

### 3. ELECTRO CHEMISTRY

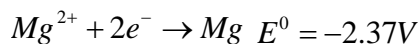
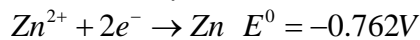
#### PREVIOUS EAMCET BITS

1. At 25°C the molar conductances at infinite dilution for the strong electrolytes NaOH, NaCl, and BaCl<sub>2</sub> are 248 × 10<sup>-4</sup>, 126 × 10<sup>-4</sup> and 280 × 10<sup>-4</sup> S m<sup>2</sup> mole<sup>-1</sup> respectively.  $\lambda_m^0$  Ba(OH)<sub>2</sub> in S m<sup>2</sup> mol<sup>-1</sup> (2009 E)
- 1) 52.4 × 10<sup>-4</sup>      2) 524 × 10<sup>-4</sup>      3) 402 × 10<sup>-4</sup>      4) 202 × 10<sup>-4</sup>

Ans : 2

Sol: 
$$\mu_{Ba(OH)_2} = \mu_{BaCl_2} + 2\mu_{NaOH} - 2\mu_{NaCl}$$
$$= 280 \times 10^{-4} + 2(248 \times 10^{-4}) - 2(126 \times 10^{-4})$$
$$= 524 \times 10^{-4}$$

2. The standard potentials at 25°C for the half reactions are given against them below (2009 M)



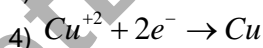
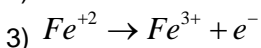
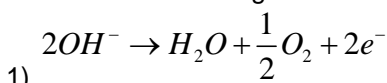
Zinc duct is added to MgCl<sub>2</sub> solution

- 1) Magnesium is precipitated      2) Zinc dissolves in the solution  
3) Zinc chloride is formed      4) No reaction takes place

Ans : 2

Sol: Zn cannot displace Mg from solution so no reaction occurs.

3. Which of the following reactions occur at the cathode? (2008 M)

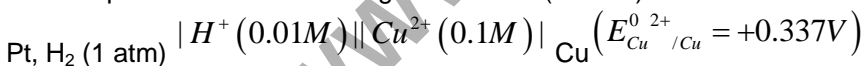


Ans : 4

Sol: At cathode reduction takes place.

Gain of electrons is reduction

4. The cell potential of the following cell at 25°C (in volts) is



(2008 M)

- 1) 0.308      2) 0.427      3) -0.308      4) 0.337

Ans : 2

Sol: 
$$E_{H_2} = E^0 + \frac{0.059}{n} \log c$$

$$E_{H_2} = 0.00 + \frac{0.06}{1} \log 10^{-2}$$

$$= 0.00 + 0.06 \times (-2)$$

$$= -0.12 V$$

$$E_{Cu} = E^0 + \frac{0.059}{n} \log c$$

$$E_{Cu} = 0.337 + \frac{0.06}{2} \log 10^{-1}$$

$$E_{Cu} = 0.337 + 0.03 \times (-1)$$

$$= 0.337 - 0.030 = 0.307V$$

$$E = E_{\text{Cathode}} - E_{\text{Anode}}$$

$$= +0.307 - (-0.12)$$

$$= +0.427 V$$

5. When electric current is passed through acidified water for 1930 seconds, 1120 ml of  $H_2$  gas is collected (at STP) at the cathode. What is the current passed in amperes?

(2008 E)

- 1) 0.05                      2) 0.50                      3) 5.0                      4) 50

Ans : 3

Sol:  $Wt \text{ of } H_2 \text{ gas} = \frac{1120}{22400} \times 2 = 0.1 \text{ gm}$

As per 1<sup>st</sup> law of Faraday

$$C = \frac{m}{et}$$

Current strength

$$\frac{0.1 \times 96500}{1 \times 1930} = 5 \text{ amp}$$

6. When same quantity of electricity is passed through aqueous  $AgNO_3$  and  $H_2SO_4$  solutions connected in series,  $5.04 \times 10^{-2}$  gram of  $H_2$  is liberated. What is the mass of silver (in grams) deposited? (Eq wts of Hydrogen = 1.008, silver = 108) (2008 E)

- 1) 54                      2) 0.54                      3) 5.4                      4)  $10^{-8}$

Ans : 3

Sol:  $\frac{wt. \text{ of } Ag \text{ deposited}}{wt. \text{ of } H_2 \text{ liberated}} = \frac{Eq. \text{ wt of } Ag}{Eq. \text{ wt of } H_2}$

$$wt. \text{ of } Ag \text{ deposited} = \frac{wt. \text{ of } H_2 \text{ liberated} \times Eq. \text{ wt of } Ag}{Eq. \text{ wt of } H_2}$$

$$\frac{5.04 \times 10^{-2} \times 108}{1.008} = 5.4 \text{ gm}$$

7. When 3.86 amperes current is passed through an electrolyte for 50 minutes, 2.4 grams of a divalent metal is deposited. The gram atomic weight of the metal (in grams) is

(2007 M)

- 1) 24                      2) 12                      3) 64                      4) 40

Ans : 4

Sol:  $m = \frac{Ect}{96500}$

$$Eq. \text{ Wt of metal } E = \frac{96500 \times m}{ct}$$

$$Eq. \text{ Wt of metal } E = \frac{96500 \times m}{ct} = 20$$

$$At. \text{ wt} = 2 \times 20 = 40$$

8. The e.m.f. of the cell  $Ni | Ni^{2+} (1M) || Cl^- (1M) | +Cl_2.Pt$  is  
( $E^0_{Ni^{2+}/Ni} = -0.25V$ ;  $E^0_{1/2Cl/Cl^-} = +1.36V$ )

(2007 E)

- 1) +1.11V                      2) -1.11 v                      3) +1.61 v                      4) -1.61 v

Ans : 3

Sol: Cell  $Ni | Ni^{2+} (1M) || Cl^- (1M) | +Cl_2.Pt$

$$EMF \text{ of the cell} = E_{\text{Cathode}} - E_{\text{Anode}}$$

Or

$$E = E_{\text{Right}} - E_{\text{left}}$$

$$= 1.36 - (-0.25)$$

$$= 1.36 + 0.25$$

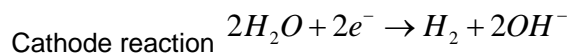
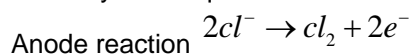
$$\text{EMF} = +1.61 \text{ V}$$

9. The pH of aqueous KCl solution is 7.0. This solution was electrolysed for a few seconds using Pt electrodes. Which of the following is correct? **(2006 M)**

- 1) The pH of solution decreases  
2) The pH of solution increases  
3)  $\text{Cl}_2$  is liberated at cathode  
4) The pH of solution remains the same

Ans : 2

Sol: Electrolysis of aqueous solution of HCl



$\therefore$  pH value increases.

10. What is the electrode potential (in V) of the following electrode at  $25^\circ\text{C}$ ?  
 $\text{Ni}^{2+} (0.1\text{M}) | \text{Ni}(s)$

$$\frac{2.303RT}{F} = 0.06$$

(Standard reduction potential of  $\text{Ni}^{2+} | \text{Ni}$  is  $-0.25 \text{ V}$ ; **(2006 M)**)

1)  $-0.25$                       2)  $-0.28$                       3)  $+0.25$                       4)  $-0.31$

Ans : 2

Sol: Nernst equation for metal electrode

$$E = E^0 + \frac{0.059}{n} \log [M^{n+}]$$

$$= -0.25 + \frac{0.6}{2} \log 0.1$$

$$= -0.25 - 0.3 = -0.28$$

11. **Assertion(A):** A current of 96.5 amperes is passed into aqueous  $\text{AgNO}_3$  solution for 100seconds. The weight of silver deposited is 10.8g (At.wt. of Ag=108).

**Reason (R):** The mass of a substance deposited during the electrolysis of an electrolyte is inversely proportional to the quantity of electricity passing through the electrolyte. **[2006 E]**

The correct answer is:

1. Both A and R are true and R is the correct explanation of A  
2. Both A and R are true and R is not the correct explanation of A  
3. A is true but R is not true  
4. A is not true but R is true

Ans : 3

$$\text{Sol: } m = \frac{108 \times 96.5 \times 100}{96500}$$

$$= 10.8 \text{ gm}$$

Reason ( R ) is wrong (not true)

12. What is the time (in sec) required for deposition all the silver present in 125ml of 1M  $\text{AgNO}_3$  solution by passing a current of 241.25 amperes? ( $1\text{F}=96500$  coulombs) **[2006 E]**

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

Ans : 2

Sol: Number of moles of  $\text{Ag}^+$  ion

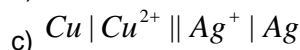
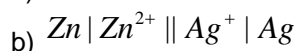
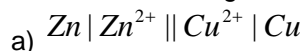
$$\text{Present in solution} = 125 \times 10^{-2} \times 1 = 0.125$$

$$\text{Number of Faradays} = 0.125$$

$$\text{Number of Coulombs} = 0.125 \times 96500$$

$$\text{Time required} = \frac{0.125 \times 96500}{241.25} = 50 \text{ sec}$$

13. The standard reduction potentials of  $Zn^{2+} | Zn, Cu^{2+} | Cu$  and  $Ag^+ | Ag$  are respectively -0.76, 0.34 and 0.8V. The following cells were constructed.



[E2006]

What is the correct order  $E^0$  cell of these cells?

- 1)  $b > c > a$       2)  $b > a > c$       3)  $a > b > c$       4)  $c > a > b$

Ans : 2

Sol: Cell (a)  $E^0 = 0.34 - 0.76 = +1.10 \text{ V}$

Cell (b)  $E^0 = 0.80 - 0.76 = +1.56 \text{ V}$

Cell (c)  $E^0 = 0.80 - 0.34 = +0.46 \text{ V}$

14. What is the electrochemical equivalent (in g coulomb<sup>-1</sup>) of silver?

(Ag = 108; F = Faraday)

[2005 M]

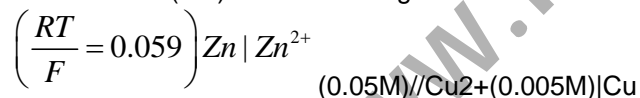
- 1) 108F      2)  $\frac{108}{F}$       3)  $\frac{F}{108}$       4)  $\frac{1}{108F}$

Ans : 2

Sol: Eq. Wt of Ag = 108

$$\text{e.c.e} = \frac{108}{F}$$

15. The standard reduction potentials of  $Zn^{2+} | Zn$  and  $Cu^{2+} | Cu$  are -0.76V and +0.34 V respectively. What is the cell e.m.f (inV) of the following cell?



[2005 M]

- 1) 1.1295      2) 1.0705      3) 1.1      4) 1.041

Ans : 2

Sol:  $E^0$  of  $Zn^{2+} | Zn = -0.76 \text{ V}$

$E^0$  of  $Cu^{2+} | Cu = +0.34 \text{ V}$

Nernst equation

$$E_{Zn} = E^0 + \frac{0.059}{n} \log c$$

$$E_{Zn} = -0.76 + \frac{0.059}{2} \log 0.05$$

$$= -0.76 + \frac{0.06}{2} \times (1,3)$$

$$= -0.76 - 0.039$$

$$E_{Zn} = -0.799V \quad E_{Cu} = E^0 + \frac{0.059}{n} \log c$$

$$E_{Cu} = 0.34 + \frac{0.06}{2} \log 0.005$$

$$= +0.34 + (0.03 \times (-2.3))$$

$$= +0.34 - 0.069$$

$$E_{Cu} = +0.271V$$

E of cell = E cathode – E anode

$$= +0.271 - (-0.799)$$

$$= +1.07 V$$

16. Which of the following is not correct?

[E2005]

- 1) Aqueous solution of NaCl is an electrolyte
- 2) The units of electrochemical equivalent are g. coulomb
- 3) In the Nernst equation, 'n' represents the number of electrons transferred in the electrode reaction
- 4) Standard reduction potential of hydrogen electrode is zero volts.

Ans : 2

Sol: Units of electrochemical equivalent are gram per coulomb

17. What is the quantity of electricity (in coulombs) required to deposit all the silver  $Ag = 108$  from 250ml of 1M AgNO<sub>3</sub> solution?

[E2005]

- 1) 2412.5
- 2) 24125
- 3) 4825.0
- 4) 48250

Ans : 2

Sol: Number of mole of AgNO<sub>3</sub> =  $\frac{250}{1000} \times 1M = \frac{1}{4}$  mole

1 mole of Ag<sup>+</sup> is deposited by 1 Faraday  
(96500 C) charge

$$\frac{1}{4} \text{ mole of Ag}^+ \text{ deposited by 1 Faraday}$$

$$\frac{96500 \times 1}{4} = 24125$$

$$= 4 \text{ coulombs}$$

18. The standard reduction potentials of Ag, Cu, Co and Zn 0.799,0.337,-0.277-0.762V respectively. Which of the following cells will have maximum cell e.m.f ?

[M 2004]

- 1)  $Zn | Zn^{2+} (1M) || Cu^{2+} (1M) | Cu$
- 2)  $Zn | Zn^{2+} (1M) || Ag^+ (1M) | Ag$
- 3)  $Cu | Cu^{2+} (1M) || Ag^+ (1M) | Ag$
- 4)  $Zn | Zn^{2+} (1M) || Co^{2+} (1M) | Co$

Ans : 2

Sol: Electrode with most negative reduction potential acts as anode ( LHS) and the electrode with most positive reduction potential acts as cathode (RHS).

Then the emf is highest.

19. The electrochemical equivalent of a metal is 'x' gram-coulomb<sup>-1</sup>. The equivalent weight of metal is

(2004 E)

- 1) x
- 2)  $x \times 96500$
- 3)  $\frac{x}{96500}$
- 4)  $1.6 \times 10^{-19} \times x$

Ans : 2

Sol:  $e.c.e = \frac{Eq. wt}{96500} = x$   
Eq. Wt = 96500 × x

20. The Cell reaction of the galvanic cell :  $Cu(s) / Cu^{2+}(aq) || Hg^2+(aq) | Hg(l)$  is

(2003 M)

- 1)  $Hg + Cu^{2+} \rightarrow Hg^{2+} + Cu$
- 2)  $Hg + Cu^{2+} \rightarrow Hg^{2+} + Cu^+$
- 3)  $Cu + Hg \rightarrow CuHg$
- 4)  $Cu + Hg^{2+} \rightarrow Cu^{2+} + Hg$

Ans : 4

Sol: Cell reaction is  $Cu(s) + Hg^{2+} \rightarrow Cu^{2+} + Hg$ 21. If the standard electrode potential of  $Cu^{2+}/Cu$  electrode is 0.34 V, what is the electrode potential at 0.01M concentration of  $Cu^{2+}$ ? (T=298 K) **(2003 M)**

- 1) 0.399 V                      2) 0.281 V                      3) 0.222 V                      4) 0.176 V

Ans : 2

Sol:

$$E = E^0 + \frac{0.059}{n} \log [M^{n+}]$$

$$= 0.34 + \frac{0.059}{2} \log [0.01]$$

$$= 0.34 + \frac{0.059}{2} \times (-2)$$

$$= 0.34 - 0.059$$

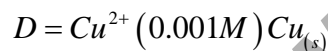
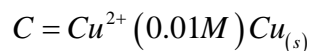
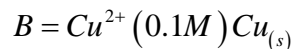
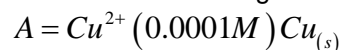
$$E = 0.281 \text{ V}$$
22. When X amperes of current is passed through molten  $AlCl_3$  for 96.5 seconds. 0.09 grams of aluminium is deposited. What is the value of X? **(2003 M)**

- 1) 10                      2) 20                      3) 30                      4) 40

Ans : 1

Sol:

$$m = \frac{Ect}{96500}$$

$$C = \frac{96500 \times 0.09}{9 \times 96.5} = 10 \text{ amp}$$
23. Consider the following four electrodes **(2002 M)**If the standard reduction potential of  $Cu^{+2} / Cu$  is +0.34 V, the reduction potentials ( in volts) of the above electrodes follow the order

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

Ans : 2

Sol: According Nernst equation

$$E = E^0 + \frac{0.059}{n} \log [M^{n+}]$$

Log  $[Cu^{2+}]$  values, A = log (0.0001) = -4

B = log (0.1) = -1

C = log (0.01) = -2

D = log (0.001) = -3

 $\therefore$  E values are in the order

B &gt; C &gt; D &gt; A

24. 0.066 gram of metal was deposited when a current of 2 amperes is passed through a metal ion solution for 100 seconds. What is the electrochemical equivalent ( in gram coulomb<sup>-1</sup>) of the metal? **(2002 E)**

- 1)
- $3.3 \times 10^{-6}$
- 2)
- $3.3 \times 10^{-4}$
- 3) 0.033                      4) 3.3

Ans : 2

$$\text{Sol: } e = \frac{m}{ct} = \frac{0.066}{2 \times 100}$$

$$= 3.3 \times 10^{-4} \text{ gm coulomb}^{-1}$$

25. What is the reduction electrode potential (in volts) of copper electrode when  $[\text{Cu}^{2+}] = 0.01\text{M}$  in a solution at 250C ? ( $E^0$  of  $\text{Cu}^{2+}/\text{Cu}$  electrode is +0.34V) **[2002 E]**
- 1) 0.3991                      2) 0.2809                      3) 0.3105                      4) 0.3695

Ans : 2

$$\text{Sol: } E = E^0 + \frac{0.059}{n} \log [M^{n+}]$$

$$= +0.34 + \frac{0.059}{2} \log [0.01]$$

$$= +0.34 + \frac{0.059}{2} \times [-2]$$

$$= +0.34 - 0.059 = 0.281 \text{ V}$$

26. One ampere of current is passed for 9650 seconds through molten  $\text{AlCl}_3$ . What is the weight in grams of Al deposited at cathode? (Atomic weight of Al=27) **(2001 E)**
- 1) 0.9                              2) 9.0                              3) 0.09                              4) 90.0

Ans : 1

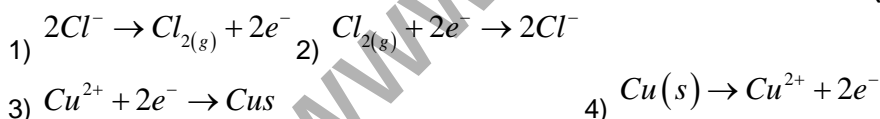
$$\text{Sol: } e = \frac{27}{3 \times 96500}$$

$c = 1$  ampere                       $t = 9650$  sec

weight of Al deposited = e.c.t

$$= \frac{27 \times 1 \times 9650}{3 \times 96500} = 0.9 \text{ gm}$$

27. Molten  $\text{CuCl}_2$  is electrolysed using platinum electrodes. The reaction occurring at anode is ..... **(2001 E)**



Ans : 1

Sol: At anode chloride ion are oxidized to chlorine gas

$$2\text{Cl}^- \rightarrow \text{Cl}_2 + 2e^-$$

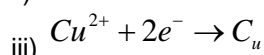
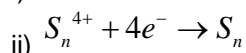
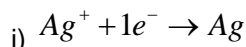
28. One faraday of electricity is passed separately through one litre of one molar aqueous solution of i)  $\text{AgNO}_3$ , ii)  $\text{SnCl}_4$  and iii)  $\text{CuSO}_4$ . The number of moles of Ag, Sn and Cu deposited at cathode are respectively

**[2001 M]**

- 1) 1.0, 0.25, 0.5                      2) 1.0, 0.5, 0.25                      3) 0.5, 1.0, 0.25                      4) 0.25, 0.5, 1.0

Ans : 1

Sol: Electrode reaction :



For 1 Faraday of electricity number of moles of ion deposited  $\alpha$

$$= \frac{1}{\text{number of electron required for 1 mole}}$$

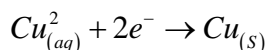
$$\therefore 1 : 0.25 : 0.5$$

29. Aqueous copper sulphate solution is electrolysed using platinum electrodes. The electrode reaction occurring at cathode is **[2001 M]**

- 1)  $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$
- 2)  $\text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$
- 3)  $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$
- 4)  $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$

Ans : 1

Sol: At cathode reduction reaction takes place (gain of electron)



30. What is the approximate quantity of electricity (in coulombs) required to deposit all the silver from 250ml of 1M  $\text{AgNO}_3$  aqueous solution ? (At. Wt. of Ag = 108) **(2000 E)**

- 1) 96500
- 2) 24125
- 3) 48250
- 4) 12062.5

Ans : 2

Sol: Amount of Ag present in 250 ml 2 M  $\text{AgNO}_3$  solution is

$$W = 1 \times 170 \times \frac{1000}{250} = 170 \times 4 = 680 \text{ gm}$$

According to Faradays First Law

$$W = \frac{\text{Eq. wt}}{96500} \times Q$$

$$Q = \frac{380 \times 96500}{180} = 24125 \text{ coulombs}$$

31. Which of the following aqueous solutions conducts electricity ? **(2000 M)**

- 1) urea
- 2) glucose
- 3) sucrose
- 4) NaCl

Ans : 4

Sol: NaCl aqueous solution carry  $\text{Na}^+$  and  $\text{Cl}^-$  ions.

