One Two academy | Ionic Equillibrium

Maximum Mark: 25 **Chemistry** **Time: 50 minutes**

Choose the correct answer: **3 x 1 = 3**

1. Which of the following fluro compounds is most likely to behave as a Lewis base?

   a) BF_3 b) PF_3 c) CF_4 d) SiF_4

2. The aqueous solution of sodium acetate, ammonium chloride, sodium nitrate are respectively.

   a) Neutral, acidic, basic b) Acidic, basic, neutral

   c) Basic, acidic, neutral d) Basic, acidic, basic

3. The pH of 10^{-5} M KOH solution will be

   a) 9 b) 5 c) 19 d) none of these

Answer Any four of the following: **4 x 2 = 8**

4. Define solubility product.
5. Explain common ion effect with an example
6. What are the limitations of Arrhenius concept?
7. Calculate the pH of 10^{-7} M HCl.
8. What are Lewis's acids and bases? Give two examples for each.

Answer any three of the following: **3 x 3 = 9**

9. Derive an expression for Ostwald's dilution law.
10. Will a precipitate be formed when 0.150 L of 0.1M Pb $(NO_3)_2$ and 0.100L of 0.2 M NaCl are mixed? K_{sp} $(PbCl_2) = 1.210 \times 10^{-5}$.
11. Write notes on buffer solution.
12. Calculate the extent of hydrolysis and the pH of 0.1 M ammonium acetate Given that $K_a = K_b = 1.8 \times 10^{-5}$.

Answer the following: **1 x 5 = 5**

13. a) Derive an expression for the hydrolysis constant and degree of hydrolysis of salt of strong acid and weak base.

    (Or)

    b) Derive Henderson – Hasselbalch equation.

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All the Best

Answer Key attached in the next page !!!

This page is to prevent seeing answer key

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Answer Key

Choose the Correct answer:

1. – b) 2. – c) 3. – a)

Answer any four of the following:

4. **Solubility product:** The solubility product of a compound is defined as the product of the molar concentration of the constituent ions, each raised to the power of its stoichiometric co – efficient in a balanced equilibrium equation.

$$X_mY_n(s) \overset{H_2O}{\rightleftharpoons} mX^{n+}(aq) + nY^{m-}(aq)$$

$$K_{sp} = [X^{n+}]^m [Y^{m-}]^n$$

5. **Common ion effect:** When a salt of a weak acid is added to the acid itself, the dissociation of the weak acid is suppressed further. For example, the addition of sodium acetate to acetic acid solution leads to the suppression in the dissociation of acetic acid which is already weakly dissociated. In this case, CH_3 COOH and CH_3 COONa have the common ion, CH_3COO^-
6. **Limitations of Arrhenius concept:**
   i. Arrhenius theory does not explain the behaviour of acids and bases in non-aqueous solvents such as acetone, Tetrahydrofuran etc...
   ii. This theory does not account for the basic nature of the substances like ammonia (NH_3) which do not possess hydroxyl group.
7. **Ans:** pH = 6.7
8. **Lewis acids and bases:** According to lewis an acid is a species that accepts an electron pair Eg: BF_3, $AlCl_3$ while base is a species that donates an electron pair Eg. NH_3, H_2O

Answer any three of the following:

9. **Ostwald's dilution law:**
   Ostwald's dilution law relates the dissociation constant of the weak acid (K_a) with its degree of dissociation (α) and the concentration (C). Degree of dissociation (α)is the fraction of the total number of moles of a substance that dissociates at equilibrium.

$$\alpha = \frac{Number\ of\ moles\ dissociated}{total\ number\ of\ moles}$$

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

The dissociation constant of acetic acid is,

$$k_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

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	CH_3COOH	H^+	CH_3COO^-
Initial number of moles	1	-	-
Degree of dissociation of CH_3COOH	α	-	-
Number of moles at equilibrium	1-α	α	α
Equilibrium concentration	(1 - α) C	α C	α C

Substituting the equilibrium concentration in equation

$$K_a = \frac{(\alpha C)(\alpha C)}{(1-\alpha)C}$$

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

We know that weak acid dissociates only to a very small extent. Compared to one, α is so small and hence in the denominator (1 - α) = 1. The above expression now becomes,

$$K_a = \alpha^2 C$$

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

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

10.

Total volume = 0.250L

$$\underset{0.1\ M}{Pb(NO_3)_2} \rightleftharpoons \underset{0.1\ M}{Pb^{2+}} + \underset{0.2\ M}{2NO_3^-}$$

Number of moles

Pb^{2+} = molarity × Volume of the solution in lit

$$= 0.1 \times 0.15$$

$$[Pb^{2+}]_{mix} = \frac{0.1 \times 0.15}{0.25} = 0.06\ M$$

$$\underset{0.2\ M}{NaCl} \rightleftharpoons \underset{0.2\ M}{Na^+} + \underset{0.2\ M}{Cl^-}$$

No.of moles $Cl^- = 0.2 \times 0.1$

$$[Cl^-]_{mix} = \frac{0.2 \times 0.1}{0.25} = 0.08\ M$$

Precipitation of $PbCl_2$ (s) occurs if

$[Pb^{2+}][Cl^-]^2 > K_{sp}$

$[Pb^{2+}][Cl^-]^2 = (0.06)(0.08)^2$

$= 3.84 \times 10^{-4}$

Since ionic product $[Pb^{2+}][Cl^-]^2 > K_{sp}$, $PbCl_2$ is precipitated

11. **Buffer solution**:

- Buffer is a solution which consists of a mixture of a weak acid and its conjugate base (or) a weak base and its conjugate acid.
- This buffer solution resists drastic changes in its pH upon addition of a small quantities of acids (or) bases, and this ability is called buffer action.
- The buffer containing carbonic acid (H_2CO_3) and its conjugate base HCO_3^- is present in our blood. There are two types of buffer solutions.

  1. **Acidic buffer solution:** a solution containing a weak acid and its salt. Example: solution containing acetic acid and sodium acetate

  2. **Basic buffer solution:** a solution containing a weak base and its salt. Example: Solution containing NH4OH and NH Cl

12. **Ans:**

h = 5.555 x 10^{-3} ; pH = 7

Answer the following:

13. a) Let us consider the reactions between a strong acid, HCl, and a weak base, NH_4OH, to produce a salt, NH_4Cl, and water

$$HCl\ (aq) + NH_4OH\ (aq) \rightleftharpoons NH_4Cl(aq) + H_2O(l)$$

$$NH_4Cl(aq) \rightarrow NH_4^+ + Cl^-(aq)$$

NH_4^+ is a strong conjugate acid of the weak base NH4OHand it has a tendency to react with OH- from water to produce unionised NH_4OH shown below.

$$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_4OH\ (aq) + H^+(aq)$$

There is no such tendency shown by Cl- and therefore [H OH + -][>]; the solution is acidic and the pH is less than 7.

Hydrolysis constant (K_b):

NH_4^+ (aq) + H_2O(l) → NH_4OH (aq) + H^+(aq)

$$K_b = \frac{[NH4+]\,[OH-]}{[NH4OH]} \qquad \ldots(1)$$

$$K_h = \frac{[NH4OH]\,[H+]}{[NH4+]} \qquad \ldots(2)$$

Eq (1) x (2)

$K_b \times K_h = K_w$

$$\mathbf{K_b = \frac{Kw}{Kb}}$$

Degree of hydrolysis: (h)

$$K_h = \frac{\boldsymbol{Ch\ x\ Ch}}{\boldsymbol{C(1-h)}} = Ch^2$$

$$h = \sqrt{\frac{Kh}{C}} = \sqrt{\frac{Kw}{Kb\ x\ C}}$$

Concentration of hydrogen ion: $[H^+]$

$$[H^+] = Ch = C\sqrt{\frac{Kw}{Kb\ x\ C}} = \sqrt{\frac{Kw\ x\ C}{Kb}}$$

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$$[H^+] = \sqrt{\frac{Kw \; x \; C}{Kb}}$$

pH of the solution

$pH = - \log [H^+] = - \log \left[\frac{C \; x \; Kw}{Kb}\right]^{1/2} = - ½ \log C – ½ \log K_w + ½ \log K_b$

$pH = ½ pK_w - ½ \log C – ½ \log K_w + ½ \log K_b$

$pH = ½ pK_w - ½ \log C - ½ pK_b$

$pH = 7 – ½ \log C – ½ pK_b$

(Or)

b) **Henderson – Hasselbalch equation:**

The concentration of hydronium ion in an acidic buffer solution depends on the ratio of the concentration of the weak acid to the concentration of its conjugate base present in the solution i.e.,

$$[H_3O^+] = K_a \frac{[acid]_{eq}}{[base]_{eq}}$$

- The weak acid is dissociated only to a small extent.
- Moreover, due to common ion effect, the dissociation is further suppressed and hence the equilibrium concentration of the acid is nearly equal to the initial concentration of the unionised acid.
- Similarly, the concentration of the conjugate base is nearly equal to the initial concentration of the added salt.

$$[H_3O^+] = K_a \frac{[acid]}{[salt]}$$

Taking logarithm on both sides of the equation

$$\log [H_3O^+] = \log K_a + \log \frac{[acid]}{[salt]}$$

reverse the sign on both sides

$$-\log [H_3O^+] = -\log K_a - \log \frac{[acid]}{[salt]}$$

We know that

$pH = - \log [H_3O^+]$ and $pK_a = -\log K_a$

$$\Rightarrow pH = pK_a - \log \frac{[acid]}{[salt]}$$

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

Similarly for a basic buffer, $pOH = pK_b + \log \frac{[salt]}{[base]}$

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