$

$$C = \frac{2.7}{9} \times 10^{-5+4}$$

$$C = 0.3 \times 10^{-1}$$

$$C = 0.03M$$

Therefore, the concentration of the acid is **0.03 M**, which corresponds to **Option B**.

---

## Question140

The solubility product of a sparingly soluble salt  $AX_2$  is  $3.2 \times 10^{-8}$ . What is its solubility in  $\text{mol dm}^{-3}$  ?

## MHT CET 2021 22th September Morning Shift

### Options:

A.  $2.8 \times 10^{-4}$

B.  $1.6 \times 10^{-5}$

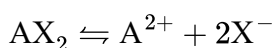
C.  $2.0 \times 10^{-3}$

D.  $4.0 \times 10^{-4}$

### Answer: C

### Solution:

To determine the solubility of the sparingly soluble salt  $AX_2$  in  $\text{mol dm}^{-3}$ , we need to use the solubility product constant ( $K_{sp}$ ). The dissociation of  $AX_2$  can be represented as follows:



Let the solubility of  $AX_2$  be  $S \text{ mol dm}^{-3}$ . At equilibrium, the concentrations of the ions will be:

$$[A^{2+}] = S$$

$$[X^-] = 2S$$

The solubility product  $K_{sp}$  is given by:

$$K_{sp} = [A^{2+}][X^-]^2$$

Substituting the equilibrium concentrations into the expression for  $K_{sp}$ , we get:

$$K_{sp} = S \cdot (2S)^2 = S \cdot 4S^2 = 4S^3$$

We are given that the solubility product of  $AX_2$  is  $3.2 \times 10^{-8}$ . Therefore, we can write:

$$4S^3 = 3.2 \times 10^{-8}$$

Solving for  $S$ , we get:

$$S^3 = \frac{3.2 \times 10^{-8}}{4} = 0.8 \times 10^{-8} = 8.0 \times 10^{-9}$$

Now, taking the cube root of both sides, we find:

$$S = \sqrt[3]{8.0 \times 10^{-9}} = 2 \times 10^{-3}$$

So, the solubility of  $AX_2$  in  $\text{mol dm}^{-3}$  is:

$$S = 2.0 \times 10^{-3} \text{ mol dm}^{-3}$$

Therefore, the correct option is:

**Option C**

$$2.0 \times 10^{-3}$$

---

## Question141

**Which of the following salt solutions is highly acidic?**

**MHT CET 2021 22th September Morning Shift**

**Options:**

- A. Ammonium acetate
- B. Ammonium cyanide
- C. Sodium chloride
- D. Ammonium nitrate

**Answer: D**

**Solution:**

$\text{NH}_4\text{NO}_3$  is a salt of strong acid,  $\text{HNO}_3$  and weak base,  $\text{NH}_4\text{OH}$ .

---

## Question142

**The  $[\text{OH}^-]$  in a solution is  $1 \times 10^{-12} \text{ mol dm}^{-3}$ . What is the concentration of  $\text{H}^+$  ions?**

**MHT CET 2021 21th September Evening Shift**

**Options:**

A.  $0.1 \text{ mol dm}^{-3}$

B.  $1.0 \text{ mol dm}^{-3}$

C.  $2.0 \text{ mol dm}^{-3}$

D.  $0.01 \text{ mol dm}^{-3}$

**Answer: D**

**Solution:**

$$[\text{OH}^-] = 1 \times 10^{-12} \text{ mol dm}^{-3}, [\text{H}^+] = ?$$

Product of hydrogen ion concentration and hydroxide ion concentration is equal to  $1.0 \times 10^{-14}$ .

$$\begin{aligned} [\text{H}^+] \times [\text{OH}^-] &= 1 \times 10^{-14} \\ \therefore [\text{H}^+] &= \frac{1 \times 10^{-14}}{1 \times 10^{-12}} = 1 \times 10^{-2} \\ &= 0.01 \text{ mol dm}^{-3} \end{aligned}$$

---

## Question143

**Which among the following salt solution in water shows pH greater than 7?**

**MHT CET 2021 21th September Evening Shift**

**Options:**

A. Sodium acetate

B. Sodium sulphate

C. Copper sulphate

D. Ammonium chloride

**Answer: A**

**Solution:**

Sodium acetate is a salt of weak acid, acetic acid and strong base, sodium hydroxide. Therefore, it is basic solution with pH greater than 7.

---

## Question144

The pH of 0.005 M KOH is 9.95. Calculate the  $[\text{OH}^-]$  ?

MHT CET 2021 21th September Evening Shift

Options:

A.  $6.71 \times 10^{-4} \text{ M}$

B.  $1.12 \times 10^{-4} \text{ M}$

C.  $4.45 \times 10^{-5} \text{ M}$

D.  $8.91 \times 10^{-5} \text{ M}$

**Answer: D**

**Solution:**

$$\text{pH} = 9.95, c = 0.05\text{M}$$

$$\text{pOH} = 14 - \text{pH} = 14 - 9.95 = 4.05$$

$$\text{pOH} = -\log_{10}[\text{OH}^-]$$

$$\log_{10}[\text{OH}^-] = -4.05$$

$$= -4 - 0.05 - 1 + 1$$

$$= -5 + 0.95 = \bar{5}.95$$

$$[\text{OH}^-] = \text{antilog } \bar{5}.95$$

$$= 8.913 \times 10^{-5}\text{M}$$

---

## Question145

Which among the following salts turns blue litmus red in it's aqueous solution?

## MHT CET 2021 21th September Morning Shift

### Options:

- A.  $\text{NH}_4\text{CN}$
- B.  $\text{NH}_4\text{F}$
- C.  $\text{CH}_3\text{COONa}$
- D.  $\text{CH}_3\text{COONH}_4$

**Answer: B**

### Solution:

$\text{NH}_4\text{CH}$  - Basic,  $\text{NH}_4\text{F}$  - Slightly acidic,  $\text{CH}_3\text{COONa}$  - Basic,  $\text{CH}_3\text{COONH}_4$  - Neutral

The solution of  $\text{NH}_4\text{F}$  is only slightly acidic and turns blue litmus red.

---

## Question146

What is the pH of 0.02 M NaOH solution?

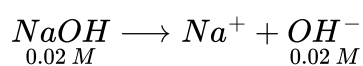
## MHT CET 2021 21th September Morning Shift

### Options:

- A. 10.3
- B. 11.3
- C. 11.7
- D. 12.3

**Answer: D**

### Solution:



$$\begin{aligned}P^{\text{OH}} &= -\log_{10} [\text{OH}^-] = -\log_{10}[0.02] \\ &= -\log_{10} (2 \times 10^{-2}) = 2 - 0.3010 \\ P^{\text{OH}} &= 1.6990 \\ \therefore P^{\text{H}} &= 14 - P^{\text{OH}} = 14 - 1.6990 = 12.301\end{aligned}$$

---

## Question147

**Dissociation constant of propionic acid is  $1.32 \times 10^{-5}$ . Calculate the degree of dissociation of acid in 0.05 M solution.**

**MHT CET 2021 21th September Morning Shift**

**Options:**

- A.  $2.6 \times 10^{-4}$
- B.  $1.61 \times 10^{-2}$
- C.  $1.90 \times 10^{-2}$
- D.  $3.5 \times 10^{-5}$

**Answer: B**

**Solution:**

$$\begin{aligned}K_a &= 1.32 \times 10^{-5}, c = 0.05\text{M}, \alpha = ? \\ \alpha &= \sqrt{\frac{K_a}{c}} = \sqrt{\frac{1.32 \times 10^{-5}}{0.05}} = \sqrt{2.64} \times 10^{-2} \\ \therefore \alpha &= 1.62 \times 10^{-2}\end{aligned}$$

---

## Question148

**Solubility of AgCl is  $7.2 \times 10^{-7}$  mol dm<sup>-3</sup>. What is it's solubility product?**

## MHT CET 2021 20th September Evening Shift

Options:

A.  $3.6 \times 10^{-13}$

B.  $7.2 \times 10^{-14}$

C.  $2.59 \times 10^{-14}$

D.  $5.18 \times 10^{-13}$

**Answer: D**

**Solution:**

$$S = 7.2 \times 10^{-7} \text{ mol dm}^{-3}, K_{\text{sp}} = ?$$

$$\text{For AgCl, } K_{\text{sp}} = S^2$$

$$= (7.2 \times 10^{-7})^2 = 5.18 \times 10^{-13}$$

---

## Question149

**A substance containing hydrogen and releasing  $\text{H}^+$  in aqueous medium is acid. Identify theory suggesting this concept, from following.**

## MHT CET 2021 20th September Evening Shift

Options:

A. Ostwald theory

B. Bronsted-Lowry theory

C. Arrhenius theory

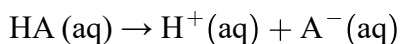
D. Lewis theory

**Answer: C**

## Solution:

The concept of a substance containing hydrogen and releasing  $H^+$  ions in an aqueous medium is primarily associated with the Arrhenius theory of acids and bases.

According to Arrhenius theory, an acid is a substance that increases the concentration of  $H^+$  ions when dissolved in water. Mathematically, this can be represented as:



Where "HA" represents the acid, " $H^+$ " is the hydrogen ion, and " $A^-$ " is the conjugate base.

Therefore, the correct answer is:

**Option C: Arrhenius theory**

---

## Question150

The solubility of AgCl in it's solution is  $1.25 \times 10^{-5} \text{ mol dm}^{-3}$ . What is solubility product of AgCl?

**MHT CET 2021 20th September Morning Shift**

**Options:**

A.  $1.56 \times 10^{-10}$

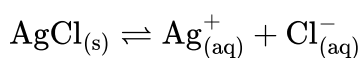
B.  $3.50 \times 10^{-6}$

C.  $1.10 \times 10^{-5}$

D.  $2.53 \times 10^{-3}$

**Answer: A**

## Solution:



Here,  $x = 1$ ,  $y = 1$

$$\therefore K_{sp} = S^2 = (1.25 \times 10^{-5})^2 = 1.56 \times 10^{-10}$$

---

## Question151

Which from following compounds accepts proton from water molecule according to Bronsted-Lowry theory?

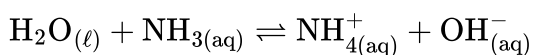
MHT CET 2021 20th September Morning Shift

Options:

- A.  $\text{NaOH}_{(\text{aq})}$
- B.  $\text{HCl}_{(\text{aq})}$
- C.  $\text{NH}_3_{(\text{aq})}$
- D.  $\text{NH}_4\text{OH}_{(\text{aq})}$

**Answer: C**

**Solution:**



## Question152

The pH of 0.1 M solution of monobasic acid is 2.34. Calculate the degree of dissociation of the acid.

MHT CET 2021 20th September Morning Shift

Options:

- A.  $3.1 \times 10^{-2}$
- B.  $4.5 \times 10^{-2}$
- C.  $2.18 \times 10^{-2}$

D.  $2.5 \times 10^{-3}$

**Answer: B**

**Solution:**

$$\text{pH} = 2.34, c = 0.1\text{M}$$

$$(i) \text{pH} = -\log_{10} [\text{H}^+]$$

$$\log_{10} [\text{H}^+] = -\text{pH} = -2.34$$

$$= -2 - 0.34 - 1 + 1$$

$$= -3 + 0.66 = \bar{3}.66$$

$$\therefore [\text{H}^+] = \text{antilog } \bar{3}.66$$

$$= 4.571 \times 10^{-3} \text{ mol dm}^{-3}$$

$$(ii) [\text{H}^+] = \alpha c$$

$$\therefore \alpha = \frac{[\text{H}^+]}{c} = \frac{4.571 \times 10^{-3}}{0.1} = 4.571 \times 10^{-2}$$

---

## Question153

**The precipitation power of an electrolyte increases with**

**MHT CET 2019 3rd May Morning Shift**

**Options:**

A. rise in temperature

B. atomic radii

C. ionic radii

D. charge of an ion

**Answer: D**

**Solution:**

The precipitation power of an electrolyte increases with charge of an ion. It can be explained on the basis of Hardy-Schulze rule. Greater the valence of the flocculating ion added, the greater is its power to cause precipitation.

---

## Question154

If the van't Hoff-factor for 0.1 M  $\text{Ba}(\text{NO}_3)_2$  solution is 2.74 , the degree of dissociation is

**MHT CET 2019 2nd May Morning Shift**

**Options:**

A. 0.87

B. 0.74

C. 0.91

D. 87

**Answer: A**

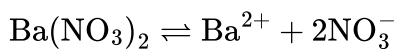
**Solution:**

Given,

Molarity = 0.1M

van't Hoff factor ( $i$ ) = 2.74

Since,  $i > 1$ , it means solute is undergoing dissociation.



Number of particles dissociated ( $n$ ) = 3

Now,  $\alpha$  ( degree of dissociation ) =  $\frac{(i-1)}{(n-1)}$

$$\begin{aligned} &= \frac{(2.74 - 1)}{(3 - 1)} \\ &= 0.87 \end{aligned}$$

---