JEE MAIN ANSWER KEY & SOLUTIONS

SUBJECT :- CHEMISTRY
CLASS :- 11th   **PAPER CODE :- CWT-7**
CHAPTER :- IONIC EQUILIBRIUM

ANSWER KEY

1.	(B)	**2.**	(C)	**3.**	(C)	**4.**	(B)	**5.**	(C)	**6.**	(A)	**7.**	(D)
8.	(C)	**9.**	(D)	**10.**	(A)	**11.**	(A)	**12.**	(D)	**13.**	(C)	**14.**	(C)
15.	(B)	**16.**	(A)	**17.**	(C)	**18.**	(D)	**19.**	(B)	**20.**	(B)	**21.**	4
22.	200	**23.**	28	**24.**	5	**25.**	2	**26.**	1	**27.**	4	**28.**	5
29.	50	**30.**	2										

SOLUTIONS

1.   (B)

Sol. $pH = pK_{a_2} + \log\dfrac{[CO_3^{2-}]}{[HCO_3^-]} \Rightarrow (11 - \log 4) +$

$\log\left(\dfrac{1}{2}\right) \Rightarrow 10.1$]

2.   (C)

Sol.   Concentration of acid = $\dfrac{N}{20}$ = 0.05 N

Out of 100 molecules, 3.5 molecules have been dissociated

$\therefore$ Out of 1 molecules the no. of dissociated molecules

$= \dfrac{35}{100} = 0.035 = \alpha$

$K_a = \dfrac{C\alpha^2}{(1-\alpha)}, K_a = \alpha^2 C$

$K_a = \dfrac{0.035 \times 0.035 \times 1}{20}$

$K_a = 6.125 \times 10^{-5}$

3.   (C)

Sol. $\underset{1-X}{NH_3} + H_2O \rightleftharpoons \underset{X}{NH_4^+} + \underset{X}{OH^-}$

$K_b = \dfrac{[NH_4^+][OH^-]}{[NH_3]} = \dfrac{X^2}{(1-X)} = X^2$

$(\because 1 - X \approx 1)$

$\therefore X = [OH^-] = \sqrt{K_b} = \sqrt{1.8 \times 10^{-5}}$

$= 4.24 \times 10^{-3} M$

4.   (B)

Sol. $HA + BOH \rightleftharpoons BA + H_2O$
           weak   strong

or $HA + B^+ + OH^- \rightleftharpoons B^+ + A^- + H_2O$ or

$HA + OH^- \rightleftharpoons A^- + H_2O$

$\therefore K \dfrac{\left[A^-\right]}{[HA]\left[OH^-\right]}$

Also for weak acid HA

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

$K_a \dfrac{\left[HA^+\right]\left[A^-\right]}{[HA]}$

$\dfrac{K_a}{K} = K_w$ or $K = \dfrac{K_a}{K_w} = \dfrac{10^{-4}}{10^{-14}} = 10^{10}$

5.   (C)

Sol. NH_4OH is a weak base and partially dissociated

$NH_4OH \rightleftharpoons NH4^+ + OH^-$

Concentration   1   0   0
before dissociation

Concentration $1 - \alpha$ α α
after dissociation

$\therefore$ [OH⁻] = Cα = 2 × 10⁻³ ×

$\dfrac{2}{100} = 4 \times 10^{-5} M$

pOH = – log[OH⁻]
= – log 4 × 10⁻⁵ = 4.4
pH = 14 – 4.4
= 9.6

6.   (A)

Sol.   [H⁺] ion concentration = 5 × 10⁻⁴ mole/litre
pH = – log [5 × 10⁻⁴]
= – (log 5 + log 10⁻⁴)
= – 0.7 + 4
= 3.3

1

7. (D)

Sol. For weak acid $K_a = \alpha^2 . C$

$$\therefore \alpha = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{4.9 \times 10^{-8}}{0.1}}$$

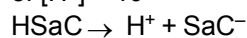= 7 × 10⁻⁴ → $= 7 \times 10^{-4}$

pH = –log H⁺ = –log α C

= – log 7 ×10⁻⁴ ×10⁻¹

= 4.1549 = 4.155

8. (C)

Sol. [HSaC] = $\dfrac{4 \times 10^{-4}}{200/1000} = 2 \times 10^{-3}M$

The dissociation of HSaC takes places in the presence

of [H⁺] = 10⁻³

HSaC → H⁺ + SaC⁻

conc. before 2×10⁻³   10⁻³   0

dissociation

In presence of H⁺ the dissociation of HSaC is almost negligible because of common ion effect. Thus at equilibrium

[HSaC] = 2 × 10⁻³, H⁺ = 10⁻³

$$K_a = \frac{\left[H^+\right]\left[SaC^-\right]}{\left[HSaC\right]}$$

$$\therefore 2 \times 10^{-12} = \frac{\left[10^{-3}\right]\left[SaC^-\right]}{\left[2 \times 10^{-3}\right]}$$

∴ [SaC⁻] = 4 × 10⁻¹² M

9. (D)

Sol. 25 ml $\dfrac{M}{5}$ NaOH will be neutralise 25 ml $\dfrac{M}{5}$ HCl.

Hence, 50 ml $\dfrac{M}{5}$ HCl get rest and to the mixing

of

both solution will give total 100 ml volume.

$N_1V_1 = N_2V_2$

50 × $\dfrac{N}{5}$ = $N_2 \times 100$

$N_2 = \dfrac{50 \times N}{5 \times 100}$ = .1 Hence, pH = 1

10. (A)

Sol. HA + H₂O ⇌ A⁻ + H₃O⁺, Ka = 1.0 × 10⁻⁶

Thus K = $\dfrac{1}{K_a}$ = 1.0 × 10⁶

11. (A)

Sol. The addition of NaOH converts equivalent amount of acetic acid into sodium acetate. Hence,Concen of acetic acid after the addition of

$$NaOH = \frac{30}{70} \times 0.2M$$

Hense, Using the expression

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

$$= -\log\left(2 \times 10^{-5}\right) + \log\left(\frac{20}{30}\right)$$

$$= 4.699 - 0.177 = 4.522$$

12. (D)

Sol. For Basic buffer solutions

$$pOH = -\log K_b + \log \frac{\left[Salt\right]}{\left[Base\right]}$$

Let a mole of NH₄Cl are added to (1.0 × 1) mole of NH₄OH in 1 litre.

Given pH = 9   ∴ pOH = 5

$$\therefore 5 = -\log 1.8 \times 10^{-5} + \log \frac{a}{1.0 \times 1}$$

or a = 1.8

13. (C)

Sol. Concentration of monobasic acid = 0.01 M

pH = 4

If the acid is completely ionised the pH of the acid would be
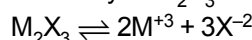
pH = – log 0.01

= – log 10⁻² = 2

So it is a weak acid. The sodium salt of a weak acid when dissolved the anions will be hydrolysed giving rise to OH⁻ ion concentration. The solution will be basic. So statement (C) is incorrect.

14. (C)

Sol. Solubility of M₂X₃ = x gm mole/litre

M₂X₃ ⇌ 2M⁺³ + 3X⁻²

∵ [M⁺³] = 2x

[X⁻²] = 3x

Solubility product K_{SP} = (2x)²·(3x)³=108 x⁵

15. (B)

Sol. (B) Solubility = $\dfrac{\text{Solubility Product}}{\text{Concentration of commonion}}$

S_1 = $\dfrac{K_{sp}}{0.04}$ = 25K_{SP}

S_2 = $\dfrac{K_{sp}}{0.01}$ = 100K_{SP}

$S_3 = \dfrac{K_{sp}}{0.05}$ = 20K_{SP}

So, $S_2 > S_1 > S_3$ Again solubility will be greatest in water.

So, $S_0 > S_2 > S_1 > S_3$

16. (A)

Sol. In 0.1 M $AgNO_3$

$AgNO_3 \rightleftharpoons Ag^+ + NO_3^-$

$AgCl \rightleftharpoons Ag^+ + Cl^-$

$K_{SP} = [Ag^+][Cl^-]$

Now $[Ag^+]$ can be taken as $[AgNO_3]$ while $[Cl^-]$ is the solubility of AgCl

$\therefore Cl = \dfrac{K_{sp}}{\left[Ag^+\right]} = \dfrac{2.8 \times 10^{-10}}{0.1}$

$\therefore$ Solubility of AgCl = 2.8×10^{-9} mole/litre

17. (C)

Sol. $K_h = \dfrac{K_w}{K_a} = \dfrac{10^{-14}}{4 \times 10^{-3}}$

= 0.25×10^{-11}

$h = \sqrt{\dfrac{K_h}{c}} = \sqrt{\dfrac{0.25 \times 10^{-11}}{0.4}}$

= 2.5×10^{-6}

18. (D)

Sol. Ammonium formate = salt of (WA + WB)

pH = $\dfrac{1}{2}$ [pK_W + pK_a – pK_b]

= $\dfrac{1}{2}$ [14 + 3.8 – 4.8] = 6.5

19. (B)

Sol. Using the expression

pH = pK_a + log $\dfrac{[Salt]}{[Acid]}$

we get, 6 = –log(1.0 × 10^{-5}) + log $\dfrac{[Salt]}{[0.02M]}$

or $\dfrac{[Salt]}{[0.02M]}$ = 10 or [Salt] = 0.2 M

20. (B)

Sol. $pH = pK_a + \log\dfrac{[Salt]}{[Acid]}$

$= -\log\left(1.0 \times 10^{-5}\right) + \log\dfrac{(0.02 - 0.01)}{(0.03 + 0.01)}$

$= 5 + \log\dfrac{1}{4} = 5 - 0.6$

$= 4.4$

21. 4

Sol. 224 moles CO_2 at STP = 0.01 moles CO_2

(i) if NaOH is limiting $\Rightarrow Na_2CO_3 + NaHCO_3$ = 0.01/L

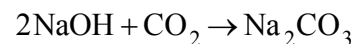(ii) if NaOH is in excess $\Rightarrow Na_2CO_3$ =100 m moles/L, NaOH = y moles

Volume of HCl used with ph as indicator = 30 m moles of HCl = 30 $\Rightarrow$ NaOH is present in excess

$\Rightarrow$ m eq of HCl = m eq of Na_2CO_3 + m eq of NaOH

30 = 10 + m eq of NaOH

$\Rightarrow$ M eq of NaOH = 20 m eq =20 m mole

Na_2CO_3 =10 m mole $\Rightarrow$ 20 m mole NaOH used in reaction

$$2NaOH + CO_2 \rightarrow Na_2CO_3$$

Total NaOH originally present = 40 m mole in 1000 ml

$\Rightarrow$ 0.04 M **Ans.**]

22. 200

Sol. $[OH^-] = 10^{-5} \times \dfrac{0.1}{0.2} = \dfrac{10^{-5}}{2}$

$[Mg^{+2}] = \dfrac{K_{sp}}{(OH^-)^2} = \dfrac{10^{-11}}{(10^{-5}/2)^2}$

= 4×10^{-1} = 0.4

moles of $MgCl_2$ = 0.2 moles
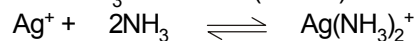
millimoles = 0.2 × 1000 = 200]

23. 28

Sol. To prevent precipitation of AgCl

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

$[Ag^+] \times 0.001 = 10^{-10}$

$[Ag^+] = \dfrac{10^{-10}}{0.001}$ = 10^{-7} M

Let a mols of NH_3 are added (V = 1L)

$Ag^+ + 2NH_3 \rightleftharpoons Ag(NH_3)_2^+$

K = 10^8

Initial conc^n 0.004   a   0

at eq^m 10^{-7} (a – 0.008) 0.004

$\dfrac{0.004}{(a - 0.008) \times 10^{-7}}$ = 10^8

$\dfrac{0.004}{(a - 0.008)^2}$ = 10

$(a - 0.008)^2$ = 0.0004 = 4×10^{-4}

a – 0.008 = 0.02

a = 0.028 mole

Millimoles of NH_3 reqd. = 28]

24. 5

Sol. $[H^+]_{HOCl} = [H^+]_{CH_3COOH}$

$\Rightarrow C \times 2 \times 10^{-6} = 0.1 \times 10^{-5}$

$\Rightarrow C = 0.5\ M$

0.5 × 10 = 5 **Ans.**

Alternative :

$[H^+] = 10^{-3}$ in HOCl

$\therefore\ \dfrac{10^{-3} \times 10^{-3}}{C - 10^{-3}} = 2 \times 10^{-6}$

C = 0.501

$\therefore\ = 5.01 = 0005$ Ans.]

25. 2

Sol. $NH_4OH \rightleftharpoons NH_4^+ + OH^-$

$\quad$ 1 $\quad\quad$ 0 $\quad\quad$ 0

c(1–α) cα cα

Given pH = 11

$\therefore [H^+] = 10^{-11}$ $\quad\quad [\therefore [H^+][OH] = 10^{-14}]$

$[OH-] = 10^{-3} = c\alpha$

Since, $\quad\quad$ c = 0.05

$\therefore\ \alpha = \dfrac{10^{-3}}{c} = \dfrac{10^{-3}}{0.05} = 2 \times 10^{-2}$ or 2%

26. 1

Sol. pH = –log10[H$^+$]

The concentration of H$^+$ ions is expressed in gm equivalent

Molarity of H_2SO_4 = 0.05

$\therefore$ Normality = 0.05 × 2 = 0.1

$\therefore$ pH = – log 0.1

or pH = 1

27. 4

Sol. We know,

$K_b(x^-) = 10^{-10}$

Also for conjugate acid – base pair

$K_{a_{HX}} \times K_{b\left(x^{-1}\right)} = 10^{-14}$

$K_{a(HX)} \times 10^{-4}$

Now [HX] = [X–]

$\quad$ (acid) (salt)

$pH = -\log K_a + \log \dfrac{[Salt]}{[Acid]}$

$= -\log 10^{-4}$

pH = 4

28. 5

Sol. We have

$pH = -\log K_a + \log \dfrac{[Salt]}{[Acid]}$

$= -\log 2 \times 10^{-5} + \log \dfrac{\dfrac{0.2 \times 1000}{100}}{\dfrac{0.1 \times 1000}{100}}$

= 4.7 + 0.3 = 5

29. 50

Sol. Let V ml of 0.1 M HCOONa be mixed to 50ml of 0.05 M HCOOH in mixture

$[HCOONa] = \dfrac{0.1 \times V}{(V + 50)}$

$[HCOOH] = \dfrac{50 \times 0.05}{(V + 50)}$

$pH = -\log K_a + \log \dfrac{[Salt]}{[Acid]}$

$4.0 = 3.7 + \log \dfrac{(0.1 \times V)/(V + 50)}{2.5/(V + 50)}, V = 50\,ml$

30. 2

Sol. $M(OH)_x \rightleftharpoons M^{x+} + XOH^-$

$\quad\quad\quad\quad$ S $\quad\quad$ XS

$K_{SP} = S.(XS)^x$ $\quad\quad$ Given S = 10^{-4}

or 4 × 10^{-12} = $10^{-4}.(X.10^{-4})x$

$10^{-4} \times 10^{-4x}.X^x$ = 4 × 10^{-12}

$X^x.(10^{-4})^{1+x}$ = 4 × 10^{-12} $\therefore$ × = 2