Ionic Equilibrium

Question1

Identify conjugate acid and conjugate base for HCO_3^- ion respectively

MHT CET 2025 5th May Evening Shift

Options:

A.

CO_3^{2-} and H_2CO_3

B.

H_2CO_3 and CO_2

C.

CO_2 and H_2CO_3

D.

H_2CO_3 and CO_3^{2-}

Answer: D

Solution:

To find conjugate acid add H^+ to given HCO_3^-.

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

Hence, conjugate acid of HCO_3^- is H_2CO_3.

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

$\therefore$ The conjugate base of HCO_3^- is CO_3^{2-}.

Question2

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

($\mathrm{pK_a}$ of $\mathrm{CH_3COOH} = 4.5$)

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Options:

A.

0.1 M

B.

0.01M

C.

1.0 M

D.

10.0 M

Answer: C

Solution:

$[\mathrm{CH_3COOH}] = [\,\mathrm{Acid}\,] = 0.1\mathrm{M},$
$[\mathrm{CH_3COONa}] = [\mathrm{Salt}] = x\mathrm{M},$
$\mathrm{pK_a} = 4.5, \mathrm{pH} = 5.5$

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

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

$\log_{10} \frac{x}{0.1} = 1$
$\frac{x}{0.1} = \text{ Antilog } (1) = 10$
$x = 10 \times 0.1 = 1\mathrm{M}$

$\therefore$ The concentration of sodium acetate

$[CH_3COONa] = 1.0M$

Question3

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

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Options:

A.

3531×10^{-7}

B.

2.081×10^{-6}

C.

3.456×10^{-8}

D.

1.25×10^{-3}

Answer: D

Solution:

Given:

- Concentration of acid, $C = 0.08\,\text{mol dm}^{-3}$
- 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 (α)

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

If total acid concentration is $C = 0.08\,\mathrm{M}$, 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:

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

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: 1.25×10^{-3} ✅

Final Answer:

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

Question4

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

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Options:

A.

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

B.

4.92×10^{-4} moldm^{-3}

C.

2.00×10^{-4} moldm^{-3}

D.

3.52×10^{-5} moldm^{-3}

Answer: C

Solution:

$B_2 A_{(s)} \rightleftharpoons 2 B^{+}_{(aq)} + A^{2-}_{(aq)}$

Here, $x = 2, y = 1$

$$\begin{aligned} & K_{sp} = x^x y^y S^{x+y} = (1)^1 (2)^2 S^{1+2} = 4 S^3 \\ \therefore \quad & 3.2 \times 10^{-11} = 4 \times (S)^3 \\ & S = \sqrt[3]{8 \times 10^{-12}} \\ & = 2.00 \times 10^{-4} \text{ moldm}^{-3} \end{aligned}$$

Question5

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

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Options:

A.

$CH_3 - COONa$

B.

$CuCl_2$

C.

KCl

D.

KCl

Answer: A

Solution:

Sodium acetate (CH_3COONa) is a salt of strong base ($NaOH$) and weak acid (CH_3COOH). Hence, it is a basic in nature and it turns red litmus blue.

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^+ 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 α dissociates:

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

because each molecule gives 2 protons upon dissociation.

$[\mathrm{H}^+] = 2 \times 0.01 \times 0.02 = 0.0004\,\mathrm{M}$

Step 3: Calculate pH

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

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

$\mathrm{pH} = 3.39794 \approx 3.397$

Final Answer:

The pH = **3.397**

Correct Option: C (3.3970) ✅

Question7

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

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Options:

A. $4 \times 10^{-6}\ \mathrm{moldm}^{-3}$

B. $4 \times 10^{-7}\ \mathrm{moldm}^{-3}$

C. $7 \times 10^{-7}\ \mathrm{moldm}^{-3}$

D. $3 \times 10^{-8}\ \mathrm{moldm}^{-3}$

Answer: C

Solution:

We are given:

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

Step 1: Dissociation Equation

$\text{AgBr (s)} \rightleftharpoons \text{Ag}^{+}(aq) + \text{Br}^{-}(aq)$

Step 2: Expression for K_{sp}

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

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

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

Step 3: Solve for s

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

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

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

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:

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

Correct option: C ✅

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 α is small, approximate denominator as 1:

$K_a \approx c\alpha^2$

Step 2: Calculate α

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

Step 3: Convert to percentage

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

✅ **Final Answer:**

Option C — 10%

Question9

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

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Options:

A. $NaNO_3$

B. CH_3COONH_4

C. KCN

D. NH_4 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 → Neutral
- Strong acid + Weak base → Acidic
- Weak acid + Strong base → Basic
- Weak acid + Weak base → Depends on the relative strengths of conjugate acid and conjugate base.

Step 2: Analyze each option

(A) $NaNO_3$

NaOH (strong base) + HNO_3 (strong acid).

→ Neutral solution. Not basic.

(B) CH_3COONH_4

CH_3COO^- (from weak acid CH_3COOH) + NH_4^+ (from weak base NH_3).

Both weak → The pH depends on relative Ka and Kb.

- CH_3COOH has Ka ~ 1.8×10^{-5}
- NH_4^+ (Ka ~ 5.6×10^{-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) → neutral.

CN^- (from weak acid HCN) → conjugate base, undergoes hydrolysis:

$CN^- + H_2O \rightleftharpoons HCN + OH^-$

→ Solution is **basic**. ✅

(D) NH_4F

NH_4^+ (from weak base NH_3) = conjugate acid → acidic.

F^- (from weak acid HF) = conjugate base → basic.

Compare Ka values: Ka(NH_4^+) ≈ 5.6×10^{-10}; Ka(HF) ≈ 7.2×10^{-4} → 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

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

Reaction:

$B + H_2O \longrightarrow BH^+ + OH^-$

Step 2: Concentration of OH⁻

Initial conc of base $= 0.01\ M$.

Dissociation $= 10\% = 0.10$.

So,

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

Step 3: pOH

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

Step 4: pH

$\mathrm{pH} = 14 - \mathrm{pOH} = 14 - 3 = 11$

✅ **Final Answer:**

$\boxed{11}$

So the correct option is **C. 11**.

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. 2×10^{-9}

B. 8×10^{-7}

C. 6×10^{-7}

D. 4×10^{-6}

Answer: B

Solution:

Step 1: Write given values

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

Step 2: Formula for dissociation constant

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

$K_a = C\alpha^2$

(Approximate because α 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

$\boxed{8 \times 10^{-7}}$

Correct Option: B

Question12

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

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Options:

A. KCN

B. Na_2CO_3

C. $NaNO_3$

D. $CuCl_2$

Answer: D

Solution:

Step 1: Consider each option.

(A) KCN

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

→ CN^- undergoes hydrolysis to form OH^-.

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

(B) Na_2CO_3

Na^+ (neutral, from strong base) and CO_3^{2-} (conjugate base of H_2CO_3, weak acid).

→ CO_3^{2-} hydrolyzes to form OH^-.

Result: Basic solution. Not acidic.

(C) $NaNO_3$

Na^+ (neutral) and NO_3^- (conjugate base of HNO_3, a strong acid).

→ No hydrolysis, solution is neutral.

Does not affect litmus.

(D) $CuCl_2$

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

- Produces H^+ ions.

→ An acidic solution is formed.

Thus $CuCl_2$ solution turns **blue litmus red**.

✅ **Correct answer: Option D — $CuCl_2$**

Question13

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

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Options:

A. 1×10^{-4}

B. 2×10^{-4}

C. 4×10^{-4}

D. 3×10^{-2}

Answer: C

Solution:

$\mathrm{AB_{2(s)}} \rightleftharpoons \mathrm{A^{2+}_{(aq)}} + 2\,\mathrm{B^{-}_{(aq)}}$
$x = 1, y = 2$
$\mathrm{K_{sp}} = x^x \mathrm{y}^y\, \mathrm{S}^{x+y} = (1)^1 (2)^2\, \mathrm{S}^{1+2} = 4\,\mathrm{S}^3$
The molar solubility S of $\mathrm{AB_2}$ is given by

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

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 $C = 0.1\,\mathrm{M}$

Step 1: Relation between K_a, α, 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

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 OH^- ion per molecule when fully dissociated.
- Given **pH = 10** → then **pOH = 14 – 10 = 4**.
- So, $[OH^-] = 10^{-\text{pOH}} = 10^{-4}$ M.

Step 2: Relating to dissociation

The base concentration is $C = 0.01$ M.

If it dissociates to produce $[OH^-]$, then:

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

where α = degree of dissociation.

$\alpha = \frac{1.0\times10^{-4}}{0.01} = \frac{1.0\times10^{-4}}{1.0\times10^{-2}} = 0.01$

Step 3: Convert to percentage

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

✅ **Final Answer:**

$\boxed{1\%}$

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

Question16

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

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Options:

A. NH_4NO_3

B. Na_2CO_3

C. $NaNO_3$

D. $CuSO_4$

Answer: B

Solution:

(A) NH_4NO_3 : Salt of strong acid and weak base: Solution is acidic.

(B) Na_2CO_3 : Salt of weak acid and strong base: Solution is basic.

(C) $NaNO_3$: Salt of strong acid and strong base: Solution is neutral.

(D) $CuSO_4$: Salt of strong acid and weak base: Solution is acidic.

Question17

Calculate the equilibrium concentration of Pb^{++} ions in a solution of PbS containing 1×10^{-11} $moldm^{-3}$ of sulphide ions.

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

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Options:

A. 4×10^{-14}

B. 4×10^{-17}

C. 8×10^{-17}

D. 8×10^{-11}

Answer: C

Solution:

Step 1: Write the dissociation equation for PbS

$\mathrm{PbS}_{(s)} \rightleftharpoons \mathrm{Pb}^{2+}_{(aq)} + \mathrm{S}^{2-}_{(aq)}$

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

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

Step 3: Substitute the given values

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

Put these values into the K_{sp} formula:

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

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

Divide both sides by 1×10^{-11}:

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

Final Answer:

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

Question18

Which among the following salts forms basic solution in water?

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Options:

A. NH_4NO_3

B. NH_4Cl

C. Na_2CO_3

D. Na_2SO_4

Answer: C

Solution:

Step 1: Analyse each option

Option A: NH_4NO_3

- This salt is formed from a **weak base (NH_4OH)** and a **strong acid (HNO_3)**.
- Solution is **acidic** due to NH_4^+ hydrolysis.

✅ Acidic, not basic.

Option B: NH_4Cl

- This salt is formed from a **weak base (NH_4OH)** and a **strong acid (HCl)**.
- Solution is **acidic** due to NH_4^+.

✅ Acidic, not basic.

Option C: Na_2CO_3

- This salt is formed from a **strong base (NaOH)** and a **weak acid (H_2CO_3)**.
- In water, CO_3^{2-} hydrolyzes:

  $CO_3^{2-} + H_2O \leftrightarrow HCO_3^- + OH^-$

- Produces **OH^- ions**, making the solution **basic**.

✅ This is **basic**.

Option D: Na_2SO_4

- This salt is formed from a **strong base (NaOH)** and a **strong acid (H_2SO_4)**.
- Hydrolysis is negligible; solution is **neutral**.

✅ Neutral.

✅ **Correct Answer:**

Option C: Na_2CO_3

Question19

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

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Options:

A. 3.5×10^{-6}

B. 2.0×10^{-5}

C. 2.3×10^{-4}

D. 2.5×10^{-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,
- α is the degree of dissociation.

Step 2. Values given:

- $C = 0.01$ 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: 2.5×10^{-5} ✅

Question20

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

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Options:

A. 1×10^{-3}

B. 1×10^{-6}

C. 1×10^{-9}

D. 1×10^{-12}

Answer: D

Solution:

We are given:

- Solubility of a sparingly soluble binary salt = 1.12×10^{-4} g/dm^3.
- Molar mass of salt = 112 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

$$\text{Molar solubility} = \frac{\text{given solubility in g/dm}^3}{\text{molar mass}} = \frac{1.12\times10^{-4}}{112}$$

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

Step 2: Write dissociation

For a binary salt AB:

$AB \rightleftharpoons A^+ + B^-$

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:**

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

Question21

Identify conjugate acid-base pair from following equilibrium reaction.

$$HSO^-_{3(aq)} + H_3O^+_{(aq)} \rightleftharpoons H_2SO_3 + H_2O$$

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Options:

A. H_2SO_3 and HSO_3^-

B. HSO_3^- and H_3O^+

C. H_2SO_3 and H_2O

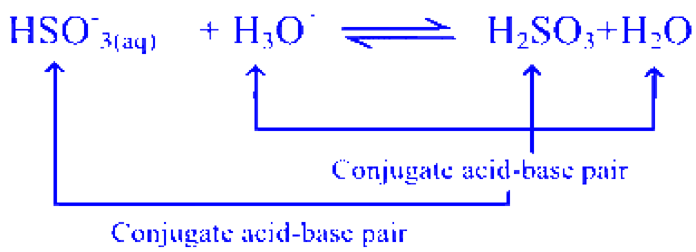
D. H_3O^+ and H_2SO_3

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.

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Options:

A. 2.25×10^{-7}

B. 3.05×10^{-7}

C. 2.5×10^{-5}

D. 3.725×10^{-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: $C = 0.001\,\mathrm{M}$
- Degree of dissociation: $\alpha = 1.5\% = 0.015$

Step 2: Expression for base dissociation constant

For a weak base:

$B + H_2O \rightleftharpoons BH^+ + OH^-$

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:**

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

Question23

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

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Options:

A. 1.69×10^{-3}

B. 7.0×10^{-3}

C. 2.45×10^{-3}

D. 6.18×10^{-3}

Answer: B

Solution:

$NiS_{(s)} \rightleftharpoons Ni^{2+}_{(aq)} + S^{2-}_{(aq)}$
$x = 1, y = 1$
$K_{sp} = x^x y^y S^{x+y} = (1)^1(1)^1 S^{1+1} = S^2$
The molar solubility S of NiS is given by
$S = \sqrt{K_{sp}}$
$= \sqrt{4.9 \times 10^{-5}}$
$S = 7.0 \times 10^{-3}\ moldm^{-3}$

Question24

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

$$HPO^{2-}_{4(aq)} + H_2O_{(\ell)} \rightleftharpoons PO^{3-}_{4(aq)} + H_3O^+_{(eq)}$$

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Options:

A. $H_3O^+$and H_2O

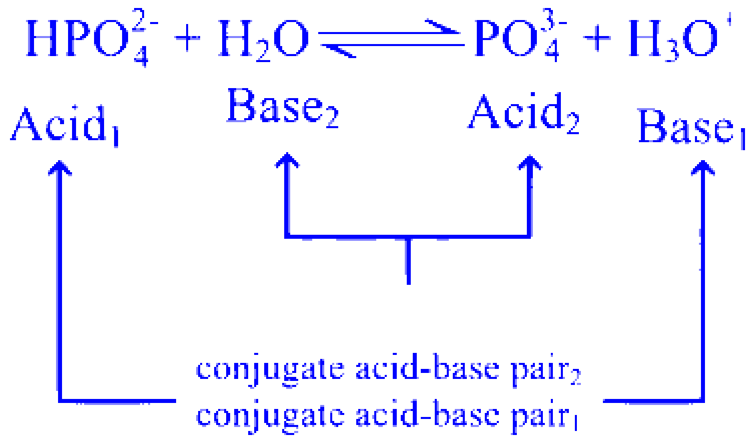
B. H_2O and HPO_4^{2-}

C. PO_4^{3-} and H_3O^+

D. $H_3O^+$and HPO_4^{2-}

Answer: A

Solution:



Question25

The solubility of sparingly soluble salt $\mathrm{AX_2}$ is 1×10^{-4} $\mathrm{moldm^{-3}}$ at 298 K . Calculate its solubility product.

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Options:

A. 2×10^{-12}

B. 4×10^{-12}

C. 2×10^{-10}

D. 4×10^{-10}

Answer: B

Solution:

Step 1: Write the dissociation reaction

$\mathrm{AX_2}(s) \rightleftharpoons \mathrm{A^{2+}}(aq) + 2\mathrm{X^-}(aq)$

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: 4×10^{-12}

Question26

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

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Options:

A. 6×10^{-5}

B. 4×10^{-5}

C. 2×10^{-5}

D. 1×10^{-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)
- α = degree of dissociation

Step 2: Insert given values

- Concentration $C = 0.1\,M$
- 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 $\left[\mathrm{OH}^{-}\right]$ ion concentration of a weak monoacidic base?

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Options:

A. $\mathrm{K_b.c}$

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

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):

$B + H_2O \rightleftharpoons BH^+ + OH^-$

Let the initial concentration of the base be c. Suppose degree of ionisation = α.

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, α is small).

So:

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

Step 3: Solve for α

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

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. 1.04×10^{-8}

B. 1.44×10^{-6}

C. 1.30×10^{-6}

D. 1.18×10^{-5}

Answer: B

Solution:

Step 1: Write the data

- Weak acid concentration: $c = 0.01\,\mathrm{M}$
- 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 6.4×10^{-5} moldm^{-3}. Calculate the value of solubility product at the same temperature?

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Options:

A. 5.06×10^{-10}

B. 4.096×10^{-9}

C. 3.05×10^{-10}

D. 2.8×10^{-9}

Answer: B

Solution:

Step 1: Write the dissociation equilibrium.

$CaCO_3(s) \leftrightharpoons Ca^{2+}(aq) + CO_3^{2-}(aq)$

Step 2: Define solubility.

Let solubility of $CaCO_3$ be $s = 6.4 \times 10^{-5}$ mol dm^{-3}.

Then at equilibrium:

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

$[CO_3^{2-}] = s = 6.4 \times 10^{-5}\,M$

Step 3: Expression for solubility product.

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

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 PbI_2 is 1.08×10^{-7}.

Calculate its solubility in moldm^{-3} at 298 K .

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Options:

A. 2.018×10^{-3}

B. 2.011×10^{-9}

C. 1.259×10^{-9}

D. 3.0×10^{-3}

Answer: D

Solution:

Given:

Solubility product of PbI_2, $K_{sp} = 1.08 \times 10^{-7}$

Let the solubility of PbI_2 be S mol/L.

Step 1: Write the dissociation equation.

$PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2\,I^-(aq)$

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

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

Step 3: Write K_{sp} expression.

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

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

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 2.7×10^{-8} as:

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

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

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

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

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

Option D is correct.

Question31

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

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Options:

A. 9.704

B. 10.301

C. 8.621

D. 8.750

Answer: B

Solution:

Given:

- Molarity of NaOH $(C) = 0.01\,\text{M}$
- Percentage dissociation $= 2\%$

Step 1: Calculate the concentration of OH^- ions formed.

NaOH dissociates as:

$\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$

The concentration of 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.

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

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 $25^\circ\mathrm{C}$,

$pH + pOH = 14$

So,

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

Final Answer:

Option B is correct.

$\boxed{10.301}$

Question32

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

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Options:

A. KCN

B. $NaNO_3$

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:

$CN^- + H_2O \rightarrow HCN + OH^-$

The production of OH^- ions makes the solution basic, which turns red litmus blue.

Other options:

- $NaNO_3$, NaCl, and 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 $\mathrm{NaOH} = 40\ \mathrm{g\ mol^{-1}}$)

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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

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

Step 2: Calculate molarity (M) of NaOH solution

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:

$\mathrm{NaOH} \rightarrow \mathrm{Na}^+ + \mathrm{OH}^-$

Therefore,

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

Step 4: Calculate pOH

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

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

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

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

So,

$\mathrm{pOH} = -(-0.6990) = 0.6990$

Step 5: Calculate pH using the relation

At 298 K,

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

$\mathrm{pH} = 14 - \mathrm{pOH} = 14 - 0.6990 = 13.3010$

Final Answer:

$\boxed{13.3010}$ (Option D)

Question34

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

What is solubility product of the salt?

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Options:

A. 1.55×10^{-8}

B. 2.56×10^{-10}

C. 3.60×10^{-8}

D. 6.41×10^{-10}

Answer: B

Solution:

Let the salt be $\mathrm{BA_2}$.

When $\mathrm{BA_2}$ dissolves in water, it dissociates as:

$\mathrm{BA_2}(s) \rightleftharpoons \mathrm{B^{2+}}(aq) + 2\mathrm{A^-}(aq)$

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

Concentrations at equilibrium:

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

Expression for solubility product K_{sp}:

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

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

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

Question35

Solubility of $\mathrm{Ca_3(PO_4)_2}$ is ' S ' $\mathrm{moldm^{-3}}$. Find solubility product.

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Options:

A. $\mathrm{S^5}$

B. $108\ \mathrm{S^5}$

C. $54\ \mathrm{S^5}$

D. $12\ \mathrm{S^5}$

Answer: B

Solution:

Step 1: Write the Dissolution Equation

When $\mathrm{Ca_3(PO_4)_2}$ dissolves in water, it breaks into ions like this:

$\mathrm{Ca_3(PO_4)_2 \rightleftharpoons 3\,Ca^{2+} + 2\,PO_4^{3-}}$

Step 2: Relate Solubility to Ion Concentrations

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

- For every 1 mole of salt that dissolves, 3 moles of $\mathrm{Ca^{2+}}$ form. So $[\mathrm{Ca^{2+}}] = 3S$.
- For every 1 mole of salt, 2 moles of $\mathrm{PO_4^{3-}}$ form. So $[\mathrm{PO_4^{3-}}] = 2S$.

Step 3: Write the Expression for Solubility Product (K_{sp})

The solubility product is:$K_{\mathrm{sp}} = [\mathrm{Ca^{2+}}]^3 \times [\mathrm{PO_4^{3-}}]^2$Substitute the values from Step 2:$K_{\mathrm{sp}} = (3S)^3 \times (2S)^2$

Step 4: Calculate and Simplify

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

Question36

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

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Options:

A. 7.08×10^{-13}

B. 3.67×10^{-13}

C. 5.89×10^{-13}

D. 5.04×10^{-13}

Answer: D

Solution:

$$\mathrm{AgBr}_{(\mathrm{s})} \rightleftharpoons \mathrm{Ag}^{+}_{(\mathrm{aq})} + \mathrm{Br}^{-}_{(\mathrm{aq})}$$

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

$$\therefore \quad \mathrm{K}_{\mathrm{sp}} = x^x y^y\, \mathrm{S}^{x+y} = (1)^1(1)^1\, \mathrm{S}^{1+1} = \mathrm{S}^2$$

$$\therefore \quad \mathrm{K}_{\mathrm{sp}} = \left(7.1 \times 10^{-7}\right)^2 = 5.04 \times 10^{-13}$$

Question37

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

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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, CH_3COOH) and its salt (sodium acetate, CH_3COONa), we use the **Henderson-Hasselbalch equation**:

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

Given:

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

Step 1: Write the equation with values:

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

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:

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

Final Answer:

$\boxed{5.44} \qquad \text{(Option C)}$

Question38

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

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Options:

A. 4×10^{-4}

B. 4×10^{-6}

C. 5×10^{-6}

D. 5×10^{-7}

Answer: D

Solution:

Given:

pH of saturated solution of $Ba(OH)_2$ = 12

Step 1: Calculate pOH

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

$12 + \mathrm{pOH} = 14$

$\mathrm{pOH} = 2$

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

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

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

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

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

$\mathrm{Ba(OH)_2}(s) \rightleftharpoons \mathrm{Ba}^{2+}(aq) + 2\mathrm{OH}^-(aq)$

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

Then,

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

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

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

So,

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

Step 4: Calculate K_{sp}

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

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

$= s \times 4s^2$

$= 4s^3$

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: $\boxed{5 \times 10^{-7}}$

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? $(\mathrm{pK_a} = 4.680)$

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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, $[HA] = 0.01$ M

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

We use the Henderson–Hasselbalch equation for buffer solutions:

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

Step 1: Substitute the values:

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

Step 2: Plug the values into the formula:

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

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:

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

Final Answer:

The pH of the buffer solution is

$$\boxed{4.981}$$

Correct option: **C (4.981)**

Question40

If pH of solution changes from 4 to 5 , then the H_3O^+ 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 H_3O^+ ion concentration:

$\mathrm{pH} = -\log_{10}[H_3O^+]$

Step 2: Let the initial H_3O^+ ion concentration (at pH 4) be $[H_3O^+]_1$:

$4 = -\log_{10}[H_3O^+]_1 \implies [H_3O^+]_1 = 10^{-4}$ mol/L

Step 3: The new concentration at pH 5, $[H_3O^+]_2$:

$5 = -\log_{10}[H_3O^+]_2 \implies [H_3O^+]_2 = 10^{-5}$ mol/L

Step 4: Compare the concentrations:

$\frac{[H_3O^+]_2}{[H_3O^+]_1} = \frac{10^{-5}}{10^{-4}} = 10^{-1} = 0.1$

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 moldm^{-3} ?

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Options:

A. 1.97×10^{-4}

B. 1.738×10^{-4}

C. 1.84×10^{-4}

D. 1.283×10^{-4}

Answer: B

Solution:

Given:

pH of vinegar, $= 3.76$

We know,

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

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

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

Substitute the value of pH:

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

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,

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

Final answer:

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

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

Question42

The solubility product of a sparingly soluble salt AX is 4.9×10^{-13}. What is its solubility in moldm^{-3} ?

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Options:

A. 2.4×10^{-13}

B. 4.9×10^{-7}

C. 7.0×10^{-7}

D. 7.0×10^{-13}

Answer: C

Solution:

Let the sparingly soluble salt be AX.

When AX dissolves in water, it dissociates as:

$AX(s) \rightleftharpoons A^{+}(aq) + X^{-}(aq)$

Let the solubility of AX be S 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,

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

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

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

Correct Answer:

Option C: 7.0×10^{-7} **mol dm**$^{-3}$

Question43

Which among the following salts is NOT hydrolysed in water?

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Options:

A. Na_2CO_3

B. NH_4CN

C. KNO_3

D. KCN

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: Na_2CO_3

- NaOH (strong base) + H_2CO_3 (weak acid)
- Salt from strong base and weak acid.
- **Hydrolysed** in water.

Option B: NH_4CN

- NH_4OH (weak base) + HCN (weak acid)
- Salt from weak base and weak acid.
- **Hydrolysed** in water.

Option C: KNO_3

- KOH (strong base) + HNO_3 (strong acid)
- Salt from strong base and strong acid.
- **NOT hydrolysed** in water.

Option D: KCN

- KOH (strong base) + HCN (weak acid)
- Salt from strong base and weak acid.
- **Hydrolysed** in water.

Final answer:

Option C: 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. CH_3COOH

B. $H_2C_2O_4$

C. HCl

D. NH_3

Answer: D

Solution:

$$\underset{\text{(Base)}}{NH_3} + \underset{\text{(acid)}}{H_2O} \rightleftharpoons \overset{+}{N}H_4 + OH^-$$

Question45

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

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Options:

A. 3.4

B. 4.5

C. 5.1

D. 5.8

Answer: A

Solution:

$$\begin{aligned}
&\mathrm{HA} \rightleftharpoons \mathrm{H}^{+}_{(\mathrm{aq})} + \mathrm{A}^{-}_{(\mathrm{aq})} \\
&[\mathrm{H}]^{+} = \alpha \mathrm{C} = \frac{2}{100} \times 0.02\ \mathrm{m} = 0.0004\ \mathrm{m} \\
&\mathrm{pH} = -\log\left[\mathrm{H}^{+}\right] = -\log(0.0004) = 3.4
\end{aligned}$$

Question46

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

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Options:

A. 3.24×10^{-10}

B. 2.44×10^{-10}

C. 1.64×10^{-10}

D. 4.00×10^{-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:

$\mathrm{BA\ (s)} \rightleftharpoons \mathrm{B^{+}(aq)} + \mathrm{A^{-}(aq)}$

The equilibrium expression for the solubility product is given by:

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

Given the solubility of BA is $1.8 \times 10^{-5}\ \mathrm{mol\ dm^{-3}}$, in a saturated solution at equilibrium, the concentration of $\mathrm{B^{+}}$ ions is $1.8 \times 10^{-5}\ \mathrm{mol\ dm^{-3}}$, and the concentration of $\mathrm{A^{-}}$ ions is also $1.8 \times 10^{-5}\ \mathrm{mol\ dm^{-3}}$, 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:** 3.24×10^{-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:

$$\mathrm{H_2SO_{4(aq)}} + \mathrm{2H_2O}_{(l)} \longrightarrow \mathrm{2H_3O^+_{(aq)}} + \mathrm{SO^{2-}_{4(aq)}}$$

$$\begin{aligned} \text{Hence, } \left[\mathrm{H_3O^+}\right] &= 2 \times \mathrm{c} = 2 \times 0.01\mathrm{M} \\ &= 2 \times 10^{-2}\mathrm{M} \end{aligned}$$

$$\begin{aligned} \mathrm{pH} &= -\log_{10}\left[\mathrm{H_3O^+}\right] \\ &= -\log_{10}\left(2 \times 10^{-2}\right) \\ &= -\log_{10} 2 - \log_{10} 10^{-2} \\ &= -\log_{10} 2 + 2 \\ &= 2 - 0.3010 \\ \mathrm{pH} &= 1.699 \end{aligned}$$

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 $\mathrm{mol dm}^{-3}$ of sparingly soluble salt BA if its solubility product 4.9×10^{-13} at same temperature.

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Options:

A. 7.0×10^{-7}

B. 7.5×10^{-7}

C. 8.0×10^{-7}

D. 4.9×10^{-7}

Answer: A

Solution:

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

$BA(s) \rightleftharpoons B^{+}(aq) + A^{-}(aq)$

The equilibrium expression of the solubility product is:

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

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

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

Thus, the K_{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:

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

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

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

Question50

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

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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 (OH^-). A strong base like NaOH dissociates completely in water:

$\mathrm{NaOH} \rightarrow \mathrm{Na}^{+} + \mathrm{OH}^{-}$

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

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

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

Substituting the concentration of hydroxide ions:

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

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

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

Thus, substituting the value of pOH:

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

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 (CH_3COOH) and sodium acetate (CH_3COONa). The equilibrium can be represented by:

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

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 NH_4OH is 4.2×10^{-2}. What is the percent dissociation of NH_4OH ?

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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 α, is given as 4.2×10^{-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:

Percent dissociation $= \alpha \times 100 = 4.2 \times 10^{-2} \times 100$

$= 4.2\%$

Therefore, the percent dissociation of NH_4OH is 4.2%.

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

Question53

Conjugate acid of NH_2^- and NH_3 are respectively

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Options:

A. NH_4OH and NH_2OH

B. NH_3 and NH_2^-

C. NH_3 and $NH_4{}^+$

D. $NH_4{}^+$ and NH_3

Answer: C

Solution:

Definition of Conjugate Acid :

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

Conjugate Acid of NH_2^- :

Starting with NH_2^- (the amide ion), add an H^+ :

$NH_2^- + H^+ \rightarrow NH_3$.

Thus, the conjugate acid of NH_2^- is NH_3.

Conjugate Acid of NH_3 :

Starting with NH_3 (ammonia), add an H^+ :

$NH_3 + H^+ \rightarrow NH_4^+$.

Thus, the conjugate acid of NH_3 is NH_4^+.

Conclusion :

Conjugate acid of NH_2^- is NH_3.

Conjugate acid of NH_3 is NH_4^+.

Correct Option :

C) NH_3 and NH_4^+

Question54

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

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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,

$$
\begin{aligned}
\mathrm{pOH} & =\mathrm{pK}_{\mathrm{b}}+\log _{10} \frac{[\text { Salt }]}{[\text { Base }]} \\
& =7.744+\log _{10} \frac{1}{0.2} \\
& =7.744+0.698
\end{aligned}
$$

$\therefore \quad \mathrm{pOH}=8.44$

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

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

Question55

The dissociation constant of a weak monobasic acid is 3.2×10^{-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 $CaCO_3$ is 7×10^{-5} $moldm^{-3}$ at $25^{\circ}C$. What is its solubility product at same temperature?

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Options:

A. 6.7×10^{-9}

B. 9.0×10^{-9}

C. 1.12×10^{-9}

D. 4.9×10^{-9}

Answer: D

Solution:

The solubility product (K_{sp}) of a sparingly soluble salt like $CaCO_3$ can be determined from its solubility. Given that the solubility of $CaCO_3$ is 7×10^{-5} mol dm^{-3}, the dissolution of $CaCO_3$ in water can be represented by the equation:

$$CaCO_3(s) \rightleftharpoons Ca^{2+}(aq) + CO_3^{2-}(aq)$$

The solubility of $CaCO_3$ provides the concentration of Ca^{2+} ions and CO_3^{2-} ions in a saturated solution:

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

$$[CO_3^{2-}] = 7 \times 10^{-5} \text{ mol dm}^{-3}$$

The solubility product expression for $CaCO_3$ is:

$$K_{sp} = [Ca^{2+}][CO_3^{2-}]$$

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 $CaCO_3$ at $25^\circ C$ is:

Option D: 4.9×10^{-9}

Question57

Which among the following is the conjugate base of $HClO_4$?

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Options:

A. ClO_4^{2-}

B. ClO_4

C. HCl

D. ClO_4^-

Answer: D

Solution:

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

$HClO_4 + H_2O \rightleftharpoons H_3O^+ + ClO_4^-$

$\therefore$ Conjugate base of $HClO_4$ is ClO_4^-.

Question58

What is pH of a centimolar solution of H_2SO_4 ?

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Options:

A. 1.7

B. 2.0

C. 3.2

D. 4.5

Answer: A

Solution:

$$\begin{aligned} [H_2SO_4] &= 2 \times C = 2 \times 0.01 M = 2 \times 10^{-2} M \\ pH &= -\log_{10}\left[H^+\right] = -\log_{10}\left[2 \times 10^{-2}\right] \\ &= -\log_{10} 2 - \log_{10} 10^{-2} \\ &= -\log_{10} 2 + 2 \\ \therefore \quad pH &= 2 - 0.3010 = 1.7 \end{aligned}$$

Question59

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

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Options:

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

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

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

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

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:

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

Where:

$pK_a = -\log(K_a)$

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

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

Step 1: Calculate pK_a.

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

$pK_a = -\log(6.6 \times 10^{-10}) \approx 9.18$

Step 2: Use the Henderson-Hasselbalch equation.

Plug in the given concentrations into the equation:

$[A^-] = 0.02\, M\, (NaCN)$

$[HA] = 0.01\, M\, (HCN)$

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

The logarithmic term becomes:

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

Thus, the pH of the solution is:

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

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

To find the hydrogen ion concentration:

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

Therefore, the concentration of H^+ ions in the solution is:

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

Question60

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

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Options:

A. 2×10^{-2}

B. 4×10^{-4}

C. 5×10^{-5}

D. 2.5×10^{-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:

$\mathrm{HA} \rightleftharpoons \mathrm{H}^+ + \mathrm{A}^-$

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

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

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

$[H^+] = 0.05 \times 0.02\,M = 0.001\,M$

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

The concentration of undissociated $[HA]$ is

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

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:

5×10^{-5}

Thus, the dissociation constant of the acid is 5×10^{-5}.

Question61

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

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Options:

A. Na_2SO_4

B. CH_3COONa

C. NH_4NO_3

D. $CuCl_2$

Answer: B

Solution:

The pH of a solution depends on hydrolysis of the salt. The salt Na_2SO_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 NH_4OH and $Cu(OH)_2$. Only CH_3COONa comes from strong base 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.

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Options:

A. 2.0×10^{-9}

B. 3.0×10^{-9}

C. 4.0×10^{-9}

D. 5.0×10^{-9}

Answer: D

Solution:

$$\mathrm{HA} \rightleftharpoons \mathrm{H}^{\oplus}_{(aq)} + \mathrm{A}^{\ominus}_{(aq)}$$
$$\alpha = 0.05\% = 0.05 \times 10^{-2}$$
$$\mathrm{K_a} = \alpha^2\mathrm{C} = \left(0.05 \times 10^{-2}\right)^2 \times 0.02\mathrm{M}$$
$$= 5.0 \times 10^{-9}$$

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 1.8×10^{-5}.

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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:

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

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

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

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

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

Substitute the value of K_a:

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

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.

Question64

Which from the following equations represents the relation between solubility ($\mathrm{mol\ L^{-1}}$) and solubility product for a salt $\mathrm{B_3A_2}$?

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Options:

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

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

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

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

Answer: A

Solution:

The equation that represents the relation between solubility (S) and solubility product (K_{sp}) for a salt $\mathrm{B_3A_2}$ can be determined by considering the dissolution equation and the corresponding expression for the solubility product:

When the salt $\mathrm{B_3A_2}$ dissolves, it dissociates as follows:

$\mathrm{B_3A_2}(s)\rightleftharpoons 3\mathrm{B}^{2+}(aq)+2\mathrm{A}^{3-}(aq)$

If S is the molar solubility, then at equilibrium:

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

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

The solubility product K_{sp} is given by:

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

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

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

Question65

Calculate solubility $\left(\mathrm{moldm}^{-3}\right)$ of a sparingly soluble electrolyte AB at 298 K if its solubility product is 1.6×10^{-5} ?

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Options:

A. 1.6×10^{-3}

B. 2.5×10^{-3}

C. 4.0×10^{-3}

D. 8.0×10^{-3}

Answer: C

Solution:

For AB,

$\mathrm{AB_{(s)}} \rightleftharpoons \mathrm{A^{+}_{(qq)}} + \mathrm{B^{-}_{(aq)}}$

Here , $x = 1, \mathrm{y} = 1$

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

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

Question66

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

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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?

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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 $\left(HCO_3^- + H_2CO_3\right)$ buffer.

Question68

Calculate the solubility of sparingly soluble salt BA in moldm^{-3} at 300 K if its solubility product is 4.9×10^{-9} at same temperature.

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Options:

A. 5.72×10^{-5}

B. 6.40×10^{-5}

C. 7.00×10^{-5}

D. 7.81×10^{-5}

Answer: C

Solution:

$$BA_{(s)} \rightleftharpoons B^+_{(aq)} + A^-_{(qq)}$$
$$x = 1, y = 1$$
$$K_{sp} = x^x y^y\ S^{x+y} = (1)^1(1)^1\ S^{1+1} = S^2$$
$$\therefore \quad S = \sqrt{K_{sp}} = \sqrt{4.9 \times 10^{-9}}$$
$$= \sqrt{49 \times 10^{-10}}$$
$$= 7.00 \times 10^{-5} M$$

Question69

Calculate the pH of buffer solution containing 0.04 M NaF and 0.02 M HF $[pK_a = 3 \cdot 142]$.

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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.

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

Question70

What is the pOH of millimolar solution of $Ca(OH)_2$?

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Options:

A. 2.7

B. 10.3

C. 12.3

D. 11.3

Answer: A

Solution:

$$\begin{aligned} & 1 \text{ mM} = 10^{-3} \text{M} \\ & Ca(OH)_2 \longrightarrow Ca^{2+} + 2OH^- \\ & 10^{-3} \text{M} \qquad\qquad 2 \times 10^{-3} \text{M} \\ \therefore \quad & pOH = -\log_{10}[OH] = -\log_{10}\left(2 \times 10^{-3}\right) = 2.7 \end{aligned}$$

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.

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Options:

A. 0.0112

B. 0.0341

C. 0.0694

D. 0.292

Answer: B

Solution:

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

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

where:

Λ_c is the molar conductivity at a given concentration c, and

Λ_0 is the molar conductivity at infinite dilution.

Given values are:

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

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

Calculate α:

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

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**.

Question72

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

$$\mathrm{HCl} + \mathrm{NH_3} \rightleftharpoons \mathrm{NH_4^+} + \mathrm{Cl^-}$$

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Options:

A. $\mathrm{Cl^-}$and $\mathrm{NH_4^+}$

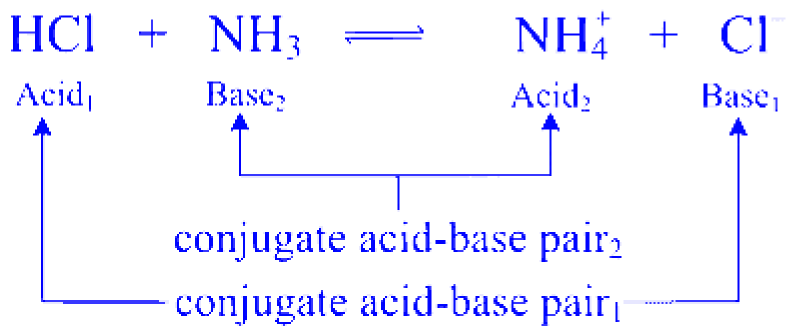
B. HCl and NH_3

C. NH_4^+ and NH_3

D. NH_4^+ and HCl

Answer: C

Solution:



Question73

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

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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 (α) by the equation:

$K_a = C\alpha^2$

Given:

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

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:

$C = \frac{1.8\times10^{-5}}{(0.03)^2}$

Calculate $(0.03)^2$:

$(0.03)^2 = 0.0009$

Now calculate the concentration C:

$C = \frac{1.8\times10^{-5}}{0.0009} = 2 \times 10^{-2}\,\mathrm{M} = 0.02\,\mathrm{M}$

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

Option B: 0.02 M

Question74

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

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Options:

A. Ammonium fluoride

B. Ammonium cyanide

C. Ammonium acetate

D. Ammonium chloride

Answer: D

Solution:

Ammonium chloride (NH_4Cl) is a salt of strong acid, HCl and weak base, NH_4OH.

Question75

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

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Options:

A. 4.810×10^{-13}

B. 5.184×10^{-13}

C. 6.454×10^{-13}

D. 5.925×10^{-13}

Answer: B

Solution:

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

$\mathrm{BA} \rightleftharpoons \mathrm{B}^{+} + \mathrm{A}^{-}$

is given by the expression:

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

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

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

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

Substituting these values into the expression for the solubility product:

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

Calculating the above expression:

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

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

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

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

Question76

Calculate the $[\mathrm{OH}]$ if pOH of solution is 4.94

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Options:

A. 2.356×10^{-5}M

B. 1.417×10^{-5}M

C. 1.881×10^{-5}M

D. 1.148×10^{-5}M

Answer: D

Solution:

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

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

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

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

Using a calculator, we find:

$[OH^-] \approx 1.148 \times 10^{-5} M$

Therefore, the correct answer is Option D: $1.148 \times 10^{-5} M$.

Question77

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

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Options:

A. $NH_4OH + NH_4Cl$

B. $C_6H_5COOH + C_6H_5COONa$

C. $HCOOH + HCOOK$

D. $CH_3COOH + CH_3COONa$

Answer: A

Solution:

Basic buffer is the mixture of weak base and its salt with strong acid. In option A, NH_4OH is weak base and NH_4Cl is salt of strong acid and weak base. Hence, $NH_4OH + NH_4Cl$ is basic buffer. Other options are examples of acidic buffer.

Question78

Calculate the pH of a buffer solution containing 0.35 M weak acid and 0.70 M of its salt with strong base if pK_a is 4.56 .

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Options:

A. 6.11

B. 3.72

C. 4.86

D. 5.65

Answer: C

Solution:

For an acidic buffer,

$$pH = pK_a + \log_{10}\frac{[\text{salt}]}{[\text{acid}]}$$

$$\therefore \quad pH = 4.56 + \log_{10}\frac{0.70}{0.35}$$

$$= 4.56 + \log_{10} 2 = 4.56 + 0.3010$$

$$= 4.861$$

Question79

Calculate the concentration of weak monobasic acid if its degree of dissociation and dissociation constant are 5.0×10^{-4} and 5.0×10^{-9} respectively.

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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 (α) and dissociation constant (K_a), the following relationship is used:

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

where:

α 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 α^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.

Question80

Calculate pH of 0.002 M KOH solution.

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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:

$KOH \rightarrow K^+ + OH^-$

This means that the concentration of hydroxide ions $[OH^-]$ will be equal to the concentration of KOH, which is 0.002 M.

Calculate pOH:

The pOH is calculated using the formula:

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

Substitute the concentration of hydroxide ions:

$pOH = -\log(0.002)$

Perform the logarithmic calculation:

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

Using logarithmic properties:

$pOH = -(\log(2) + \log(10^{-3}))$

$pOH = -(\log(2) - 3)$

$pOH = 3 - \log(2)$

Approximate value of $\log(2) \approx 0.301$

$pOH \approx 3 - 0.301 = 2.699$

Calculate pH:

Use the relationship between pH and pOH, given by:

$pH + pOH = 14$

Substitute the calculated pOH:

$pH = 14 - pOH$

$pH = 14 - 2.699$

$pH \approx 11.301$

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 ' α ' for 0.1 M acetic acid $\left(K_a = 1.0 \times 10^{-5}\right)$

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Options:

A. 10^{-2}

B. 10^{-3}

C. 10^{-4}

D. 10^{-5}

Answer: A

Solution:

For a weak acid HA, α is very small, or $(1 - \alpha) \cong 1$

$\therefore \quad \alpha = \sqrt{\frac{K_a}{c}} = \sqrt{\frac{1.0 \times 10^{-5}}{0.1}} = 10^{-2}$

Question82

Calculate $\left[H_3O^+\right]$in 0.02 M solution of monobasic acid if dissociation constant is 1.8×10^{-5}.

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Options:

A. $3.0 \times 10^{-4} \mathrm{M}$

B. $6.0 \times 10^{-4} \mathrm{M}$

C. $2.0 \times 10^{-4} \mathrm{M}$

D. $4.0 \times 10^{-4} \mathrm{M}$

Answer: B

Solution:

To calculate the concentration of hydronium ions $[\mathrm{H_3O^+}]$ in a 0.02 M solution of a monobasic acid, given that the dissociation constant (K_a) is 1.8×10^{-5}, we can use the expression for the dissociation of a weak acid:

For a monobasic acid, HA, which dissociates as:

$\mathrm{HA} \rightleftharpoons \mathrm{H^+} + \mathrm{A^-}$

The expression for the dissociation constant K_a is:

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

Assuming x is the concentration of $\mathrm{H^+}$ ions that dissociate, we have:

$[\mathrm{H^+}] = x$

$[\mathrm{A^-}] = x$

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:

$x=\sqrt{3.6\times10^{-7}}$

$x\approx6.0\times10^{-4}$

Thus, the concentration of hydronium ions $[H_3O^+]$ in the solution is:

Option B:

6.0×10^{-4} M

Question83

Which of the following equations represents the relation between solubility and solubility product for salt BA_3?

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Options:

A. $S=\left(\frac{K_{sp}}{27}\right)^{1/4}$

B. $S=\left(27\times K_{\text{sp}}\ \right)^{1/4}$

C. $S=\left(\frac{K_{sp}}{4}\right)^{1/4}$

D. $\mathrm{S}=(4\times\mathrm{K}_{\text{sp}})^{1/4}$

Answer: A

Solution:

For a salt with the formula BA_3, it dissociates in water according to the following equation:

$BA_3(s)\rightarrow B^{3+}(aq)+3A^-(aq)$

If the solubility of the salt BA_3 is S, then the concentration of B^{3+} ions will be S and the concentration of A^- ions will be $3S$.

The solubility product K_{sp} for this dissociation can be expressed as:

$K_{sp}=[B^{3+}][A^-]^3=(S)(3S)^3$

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 $\mathrm{pK_a}$ is 4.2.

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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:

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

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:

$\mathrm{pH} = 4.2 + \log\left(\frac{0.054}{0.027}\right)$

Calculate the ratio:

$\frac{0.054}{0.027} = 2$

Therefore, the equation becomes:

$\mathrm{pH} = 4.2 + \log(2)$

The logarithm of 2 is approximately 0.301:

$\log(2) \approx 0.301$

Now substitute this value back into the equation:

$\mathrm{pH} = 4.2 + 0.301 = 4.501$

Thus, the pH of the buffer solution is approximately 4.5, which corresponds to Option A.

Question85

Calculate $[\mathrm{H_3O^+}]$ of a monobasic acid if it is 0.04% dissociated in 0.05 M solution.

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Options:

A. 1×10^{-5}

B. 1.5×10^{-5}

C. 2.0×10^{-5}

D. 3.0×10^{-5}

Answer: C

Solution:

To find the concentration of hydronium ions, $[\mathrm{H_3O^+}]$, 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:

$HA \rightarrow H^+ + A^-$

Given a $0.05\,\mathrm{M}$ solution with 0.04% dissociation, calculate the resulting concentration of $[\mathrm{H_3O^+}]$:

Convert the percentage dissociation into a decimal:

$0.04\% = \frac{0.04}{100} = 0.0004$

Calculate the concentration of $[\mathrm{H_3O^+}]$:

$[\mathrm{H_3O^+}] = 0.0004 \times 0.05\,\mathrm{M}$

$= 2.0 \times 10^{-5}\,\mathrm{M}$

Thus, the concentration of hydronium ions $[\mathrm{H_3O^+}]$ in the solution is $2.0 \times 10^{-5}\,\mathrm{M}$, corresponding to option C.

Question86

Calculate the solubility product of sparingly soluble salt BA at 300 K if its solubility is $9.1 \times 10^{-3} \mathrm{mol dm}^{-3}$ at same temperature.

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Options:

A. 9.635×10^{-5}

B. 9.012×10^{-5}

C. 8.281×10^{-5}

D. 7.816×10^{-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:

$\mathrm{BA}\,(s) \rightleftharpoons \mathrm{B}^+\,(aq) + \mathrm{A}^-\,(aq)$

For each mole of BA that dissolves, it dissociates into one mole of B^+ and one mole of A^- ions. If the solubility of BA is 9.1×10^{-3} mol dm^{-3}, then the concentrations of both B^+ and A^- ions in the solution at equilibrium will also be 9.1×10^{-3} 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):

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

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 8.281×10^{-5}, which corresponds to Option C.

Question87

Which from following species acts as base [1], according to Bronsted-Lowry theory?

$\mathrm{HCl} + \mathrm{NH_3}$ 目的 $\mathrm{NH_4^+} + \mathrm{Cl^-}$

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Options:

A. $\mathrm{Cl^-}$

B. $\mathrm{NH_3}$

C. NH_4^+

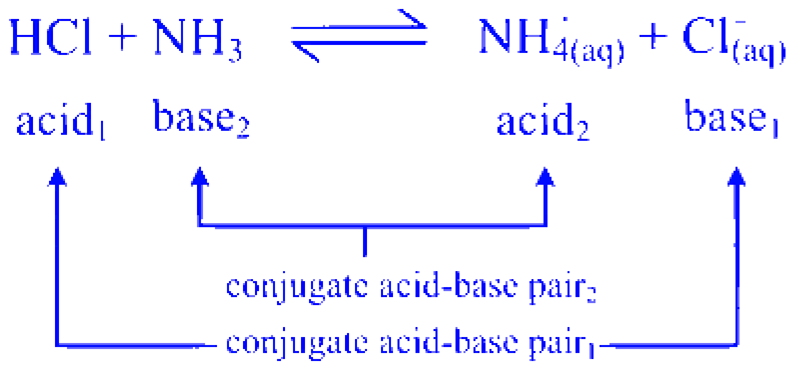
D. HCl

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 $\left(\mathrm{pK_a} = 6.2\right)$.

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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:

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

Given that:

pH = 7.2

$\mathrm{pK}_a = 6.2$

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 $0.4 \mathrm{~M} \mathrm{NH}_4\mathrm{OH}$ and $0.5 \mathrm{~M} \mathrm{NH}_4\mathrm{Cl}$ solutions. $(\mathrm{pK}_{\mathrm{b}} = 4.730)$

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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:

$$\mathrm{pOH} = \mathrm{pK}_b + \log\left(\frac{[\text{Salt}\ ]}{[\text{Base}\ ]}\right)$$

Given:

$\mathrm{pK}_b = 4.730$

$[\text{Base}] = [\mathrm{NH_4OH}] = 0.4\,\mathrm{M}$

$[\text{Salt}] = [\mathrm{NH_4Cl}] = 0.5\,\mathrm{M}$

Substitute these values into the Henderson-Hasselbalch equation:

$$\mathrm{pOH} = 4.730 + \log\left(\frac{0.5}{0.4}\right)$$

Calculate the logarithmic term:

$\frac{0.5}{0.4} = 1.25$

$\log(1.25) \approx 0.0969$

Now, substitute back into the equation:

$\mathrm{pOH} = 4.730 + 0.0969$

$\mathrm{pOH} \approx 4.8269$

Rounding to the appropriate significant figures, we get:

$\mathrm{pOH} \approx 4.83$

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?

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Options:

A. NH_4CN

B. NH_4Cl

C. NH_4NO_3

D. $NaNO_3$

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: NH_4CN

When dissolved in water, NH_4CN dissociates into NH_4^+ (the ammonium ion) and CN^- (the cyanide ion). The ammonium ion NH_4^+ is a weak acid, while CN^- is the conjugate base of the weak acid hydrocyanic acid (HCN). The CN^- ion will hydrolyze in water, producing OH^- ions and leading to a basic solution. Thus, this solution can turn red litmus blue.

Option B: NH_4Cl

When NH_4Cl is dissolved in water, it dissociates into NH_4^+ and Cl^-. Here, NH_4^+ acts as a weak acid, but 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: NH_4NO_3

Upon dissolution, NH_4NO_3 dissociates into NH_4^+ and NO_3^-. Just like Cl^-, the NO_3^- ion, originating from a strong acid (HNO_3), is neutral. Hence, the solution is slightly acidic due to the presence of NH_4^+ and would not turn red litmus blue.

Option D: $NaNO_3$

This salt dissociates into Na^+ and NO_3^-. Both ions come from strong parents; Na^+ from NaOH (strong base) and NO_3^- from 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:** NH_4CN.

Question91

The solubility product of $\mathrm{PbCl_2}$ at $298\ \mathrm{K}$ is 3.2×10^{-5}. What is its solubility in $\mathrm{moldm^{-3}}$?

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Options:

A. 8×10^{-6}

B. 2×10^{-2}

C. 5.6×10^{-3}

D. 5.0×10^{-2}

Answer: B

Solution:

$$\begin{aligned}
& \mathrm{K_{sp}} = 3.2 \times 10^{-5} \\
& \mathrm{PbCl_{2(s)}} \rightleftharpoons \mathrm{Pb^{2+}_{(aq)}} + 2\mathrm{Cl^{-}_{(aq)}} \\
& \therefore \quad \mathrm{K_{sp}} = 4\,\mathrm{S}^3 \\
& \therefore \quad 4\,\mathrm{S}^3 = 3.2 \times 10^{-5} \\
& \mathrm{S} = \sqrt[3]{\frac{3.2 \times 10^{-5}}{4}} \\
& = \sqrt[3]{8 \times 10^{-6}} \\
& = 2 \times 10^{-2}\ \mathrm{mol\ dm^{-3}}
\end{aligned}$$

Question92

Identify base$_2$ for following equation according to Bronsted-Lowry theory.

$$\mathrm{HCl_{(aq)}} + \mathrm{H_2O}_{(l)} \rightleftharpoons \mathrm{H_3O^{+}_{(aq)}} + \mathrm{Cl^{-}_{(aq)}}$$

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Options:

A. $H_3O^+_{(aq)}$

B. $H_2O_{(l)}$

C. $Cl^-_{(aq)}$

D. $HCl_{(aq)}$

Answer: B

Solution:

The Bronsted-Lowry theory defines an acid as a substance that can donate a proton (H^+ ion), and a base as a substance that can accept a proton. For the given chemical reaction:

$HCl_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$

In this reaction, hydrochloric acid (HCl) donates a proton to water (H_2O), which acts as a base and accepts the proton, forming hydronium ion (H_3O^+) and chloride ion (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.
- H_2O is a Bronsted-Lowry base because it accepts a proton from HCl.
- H_3O^+ is the conjugate acid formed after H_2O accepts the proton.
- 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: $H_2O_{(l)}$

Question93

An organic monobasic acid has dissociation constant 2.25×10^{-6}. What is percent dissociation in its 0.01 M solution?

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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:

$HA \rightleftharpoons H^+ + A^-$

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 (α) 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 α 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 α, 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:

$\text{Percent dissociation} = \alpha \times 100\%$

Percent dissociation $= 1.5 \times 10^{-2} \times 100\%$

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 $0.01\ \mathrm{M}$ strong dibasic acid.

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Options:

A. 5.5

B. 2.5

C. 2.0

D. 1.7

Answer: D

Solution:

$$
\begin{aligned}
&\left[\mathrm{H_3O^+}\right] = 2 \times \mathrm{c} = 2 \times 0.01\mathrm{M} = 2 \times 10^{-2}\mathrm{M} \\
&\mathrm{pH} = -\log_{10}\left[\mathrm{H_3O^+}\right] \\
&= -\log_{10}\left[2 \times 10^{-2}\right] \\
&= -\log_{10} 2 - \log_{10} 10^{-2} \\
&= -\log_{10} 2 + 2 \\
&= 2 - 0.3010 \\
&\mathrm{pH} = 1.7
\end{aligned}
$$

When $\mathrm{H_3O^+}$ concentration is $2 \times 10^{-2}\mathrm{M}$, the pH should be less than 2. Only option (D) is valid.

Question95

Calculate dissociation constant of $0.001\mathrm{M}$ weak monoacidic base undergoing 2% dissociation.

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Options:

A. 4×10^{-7}

B. 2×10^{-6}

C. 2×10^{-7}

D. 1×10^{-7}

Answer: A

Solution:

$$\alpha = \frac{\text{Percent dissociation}}{100} = \frac{2}{100} = 0.02$$
$$\mathrm{K_a} = \alpha^2 \mathrm{c} = (0.02)^2 \times 0.001 = 4 \times 10^{-7}$$

Question96

What is the pH of solution containing $4.62 \times 10^{-4} \mathrm{MH}^{+}$ ions?

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Options:

A. 8.62

B. 4.64

C. 5.66

D. 3.34

Answer: D

Solution:

$\left[H^{+}\right]=4.62 \times 10^{-4} M$
$pH=\ ?$
Using, $pH=-\log \left[H^{+}\right]=-\log \left[4.62 \times 10^{-4}\right]$
$=-\log [4.62]-\log \left[10^{-4}\right]=-0.66+4$
$=3.34$

Question97

Which among the following salt solution in water is acidic in nature?

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Options:

A. $CuCl_2$

B. NH_4CN

C. KCN

D. CH_3COONa

Answer: A

Solution:

$CuCl_2$ is an acidic in water as Cl^{-} do not hydrolyse. Since, it is an anion of strong acid, HCl. Cu^{2+} will hydrolyse to form $Cu(OH)^{-}$.

$Cu^{2+}+H_2O(aq) \rightleftharpoons Cu(OH)^{-}(aq)+H^{+}(aq)$ which is an acidic solution. So, the solution of $CuCl_2$ in water is acidic.

Question98

What is the molar concentration of acetic acid if value of it's, dissociation constant is 1.8×10^{-5} and degree of dissociation is 0.02 ?

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Options:

A. 4.6×10^{-3} M

B. 4.5×10^{-2} M

C. 4.0×10^{-4} M

D. 3.6×10^{-2} M

Answer: B

Solution:

$k = 1.8 \times 10^{-5}$
$\alpha = 0.02$
$C = ?$

For acetic acid, dissociation is given as

$$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$$

	CH_3COOH	CH_3COO^-	H^+
Initial conc.	C	0	0
Conc. at equilibrium	$C - C\alpha$	$C\alpha$	$C\alpha$

$$\therefore k = \frac{\alpha^2 C^2}{(1-\alpha)C} = \frac{\alpha^2 C}{(1-\alpha)} \quad [\alpha <<< 1 \Rightarrow (1-\alpha) \approx 1]$$

$$1.8 \times 10^{-5} = \frac{(0.02)^2 C}{1}$$

$$\text{or} \quad C = 4.5 \times 10^{-2} \mathrm{M}$$

Question99

A buffer solution is prepared by mixing equimolar acetic acid and sodium acetate. If 'K_d' of acetic acid is 1.78×10^{-5}, find the pH of buffer solution.

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Options:

A. 4.75

B. 8.9

C. 9.4

D. 2.6

Answer: A

Solution:

For acidic buffer,

$$\mathrm{pH} = \mathrm{pK_a} + \log_{10}\frac{[\text{ Salt }]}{[\text{ Acid }]} = \mathrm{pK_a} + \log_{10}\frac{1}{1} = \mathrm{pK_a}$$

$$\begin{aligned} \mathrm{pK_a} &= -\log_{10}\ \mathrm{K_a} = -\log_{10}\left(1.78 \times 10^{-5}\right) \\ &= -\left(\log_{10} 1.78 + \log_{10} 10^{-5}\right) \\ &= -(0.25 - 5) = 4.75 \end{aligned}$$

Therefore, pH of buffer solution $= 4.75$

Question100

The solubility product of $\mathrm{Mg(OH)_2}$ is 1.8×10^{-11} at $298\ \mathrm{K}$. What is its solubility in $\mathrm{mol\ dm^{-3}}$?

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Options:

A. 1.650×10^{-4}

B. 2.120×10^{-4}

C. 3.184×10^{-4}

D. 4.550×10^{-4}

Answer: A

Solution:

$$\mathrm{Mg(OH)_{2(s)}} \rightleftharpoons \mathrm{Mg^{2+}_{(aq)}} + 2\mathrm{OH^-_{(aq)}}$$

Here, $x = 1, \mathrm{y} = 2$

$\therefore \quad \mathrm{K_{sp}} = x^x \mathrm{y^y S^{\mathit{x}+y}} = (1)^1(2)^2\mathrm{S}^{1+2} = 4\mathrm{S}^3$

$\therefore \quad \mathrm{S} = \sqrt[3]{\frac{\mathrm{K_{sp}}}{4}} = \sqrt[3]{\frac{1.8\times10^{-11}}{4}}$

$= \sqrt[3]{4.5 \times 10^{-12}} = 1.650 \times 10^{-4}$

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 $\mathrm{K_{sp}}$ is solubility product of $\mathrm{Al(OH)_3}$, its solubility is expressed by formula,

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Options:

A. $\sqrt[3]{\frac{4}{\mathrm{K_{sp}}}}$

B. $\sqrt[3]{\frac{K_{sp}}{4}}$

C. $\sqrt[4]{\frac{K_{sp}}{27}}$

D. $\sqrt[4]{\mathrm{K_{sp}} \times 27}$

Answer: C

Solution:

$$\mathrm{Al(OH)_{3(s)} \rightleftharpoons Al^{3+}_{(aq)} + 3OH^-_{(aq)}}$$

Here, $x = 1, \mathrm{y} = 3$

$\therefore \quad \mathrm{K_{sp}} = x^x \mathrm{y^y S^{x+y}} = (1)^1(3)^3\,\mathrm{S}^{1+3} = 27\,\mathrm{S}^4$

$\therefore \quad \mathrm{S} = \sqrt[4]{\frac{\mathrm{K_{sp}}}{27}}$

Question102

What is the solubility of $\mathrm{AgCl_{(s)}}$ if its solubility product is 1.6×10^{-10} ?

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Options:

A. $1.26 \times 10^{-5}\mathrm{M}$

B. $1.00 \times 10^{-9}\mathrm{M}$

C. $2.6 \times 10^{-5}\mathrm{M}$

D. $1.56 \times 10^{-9}\mathrm{M}$

Answer: A

Solution:

$\mathrm{AgCl_{(s)} \rightleftharpoons Ag^+_{(sq)} + Cl^-_{(sq)}}$

$x = 1, \mathrm{y} = 1$

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

$\mathrm{S} = \sqrt{\mathrm{K_{sp}}} = \sqrt{1.6 \times 10^{-10}}$

$= 1.26 \times 10^{-5}\mathrm{M}$

Question103

A buffer solution is prepared by mixing $0.01 \mathrm{~M}$ weak acid and $0.05 \mathrm{~M}$ solution of a salt of weak acid and strong base. What is the pH of buffer solution? $(\mathrm{pKa}=4.74)$

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Options:

A. 3.34

B. 4.80

C. 5.44

D. 6.93

Answer: C

Solution:

$$
\begin{aligned}
\mathrm{pH} & =\mathrm{pK}_{\mathrm{a}}+\log _{10} \frac{[\text { salt }]}{[\text { acid }]} \\
& =4.74+\log _{10} \frac{0.05}{0.01} \\
& =4.74+0.7=5.44
\end{aligned}
$$

Question104

Which among the following is NOT an example of salt of strong acid and weak base?

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Options:

A. NH_4Cl

B. NH_4NO_3

C. $CuSO_4$

D. Na_2SO_4

Answer: D

Solution:

Na_2SO_4 is a salt of strong base (NaOH) and strong acid (H_2SO_4).

Question105

Find $[OH]$ if a monoacidic base is 3% ionised in its $0.04\ \mathrm{M}$ solution.

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Options:

A. $3.1 \times 10^{-2}\ \mathrm{mol\ L}^{-1}$

B. $4.5 \times 10^{-3}\ \mathrm{mol\ L}^{-1}$

C. $9.0 \times 10^{-2}\ \mathrm{mol\ L}^{-1}$

D. $1.2 \times 10^{-3}\ \mathrm{mol\ L}^{-1}$

Answer: D

Solution:

For a monoacidic base,

$$\left[\mathrm{OH}^{-}\right]=\mathrm{c} \times \alpha$$
$$\left[\mathrm{OH}^{-}\right]=0.04 \times \frac{3}{100}$$
$$\left[\mathrm{OH}^{-}\right]=1.2 \times 10^{-3}\ \mathrm{mol\ L}^{-1}$$

Question106

What is the expression for solubility product of silver chromate if it's solubility is expressed as $\mathrm{S\ mol\ L^{-1}}$?

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Options:

A. $2\ S^2$

B. $3\ S^3$

C. $4\ S^3$

D. $27\ S^4$

Answer: C

Solution:

$$\mathrm{Ag_2CrO_{4(s)} \rightleftharpoons 2Ag^+_{(aq)} + CrO^{2-}_{4(aq)}}$$

$$\text{Here, } x = 2, \mathrm{y} = 1$$

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

Question107

What is pH of solution containing $50\ \mathrm{mL}$ each of $0.1\ \mathrm{M}$ sodium acetate and $0.01\ \mathrm{M}$ acetic acid? $(\mathrm{pK_a CH_3COOH} = 4.50)$

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Options:

A. 2.5

B. 3.5

C. 4.5

D. 5.5

Answer: D

Solution:

For acidic buffer solution,

$$pH = pK_a + \log_{10} \frac{[\text{ Salt }]}{[\text{ Acid }]};$$
$$pH = 4.50 + \log_{10} \frac{0.1}{0.01};$$
$$pH = 4.50 + \log 10; pH = 5.50$$

For acidic buffer, if [Salt] > [Acid], then $pH > pK_a$ of acid. Hence, only option (D) is valid.

Question108

Identify the salt that undergoes hydrolysis and forms acidic solution from following.

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Options:

A. Na_2CO_3

B. NH_4NO_3

C. NH_4CN

D. KCN

Answer: B

Solution:

(A) Na_2CO_3 : Salt of weak acid and strong base: Solution is basic.

(B) NH_4NO_3 : Salt of strong acid and weak base: Solution is acidic.

(C) NH_4CN : 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.

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Options:

A. 5.5×10^{-5}

B. 4.0×10^{-5}

C. 1.8×10^{-5}

D. 1.0×10^{-5}

Answer: D

Solution:

Percent dissociation $= 1.42\%$

$\therefore \quad \alpha = 0.0142$

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 $25^{\circ}C$ $\left[\Lambda_c = 15.0 \Omega^{-1}\ cm^2\ mol^{-1}\right.$ and $\left.\Lambda_0 = 300 \Omega^{-1}\ cm^2\ mol^{-1}\right]$

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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?

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Options:

A. 2.30

B. 12.6

C. 11.7

D. 3.2

Answer: C

Solution:

$$\underset{0.005\ \mathrm{M}}{\mathrm{NaOH}} \longrightarrow \underset{0.005\ \mathrm{M}}{\mathrm{Na}^{+}} + \underset{0.005\ \mathrm{M}}{\mathrm{OH}^{-}}$$

$$
\begin{aligned}
\mathrm{pOH} &= -\log_{10}[\mathrm{OH}] \\
&= -\log_{10}[0.005] \\
&= -\log_{10}\left(5 \times 10^{-3}\right) \\
&= -\left[\log 10^{-3} + \log 5\right] \\
&= -[-3 + 0.699] \\
&= -[-2.301] \\
&= 2.301 \\
\mathrm{pOH} + \mathrm{pH} &= 14 \\
\therefore \quad \mathrm{pH} &= 14 - 2.301 \\
&= 11.699 \\
&\approx 11.7
\end{aligned}
$$

Question112

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

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Options:

A. NH_4NO_3

B. NH_4Cl

C. NH_4CN

D. NH_4F

Answer: C

Solution:

NH_4CN is the salt of weak acid HCN $\left(K_a = 4.0 \times 10^{-10}\right)$ and weak base NH_4OH $\left(K_b = 1.8 \times 10^{-5}\right)$ showing that $K_a < K_b$. If $K_a < K_b$, the aqueous solution of the salt will be basic. Thus, the solution of NH_4CN is basic and turns red litmus blue.

Question113

What is the concentration of OH^{-} ion in a solution containing $0.05 \mathrm{~M} \mathrm{~H}^{+}$ ions?

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Options:

A. $2.5 \times 10^{-13} \mathrm{M}$

B. $5.0 \times 10^{-2} \mathrm{M}$

C. $2.0 \times 10^{-13} \mathrm{M}$

D. $4.2 \times 10^{-12} \mathrm{M}$

Answer: C

Solution:

$$
\begin{aligned}
& \mathrm{K}_{\mathrm{w}}=\left[\mathrm{H}_3 \mathrm{O}^{+}\right]\left[\mathrm{OH}^{-}\right] \\
\therefore \quad & 10^{-14}=0.05 \times\left[\mathrm{OH}^{-}\right] \\
& [\mathrm{OH}]=\frac{10^{-14}}{0.05} \\
& \left[\mathrm{OH}^{-}\right]=2.0 \times 10^{-13} \mathrm{M}
\end{aligned}
$$

Question114

Calculate the pH of $1.36 \times 10^{-2} \mathrm{M}$ solution of perchloric acid.

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Options:

A. 1.43

B. 1.86

C. 2.43

D. 2.86

Answer: B

Solution:

Perchloric acid is a strong monobasic acid.

Hence, $\left[H_3O^+\right] = 1.36 \times 10^{-2} M$

$$\begin{aligned} \therefore \quad pH &= -\log_{10}\left[H_3O^+\right] \\ &= -\log_{10}\left[1.36 \times 10^{-2}\right] \\ &= -\log_{10} 1.36 - \log_{10} 10^{-2} \\ &= -\log_{10} 1.36 + 2 \\ &= 2 - 0.1335 \\ \therefore \quad pH &= 1.86 \end{aligned}$$

Question115

Find solubility in terms of $\mathrm{mol\ L^{-1}}$ if solubility product of silver bromide is 6.4×10^{-13}.

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Options:

A. $4.0 \times 10^{-5}\ \mathrm{mol\ L^{-1}}$

B. $8.0 \times 10^{-7}\ \mathrm{mol\ L^{-1}}$

C. $7.5 \times 10^{-5}\ \mathrm{mol\ L^{-1}}$

D. $6.4 \times 10^{-4}\ \mathrm{mol\ L^{-1}}$

Answer: B

Solution:

$AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$

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

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

$\therefore \quad S=\sqrt{K_{sp}}=\sqrt{6.4 \times 10^{-13}}=\sqrt{64 \times 10^{-14}}$

$=8 \times 10^{-7} \text{ mol L}^{-1}$

Question116

Calculate the concentration of H^{+} ions in a solution if pOH is 11.

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Options:

A. 10^{-11} M

B. 10^{-8} M

C. 10^{-6} M

D. 10^{-3} M

Answer: D

Solution:

$$\begin{aligned} & \mathrm{pOH}=11 \\ & \mathrm{pH}+\mathrm{pOH}=14 \\ \therefore \quad & \mathrm{pH}=14-\mathrm{pOH}=14-11=3 \\ & \mathrm{pH}=-\log_{10}[\mathrm{H}^{+}] \\ \therefore \quad & [\mathrm{H}^{+}]=10^{-\mathrm{pH}}=10^{-3}\mathrm{M} \end{aligned}$$

Question117

Acetic acid dissociated to 1.20% in its $0.01\ \mathrm{M}$ solution. What is the value of its dissociation constant?

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Options:

A. 2.20×10^{-2}

B. 1.60×10^{-4}

C. 1.44×10^{-6}

D. 2.40×10^{-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 7.0×10^{-9}.

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Options:

A. 1.21×10^{-3} mol L^{-1}

B. 3.228×10^{-3} mol L^{-1}

C. 2.831×10^{-3} mol L^{-1}

D. 1.811×10^{-3} mol L^{-1}

Answer: A

Solution:

For PbI_2,

$PbI_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2I^{-}_{(aq)}$

$x = 1, y = 2$

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

$\therefore \quad S = \sqrt[3]{\frac{Ksp}{4}} = \sqrt[3]{\frac{7.0\times10^{-9}}{4}} = 1.21 \times 10^{-3}\ \mathrm{mol\ L^{-1}}$

Question119

What is the pH of a solution containing $2.2 \times 10^{-6}\mathrm{M}$ hydrogen ions?

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Options:

A. 6.34

B. 5.66

C. 4.34

D. 3.80

Answer: B

Solution:

$$\begin{aligned} \mathrm{pH} = -\log_{10}\left[H_3O^{+}\right] &= -\log_{10}\left[2.2 \times 10^{-6}\right] \\ &= 6 - \log_{10}(2.2) = 6 - 0.34 \\ &= 5.66 \end{aligned}$$

Question120

What is the concentration of $\left[H_3O^{+}\right]$ ion in $\mathrm{mol\ L^{-1}}$ of $0.001\ \mathrm{M}$ acetic acid $(\alpha = 0.134)$?

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Options:

A. 1.34×10^{-4}

B. 1.54×10^{-4}

C. 1.80×10^{-4}

D. 1.70×10^{-4}

Answer: A

Solution:

$$\begin{aligned} \left[H_3O^{+}\right] &= \alpha \times c = 0.134 \times 0.001 \\ &= 1.34 \times 10^{-4} \text{ mol dm}^{-3} \end{aligned}$$

Question121

A buffer solution is prepared by mixing 0.2 M sodium acetate and 0.1 M acetic acid. If pK_a for acetic acid is 4.7, find the pH.

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Options:

A. 3.0

B. 4.0

C. 5.0

D. 2.0

Answer: C

Solution:

Using Henderson's equation,

$$pH = pK_a + \log_{10} \frac{[\text{Salt}]}{[\text{Acid}]}$$
$$pH = 4.7 + \log_{10} \frac{0.2}{0.1}$$
$$pH = 4.7 + \log_{10} 2 = 4.7 + 0.3010 = 5.001$$

For acidic buffer, if [Salt] > [Acid], then $pH > pK_a$ of acid. Hence, only option (C) is valid.

Question122

Which salt from following forms aqueous solution having pH less than 7 ?

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Options:

A. CH_3COONa

B. Na_2SO_4

C. $CuSO_4$

D. Na_2CO_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). $CuSO_4$ is salt of strong acid H_2SO_4 and weak base $Cu(OH)_2$.

Question123

Solubility of a salt $A_2 B_3$ is 1×10^{-3} mol dm^{-3}. What is the value of its solubility product?

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Options:

A. 1.08×10^{-13}

B. 8.1×10^{-15}

C. 2.7×10^{-15}

D. 2.0×10^{-13}

Answer: A

Solution:

$A_2 B_{3(s)} \rightleftharpoons 2 A^{3+}_{(aq)} + 3 B^{2-}_{(aq)}$

$$\begin{aligned} &\text{Here, } x = 2, \mathrm{y} = 3 \\ &\mathrm{K_{sp}} = x^x \mathrm{y^y S^{x+y}} \\ &\quad = (2)^2 (3)^3 \mathrm{S}^{2+3} \\ &\quad = 108 \mathrm{~S}^5 \\ &\mathrm{K_{sp}} = 108 \times \left(1 \times 10^{-3}\right)^5 \\ &\quad = 108 \times 10^{-15} \\ &\quad = 1.08 \times 10^{-13} \end{aligned}$$

Question124

Which of the following aqueous solutions of salts will have highest pH value?

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Options:

A. CH_3COONH_4

B. Na_2CO_3

C. NH_4Cl

D. NaCl

Answer: B

Solution:

Na_2CO_3 is the only alkaline solution having highest pH value. It is salt of weak acid (H_2CO_3) and strong base (NaOH).

Question125

What is the pH of $2 \times 10^{-3}\mathrm{M}$ solution of monoacidic weak base if it ionises to the extent of 5% ?

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Options:

A. 14

B. 6

C. 4

D. 2

Answer: C

Solution:

$\left[H^{+}\right] = c\alpha = 2 \times 10^{-3} \times 5 \times 10^{-2} = 10^{-4}\mathrm{M}$
$\therefore \mathrm{pH} = 4$

Question126

The solubility of sparingly soluble salt $\mathrm{AB_2}$ is $1.0 \times 10^{-4}\ \mathrm{mol\ dm^{-3}}$. What is it's solubility product?

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Options:

A. 2×10^{-12}

B. 4×10^{-8}

C. 4×10^{-12}

D. 2×10^{-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:

$AB_2(s) \rightleftharpoons A^{2+}(aq) + 2B^{-}(aq)$

Given that the solubility of AB_2 is 1.0×10^{-4} mol dm^{-3}, let's denote the solubility by s.

Thus, $s = 1.0 \times 10^{-4}$ mol dm^{-3}. At equilibrium:

- The concentration of A^{2+} ions will be $s = 1.0 \times 10^{-4}$ mol dm^{-3}.

- The concentration of B^{-} ions will be $2s = 2 \times 10^{-4}$ 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 4×10^{-12}. Hence, the correct answer is:

Option C

4×10^{-12}

Question127

What is the pH of $0.005 \mathrm{~M~H_2SO_4}$ solution?

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Options:

A. 5.0

B. 2.3

C. 3.3

D. 2.0

Answer: D

Solution:

$$\underset{0.005 \mathrm{~M}}{\mathrm{H_2SO_4}} \longrightarrow \underset{2 \times 0.005=0.01 \mathrm{~M}}{2 \mathrm{H}^{+}}+\underset{0.005 \mathrm{~M}}{\mathrm{SO}_4^{2-}}$$

$$\begin{aligned} \mathrm{pH} & =-\log_{10}\left[\mathrm{H}^{+}\right] \\ & =-\log_{10}\left[10^{-2}\right]=2.0 \end{aligned}$$

Question128

The solubility product expression for $\mathrm{Ca_3(PO_4)_2}$ is represented as

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Options:

A. $\mathrm{K_{sp}}=\left[\mathrm{Ca}^{2+}\right]^2\left[\mathrm{PO_4}^3\right]^2$

B. $\mathrm{K_{sp}}=\left[\mathrm{Ca}^{2+}\right]^3\left[\mathrm{PO_4}^3\right]^2$

C. $K_{sp} = [Ca^{2+}][PO_4{}^{2}]^3$

D. $K_{sp} = [Ca^{2+}][PO_4{}^{2}]$

Answer: B

Solution:

To determine the correct solubility product expression for $Ca_3(PO_4)_2$, we need to start by understanding its dissociation in water. When $Ca_3(PO_4)_2$ dissolves, it dissociates into calcium ions (Ca^{2+}) and phosphate ions (PO_4^{3-}) according to the following equation:

$Ca_3(PO_4)_2 \rightleftharpoons 3Ca^{2+} + 2PO_4^{3-}$

From this dissociation equation, for every 1 mole of $Ca_3(PO_4)_2$ that dissolves, 3 moles of Ca^{2+} ions and 2 moles of PO_4^{3-} ions are produced.

The solubility product constant, K_{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 $Ca_3(PO_4)_2$ is:

$K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2$

Thus, the correct option is:

Option B

$K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2$

Question129

Identify conjugate acid-base pair in the following reaction.

$$HCl_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + Cl^-_{(aq)}$$

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Options:

A. $H_3O^+_{(aq)}$ and $Cl^-_{(aq)}$

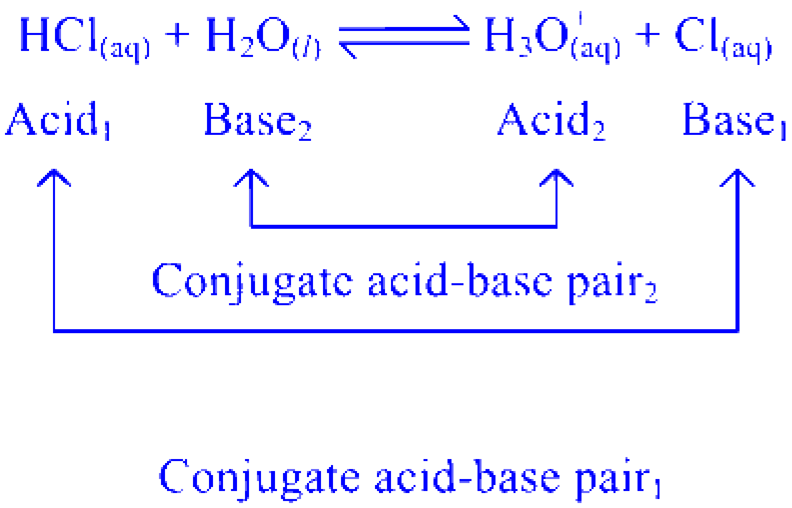
B. $H_3O^+_{(aq)}$ and $H_2O_{(l)}$

C. $HCl_{(aq)}$ and $H_2O_{(l)}$

D. $Cl^-_{(aq)}$ and $H_2O_{(l)}$

Answer: B

Solution:



Question130

Which among the following salts undergoes hydrolysis?

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Options:

A. Na_2SO_4

B. KCl

C. NH_4Cl

D. KNO_3

Answer: C

Solution:

NH_4Cl is a salt of weak base, NH_4OH and strong acid, HCl. Hence, it undergoes hydrolysis. While Na_2SO_4, KCl and 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?

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Options:

A. 1.4×10^{-3}

B. 1.8×10^{-5}

C. 1.6×10^{-3}

D. 1.34×10^{-5}

Answer: B

Solution:

$$c = 0.1 \text{ M}, \propto = 1.34\% = \frac{1.34}{100} = 1.34 \times 10^{-2}$$

$$\begin{aligned} \mathrm{K_a} &= \alpha^2 \mathrm{c} \\ &= \left(1.34 \times 10^{-2}\right)^2 \times 0.1 \\ &= 1.8 \times 10^{-5} \end{aligned}$$

Question132

pH of soft drink is 3.6. Calculate the concentration of hydrogen ions in it.

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Options:

A. $2.51 \times 10^{-4} \mathrm{M}$

B. $2.3 \times 10^{-3} \mathrm{M}$

C. $2.0 \times 10^{-3} \mathrm{M}$

D. $2.81 \times 10^{-4} \mathrm{M}$

Answer: A

Solution:

$$\begin{aligned}
&\mathrm{pH}=3.6,\left[\mathrm{H}^{+}\right]=?\\
&\mathrm{pH}=-\log_{10}\left[\mathrm{H}^{+}\right]\\
&\therefore \log_{10}\left[\mathrm{H}^{+}\right]=-\mathrm{pH}\\
&=-3.6\\
&=-3-0.6-1+1\\
&=-4+0.4 \quad=\overline{4}.4\\
&\therefore \quad \left[\mathrm{H}^{+}\right] \quad=\operatorname{Antilog}(\overline{4}.4)=2.512 \times 10^{-4} \mathrm{M}
\end{aligned}$$

Question133

Which among the following salt solution in water shows pH less than 7 ?

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Options:

A. $CuCl_2$

B. CH_3COONH_4

C. Na_2CO_3

D. KNO_3

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. **$CuCl_2$ (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 $CuCl_2$ in water will produce H_3O^+ ions, resulting in an acidic solution.

2. **CH_3COONH_4 (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_2CO_3 (Sodium carbonate)**: Sodium carbonate comes from a strong base (sodium hydroxide) and a weak acid (carbonic acid). The hydrolysis will produce OH^- ions, resulting in a basic solution with pH greater than 7.

4. **KNO_3 (Potassium nitrate)**: This salt is derived from a strong base (potassium hydroxide) and a strong acid (nitric acid). The solution of KNO_3 is neutral with pH around 7.

Given the above analysis, the salt solution in water that shows a pH less than 7 is **$CuCl_2$**.

The correct answer is **Option A: $CuCl_2$.**

Question134

The solubility of $\mathrm{Ag_2C_2O_4}$ is 2×10^{-4} mol L^{-1} at 298 K. What is it's solubility product?

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Options:

A. 1.6×10^{-6}

B. 3.2×10^{-11}

C. 1.6×10^{-11}

D. 3.2×10^{-6}

Answer: B

Solution:

To determine the solubility product (K_{sp}) of $Ag_2C_2O_4$, we start by writing the dissociation equation:

$Ag_2C_2O_4(s) \leftrightarrow 2Ag^+(aq) + C_2O_4^{2-}(aq)$

Given that the solubility of $Ag_2C_2O_4$ is 2×10^{-4} mol L^{-1} at 298 K, we can say that the concentration of $Ag_2C_2O_4$ that dissolves is 2×10^{-4} mol L^{-1}.

From the dissociation equation, for each mole of $Ag_2C_2O_4$ that dissolves, we get 2 moles of Ag^+ and 1 mole of $C_2O_4^{2-}$.

Therefore, the concentrations at equilibrium are:

$[Ag^+] = 2 \times (2 \times 10^{-4}) = 4 \times 10^{-4}$ mol L^{-1}

$[C_2O_4^{2-}] = 2 \times 10^{-4}$ mol L^{-1}

The solubility product (K_{sp}) is given by:

$K_{sp} = [Ag^+]^2[C_2O_4^{2-}]$

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 $Ag_2C_2O_4$ is 3.2×10^{-11}. The correct option is:

Option B 3.2×10^{-11}

Question135

A weak monobasic acid is 3.0% dissociated in it's 0.04 M solution. What is the dissociation constant of acid?

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Options:

A. 9×10^{-4}

B. 3.6×10^{-5}

C. 3×10^{-2}

D. 4×10^{-2}

Answer: B

Solution:

To determine the dissociation constant (Ka) of the weak monobasic acid, we need to proceed with the following steps:

Given:

The degree of dissociation α 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 α^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: 3.6×10^{-5}

Question136

Solubility product of AgBr is 4.9×10^{-13}. What is its solubility?

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Options:

A. 2.4×10^{-7} mol dm^{-3}

B. 3.2×10^{-7} mol dm^{-3}

C. 4.9×10^{-7} mol dm^{-3}

D. 7.0×10^{-7} mol dm^{-3}

Answer: D

Solution:

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

$AgBr \rightleftharpoons Ag^{+} + Br^{-}$

$\therefore K_{sp} = S^2 \quad \therefore S = \sqrt{K_{sp}}$

$\therefore S = \sqrt{49 \times 10^{-14}} = 7.0 \times 10^{-7}$ mol dm^{-3}

Question137

Which of the following is NOT a correct mathematical equation for Ostwald dilution law?

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Options:

A. $\propto = \sqrt{\frac{K_a}{c}}$

B. $K = \frac{\alpha^2}{V}$

C. $K = \propto^2 c$

D. $\propto = \sqrt{\frac{K_a}{V}}$

Answer: D

Solution:

$\propto = \sqrt{K_a \cdot V}$

Question138

A weak monobasic acid is 10% dissociated in 0.05 M solution. What is its percentage dissociation in 0.10 M solution?

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Options:

A. 5.27%

B. 7.17%

C. 10.3%

D. 4.5%

Answer: B

Solution:

$\propto_1 = 10\% = \frac{10}{100} = 0.1, c_1 = 0.05\mathrm{M}$
$c_2 = 0.10\mathrm{M}, \propto_2 = ?$
$\mathrm{K_a} = \propto_1^2 \mathrm{c}_1 = \propto_2^2 \mathrm{c}_2$
$\therefore (0.1)^2 \times 0.05 = \propto_2^2 \times 0.10$
$\therefore \propto_2^2 = \frac{10^{-2} \times 5 \times 10^{-2}}{10^{-1}} = 50 \times 10^{-4}$
$\therefore \propto_2 = \sqrt{50 \times 10^{-4}} = 7.17 \times 10^{-2} = 7.17\%$

Question139

The dissociation constant of weak monobasic acid is 2.7×10^{-5}. If degree of dissociation of acid is 3×10^{-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 (α), 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:

Dissociation constant, $K_a = 2.7 \times 10^{-5}$

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 9×10^{-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 3.2×10^{-8}. What is it's solubility in mol dm^{-3} ?

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Options:

A. 2.8×10^{-4}

B. 1.6×10^{-5}

C. 2.0×10^{-3}

D. 4.0×10^{-4}

Answer: C

Solution:

To determine the solubility of the sparingly soluble salt AX_2 in mol dm^{-3}, we need to use the solubility product constant (K_{sp}). The dissociation of AX_2 can be represented as follows:

$AX_2 \leftrightharpoons A^{2+} + 2X^-$

Let the solubility of AX_2 be S 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 3.2×10^{-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 mol dm^{-3} is:

$S = 2.0 \times 10^{-3}$ mol dm^{-3}

Therefore, the correct option is:

Option C

2.0×10^{-3}

Question141

Which of the following salt solutions is highly acidic?

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Options:

A. Ammonium acetate

B. Ammonium cynide

C. Sodium chloride

D. Ammonium nitrate

Answer: D

Solution:

NH_4NO_3 is a salt of strong acid, HNO_3 and weak base, NH_4OH.

Question142

The $\left[\mathrm{OH}^{-}\right]$ in a solution is 1×10^{-12} mol dm^{-3}. What is the concentration of H^{+} ions?

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Options:

A. 0.1 mol dm^{-3}

B. 1.0 mol dm^{-3}

C. 2.0 mol dm^{-3}

D. 0.01 mol dm^{-3}

Answer: D

Solution:

$\left[OH^{-}\right] = 1 \times 10^{-12} \text{ mol dm}^{-3}, [H^{+}] = ?$

Product of hydrogen ion concentration and hydroxide ion concentration is equal to 1.0×10^{-14}.

$$\left[H^{+}\right] \times \left[OH^{-}\right] = 1 \times 10^{-14}$$

$$\begin{aligned} \therefore \quad \left[H^{+}\right] &= \frac{1 \times 10^{-14}}{1 \times 10^{-12}} = 1 \times 10^{-2} \\ &= 0.01 \text{ mol dm}^{-3} \end{aligned}$$

Question143

Which among the following salt solution in water shows pH greater than 7?

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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 $[OH^-]$?

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Options:

A. 6.71×10^{-4} M

B. 1.12×10^{-4} M

C. 4.45×10^{-5} M

D. 8.91×10^{-5} M

Answer: D

Solution:

$$\begin{aligned} \mathrm{pH} &= 9.95, \mathrm{c} = 0.05\mathrm{M} \\ \mathrm{pOH} &= 14 - \mathrm{pH} = 14 - 9.95 = 4.05 \\ \mathrm{pOH} &= -\log_{10}[\mathrm{OH}] \\ \log_{10}[\mathrm{OH}] &= -4.05 \\ &= -4 - 0.05 - 1 + 1 \\ &= -5 + 0.95 = \bar{5}.95 \\ \left[\mathrm{OH}^-\right] &= \text{antilog}\, \bar{5}.95 \\ &= 8.913 \times 10^{-5}\mathrm{M} \end{aligned}$$

Question145

Which among the following salts turns blue litmus red in it's aqueous solution?

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Options:

A. NH_4CN

B. NH_4F

C. CH_3COONa

D. CH_3COONH_4

Answer: B

Solution:

NH_4CH - Basic, NH_4F - Slightly acidic, CH_3COONa - Basic, CH_3COONH_4 - Neutral

The solution of NH_4F is only slightly acidic and turns blue litmus red.

Question146

What is the pH of 0.02 M NaOH solution?

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Options:

A. 10.3

B. 11.3

C. 11.7

D. 12.3

Answer: D

Solution:

$$\underset{0.02\ M}{NaOH} \longrightarrow Na^{+} + \underset{0.02\ M}{OH^{-}}$$

$$P^{OH} = -\log_{10}\left[OH^-\right] = -\log_{10}[0.02]$$
$$= -\log_{10}\left(2 \times 10^{-2}\right) = 2 - 0.3010$$
$$P^{OH} = 1.6990$$
$$\therefore \quad P^H = 14 - P^H = 14 - 1.6990 = 12.301$$

Question147

Dissociation constant of propionic acid is 1.32×10^{-5}. Calculate the degree of dissociation of acid in 0.05 M solution.

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Options:

A. 2.6×10^{-4}

B. 1.61×10^{-2}

C. 1.90×10^{-2}

D. 3.5×10^{-5}

Answer: B

Solution:

$$K_a = 1.32 \times 10^{-5}, c = 0.05M, \propto \; =?$$
$$\propto \; = \sqrt{\frac{Ka}{c}} = \sqrt{\frac{1.32 \times 10^{-5}}{0.05}} = \sqrt{2.64} \times 10^{-2}$$
$$\therefore \; \propto \; = 1.62 \times 10^{-2}$$

Question148

Solubility of AgCl is 7.2×10^{-7} mol dm^{-3}. What is it's solubility product?

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Options:

A. 3.6×10^{-13}

B. 7.2×10^{-14}

C. 2.59×10^{-14}

D. 5.18×10^{-13}

Answer: D

Solution:

$$\begin{aligned} & \mathrm{S}=7.2 \times 10^{-7} \mathrm{~mol~dm}^{-3}, \mathrm{~K}_{\mathrm{sp}}=? \\ & \text { For AgCl, } \mathrm{K}_{\mathrm{sp}}=\mathrm{S}^2 \\ & =\left(7.2 \times 10^{-7}\right)^2=5.18 \times 10^{-13} \end{aligned}$$

Question149

A substance containing hydrogen and releasing H^+ in aqueous medium is acid. Identify theory suggesting this concept, from following.

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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:

$HA\ (aq) \rightarrow H^+(aq) + A^-(aq)$

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 1.25×10^{-5} mol dm^{-3}. What is solubility product of AgCl?

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Options:

A. 1.56×10^{-10}

B. 3.50×10^{-6}

C. 1.10×10^{-5}

D. 2.53×10^{-3}

Answer: A

Solution:

$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$

Here, x = 1, y = 1

$\therefore Ksp = S^2 = (1.25 \times 10^{-5})^2 = 1.56 \times 10^{-10}$

Question151

Which from following compounds accepts proton from water molecule according to Bronsted-Lowry theory?

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Options:

A. $NaOH_{(aq)}$

B. $HCl_{(aq)}$

C. $NH_{3(aq)}$

D. $NH_4OH_{(aq)}$

Answer: C

Solution:

$H_2O_{(\ell)} + NH_{3(aq)} \rightleftharpoons NH^+_{4(aq)} + OH^-_{(aq)}$

Question152

The pH of 0.1 M solution of monobasic acid is 2.34. Calculate the degree of dissociation of the acid.

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Options:

A. 3.1×10^{-2}

B. 4.5×10^{-2}

C. 2.18×10^{-2}

D. 2.5×10^{-3}

Answer: B

Solution:

$\mathrm{pH} = 2.34, \mathrm{c} = 0.1\mathrm{M}$

(i) $\mathrm{pH} = -\log_{10}\left[\mathrm{H}^{+}\right]$
$\log_{10}\left[\mathrm{H}^{+}\right] = -\mathrm{pH} = -2.34$
$= -2 - 0.34 - 1 + 1$
$= -3 + 0.66 = \overline{3}.66$
$\therefore \left[\mathrm{H}^{+}\right] = \text{antilog } \overline{3} \cdot 66$
$= 4.571 \times 10^{-3} \text{ mol dm}^{-3}$

(ii) $[\mathrm{H}^{+}] = \propto \mathrm{c}$

$\therefore \quad \propto \ = \frac{[\mathrm{H}^{+}]}{\mathrm{c}} = \frac{4.571 \times 10^{-3}}{0.1} = 4.571 \times 10^{-2}$

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 $0.1 \mathrm{~M} \mathrm{Ba}(\mathrm{NO}_3)_2$ 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 $= 0.1\mathrm{M}$

vant Hoff factor $(i) = 2.74$

Since, $i > 1$, it means solute is undergoing dissociation.

$\mathrm{Ba}(\mathrm{NO}_3)_2 \rightleftharpoons \mathrm{Ba}^{2+} + 2\mathrm{NO}_3^-$

Number of particles dissociated $(n) = 3$

Now, $\alpha($ degree of dissociation $) = \frac{(i-1)}{(n-1)}$

$$= \frac{(2.74-1)}{(3-1)}$$
$= 0.87$
