Ionic Equilibrium

Question1

The molar solubility of $\mathrm{PbI_2}$ in $0.2\mathrm{M}\,\mathrm{Pb(NO_3)_2}$ solution in terms of K_{sp} (solubility product) is

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Options:

A.

$\left(\frac{K_{sp}}{0.2}\right)^{1/2}$

B.

$\left(\frac{K_{sp}}{0.4}\right)^{1/4}$

C.

$\left(\frac{K_{sp}}{0.8}\right)^{1/2}$

D.

$\left(\frac{K_{sp}}{0.8}\right)^{1/3}$

Answer: C

Solution:

The dissociation equation of $\mathrm{PbI_2}$ is,

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

Let s be molar solubility of $\mathrm{PbI_2}$

Initial $\left[\mathrm{Pb}^{2+}\right] = 0.2\mathrm{M}$

At equilibrium $\left[\mathrm{Pb}^{2+}\right] = (0.2 + s)\mathrm{M}$

At equilibrium $[\mathrm{I}^{-}] = 2\ \mathrm{sM}$

$K_{sp} = \left[\mathrm{Pb}^{2+}\right][\mathrm{I}^{-}]^2$

$= (0.2 + s)(2s)^2 = (0.2)\left(4s^2\right)$

Since s is very small compared to 0.2 there, neglecting it

$$= 0.8s^2$$

$$s = \left(\frac{K_{sp}}{0.8}\right)^{1/2}$$

Question2

The percentage of ionisation of 1 L of xM acetic acid is 4.242 and is called solution " A ". The percentage of ionisation of 1 L of yM acetic acid is 3 and is called solution " B ". Solution " A " is mixed with solution " B ". What is the concentration of acetic acid in the resultant solution? (K_a of acetic acid $= 1.8 \times 10^{-5}$)

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Options:

A.

0.05 M

B.

0.015 M

C.

0.02 M

D.

0.15 M

Answer: B

Solution:

Convert the percentage ionisation to decimal

$$\alpha_A = \frac{4.242}{100} = 0.04242$$

$$\alpha = \sqrt{\frac{K_\alpha}{C}}$$

$$\Rightarrow \quad C_A = \frac{K_\alpha}{\alpha_A^2}$$

$$C_A = \frac{1.8 \times 10^{-5}}{(0.04242^2}, C_A = 0.01\mathrm{M}$$

$$\alpha_B = \frac{3}{100} = 0.03$$

$$C_B = \frac{K_\alpha}{\alpha_B^2} = \frac{1.8 \times 10^{-5}}{(0.03)^2} = 0.02\mathrm{M}$$

Moles from solution A,

$$\begin{aligned} n_A &= C_A \times V_A = 0.01\mathrm{M} \times 1\ \mathrm{L} = 0.01\ \mathrm{mol} \\ n_B &= c_B \times V_B = 0.02\mathrm{M} \times 1\ \mathrm{L} = 0.02\ \mathrm{mol} \\ n_{\text{Total}} &= n_A + n_B = 0.03\ \mathrm{mol} \\ V_{\text{Total}} &= V_A + V_B = 1 + 1 = 2\ \mathrm{L} \\ c_{\text{Final}} &= \frac{n_{\text{Total}}}{V_{\text{Total}}} = \frac{0.03}{2\ \mathrm{L}} = 0.015\mathrm{M} \end{aligned}$$

Question3

When 30 mL of $0.2\mathrm{M}\,\mathrm{NH_4OH}$ is added to 30 mL of $2\mathrm{M}\,\mathrm{NH_4Cl}$ solution. If the pH of the buffer formed is 8.2, what is the $\mathrm{p}K_b$ of $\mathrm{NH_4OH}$?

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Options:

A.

5.8

B.

7.2

C.

6.8

D.

4.8

Answer: D

Solution:

pOH of buffer solution

$p\text{OH} = 14 - \text{pH} \Rightarrow \text{pH} = 14 - 8.2 = 5.8$

Moles of $\text{NH}_4\text{OH} = 0.2 \times 0.030$

$= 0.006$ mol

Moles of $\text{NH}_4\text{Cl} = 2 \times 0.030 = 0.060$ mol

Total volume $= 30\ \text{mL} + 30\ \text{mL} = 0.060\text{I}$

Concentration of conjugate base and acid in buffer

$$[\text{NH}_4\text{OH}] = \frac{0.006\ \text{mol}}{0.060\ \text{L}} = 0.1\text{M}$$
$$[\text{NH}_4\text{Cl}] = \frac{0.060\ \text{mol}}{0.060\ \text{L}} = 1\text{M}$$

Using Henderson equation

$$\text{pOH} = \text{p}K_b + \log \frac{[\text{NH}_4\text{Cl}]}{[\text{NH}_4\text{OH}]}$$
$$5.8 = \text{p}K_b + \log \left(\frac{1}{01}\right)$$
$$\text{p}K_b = 4.8$$

Question4

At 25°C, K_a of formic acid is 1.8×10^{-4}. What is the K_b of HCOO^-?

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Options:

A.

1.8×10^{-10}

B.

5.55×10^{-4}

C.

5.55×10^{-11}

D.

5.55×10^{-12}

Answer: C

Solution:

The ion product of water (K_w) is the product of the acid dissociation constant (K_a) and the base dissociation constant (K_b) :

$K_a \times K_b = K_w$

So, you can solve for K_b using:

$$K_b = \frac{K_w}{K_a}$$

We know:

$K_w = 1.0 \times 10^{-14}$ at 25°C

$K_a = 1.8 \times 10^{-4}$

Plug these values into the formula:

$$K_b = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-4}}$$

$K_b = 5.55 \times 10^{-11}$

Question5

At $25^{\circ}\mathrm{C}$, the percentage of ionisation of ' x ' M acetic acid is 4.242 . What is the value of x ? $\left(K_{\mathrm{a}} = 1.8 \times 10^{-5}\right)$

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Options:

A.

0.05

B.

0.04

C.

0.02

D.

0.01

Answer: D

Solution:

x be the initial concentration of acetic acid.

Let α be degree of ionisation $= 0.04242$

Concentration of ionised acid $= \alpha x$

$$\begin{aligned} k_a &= \frac{[\mathrm{H}^+][A^-]}{[\mathrm{H}A]} = \frac{\alpha x(\alpha x)}{x - \alpha x} \\ k_a &= \frac{\alpha^2 x}{1-\alpha} \\ \Rightarrow \quad x &= \frac{k_a(1-\alpha)}{\alpha^2} \qquad \ldots (i) \end{aligned}$$

Substituting the values

$x = \frac{1.8 \times 10^{-5}(1-0.04242)}{(0.04242)^2} \simeq 0.01$

Question6

At 25°C, the percentage of ionisation of xM acetic acid is 4.242 . What is the pH of the acetic acid solution?

$(\log 4.242 = 0.6275); (\log 0.04242 = -1.372)$
$\left(K_{\mathrm{a}} = 1.8 \times 10^{-5}\right)$

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Options:

A.

3.37

B.

1.70

C.

1.37

D.

2.37

Answer: A

Solution:

(a) Percentage ionisation

$= \frac{[\mathrm{H}^+]}{\text{Initial concentration of acid}} \times 100\%$

$$4.242=\frac{[\mathrm{H}^{+}]}{x}\times 100$$

$$[\mathrm{H}^{+}]=\frac{4.242\times x}{100}=0.04242x$$

$$[\mathrm{CH_3COO^-}]=0.04242x$$

[∵ % ionisation is small]

$$\begin{aligned}[\mathrm{CH_3COOH}]&=x-0.04242x\\&=0.95758x\end{aligned}$$

For equilibrium reaction,

$$\mathrm{CH_3COOH}\rightleftharpoons\mathrm{CH_3COO^-}+\mathrm{H^+}$$

$$\begin{aligned}K_a&=\frac{[\mathrm{CH_3COO^-}][\mathrm{H^+}]}{[\mathrm{CH_3COOH}]}\\1.8\times10^{-5}&=\frac{0.04242x\times0.04242x}{0.95758\mathrm{x}}\\x&=0.009579\\[\mathrm{H^+}]&=0.04242\times0.009579\\&=0.00040693\\\mathrm{pH}&=-\log[\mathrm{H^+}]\\&=-\log[0.00040693]=3.37\end{aligned}$$

Question7

At $T(\mathrm{K})$, the solubility product of AgBr is 4×10^{-13}. What is its solubility in 0.1 M KBr solution?

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Options:

A.

$2\times10^{-6}\mathrm{M}$

B.

$4\times10^{-10}\mathrm{M}$

C.

$4\times10^{-12}\mathrm{M}$

D.

$4\times10^{-14}\mathrm{M}$

Answer: C

Solution:

Dissociation equation of AgBr

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

Let solubility of AgBr $= s$

$$\begin{aligned} \mathrm{Ag}^{+} &= s(\text{ concentration }), \mathrm{Br}^{-} = 0.1 + s \\ K_{\mathrm{sp}} &= \left[\mathrm{Ag}^{+}\right]\left[\mathrm{Br}^{-}\right] \\ 4 \times 10^{-13} &= s \times (0.1 + s) \\ s &= 4 \times 10^{-12} \mathrm{M} \end{aligned}$$

Question8

At $27^{\circ}\mathrm{C}$, the degree of dissociation of $\mathrm{H}A$ (weak acid) in 0.5 M of its solution is 1%. The concentrations of $\mathrm{H}_3\mathrm{O}^{+}$, $A^{-}$and $\mathrm{H}A$ at equilibrium (in molL^{-1}) are respectively.

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Options:

A. $0.005, 0.005, 0.495$

B. $0.05, 0.05, 0.45$

C. $0.01, 0.01, 0.49$

D. $0.005, 0.495, 0.005$

Answer: A

Solution:

To determine the concentrations of $\mathrm{H}_3\mathrm{O}^{+}$, A^{-}, and $\mathrm{H}A$ at equilibrium, we analyze the dissociation of the weak acid $\mathrm{H}A$ in a 0.5 M solution where its degree of dissociation, α, is 1%.

The dissociation reaction is as follows:

$\mathrm{H}A + \mathrm{H}_2\mathrm{O} \rightleftharpoons \mathrm{H}_3\mathrm{O}^{+} + A^{-}$

At equilibrium:

The concentration of undissociated $\mathrm{H}A$ is $(1-\alpha)C$.

The concentrations of $\mathrm{H}_3\mathrm{O}^{+}$ and A^{-}, which are produced in equal amounts due to the stoichiometry of the reaction, are both αC.

Given that:

$C = 0.5\,\mathrm{M}$

$\alpha = \frac{1}{100} = 0.01$

We calculate:

$[\mathrm{H_3O^+}] = [A^-] = \alpha C = 0.01 \times 0.5 = 0.005\,\mathrm{M}$

The concentration of $\mathrm{H}A$ at equilibrium is:

$[\mathrm{H}A] = (1-\alpha)C = (1-0.01) \times 0.5 = 0.495\,\mathrm{M}$

Thus, the concentrations at equilibrium are:

$[\mathrm{H_3O^+}] = 0.005\,\mathrm{M}$

$[A^-] = 0.005\,\mathrm{M}$

$[\mathrm{H}A] = 0.495\,\mathrm{M}$

Question9

100 mL of $0.1 \mathrm{MH}A$ (weak acid) and 100 mL of 0.2 M $\mathrm{Na}A$ are mixed. What is the pH of resultant solution? (K_{a} of $\mathrm{H}A$ is 10^{-5} : $\log 2 = 0.3$)

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Options:

A. 4.7

B. 5.0

C. 5.3

D. 4.0

Answer: C

Solution:

When 100 mL of 0.1 M HA (a weak acid) is mixed with 100 mL of 0.2 M NaA (its conjugate base), we can determine the pH of the solution using the Henderson-Hasselbalch equation. This equation is suitable for calculating the pH of a buffer solution and is given by:

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

Given:

The acid dissociation constant, $K_a = 10^{-5}$

The concentration of acid HA is $0.1\,\mathrm{M}$

The concentration of salt NaA is $0.2\,\mathrm{M}$

First, calculate the $\mathrm{p}K_a$:

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

Now, use the Henderson-Hasselbalch equation:

$\mathrm{pH} = 5 + \log\left(\frac{0.2}{0.1}\right)$

$\mathrm{pH} = 5 + \log(2)$

Given that $\log(2) = 0.3$:

$\mathrm{pH} = 5 + 0.3 = 5.3$

Therefore, the pH of the resultant solution is 5.3.

Question10

Which of the following when added to 20 mL of a 0.01 M solution of HCl would decrease its pH ?

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Options:

A. 20 mL of 0.02 M HCl

B. 20 mL of 0.005 M HCl

C. 20 mL of 0.01 M HCl

D. 40 mL of 0.005 M HCl

Answer: A

Solution:

Original solution,

$[\mathrm{H}^+] = 0.01\mathrm{M} = 10 \times 10^{-3}\mathrm{M}$

and $\quad \mathrm{pH} = 2$

$\Rightarrow [\mathrm{H}^+]$and pH of the mixture as in,

$$(a)\ [\mathrm{H}^+] = \frac{20 \times 0.01 + 20 \times 0.02}{40}$$
$$= 15 \times 10^{-3}\mathrm{M}, \mathrm{pH} = 1.82$$
$$[\because [\mathrm{HCl}] = [\mathrm{H}^+]]$$

$$(b)\ [\mathrm{H}^+] = \frac{20 \times 0.01 + 20 \times 0.005}{40}$$
$$= 7.5 \times 10^{-3}\mathrm{M}; \mathrm{pH} = 2.12$$

$$(c)\ [\mathrm{H}^+] = \frac{20 \times 0.01 + 20 \times 0.01}{40}$$
$$= 10 \times 10^{-3}\mathrm{M}; \mathrm{pH} = 2$$

$$(d)\quad [\mathrm{H}^{+}] = \frac{20 \times 0.01 + 40 \times 0.005}{60}$$
$$= 6.67 \times 10^{-3}; \mathrm{pH} = 2.17$$

So, decrease in pH , i.e., $\mathrm{pH} < 2$ takes place in option (a) only.

Question11

The solubility of barium phosphate of molar mass ' M ' gmol^{-1} in water is x g per 100 mL at 298 K . Its solubility product is $1.08 \times \left(\frac{x}{M}\right)^{a} \times (10)^{b}$. The values of a and b respectively are

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Options:

A. 7,5

B. 5,7

C. 5,5

D. 7,7

Answer: B

Solution:

The solubility of barium phosphate, with molar mass M, in water is x g per 100 mL at 298 K. To convert this to moles per liter, the solubility is expressed as $\frac{10 \times x}{M}$ mol/L.

The solubility product (K_{sp}) expression for $\mathrm{Ba}_3(\mathrm{PO}_4)_2$ is given by:

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

Substituting the molar solubility into this expression, we have:

$$K_{\mathrm{sp}} = \left(3 \times \frac{10 \times x}{M}\right)^3\left(\frac{2 \times 10 \times x}{M}\right)^2$$

Simplifying this:

$$K_{\mathrm{sp}} = 108\left(\frac{10 \times x}{M}\right)^5 = 1.08 \times 10^7\left(\frac{x}{M}\right)^5$$

From this, the values of a and b in the solubility product expression $1.08 \times \left(\frac{x}{M}\right)^a \times 10^b$ are 5 and 7, respectively.

Question12

At $T(\mathrm{~K})$, the solubility product of AX is 10^{-10}. What is the molar solubility of AX in 0.1 MHX ?

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Options:

A. 10^{-5}

B. 10^{-10}

C. 10^{-9}

D. 10^{-8}

Answer: C

Solution:

Initial concentration at equilibrium

	$X \longrightarrow$	A^{+}	$+X^{-}$
Initial conc.	a	0	0
At equilibrium	$a-S$	S	$S+0.1$

$$K_{\mathrm{sp}} = [A^{+}][X^{-}] = 10^{-10}$$

$$10^{-10} = S[S \times 0.1]$$

As $K_{sp} \lll 1$ then, $10^{-10} = S \times 0]$

$$S = 10^{-9}\mathrm{M}$$

Question13

Which of the following will make a basic buffer solution?

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Options:

A. 100 mL of $0.1\mathrm{M}\,\mathrm{CH_3COOH}$ + 100 mL of 0.1 M NaOH

B. 100 mL of $0.1\mathrm{M}\,\mathrm{HCl}$ + 100 mL of 0.1 M NaOH

C. 50 mL of $0.1\mathrm{M}\,\mathrm{KOH}$ + 25 mL of $0.1\mathrm{M}\,\mathrm{CH_3COOH}$

D. 100 mL of $0.1\mathrm{M}\,\mathrm{HCl}$ + 200 mL of $0.1\mathrm{M}\,\mathrm{NH_4OH}$

Answer: D

Solution:

A basic buffer solution consists of a weak base and the salt of its conjugate strong acid.

In this specific reaction, we have:

100 mL of 0.1 M HCl and 200 mL of 0.1 M NH_4OH.

First, calculate the millimoles:

HCl: $n = 100\,\text{mL} \times 0.1\,\text{M} = 10\,\text{millimoles}$

NH_4OH: $n = 200\,\text{mL} \times 0.1\,\text{M} = 20\,\text{millimoles}$

The balanced chemical reaction is:

$\text{HCl}(10\,\text{millimoles}) + \text{NH}_4\text{OH}(20\,\text{millimoles}) \rightarrow \text{NH}_4\text{Cl}(10\,\text{millimoles}) + \text{H}_2\text{O}$

After the reaction, 10 millimoles of NH_4OH remain, along with 10 millimoles of NH_4Cl. As NH_4OH is a weak base and NH_4Cl is its conjugate acid's salt, this mixture forms a basic buffer solution. Therefore, the solution will effectively resist changes in pH upon the addition of small amounts of acids or bases. Thus, the formation of a basic buffer makes the scenario correct.

Question14

At 27°C, $100\ \text{mL}$ of 0.4 M HCl is mixed with 100 mL of 0.5 M NaOH solution. To the resultant solution, 800 mL of distilled water is added. What is the pH of final solution?

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Options:

A. 12

B. 2

C. 1.3

D. 1.0

Answer: A

Solution:

First step is to determine the number of moles of HCl and NaOH before mixing.

$\text{HCl} = 0.4\text{M} \times 0.1\ \text{L} = 0.04\ \text{mol}$
$\text{NaOH} = 0.5\text{M} \times 0.1\ \text{L} = 0.05\ \text{mol}$

Since, HCl and NaOH react in a $1:1$ molar ratio.

$\text{NaOH} + \text{HCl} \longrightarrow \text{NaCl} + \text{H}_2\text{O}$

0.04 moles of HCl will react with 0.04 moles of NaOH

$$\begin{aligned}\text{Remaining NaOH} &= 0.05 - 0.04 \\ &= 0.01 \text{ mol}\end{aligned}$$

Final volume after adding 800 mL of distilled water

$200 \text{ mL} + 800 \text{ mL} = 1 \text{ L}$

Concentration of NaOH after dilution

$$= \frac{0.01}{1} = 0.01\text{M}$$

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

Concentration of OH^- is equal to concentration of $\text{NaOH} = 0.01\text{M}$

$$\begin{aligned}&\text{pOH} = -\log\left[\text{OH}^-\right] \\ &\text{pOH} = -\log[0.01] \\ &\text{pOH} = 2; \text{pH} = 14 - 2 = 12\end{aligned}$$

Question15

What is the conjugate base of chloric acid?

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Options:

A. ClO_4^-

B. ClO^-

C. ClO_2^-

D. ClO_3^-

Answer: D

Solution:

The conjugate base of chloric acid is formed when chloric acid donates a proton (H^+).

$\text{HClO}_3 \rightleftharpoons \text{H}^+ + \text{ClO}_3^-$

Thus, the conjugate base of chloric acid is ClO_3^-.

Question16

The pH of 0.01 N lime water is

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Options:

A. 13.09

B. 10

C. 12

D. 9.8

Answer: C

Solution:

Lime water is $\mathrm{Ca(OH)_2}$. Acidity of lime water is 2.

$$\text{So, Normality } = \text{ Acidity } \times \text{ Molarity}$$
$$\Rightarrow \text{ Molarity } = \frac{0.01}{2}\mathrm{M}$$

Concentration of $\left[\mathrm{OH}^{-}\right]$ ions, $\left[\mathrm{OH}^{-}\right] = 2 \times \frac{0.01}{2} = 0.01\mathrm{M}$

Now, $\mathrm{pOH} = -\log\left[\mathrm{OH}^{-}\right]$

$$\begin{aligned} & \mathrm{pOH} = -\log 10^{-2} \\ \Rightarrow \quad & \mathrm{pOH} = 2\log 10 = 2 \end{aligned}$$
$$\mathrm{pH} + \mathrm{pOH} = 14$$
$$\Rightarrow \quad \mathrm{pH} = 14 - \mathrm{pOH} = 14 - 2$$

Thus, $\mathrm{pH} = 12$

Question17

Equal volumes of 0.5 N acetic acid and 0.5 N sodium acetate are mixed. What is the pH of resultant solution? ($\mathrm{p}K_a$ of acetic acid $= 4.75$)

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Options:

A. 4.85

B. 4.65

C. 4.75

D. 7.0

Answer: C

Solution:

Given, concentration of acetic acid,

$[\mathrm{CH_3COOH}] = 0.5\ \mathrm{N}$

Concentration of sodium acetate,

$[\mathrm{CH_3COONa}] = 0.5\ \mathrm{N}$

$\mathrm{p}K_a$ of $\mathrm{CH_3COOH} = 4.75$

Using Henderson-Hasselbalch equation,

$$\mathrm{pH} = \mathrm{pK}_a + \log \frac{[\mathrm{CH_3COONa}]}{[\mathrm{CH_3COOH}]}$$

$$\Rightarrow \mathrm{pH} = 4.75 + \log \frac{0.5}{0.5}$$

So, $\mathrm{pH} = 4.75 + \log 1 = 4.75 + 0 = 4.75$

Question18

Match the following columns.

	Column I (Species)		Column II (Conjugate acid)
A.	NH_3	1.	H_2CO_3
B.	HCO_3^-	2.	H_3O^+
C.	H_2O	3.	NH_4^+
D.	HSO_4^-	4.	H_2SO_4

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Options:

A. A - 3, B - 1, C - 2, D - 4

B. A - 1, B - 3, C - 2, D - 4

C. A - 3, B - 2, C - 1, D - 4

D. A - 3, B - 2, C - 4, D - 1

Answer: A

Solution:

Conjugate acid is a chemical compound formed when an acid donates a proton $[H^+]$ to a base.

Species	**Conjugated acid**
NH_3	NH_4^+
HCO_3^-	H_2CO_3
H_2O	H_3O^+
HSO_4^-	H_2SO_4

Question19

The pH of 0.1 M solution of acetic acid will be[degree of dissociation of acetic acid is0.0132]

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Options:

A. 4.32

B. 3.14

C. 1.14

D. 2.88

Answer: D

Solution:

- Concentration of acetic acid = 0.1 M
- Degree of dissociation (α) = 0.0132

We need to find the **pH**.

Step 1: Find the hydronium ion concentration

$[\mathrm{H}^+] = C \times \alpha = 0.1 \times 0.0132 = 0.00132\ \mathrm{M}$

Step 2: Calculate pH

$\mathrm{pH} = -\log_{10}([\mathrm{H}^+]) = -\log_{10}(0.00132)$

$\mathrm{pH} = 2.88$

✅ **Final Answer:**

Option D — 2.88

Question20

The solubility of $\mathrm{AgBr}(s)$, having solubility product 5×10^{-10} in $0.2\ \mathrm{M}$ NaBr solution equals

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Options:

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

B. $25 \times 10^{-10}\ \mathrm{M}$

C. $0.5\ \mathrm{M}$

D. $0.002\ \mathrm{M}$

Answer: B

Solution:

Given,

$$\begin{aligned}
[K_{\mathrm{sp}}]_{\mathrm{AgBr}} &= 5 \times 10^{-10} \\
[\mathrm{NaBr}] &= 0.2\mathrm{M} \\
[\mathrm{Na}^+] &= [\mathrm{Br}^-] = 0.2\mathrm{M} \\
[\mathrm{Ag}^+][\mathrm{Br}^-] &= [\mathrm{K}_{\mathrm{sp}}]_{\mathrm{AgBr}} \\
[\mathrm{Ag}^+](0.2) &= 5 \times 10^{-10} \\
[\mathrm{Ag}^+] &= \frac{5 \times 10^{-10}}{0.2} = 25 \times 10^{-10}\mathrm{M}
\end{aligned}$$

Question21

Calculate the pOH of 0.10 M HCl solution.

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Options:

A. 13

B. 1

C. 7

D. 14

Answer: A

Solution:

HCl is the strongest acid, it dissociate into ions,

$$\underset{\substack{0.10\\ \text{mole}}}{\mathrm{HCl}} \longrightarrow \underset{\substack{0.10\\ \text{mole}}}{\mathrm{H}^{+}} + \underset{\substack{0.10\\ \text{mole}}}{\mathrm{Cl}^{-}}$$

We know that $\mathrm{pH} = -\log\left[\mathrm{H}^{+}\right]$

$\mathrm{pH} = -\log[0.10] \Rightarrow \mathrm{pH} = 1$

Using formula of $\mathrm{pH} + \mathrm{pOH} = 14$ at 273 K,

$1 + \mathrm{pOH} = 14$

$\mathrm{pOH} = 14 - 1 = 13$

Question22

Which among the following pairs is not an acidic buffer?

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Options:

A. HCOOH and HCOOK

B. $\mathrm{HClO_4}$ and $\mathrm{NaClO_4}$

C. $\mathrm{C_6H_5COOH}$ and $\mathrm{C_6H_5COONa}$

D. HCN and KCN

Answer: B

Solution:

Acidic buffer $\Rightarrow$ weak acid + salt with strong base. An acidic buffer contains weak acid and its salt with a strong base.

$HClO_4$ and $NaClO_4$ contains strong acid $HClO_4$ and its salt with strong base NaOH. $HClO_4$ and $NaClO_4$ are not an acidic buffer because strong acid with its salt cannot form buffer solution.

So, $HClO_4$ and $NaClO_4$ cannot form acidic buffer.
