# Ionic Equilibrium

## SET Self Evaluation Test - 9

**1.** The most important buffer in the blood consists of **[BHU 1981]**

(a) $HCl$ and $Cl^{\oplus}$ (b) $H_2CO_3$ and $HCO_3^{\ominus}$

(c) $H_2CO_3$ and $Cl^{\ominus}$ (d) $HCl$ and $HCO_3^{\ominus}$

**2.** The solubility product of $AgI$ at $25°C$ is $1.0 \times 10^{-16} mol^2 L^{-2}$. The solubility if $AgI$ in $10^{-4} N$ solution of $KI$ at $25°C$ is approximately (in mol $l^{-1}$) **[CBSE PMT 2003]**

(a) $1.0 \times 10^{-8}$ (b) $1.0 \times 10^{-16}$

(c) $1.0 \times 10^{-12}$ (d) $1.0 \times 10^{-10}$

**3.** The pH of the solution: $5\ mL$ of $\frac{M}{5}$, $HCl + 10\ mL$ of $\frac{M}{10}$ $NaOH$ is **[MH CET 2004]**

(a) 5 (b) 3

(c) 7 (d) 8

**4.** Given that the dissociation constant for $H_2O$ is $K_w = 1 \times 10^{-14} mole^2 / litre^2$. What is the $pH$ of a 0.001 *molar KOH* solution **[UPSEAT 2000; MP PET 2001]**

(a) $10^{-11}$ (b) $10^{-3}$

(c) 3 (d) 11

**5.** The $pH$ of 0.1 $M$ solution of the following salts increases in the order **[Pb. CET 2004]**

(a) $NaCl < NH_4Cl < NaCN < HCl$

(b) $HCl < NH_4Cl < NaCl < NaCN$

(c) $NaCN < NH_4Cl < NaCl < HCl$

(d) $HCl < NaCl < NaCN < NH_4Cl$

**6.** The degree of hydrolysis in hydrolytic equilibrum $A^- + H_2O \rightleftharpoons HA + OH^-$ at salt concentration of 0.001 $M$ is $\left(K_a = 1 \times 10^{-5}\right)$ **[UPSEAT 2004]**

(a) $1 \times 10^{-3}$ (b) $1 \times 10^{-4}$

(c) $5 \times 10^{-4}$ (d) $1 \times 10^{-6}$

**7.** If $pK_b$ for fluoride ion at $25^{o}C$ is 10.83, the ionisation constant of hydrofluoric acid in water at this temperature is **[IIT 1997]**

(a) $1.74 \times 10^{-3}$ (b) $3.52 \times 10^{-3}$

(c) $6.75 \times 10^{-4}$ (d) $5.38 \times 10^{-2}$

**8.** If the hydrogen ion concentration of a given solution is $5.5 \times 10^{-3} mol\ litre^{-1}$, the $pH$ of the solution will be **[AMU 1985]**

(a) 2.26 (b) 3.40

(c) 3.75 (d) 2.76

**9.** Henderson's equation is $pH = pK_a + \log \frac{[\text{salt}]}{[\text{acid}]}$. If the acid gets half neutralized the value of $pH$ will be : $[pK_a = 4.30]$ **[RPMT 2000]**

(a) 4.3 (b) 2.15

(c) 8.60 (d) 7

**10.** The $pH$ of a $0.01 M$ solution of acetic acid having degree of dissociation 12.5% is **[JIPMER 2000]**

(a) 5.623 (b) 2.903

(c) 3.723 (d) 4.509

**11.** Which of the following solutions will have $pH$ close to 1.0 **[IIT 1992; MP PET 1993; AMU 1999]**

(a) $100\ ml$ of $\frac{M}{10} HCl + 100\ ml$ of $\frac{M}{10} NaOH$

(b) $55\ ml$ of $\frac{M}{10} HCl + 45\ ml$ of $\frac{M}{10} NaOH$

(c) $10\ ml$ of $\frac{M}{10} HCl + 90\ ml$ of $\frac{M}{10} NaOH$

(d) $75\ ml$ of $\frac{M}{5} HCl + 25\ ml$ of $\frac{M}{5} NaOH$

**12.** In which of the following solvents will $AgBr$ have the highest solubility **[CBSE PMT 1992]**

(a) $10^{-3} M\ NaBr$ (b) $10^{-3} M\ NH_4OH$

(c) Pure water (d) $10^{-3} M\ HBr$

**13.** How many grams of $CaC_2O_4$ will dissolve in distilled water to make one litre of saturated solution ? (Solubility product of $CaC_2O_4$ is $2.5 \times 10^{-9} mole^2 litre^{-2}$ and its molecular weight is 128) **[MP PET 1993; MP PMT 2000]**

(a) 0.0064 *gm* (b) 0.0128 *gm*

(c) 0.0032 *gm* (d) 0.0640 *gm*

**14.** The solubility product of $CuS$, $Ag_2S$, $HgS$ are $10^{-31}$, $10^{-44}$, $10^{-54}$ respectively. The solubilities of these sulphides are in the order

(a) $Ag_2S > CuS > HgS$ (b) $Ag_2S > HgS > Cus$

(c) $HgS > Ag_2S > Cus$ (d) $CuS > Ag_2S > HgS$

**15.** The solubility product constant $K_{sp}$ of $Mg(OH)_2$ is $9.0 \times 10^{-12}$. If a solution is $0.010\ M$ with respect to $Mg^{2+}$ ion, what is the maximum hydroxide ion concentration which could be present without causing the precipitation of $Mg(OH)_2$

(a) $1.5 \times 10^{-7}\ M$ (b) $3.0 \times 10^{-7}\ M$

(c) $1.5 \times 10^{-5}\ M$ (d) $3.0 \times 10^{-5}\ M$

**16.** If the $K_b$ value in the hydrolysis reaction $B^+ + H_2O \rightleftharpoons BOH + H^+$ is $1.0 \times 10^{-6}$, then the hydrolysis constant of the salt would be **[Roorkee Qualifying 1998]**

(a) $1.0 \times 10^{-6}$ (b) $1.0 \times 10^{-7}$

(c) $1.0 \times 10^{-8}$ (d) $1.0 \times 10^{-9}$

**17.** For a sparingly soluble salt $A_pB_q$, the relationship of its solubility product $(L_S)$ with its solubility $(S)$ is

**[IIT Screening 2001]**

(a) $L_s = S^{p+q}.p^p.q^q$ (b) $L_s = S^{p+q}.p^q.q^p$

(c) $L_s = S^{pq}.p^p.q^q$ (d) $L_s = S^{pq}.(p.q)^{p+q}$

**18.** Arrange $NH_4^+, H_2O, H_3O^+, HF$ and $OH^-$ in increasing order of acidic nature **[BVP 2003]**

(a) $H_3O^+ < NH_4^+ < HF < OH^- < H_2O$

(b) $NH_4^+ < HF < H_3O^+ < H_2O < OH^-$

(c) $OH^- < H_2O < NH_4^+ < HF < H_3O^+$

(d) $H_3O^+ > HF > H_2O > NH_4^+ > OH^-$

**19.** How many grams of $CaC_2O_4$ (molecular weight = 128) on dissolving in distilled water will give a saturated solution $[K_{sp}(CaC_2O_4) = 2.5 \times 10^{-9}\ mol^2 l^{-2}]$ **[KCET 2003]**

(a) 0.0064 *g* (b) 0.1280 *g*

(c) 0.0128 *g* (d) 1.2800 *g*

**20.** If the concentration of $CrO_4^-$ ions in a saturated solution of silver chromate is $2 \times 10^{-4}$. Solubility product of silver chromate will be

(a) $4 \times 10^{-8}$ (b) $8 \times 10^{-12}$

(c) $12 \times 10^{-12}$ (d) $32 \times 10^{-12}$

**21.** According to Bronsted-Lowry concept, the correct order of relative strength of bases follows the order

**[Pb. PMT 2001]**

**[CBSE PMT 1997]**

(a) $CH_3COO^- > Cl^- > OH^-$

(b) $CH_3COO^- > OH^- > Cl^-$

(c) $OH^- > CH_3COO^- > Cl^-$

(d) $OH^- > Cl^- > CH_3COO^-$

**22.** $H_2SO_4^- + OH^- \rightarrow SO_4^{2-} + H_2O$ Which is correct about conjugate acid base pair **[JEE Orissa 2004]**

**[Pb. PMT 1998]**

(a) $HSO_4^{2-}$ is conjugate acid of base $SO_4^{2-}$

(b) $HSO_4^-$ is conjugate base of acid $SO_4^{2-}$

(c) $SO_4^-$ is conjugate acid of base $HSO_4^-$

(d) None of these

**23.** Which may be added to one litre of water to act as a buffer

**[JIPMER 2000]**

(a) One mole of $HC_2H_3O_2$ and 0.5 mole of $NaOH$

(b) One mole of $NH_4Cl$ and one mole of $HCl$

(c) One mole of $NH_4OH$ and one mole of $NaOH$

(d) One mole of $HC_2H_3O_2$ and one mole of $HCl$

**24.** Which of the following base is weakest **[DCE 2003]**

(a) $NH_4OH : K_b = 1.6 \times 10^{-6}$

(b) $C_6H_5NH_2 : K_b = 3.8 \times 10^{-10}$

(c) $C_2H_5NH_2 : K_b = 5.6 \times 10^{-4}$

(d) $C_6H_7N : K_b = 6.3 \times 10^{-10}$

**25.** $HClO$ is a weak acid. The concentration of $H^+$ ions in $0.1\,M$ solution of $HClO\,(K_a = 5 \times 10^{-8})$ will be equal to

**[CPMT 1993]**

(a) $7.07 \times 10^{-5}\ m$ (b) $5 \times 10^{-9}\ m$

(c) $5 \times 10^{-7}\ m$ (d) $7 \times 10^{-4}\ m$

**26.** Upto what $pH$ must a solution containing a precipitate of $Cr(OH)_3$ be adjusted so that all of precipitate dissolves

(When $Cr^{3+} = 0.1 mol/l, K_{sp} = 6 \times 10^{-31}$) **[MP PET 2003]**

**[MP PET 1992; CPMT 1993]**

(a) Upto 4.4 (b) Upto 4.1

(c) Upto 4.2 (d) Upto 4.0

**27.** $NH_4Cl$ is acidic, because **[JEE Orissa 2004]**

(a) On hydrolysis $NH_4Cl$ gives weak base $NH_4OH$ and strong acid $HCl$

(b) Nitrogen donates a pair of electron

(c) It is a salt of weak acid and strong base

(d) On hydrolysis $NH_4Cl$ gives strong base and weak acid

**28.** A solution of weak acid $HA$ containing 0.01 *moles* of acid *per litre* of solutions has $pH = 4$. The percentage degree of ionisation of the acid and the ionisation constant of acid are respectively **[UPSEAT 2001]**

(a) $1\%, 10^{-6}$ (b) $0.01\%, 10^{-4}$

(c) $1\%, 10^{-4}$ (d) $0.01\%, 10^{-6}$

**29.** The $pH$ of a buffer solution containg 0.2 *mole per litre* $CH_3COONa$ and 1.5 *mole per litre* $CH_3COOH$ is ($Ka$ for acetic acid is $1.8 \times 10^{-5}$) **[CPMT 2001]**

(a) 4.87 (b) 5.8

(c) 2.4 (d) 9.2

**30.** 100 *mL* of $0.04\ N\ HCl$ aqueous solution is mixsed with 100 *mL* of 0.02 *N NaOH* solution. The *pH* of the resulting solution is **[UPSEAT 2004]**

(a) 1.0 (b) 1.7

(c) 2.0 (d) 2.3

**31.** An alcoholic drink substance $pH = 4.7$ then $OH$ ion concentration of this solution is $(K_w = 10^{-14} mol^2/l^2)$

**[RPMT 2002]**

(a) $3 \times 10^{-10}$ (b) $5 \times 10^{-10}$

(c) $1 \times 10^{-10}$ (d) $5 \times 10^{-8}$

**32.** In its 0.2 *M* solution, an acid ionises to an extent of 60%. Its hydrogen ion concentration is

(a) 0.6 *M* (b) 0.2 *M*

(c) 0.12 *M* (d) None of these

**33.** *p*H of $0.1\ M$ $NH_3$ aqueous solution is

$(K_b = 1.8 \times 10^{-5})$ **[UPSEAT 2004]**

(a) 11.13 (b) 12.5

(c) 13.42 (d) 11.55

**34.** 40 *mg* of pure sodium hydroxide is dissolved in 10 litres of distilled water. The *p*H of the solution is

**[Kerala PMT 2004]**

(a) 9.0 (b) 10

(c) 11 (d) 12

(e) 8

**35.** Solubility of $PbI_2$ is 0.005 *M*. Then, the solubility product of $PbI_2$ is **[BVP 2004]**

(a) $6.8 \times 10^{-6}$

(b) $6.8 \times 10^{6}$

(c) $2.2 \times 10^{-9}$

(d) None of these

**36.** A monoprotic acid in a $0.1\ M$ solution ionizes to 0.001%. Its ionisation constant is

**[MP PET 1985,88,99; MP PMT1988; CPMT 2003]**

(a) $1.0 \times 10^{-3}$ (b) $1.0 \times 10^{-6}$

(c) $1.0 \times 10^{-8}$ (d) $1.0 \times 10^{-11}$

**37.** Select the $pK_a$ value of the strongest acid from the following

**[KCET 2004]**

(a) 1.0 (b) 3.0

(c) 2.0 (d) 4.5

**38.** At 90°*C*, pure water has $H_3O^+$ ion concentration of $10^{-6} mol/L^{-1}$. The $K_w$ at 90°*C* is **[DCE 2004]**

(a) $10^{-6}$ (b) $10^{-14}$

(c) $10^{-12}$ (d) $10^{-8}$

**39.** By adding 20 $ml$ $0.1\ N\ HCl$ to $20\ ml\ 0.1\ N\ KOH$, the $pH$ of the obtained solution will be **[CPMT 1975, 86, 93]**

(a) 0 (b) 7

(c) 2 (d) 9

# AS Answers and Solutions

(SET -9)

**1.** (b) Blood consists of $H_2CO_3 + HCO_3^-$ buffer solution.

**2.** (c) $AgI \rightleftharpoons \underset{(s)}{Ag^+} + \underset{(s)}{I^-}$ ; $K_{sp} = S^2 = 10^{-4} \times S$

$$S = \frac{1.0\times10^{-16}}{10^{-4}} = 1\times10^{-12}\,\frac{mol^2}{l^2}$$

**3.** (c) Milliequivalents of $HCl = 5\times\frac{1}{5} = 1$

Milliequivalents of $NaOH = 10\times\frac{1}{10} = 1$

$\therefore\ 5ml\ \frac{M}{5}HCl = 10\ ml\ \frac{M}{5}HCl$

Hence the solution will be neutral *i.e.*, $pH = 7$.

**4.** (d) $pH = 14 - pOH = 14 - 3 = 11$

**5.** (b) $HCl$ is strong acid. In its $.1M$ solution, $[H^+] = 0.1M$ and hence, $pH = 1$

$NH_4Cl_{(aq)}$ hydrolyses in solution and give acidic solution which is less acidic than $.1M\ HCl$. $NaCl$ is not hydrolysed in aqueous solutions. Its $pH = 7$ $NaCN$ undergoes hydrolysis in solution to give alkaline solution. So that $pH$ increases in the order, $HCl < NH_4Cl < NaCl < NaCN$

**6.** (a) $K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{1\times10^{-5}} = 10^{-9}$

$K_h = \alpha^2 C$ ; $\alpha = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{1\times10^{-9}}{.001}} = 1\times10^{-3}$

**7.** (c) $K_a \times K_b = K_w$

$\therefore K_a = \frac{K_w}{K_b} = \frac{10^{-14}}{1.48\times10^{-11}} = 6.75\times10^{-4}$

**8.** (a) $[H^+] = 5.5\times10^{-3}$ *mole/litre*

$pH = -\log[H^+]$ ; $pH = -\log[5.5\times10^{-3}]$ ; $pH = 2.26$

**9.** (a) $pH = pK_a + \log\frac{[\text{Salt}]}{[\text{Acid}]}$

$pH = 4.3 + \log\frac{\frac{1}{2}}{\frac{1}{2}} = 4.3 + \log 1$ ; $pH = 4.3 + 0 = 4.3$

**10.** (b) $[H^+] = C\alpha = 0.01\times\frac{1.25}{100}$

$H^+ = 1.25\times10^{-3}$ ; $pH = \text{between}\ \ 2\ \text{or}\ 3 = 2.90$

**11.** (d) M.eq. of $HCl = \frac{1}{5}\times75 = 15$

M.eq. of $NaOH = 25\times\frac{1}{5} = 5$

Total No. of eq. $= 15 - 5 = 10$

Total volume = 100

Normality $= \frac{10}{100} = \frac{1}{10}$, $[H^+] = 10^{-1}\,M$

**12.** (b) $AgBr$ are not dissolved in $NaBr$ and $HBr$ due to common ion effect. And pure water is a neutral solvent. They do not have ions.

**13.** (a) $CaC_2O_4$ is a binary electrolyte. Then solubility is

$S = \sqrt{K_{sp}} = \sqrt{2.5\times10^{-9}}$

$= 5\times10^{-5}$ *mole/l.* $= 0.0064$ *gm/l.*

**14.** (a)

**15.** (d) $\underset{K_{sp}}{Mg(OH)_2}\ \underset{=}{\ }\ \rightleftharpoons\ \underset{S}{Mg^{++}} + \underset{(2S)^2}{2OH^-}$

$K_{sp} = S\times4S^2$

$\frac{K_{sp}}{S\times4} = S^2 = \frac{9\times10^{-12}}{.010\times4} = 2.25\times10^{-10}$

$S = \sqrt{2.25 \times 10^{-10}} = 1.5 \times 10^{-5}\ m/l$

**16.** (c) For hydrolysis of $B^+$; $K_H = \frac{K_w}{K_b} = \frac{10^{-14}}{10^{-6}} = 10^{-8}$.

**17.** (a) $A_pB_q \rightleftharpoons pA^{1+} + qB^{p-}$

$L_s = [A^{q+}]^p[B^{p-}]^q = (p \times S)^p(q \times S)^q = S^{p+q}.p^p.q^q$.

**18.** (c) $H_3O^+ > HF > NH_4^+ > H_2O > OH^-$.

Acidic nature is decreasing order.

**19.** (a) Solubility of $CaC_2O_4 = \sqrt{K_{sp}} = \sqrt{2.5 \times 10^{-9}}$

$= 5 \times 10^{-5} molL^{-1}$

$= 5 \times 10^{-5} \times 128 = 640 \times 10^{-5} = 0.0064\ g$

**20.** (d) $K_{sp}$ of $Ag_2CrO_4 = [Ag^+]^2\,[Cro_4^{--}]$

$CrO_4^{--} = 2 \times 10^{-4}$ then $Ag^+ = 2 \times 2 \times 10^{-4}$

$K_{sp} = (4 \times 10^{-4})^2\,(2 \times 10^{-4}) = 32 \times 10^{-12}$

**21.** (c) Relative strength of bases can be shown by their conjugated acids.

Conjugate acid of $OH^-$ is $H_2O$ which is a weak acid conjugate acid of $CH_3COO^-$ is $CH_3COOH$ which is stronger than $H_2O$. while conjugate acid of $Cl^-$ is $HCl$ which is strongest out of there. so the order of relative strength of bases is $OH^- > CH_3COO^- > Cl^-$.

**22.** (a) $\underset{\text{Conjugate acid}}{HSO_4^- + OH^-} \to \underset{\text{Conjugate base}}{SO_4^{2-} + H_2O}$

**23.** (a) One mole oxalic acid & 0.5 mole of $NaOH$ will make.

**24.** (b) Smallest value of $K_b$ indicates that aniline $(C_2H_5NH_2)$ is the weakest base.

**25.** (a) $[H^+]^2 = C.\alpha = 0.1 \times 5 \times 10^{-8}$

$H^+ = \sqrt{5 \times 10^{-9}} = 7.07 \times 10^{-5}\ M.$

**26.** (d) $K_{sp} = [Cr^{3+}][OH^-]^3$

$[OH]^{-3} = K_{sp/Cr^{3+}} = \frac{6 \times 10^{31}}{1 \times 10^{-1}} = 6 \times 10^{-30}$

$[OH]^- = 1.8 \times 10^{-10}$

$pOH = (\log 1.8 + \log 10^{10}) = 10 + 0.25 + 1 = 11.25$

$pH = 14 - 11.25 = 2.27$

**27.** (a) $NH_3Cl + H_2O \rightleftharpoons NH_4OH + HCl$

$NH_4Cl$ is a salt of weak base & strong acid so solution will be acidic.

**28.** (a) $H^+ = C\alpha$

$\alpha = \frac{H^+}{C} = \frac{10^{-4}}{10^{-2}} = 10^{-6}$

**29.** (a) $pH = -\log K_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$

$= -\log\ [1.8 \times 10^{-5}] + \log \frac{0.2}{0.1} = 4.87$

**30.** (c) $N_1V_1 = .04 \times 100 = 4$

$N_2V_2 = .02 \times 100 = 2$

$N_1V_1 - N_2V_2 = N_3V_3$

$4 - 2 = N_3 \times 200$, $N_3 = 10^{-2}\ M$

$pH = \log 10 \frac{1}{H^+} = \log 10 \frac{1}{10^{-2}} = 2$.

**31.** (b) $pH = 4.7$

$pH + pOH = 14$ ; $pH = 14 - 4.7$ ; $pOH = 9.3$

$[OH^-] =$ Antilog $[-pOH] =$ Antilog $[-9.3]$

$[OH^-] = 5 \times 10^{-10}$

***

**32.** (c) $[H^+] = C.\alpha,\ = 0.2 \times 0.60 = 0.12\ M$

**33.** (a) $NH_4OH \rightleftharpoons NH_4^+ + OH^-$

$K_b = C\alpha^2$ ; $\frac{1.8 \times 10^{-5}}{.1} = \alpha^2$ ; $\alpha = 1.34 \times 10^{-3}$

$[OH^-] = \alpha\ .\ C = 1.34 \times 10^{-3} \times .1$

$pOH = \log 10 \frac{1}{1.34 \times 10^{-4}}$ ; $pOH = 2.87$

$pH + pOH = 14$ ; $pH + 2.87 = 14$

$pH = 14 - 2.87$ ; $pH = 11.13$

**34.** (b) $M = \frac{\text{Solute in 1 litre solution}}{\text{Molecular weight of solute}}$

$= \frac{40 \times 10^{-3}}{40} \times \frac{1}{10} = 10^{-4}\ M$

$pOH = \log 10 \frac{1}{[OH^-]} = \log 10 \frac{1}{10^{-4}} = 4$

$pH + pOH = 14$ ; $pH + 4 = 14 \Rightarrow pH = 10$ .

**35.** (d) $PbI_2 \to \underset{x}{Pb} + \underset{2x}{I_2}$

$K_{sp} = 4x^3 = 4(.005)^3 = 4 \times .005 \times .005 = .4 \times 10^{-6}$.

**36.** (d) $\therefore$ Monoprotic acid $HA$

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

Ionisation constant = ?

$$\alpha = 0.001\,\% = \frac{0.001}{100} = 10^{-5}$$

$$K = \frac{\alpha^2}{V} = \frac{[10^{-5}]^2}{10} = 10^{-11}\,.$$

**37.** (a) $pKa <<$ then strongest acid

$pKa >>$ then weak acid

$$pKa \propto \frac{1}{\text{Acidic strength}}$$

**38.** (c) $\underset{10^{-6}}{H_3O^+} \rightarrow \underset{10^{-6}}{H_2O} + \underset{10^{-6}}{H^+}$

$$K_w = [H_2O][H^+] = [10^{-6}][10^{-6}] = 10^{-12}$$

**39.** (b) Neutralization reaction will takes place and form salt of strong acid and strong base. Which does not hydrolysed and thus $pH = 7$.