

Ionic Equilibrium

Question1

The molar solubility of PbI_2 in $0.2\text{M Pb(NO}_3)_2$ 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.

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

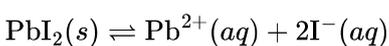
D.

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

Answer: C

Solution:

The dissociation equation of PbI_2 is,



Let s be molar solubility of PbI_2

$$\text{Initial } [\text{Pb}^{2+}] = 0.2\text{M}$$

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

$$\text{At equilibrium } [\text{I}^-] = 2s\text{M}$$

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



$$= (0.2 + s)(2s)^2 = (0.2)(4s^2)$$

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 x M acetic acid is 4.242 and is called solution " A ". The percentage of ionisation of 1 L of y M 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 = 1.8×10^{-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

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

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

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

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

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

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

Moles from solution A,

$$n_A = C_A \times V_A = 0.01\text{M} \times 1\text{ L} = 0.01\text{ mol}$$

$$n_B = c_B \times V_B = 0.02\text{M} \times 1\text{ L} = 0.02\text{ mol}$$

$$n_{\text{Total}} = n_A + n_B = 0.03\text{ mol}$$

$$V_{\text{Total}} = V_A + V_B = 1 + 1 = 2\text{ L}$$

$$c_{\text{Final}} = \frac{n_{\text{Total}}}{V_{\text{Total}}} = \frac{0.03}{2\text{ L}} = 0.015\text{M}$$

Question3

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

$$pOH = 14 - pH \Rightarrow pH = 14 - 8.2 = 5.8$$

$$\text{Moles of } NH_4OH = 0.2 \times 0.030$$

$$= 0.006 \text{ mol}$$

$$\text{Moles of } NH_4Cl = 2 \times 0.030 = 0.060 \text{ mol}$$

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

Concentration of conjugate base and acid in buffer

$$[NH_4OH] = \frac{0.006 \text{ mol}}{0.060 \text{ L}} = 0.1 \text{ M}$$

$$[NH_4Cl] = \frac{0.060 \text{ mol}}{0.060 \text{ L}} = 1 \text{ M}$$

Using Henderson equation

$$pOH = pK_b + \log \frac{[NH_4Cl]}{[NH_4OH]}$$

$$5.8 = pK_b + \log \left(\frac{1}{0.1} \right)$$

$$pK_b = 4.8$$

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.

$$1.8 \times 10^{-10}$$

B.

$$5.55 \times 10^{-4}$$

C.

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

D.

$$5.55 \times 10^{-12}$$

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:

$$K_w = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

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

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°C , the percentage of ionisation of ' x ' M acetic acid is 4.242 . What is the value of x ? ($K_a = 1.8 \times 10^{-5}$)

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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 = αx

$$k_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{\alpha x(\alpha x)}{x - \alpha x}$$

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

$$\Rightarrow x = \frac{k_a(1 - \alpha)}{\alpha^2} \quad \dots (i)$$

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 x M acetic acid is 4.242 . What is the pH of the acetic acid solution?

$$(\log 4.242 = 0.6275); (\log 0.04242 = -1.372)$$

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

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Options:

A.

3.37

B.

1.70

C.

1.37

D.

2.37

Answer: A

Solution:

(a) Percentage ionisation

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

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

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

$$[\text{CH}_3\text{COO}^-] = 0.04242x$$

[∵ % ionisation is small]

$$\begin{aligned}[\text{CH}_3\text{COOH}] &= x - 0.04242x \\ &= 0.95758x\end{aligned}$$

For equilibrium reaction,



$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$$

$$1.8 \times 10^{-5} = \frac{0.04242x \times 0.04242x}{0.95758x}$$

$$x = 0.009579$$

$$\begin{aligned}[\text{H}^+] &= 0.04242 \times 0.009579 \\ &= 0.00040693\end{aligned}$$

$$\begin{aligned}\text{pH} &= -\log [\text{H}^+] \\ &= -\log[0.00040693] = 3.37\end{aligned}$$

Question7

At $T(K)$, 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.

$$2 \times 10^{-6}\text{M}$$

B.

$$4 \times 10^{-10}\text{M}$$

C.

$$4 \times 10^{-12}\text{M}$$

D.

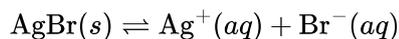
$$4 \times 10^{-14}\text{M}$$

Answer: C



Solution:

Dissociation equation of AgBr



Let solubility of AgBr = s

$$\text{Ag}^+ = s(\text{ concentration}), \text{Br}^- = 0.1 + s$$

$$K_{\text{sp}} = [\text{Ag}^+][\text{Br}^-]$$

$$4 \times 10^{-13} = s \times (0.1 + s)$$

$$s = 4 \times 10^{-12} \text{M}$$

Question8

At 27°C, the degree of dissociation of HA (weak acid) in 0.5 M of its solution is 1%. The concentrations of H_3O^+ , A^- and HA 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

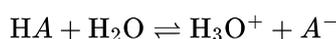
D. 0.005, 0.495, 0.005

Answer: A

Solution:

To determine the concentrations of H_3O^+ , A^- , and HA at equilibrium, we analyze the dissociation of the weak acid HA in a 0.5 M solution where its degree of dissociation, α , is 1%.

The dissociation reaction is as follows:



At equilibrium:

The concentration of undissociated HA is $(1 - \alpha)C$.

The concentrations of H_3O^+ and A^- , which are produced in equal amounts due to the stoichiometry of the reaction, are both αC .

Given that:

$$C = 0.5 \text{ M}$$

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



We calculate:

$$[\text{H}_3\text{O}^+] = [\text{A}^-] = \alpha C = 0.01 \times 0.5 = 0.005 \text{ M}$$

The concentration of HA at equilibrium is:

$$[\text{HA}] = (1 - \alpha)C = (1 - 0.01) \times 0.5 = 0.495 \text{ M}$$

Thus, the concentrations at equilibrium are:

$$[\text{H}_3\text{O}^+] = 0.005 \text{ M}$$

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

$$[\text{HA}] = 0.495 \text{ M}$$

Question9

100 mL of 0.1MHA (weak acid) and 100 mL of 0.2 M NaA are mixed. What is the pH of resultant solution? (K_a of HA 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:

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

Given:

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

The concentration of acid HA is 0.1 M

The concentration of salt NaA is 0.2 M

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

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

Now, use the Henderson-Hasselbalch equation:

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

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

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

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

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,

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

and $\text{pH} = 2$

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

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

$$\begin{aligned} (b) \quad [\text{H}^+] &= \frac{20 \times 0.01 + 20 \times 0.005}{40} \\ &= 7.5 \times 10^{-3}\text{M}; \text{pH} = 2.12 \end{aligned}$$

$$\begin{aligned} (c) \quad [\text{H}^+] &= \frac{20 \times 0.01 + 20 \times 0.01}{40} \\ &= 10 \times 10^{-3}\text{M}; \text{pH} = 2 \end{aligned}$$



$$(d) \quad [\text{H}^+] = \frac{20 \times 0.01 + 40 \times 0.005}{60}$$

$$= 6.67 \times 10^{-3}; \text{pH} = 2.17$$

So, decrease in pH, i.e., $\text{pH} < 2$ 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 $\text{Ba}_3(\text{PO}_4)_2$ is given by:

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

Substituting the molar solubility into this expression, we have:

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

Simplifying this:

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

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

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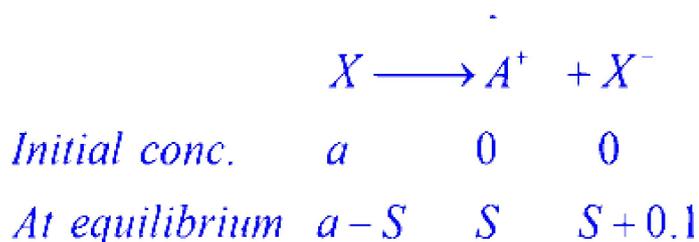
Options:

- A. 10^{-5}
- B. 10^{-10}
- C. 10^{-9}
- D. 10^{-8}

Answer: C

Solution:

Initial concentration at equilibrium



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

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

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

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

Question13

Which of the following will make a basic buffer solution?

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Options:

- A. 100 mL of $0.1 \text{ MCH}_3\text{COOH}$ + 100 mL of 0.1 M NaOH
- B. 100 mL of 0.1 M HCl + 100 mL of 0.1 M NaOH
- C. 50 mL of 0.1 M KOH + 25 mL of $0.1 \text{ MCH}_3\text{COOH}$
- D. 100 mL of 0.1 M HCl + 200 mL of $0.1 \text{ MNH}_4\text{OH}$



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 NH₄OH.

First, calculate the millimoles:

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

NH₄OH: $n = 200 \text{ mL} \times 0.1 \text{ M} = 20 \text{ millimoles}$

The balanced chemical reaction is:



After the reaction, 10 millimoles of NH₄OH remain, along with 10 millimoles of NH₄Cl. As NH₄OH is a weak base and NH₄Cl 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 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.

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

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.



0.04 moles of HCl will react with 0.04 moles of NaOH

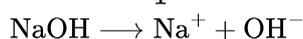
$$\begin{aligned}\text{Remaining NaOH} &= 0.05 - 0.04 \\ &= 0.01 \text{ mol}\end{aligned}$$

Final volume after adding 800 mL of distilled water

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

Concentration of NaOH after dilution

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



Concentration of OH^- is equal to concentration of NaOH = 0.01M

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

$$\text{pOH} = -\log[0.01]$$

$$\text{pOH} = 2; \text{pH} = 14 - 2 = 12$$

Question15

What is the conjugate base of chloric acid?

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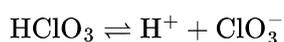
Options:



Answer: D

Solution:

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



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

So, Normality = Acidity \times Molarity

$$\Rightarrow \text{Molarity} = \frac{0.01}{2} \text{M}$$

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

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

$$\text{pOH} = -\log 10^{-2}$$

$$\Rightarrow \text{pOH} = 2 \log 10 = 2$$

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

$$\Rightarrow \text{pH} = 14 - \text{pOH} = 14 - 2$$

Thus, $\text{pH} = 12$

Question17

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

AP EAPCET 2022 - 4th July Morning Shift

Options:

A. 4.85

B. 4.65

C. 4.75



D. 7.0

Answer: C

Solution:

Given, concentration of acetic acid,



Concentration of sodium acetate,



pK_a of $\text{CH}_3\text{COOH} = 4.75$

Using Henderson-Hasselbalch equation,

$$\text{pH} = pK_a + \log \frac{[\text{CH}_3\text{COONa}]}{[\text{CH}_3\text{COOH}]}$$
$$\Rightarrow \text{pH} = 4.75 + \log \frac{0.5}{0.5}$$

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

Question 18

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 is 0.0132]

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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

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

Step 2: Calculate pH



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

$$\text{pH} = 2.88$$

✔ **Final Answer:**

Option D — 2.88

Question20

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

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Options:

A. 5×10^{-10} M

B. 25×10^{-10} M

C. 0.5 M

D. 0.002 M

Answer: B

Solution:

Given,

$$[K_{\text{sp}}]_{\text{AgBr}} = 5 \times 10^{-10}$$

$$[\text{NaBr}] = 0.2\text{M}$$

$$[\text{Na}^+] = [\text{Br}^-] = 0.2\text{M}$$

$$[\text{Ag}^+][\text{Br}^-] = [K_{\text{sp}}]_{\text{AgBr}}$$

$$[\text{Ag}^+](0.2) = 5 \times 10^{-10}$$

$$[\text{Ag}^+] = \frac{5 \times 10^{-10}}{0.2} = 25 \times 10^{-10}\text{M}$$

Question21

Calculate the pOH of 0.10 M HCl solution.

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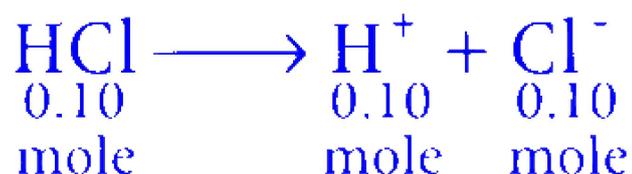
Options:

- A. 13
- B. 1
- C. 7
- D. 14

Answer: A

Solution:

HCl is the strongest acid, it dissociate into ions,



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

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

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

$$1 + \text{pOH} = 14$$

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

Question22

Which among the following pairs is not an acidic buffer?

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Options:

- A. HCOOH and HCOOK
- B. HClO₄ and NaClO₄
- C. C₆H₅COOH and C₆H₅COONa
- 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.

