Electrochemistry - CET

1. What is the time taken in seconds required for depositing all the silver present in 125 ml of 1 M $AgNO_3$ solution by passing a current of 241.25A ? (1F = 96500 C)

1) 10 2) 50 3) 1000 4) 100

Ans: 2 50

Soln : For $AgNO_3$ solution 1 M = 1 N

No. of g.equivalents of silver in 125ml of 1N $AgNO_3$ solution

$$\frac{125\times1}{1000}=0.125 \text{ g.equivalents}$$

For the deposition of 1g equivalent of silver 96500C of electric current is required.

$\therefore$ For the deposition of 0.125 g.eq. of Ag current required is 96500×0.125 C

$Q = I\,t \qquad 96500 \times 0.125 = 241.25 \times t$

$$t = \frac{96500 \times 0.125}{241.25} = 50 \text{ seconds}$$

2. The volume of H_2 obtained at STP when Mg obtained by passing a current of 0.5 amp through molten $MgCl_2$ for 32.2 minutes is treated with excess of dil.HCl is approximately

1) 56 ml 2) 5.6 ml 3) 28 ml 4) 112 ml

Ans: 4 112 ml

Solution: Coulombs of electricity passed $Q = I\,t$

$$= 0.5 \times 32.2 \times 60$$

$$= 966\ C$$

96500 C of electric current deposits 1g eq of magnesium which with dil HCl gives 1 g eq of H_2 = 11200 ml

$\therefore$ 966 C of electric current gives

$$\frac{966 \times 11200}{96500} = 112\ ml$$

3. What is the ratio of weights of iron liberated at cathodes when the same current is passed through two solutions of ferric and ferrous salts arranged in series for a given time interval?

1) 3 : 2   2) 2 : 3   3) 1 : 3   4) 1 : 1

Ans: 2 2 : 3

Solution: Let the at. Mass of iron be x

Equivalent masses of Fe^{+3} and Fe^{+2} are

$\frac{x}{3}$ and $\frac{x}{2}$

On passing same current through both the solutions the ratio of masses of iron deposited at cathode will be

$$\frac{x}{3} : \frac{x}{2} \qquad \text{i.e, } 2 : 3$$

4. 1 mole of Al is deposited by x coulomb of electricity passing through molten aluminium nitrate. The number of moles of silver deposited by x coulomb of electricity from silver nitrate solution is

1) 3 2) 4 3) 2 4) 1

Ans: 1 3

x coulomb deposits 1 mole of Al $\equiv$ 3 eq of Al

$\therefore$ Silver deposited = 3 eq = 3 mole

(Since silver is monovalent no of moles = no of equivalent)

5. The specific conductance of 0.01 M NaCl solution is 0.12 Sm^{-1}. Its molar conductance in $Sm^2 \, mol^{-1}$ is

1) 1.2×10^{-4}

2) 1.2×10^{-3}

3) 1.2×10^{-2}

4) 120

Ans: 3 **1.2×10^{-2}**

Solution :

$$\mu = \frac{k}{1000C} = \frac{0.12}{1000\times0.01}$$

$$=1.2\times10^{-2}\ Sm^{2}\ mol^{-1}$$

6. The relationship between molar conductance (μ) and equivalent conductance (λ) for Na_2SO_4 is

1) $\mu = 2\lambda$
2) $\lambda = 2\mu$
3) $\lambda = \mu$
4) $\lambda = 3\mu$

Ans: 1 $\mu = 2\,\lambda$

Solution :

For Na_2SO_4

$$\textbf{Eq mass} = \frac{\text{Mol. mass}}{\text{Total no. of + ve charges on the cation}}$$

$$= \frac{\text{Mol.mass}}{2}$$

$$\therefore \qquad n = \frac{\text{Mol.mass}}{\text{Eq. mass}} = 2$$

$$\text{Equivalent conductance} = \frac{\text{Molar conductance}}{\text{n}}$$

$$\lambda = \frac{\mu}{2} \quad or \quad \mu = 2\lambda$$

7. λ_∞ NH_4Cl = 130 S cm^2 eq^{-1}; λ_∞ NaOH = 220 S cm^{2}eq^{-1} λ_∞ NaCl = 110.0 S cm^2 eq^{-1} If λ_m of NH_4OH at a given concentration is 12 S cm^2 eq^{-1}, what is its percentage dissociation?

1) 1% 2) 2% 3) 3% 4) 5%

Ans : 4 5%

$$\lambda_\infty NH_4OH = \lambda_\infty NH_4Cl + \lambda_\infty NaOH - \lambda_\infty NaCl$$

$$= 130+220-110 = 240 \text{ S cm}^2 \text{ eq}^{-1}$$

$$\alpha\ \% = \frac{\lambda_c}{\lambda_\infty} \times 100 = \frac{12}{240} \times 100 = 5\%$$

8. In the reaction $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

Which of the following sets represents Bronsted acid?

1) NH_3 and H_2O

2) NH_4^+ and OH^-

3) NH_3 and OH^-

4) H_2O and NH_4^+

Ans (4)

9. The conjugate base of NH_2^- is

1) NH_3 2) NH^{-2} 3) NH_4^+ 4) N_3^{-1}

Ans : 2 NH^{-2}

$$\underset{\text{Acid}}{NH_2^-} \xrightarrow{-H^+} \underset{\text{Base}}{NH^{-2}}$$

10. Which equilibrium can be described as Lewis acid base reaction but not Brownsted acid-base reaction ?

1) $H_2O + CH_3COOH \rightleftharpoons H_3O^+ + CH_3COO^-$

2) $2NH_3 + H_2SO_4 \rightleftharpoons 2NH_4^+ + SO_4^{-2}$

3) $NH_3 + CH_3COOH \rightleftharpoons NH_4^+ + CH_3COO^-$

4) $[Cu(H_2O)_4]^{2+} + 4NH_3 \rightleftharpoons [Cu(NH_3)_4]^{2+} + 4H_2O$

Ans : 4

11. The strongest Bronsted base among the following anion is

1) ClO_2^- 2) ClO_4^- 3) ClO^- 4) ClO_3^-

Ans : 3 ClO^-

Solution:

Acid strength

$$HClO < HClO_2 < HClO_3 < HClO_4$$

Strength of its conjugate base

$$ClO^- > ClO_2^- > ClO_3^- > ClO_4^-$$

12. Dissociation constants of HCOOH and CH_3COOH at certain temperature are 1.8×10^{-4} and 1.8×10^{-5} respectively. At what concentration would CH_3COOH have the same $[H^+]$ concentration as 0.01M HCOOH

1) 1×10^{-4} M
2) 1×10^{-5} M
3) 0.1 M
4) 0.01 M

Ans: 3 0.1 M

For weak acid [H+] = $\sqrt{K_a C}$

$$\left(\sqrt{K_a C}\right)_{CH_3COOH} = \left(\sqrt{K_a C}\right)_{HCOOH}$$

$$\sqrt{1.8\times10^{-5}\times C} = \sqrt{1.8\times10^{-4}\times0.01}$$

$$C_{CH_3COOH} = \frac{1.8\times10^{-4}\times0.01}{1.8\times10^{-5}} = 0.1M$$

13. Aqueous solution of CH_3COOH contains

1) CH_3COOH; H^+

2) CH_3COO^-, H_3O^+, CH_3COOH

3) CH_3COO^-, H_3O^+, H^+

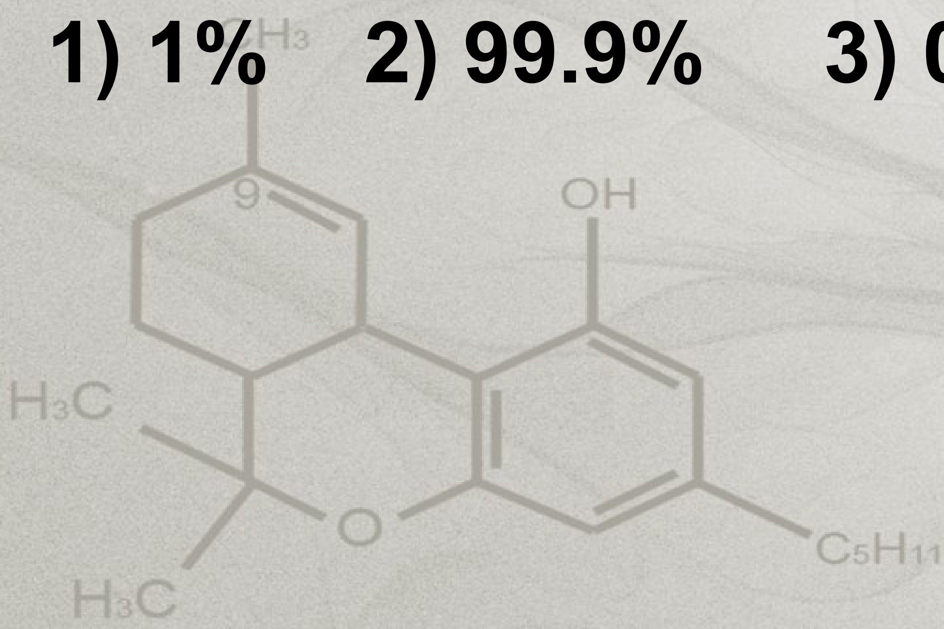
4) CH_3COOH, CH_3COO^- , H^+

Ans: 2 CH_3COO^-, H_3O^+, CH_3COOH

CH_3COOH is a weak acid that ionizes partially

14. A monoprotonic acid in 0.1 M solution has $K_a = 1.0\times10^{-5}$. The degree of dissociation of the acid is

1) 1% 2) 99.9% 3) 0.1% 4) 99%



Ans : 1 1%

$K_a = C\alpha^2$

$$\alpha = \sqrt{\frac{K_a}{C}} \times 100 = \sqrt{\frac{1.0\times10^{-5}}{0.1}} \times 100 = 1\%$$

15. At 25°C the dissociation constants of CH_3COOH and NH_4OH are almost same (10^{-5}). If pH of some acetic acid is 3, the pH of the solution of NH_4OH of same concentration at the same temperature would be

1) 3.0 2) 4.0 3) 10.0 4) 11.0

Ans: 4 11.0

Solution :

$CH_3COOH \qquad NH_4OH$

When conc. Of both solutions are same

$K_a = K_b$

$[H^+] = [OH^-]$

$p^H = p^{OH} = 3$

$\therefore p^H$ of $NH_4OH = 14 - p^{OH}$

$= 14 - 3 = 11$

16. 0.023 g of sodium metal is reacted with 100 cm^3 of water. The pH of the resulting solution is

1) 8 2) 10 3) 9 4) 12

Ans: 4 12

Solution : $Na + H_2O \longrightarrow NaOH + \frac{1}{2} H_2$

23g **40g**

0.023g **0.04g**

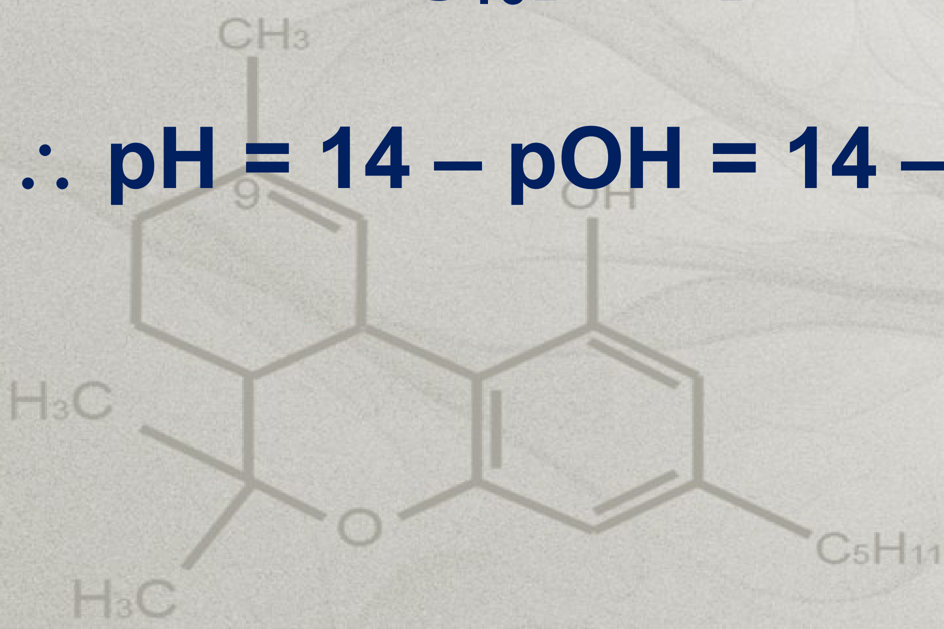
No. of moles of NaOH produced =

$$\frac{\text{mass in g}}{\text{mol mass}} = \frac{0.04}{40} = 10^{-3}$$

$$\left[NaOH\right]=\frac{no\ of\ moles}{vol\ in\ lit}=\frac{10^{-3}}{0.1}=10^{-2}=\left[OH^{-}\right]$$

$pOH = -\log_{10}[OH^{-}] = -\log_{10}10^{-2} = 2$

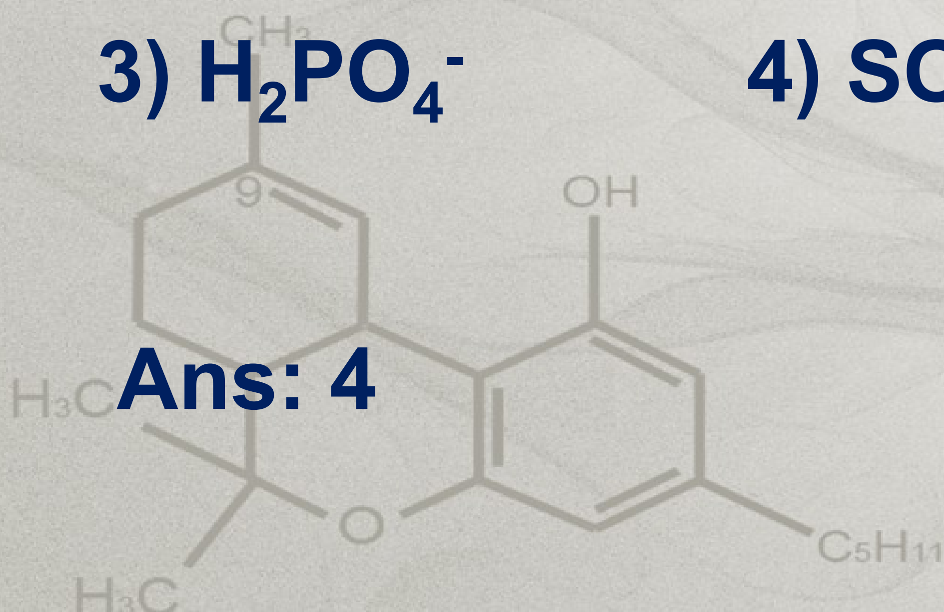
$\therefore\ pH = 14 - pOH = 14 - 2 = 12$



17. Which one of the following is not an amphoteric substance?

1) H_2O 2) HSO_4^-

3) $H_2PO_4^-$ 4) SO_4^{-2}



Ans: 4

18. Hydrogen ion concentration of an aqueous solution is 1×10^{-4} M. The solution is diluted with equal volume of water. Hydroxyl ion concentration of the resultant solution in mol dm^{-3} is

1) 1×10^{-6} 2) 1×10^{-8}

3) 0.5×10^{-10} 4) 2×10^{-10}

Ans: 4 2×10^{-10}

Solution: $[H^{+}]$ after dilution = $\dfrac{1\times10^{-4}}{2}$

$$[OH^{-}]=\frac{1\times10^{-14}}{\left[H^{+}\right]}=\frac{1\times10^{-14}}{1\times10^{-4}}\times2=2\times10^{-10}\ M$$

19. The pH of HCl is 5. If 1 ml of this solution is diluted to 1000 ml, the pH of the resulting solution is

1) 8 2) -8

3) between 6 – 7 4) between 7 - 8

Ans: 3 between 6 – 7

Solution : pH of the given HCl is 5

$[H^+] = 10^{-5}\ M = [HCl]$

Molarity of the diluted HCl

$$M_1V_1 = M_2V_2$$

$$1 \times 10^{-5} = M_2 \times 1000$$

$$M_2 = \frac{10^{-5}}{1000} = 10^{-8}\ M$$

$[H^+]$ in such a solution

$= 10^{-8}$ from HCl $+ 10^{-7}$ from water

$= 10^{-8} + 10\times10^{-8} = 11\times10^{-8} = 1.1\times10^{-7}$

pH of the solution $= -\log_{10}1.1\times10^{-7}$

$= 7 - \log 1.1$

$= 7 - 0.0414 = 6.96$

20. What is the pH of the solution obtained by mixing 250 cm^3 of a solution of pH = 3 and 750cm^3 of the solution with pH = 5

1) 4.5 2) 4 3) 3.3 4) 3.6

Ans: 4   3.6

Solution   pH = 3 $[H^+] = 10^{-3}$ M

pH = 5 $[H^+] = 10^{-5}$ M

When the two solutions are mixed in the said ratio

$$\left[H^+\right] = \frac{250\times10^{-3}+750\times10^{-5}}{1000}$$

$$= \frac{25000 \times 10^{-5} + 750 \times 10^{-5}}{1000}$$

$$= \frac{25750 \times 10^{-5}}{1000} = 2.575 \times 10^{-4}$$

$$\mathbf{pH = -\log_{10} 2.575 \times 10^{-4} = 4 - \log 2.575 \approx 3.6}$$

21. pKa of two acids A and B are 4 and 5. The strength of these two acids are related as

1) Strength of A : Strength of B = 4 : 5

2) Acid A is 3.3 times stronger than acid B

3) Acid B is 10 times stronger than acid A

4) The strength of two acids cannot be compared

Ans: 2

Solution :

Acid A $pK_a = 4 \quad \therefore K_a = 10^{-4}$

Acid B $pK_a = 5 \qquad K_a = 10^{-5}$

$$\frac{\text{Strength of Acid A}}{\text{Strength of Acid B}} = \sqrt{\frac{K_a \text{ for acid A}}{K_a \text{ for acid B}}}$$

$$= \sqrt{\frac{10^{-4}}{10^{-5}}} = \sqrt{10} = 3.3$$

22. A buffer solution is prepared by mixing 10 ml of 0.1 M acetic acid and 20 ml of 0.5 M sodium acetate and then diluted to 100 ml with distilled water. If the pKa of acetic acid is 4.76 what is the pH of the buffer solution prepared?

1) 4.76   2) 3.76   3) 5.76   4) 5.21

Ans: 3   5.76

Solution: [Acid] in the solution

$$M_1V_1 = M_2V_2$$

$$0.1 \times 10 = M_2 \times 100 \qquad M_2 = \frac{10 \times 0.1}{100} = 0.01M$$

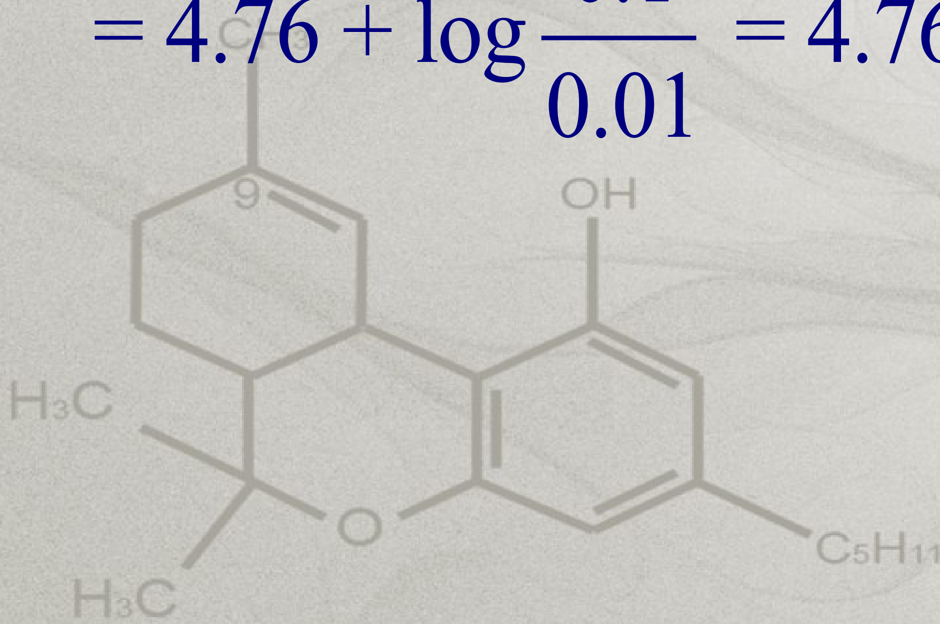
[Salt] in the solution

$$M_1V_1 = M_2V_2$$

$$0.5 \times 20 = M_2 \times 100 \qquad M_2 = \frac{0.5 \times 20}{100} = 0.1M$$

$$pH = pKa + \log\frac{[salt]}{[Acid]}$$

$$= 4.76 + \log\frac{0.1}{0.01} = 4.76 + \log 10 = 5.76$$



23. A buffer solution prepared by mixing 0.1 M NH_4OH and 0.1 M NH_4Cl in equal volumes has a pH of 9.25. pK_b of NH_4OH is

1) 9.25 2) 4.75 3) 3.75 4) 8.25

Ans: 2 4.75

Solution For a basic buffer

$$pOH = pK_b + \log \frac{[salt]}{[Base]}$$

pH = 9.25 $\therefore$ **pOH = 14 – 9.25 = 4.75**

$$\log \frac{[salt]}{[Base]} = \log 1 = 0$$

$\therefore$ **4.75 = pK_b**

24. What happens to the pH of a buffer solution of a mixture of NH_4Cl and NH_4OH when a further quantity of NH_4Cl is added

1) Decrease

2) Increases

3) Remains unchanged

4) May increase or decrease

Ans : 1 Decrease

For a basic buffer

$$pOH = pK_b + \log\frac{[\text{salt}]}{[\text{Base}]}$$

If only [Salt] is increased, pOH increases and pH decreases

25. NH_4OH is a weak base but it becomes still weaker in the aqueous solution of

1) 0.1 M HCl
2) 0.1 M NH_4Cl
3) 0.1 M H_2SO_4
4) 0.1 M CH_3COOH

Ans: 2 0.1 M NH_4Cl

Due to common ion effect NH_4Cl suppresses the degree of dissociation of NH_4OH making it a still weaker base

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

$$NH_4Cl \longrightarrow NH_4^+ + Cl^-$$

26. The precipitate of CaF_2 (K_{sp} = 1.7×10^{-10}) is obtained when equal volumes of the following are mixed

1) 10^{-4} M Ca^{+2} + 10^{-4} M F^-

2) 10^{-2} M Ca^{+2} + 10^{-3} M F^-

3) 10^{-5} M Ca^{+2} + 10^{-5} M F^-

4) 10^{-3} M Ca^{+2} + 10^{-5} M F^-

Ans: 2 10^{-2} M Ca^{+2} + 10^{-3} M F^{-}

For precipitate to take place Ksp < I.P

I.P $$\left[Ca^{+2}\right]\left[F^{-}\right]^{2}=\left(\frac{10^{-2}\times v}{2v}\right)\left(\frac{10^{-5}\times v}{2v}\right)^{2}$$

$$= 1.25\times10^{-9}$$

Which is greater than K_{sp} of CaF_2(1.7×10^{-10})

Hence precipitation takes place.

27. The solution of $CuSO_4$ in which copper plate is immersed, is diluted to 10 times, the reduction electrode potential

1) Increased by 0.030V

2) Decreased by 0.030V

3) Increased by 0.059V

4) Decreased by 0.059V

Ans: 2 Decreased by 0.030V

Solution: $$E_{Cu} = E^{o}_{Cu} + \frac{0.059}{2}\log\left[Cu^{+2}\right]$$

If $$\left[Cu^{+2}\right] = \frac{1}{10}$$

$$E_{Cu} = E^{o}_{Cu} + \frac{0.059}{2}\log\frac{1}{10}$$

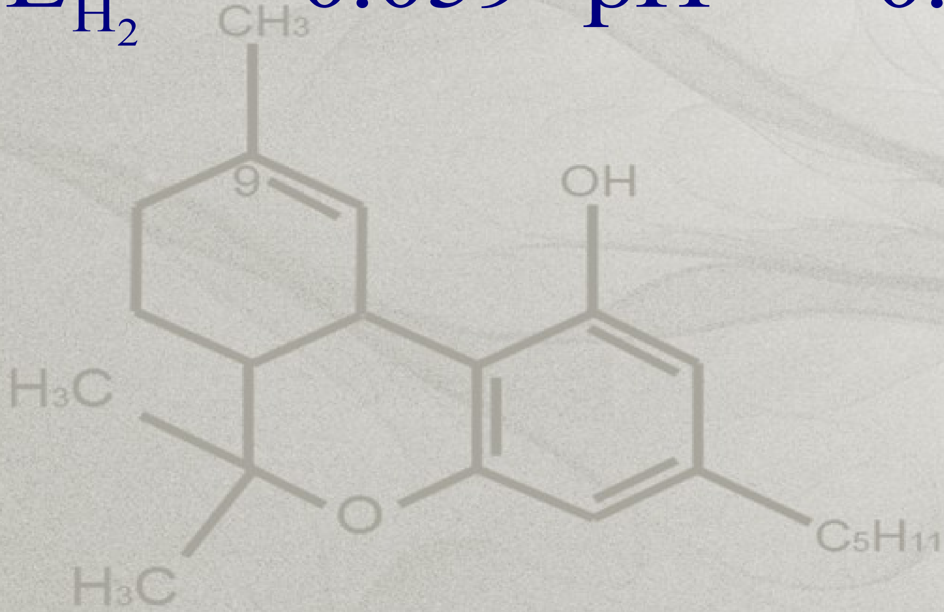
$$= E^{o}_{Cu} - 0.030V$$

28. The hydrogen electrode is dipped in a solution of pH 3 at 25°C. The potential would be

1) 0.087V
2) 0.177V
3) -0.177V
4) 0.059V

Ans: 3 -0.177V

$$E_{H_2} = -\,0.059 \times pH = -\,0.059 \times 3 = -\,0.177V$$



29. What is the EMF of the cell?

$Zn_{(s)} / Zn^{2+}(0.1M) \,||\, Sn^{2+}(0.001M) / Sn_{(s)}$

Given $E^o\ Sn^{2+}/Sn = -0.14V$ and
$E^o\ Zn^{2+}/Zn = -0.76V$

1) 0.62V 2) 0.56V 3) 1.12V 4) 0.31V

Ans: 2 0.56V

Solution: EMF = $E_{right} - E_{left}$

$$= E_{Sn^{2+}/Sn} - E_{Zn^{2+}/Zn}$$

Applying Nernst equation to each electrode

$$EMF = \left(E^{o}_{Sn^{2+}/Sn} + \frac{0.059}{2}\log_{10}\left[Sn^{2+}\right] \right) - \left(E^{o}_{Zn^{2+}/Zn} + \frac{0.059}{2}\log_{10}\left[Zn^{2+}\right] \right)$$

$$= E^{o}{}_{Sn^{2+}/Sn} - E^{o}{}_{Zn^{+2}/Zn} + \frac{0.059}{2}\log_{10}\frac{\left[Sn^{2+}\right]}{\left[Zn^{2+}\right]}$$

$$= -0.14 - (-0.76) + \frac{0.059}{2}\log_{10}\frac{0.001}{0.1}$$

$$= +\,0.62 + \frac{0.059}{2}\ \log\ 10^{-2}$$

$$\mathbf{= +0.62 - 0.059 = 0.56V}$$

30. The equilibrium constant for the reaction

$Cu + 2Ag^{+}(aq) \longrightarrow Cu^{2+}(aq) + 2Ag$;

$E^{o} = 0.46V$ at 298K is

1) 2.0×10^{10}
2) 4.0×10^{10}
3) 4.0×10^{15}
4) 2.4×10^{10}

Ans: 3 **4.0×10^{15}**

$$E^o = \frac{0.059}{n} \log Kc$$

$$0.46 = \frac{0.059}{2} \log Kc$$

$$\log Kc = \frac{0.46 \times 2}{0.059} = 15.59$$

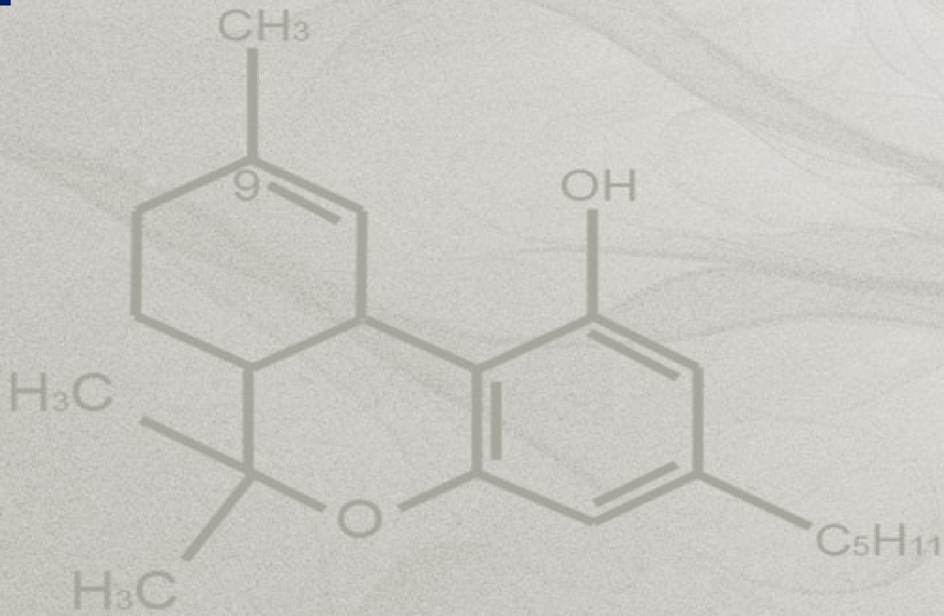
Taking the antilog $Kc = 3.9 \times 10^{15}$

31. The standard reduction electrode potentials of three metals A, B and C are +0.5V, -3.0V and -1.2V respectively. The reducing power of these metals are

1) B > C > A
2) A > B > C
3) C > B > A
4) A > C > B

Ans: 1 B > C > A

Smaller the SRP more is the reducing power



32. Which one of the following solutions when stirred with a copper spoon turns blue?

1) $Al(NO_3)_3$
2) $AgNO_3$
3) $LiNO_3$
4) $Zn(NO_3)_2$

Ans: 2 $AgNO_3$

Li/Li^{+}	- 3.04V
Al/Al^{3+}	- 1.66V
Zn/Zn^{2+}	- 0.76V
Cu/Cu^{2+}	+ 0.34V
Ag/Ag^{+}	+ 0.8V

Only Ag is below copper in electrochemical series but Al, Li & Zn are above copper. Copper can display only Ag from $AgNO_3$ solution but not other metals Al, Li and Zn from their salt solution

$$Cu + 2AgNO_3 \longrightarrow \underset{\text{Blue}}{Cu(NO_3)_2} + 2Ag$$

34. One molar aqueous solution of each $Cu(NO_3)_2$, $AgNO_3$, $Hg(NO_3)_2$ and $Mg(NO_3)_2$ is being electrolysed by using inert electrodes. The values of standard electrode potentials are

$Ag/Ag^+ = +0.80V$; $Hg/Hg^{2+} = +0.79V$,

$Cu/Cu^{2+} =+0.34V$ and $Mg/Mg^{2+} = -2.37V$

With increasing voltage, the sequence of deposition of metals on the cathode will be

1) Ag, Hg, Cu and Mg 2) Mg, Cu, Hg and Ag

3) Ag, Hg and Cu 4) Cu Hg and Ag

Ans: 3 Ag Hg and Cu

Solution: Electrolysis of aqueous solution of a metal salt above certain concentration will liberate metal at cathode if its reduction potential is more than – 0.83 the reduction potential water

Mg metal can not be deposited by electrolysing aqueous solution of $Mg(NO_3)_2$

Ions with higher values of SRP is discharged first at the cathode

Ag ; Hg ; Cu

$\longrightarrow$

Decreasing order of deposition

35. Which of the following salts when dissolved in water gives a pH of greater than 7?

1) NH_4Cl
2) CH_3COONa
3) CH_3COONH_4
4) $NaCl$

Ans:2 CH_3COONa

1) Salt of Strong acid & weak base

$$NH_4Cl + H_2O \rightleftharpoons NH_4OH + HCl$$

weak base St acid

Solution is acidicpH < 7

2) Salt of weak acid and strong base

$$CH_3COONa + H_2O \rightleftharpoons CH_3COOH + NaOH$$

wk acid st base

Soln is basic pH > 7

3) Salt of weak acid and wk base

$$CH_3COONH_4 + H_2O \rightleftharpoons \underset{\text{wk acid}}{CH_3COOH} + \underset{\text{wk base}}{NH_4OH}$$

Since Ka of CH_3COOH = Kb of NH_4OH the soln is neutral pH = 7

4) Salts of st.acid & st.base like NaCl, KNO_3 do not undergo hydrolysis and their solutions are neutral with pH = 7

36.Corrosion of iron is essentially an electrochemical phenomenon where the cell reactions are

1) Fe is oxidised to Fe^{+3} and H_2O is reduced to O_2^{-2}
2) Fe is oxidised toFe^{2+} and dissolved oxygen in water is reduced to OH^{-}
3) Fe is oxidised to Fe^{2+} and H_2O is reduced to $O_2^{\ -}$
4) Fe is oxidised to Fe^{+2} and H_2O is reduced to O_2

Ans:2

Solution

At anode $Fe \xrightarrow{oxdn} Fe^{+2} + 2\bar{e}$

At cathode $H_2O + \frac{1}{2}O_2 + 2\bar{e} \longrightarrow 2OH^-$

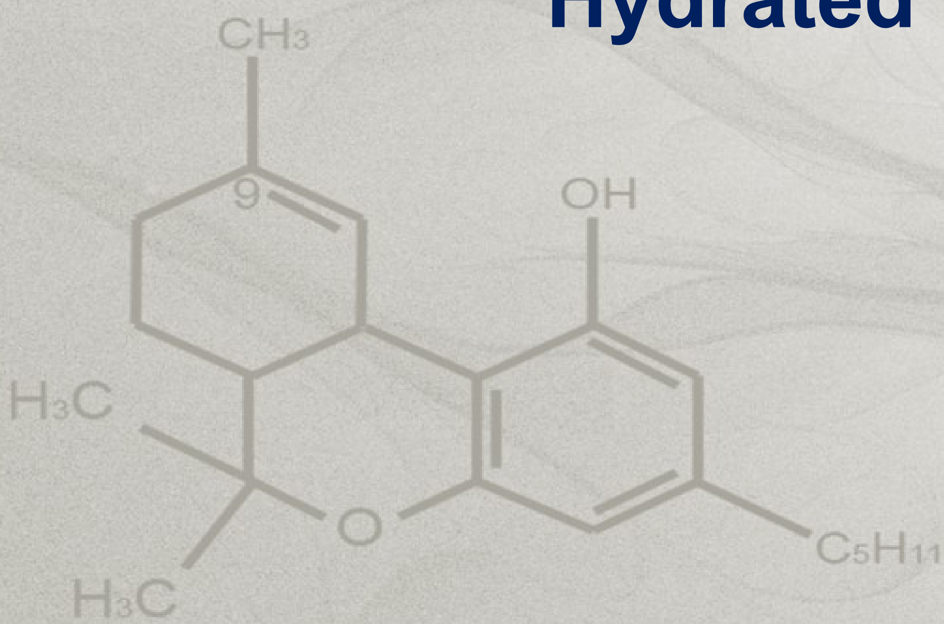
$Fe^{2+} + 2OH^- \longrightarrow Fe(OH)_2$

$2Fe(OH)_2 + H_2O + \frac{1}{2}O_2 \longrightarrow 2Fe(OH)_3$

$$2Fe(OH)_3 \longrightarrow Fe_2O_3 + 3H_2O$$

$$Fe_2O_3 + xH_2O \longrightarrow Fe_2O_3 \, . \, xH_2O$$

Hydrated ferric oxide (Rust)



37. The cell reaction for the given cell is spontaneous if

$Pt(H_2=P_1)/ H+ (1M) \;||\; H^+(1M) / Pt(H_2 = P_2)$

1) $P_1 > P_2$
2) $P_1 < P_2$
3) $P_1 = P_2$
4) $P_1 = 1$ atm

Reduction reaction taking place at the right side electrode is

$$2H^{+} + 2\bar{e} \longrightarrow H_2$$

$$E_{H_2(R)} = E^{o}_{H_2} + \frac{0.059}{2}\log_{10}\frac{\left[H^{+}\right]^{2}\ 1\mathrm{M}}{H_2(P_2 atm)}$$

Similarly

$$E_{H_2(L)} = E^{o}_{H_2} + \frac{0.059}{2}\log_{10}\frac{\left[H^{+}\right]^{2}\ 1\mathrm{M}}{H_2(P_1 atm)}$$

EMF of the cell = E $_{right}$ – E $_{left}$

$$=\frac{0.059}{2}\ \log_{10}\frac{\left[H^{+}\right]^{2}1M}{\left[H_2P_2\ atm\right]}\times\frac{\left[H_2P_1\ atm\right]}{\left[H^{+}\right]^{2}1M}$$

$$=\frac{0.059}{2}\log_{10}\frac{P_1}{P_2}$$

For a cell reaction to be spontaneous

$\Delta G^{o}=-ve$ **or E$_{cell}$ is positive which is possible only when P$_1$ > P$_2$**

38. The standard reduction potentials of Cu^{2+}/Cu and Cu^{2+}/Cu^{+} are 0.337 and 0.153v respectively. The standard electrode potential of Cu^{+}/Cu half cell is

1) 0.184v 2) 0.827v

3) 0.521v 4) 0.490v

Ans : 3 0.521V

$$Cu^{2+} + 2\bar{e} \longrightarrow Cu;\ E^{o} = 0.337V$$

$$\therefore \Delta G^{o} = -\,2 \times 0.337 \times F \quad \text{--- (1)}$$

$$Cu^{2+} + \bar{e} \longrightarrow Cu^{+};\ E^{o} = 0.153V$$

$$\therefore \Delta G^{o} = -\,1 \times 0.153 \times F \quad \text{--- (2)}$$

Eqn (1) – (2)

$$Cu^{+} + \bar{e} \longrightarrow Cu \qquad \therefore \Delta G^{o} = -\,0.521F$$

$$-\Delta G^{o} = nFE^{o}$$

$$0.521F = 1 \times F \times E$$

$$\therefore E = 0.521V$$

39. K_{sp} of CuS, Ag_2S and HgS are 10^{-31} , 10^{-44} and 10^{-54} respectively. Select the correct order for their solubility in water

1) Ag_2S > HgS > CuS

2) HgS > CuS > Ag_2S

3) HgS > Ag_2S > CuS

4) Ag_2S > CuS > AgS

Ans: 4 Ag_2S > CuS > AgS

For HgS & CuS $S = \sqrt{Ks}$

For HgS; $S = \sqrt{Ks} = \sqrt{10^{-54}}$ **= 10^{-27} M**

CuS $S = \sqrt{Ks} = \sqrt{10^{-31}} = \sqrt{10 \times 10^{-32}}$ **= 3.3×10^{-16}M**

For Ag_2S $S = 3\sqrt{\frac{Ks}{4}} = 3\sqrt{10^{-44}} = 3\sqrt{10 \times 10^{-45}}$

$\approx 2.15 \times 10^{-15}$

40. The pH at which $Mg(OH)_2$ begins to precipitate from a solution containing 0.10M Mg^{+2} ions (K_{sp} of $Mg(OH)_2 = 1 \times 10^{-11}$) is

1) 5 2) 9 3) 4 4) 10

Ans: 2 9

Solution: When $Mg(OH)_2$ starts precipitation then I.P

$[Mg^{2+}][OH^-]^2$ > Ksp of $Mg(OH)_2$

$(0.1)[OH^-]^2 > 1\times10^{-11}$

$$\left[OH^-\right]^2 > \frac{10^{-11}}{0.1} = 10^{-10}$$

$[OH^-] > 10^{-5}$

Taking the -log10 of both sides

pOH > 5

14- pOH > 14 – 5

pH > 9