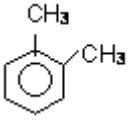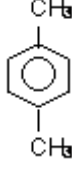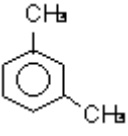1. The resistance of 1 N solution of acetic acid is 250 Ohm, when measured in a cell of cell constant $1.15 cm^{-1}$. The equivalent conductance (in $Ohm^{-1} cm^2 equiv^{-1}$) of 1 N acetic acid is.
   (a) 4.6   (b) 9.2   (c) 18.4   (d) 0.023

2. Give that at $25^o C$, for
   $Cr^{3+}(aq) + e^- \rightarrow Cr^{2+}(aq), E^0 = -0.424 V$
   $Cr^{2+}(aq) + 2e^- \rightarrow Cr(s), E^0 = -0.900 V$

   find E^0 at $25^0 C$ for $Cr^{3+}(aq) + 3e \rightarrow Cr(s)$
   (a) – 0.741V   (b) + 0.741 V
   (c) – 1.324 V   (d) – 0.476 V

3. Terepthalic acid is obtained by oxidation of which of following compounds.

   (a)   (b)

   (c)   (d)

4. The most basic compound among the following is
   (a) Benzylamine   (b) Aniline
   (c) Acetanilide   (d) P-nitroaniline

5. A solution containing one mole per litre of each $Cu(NO_3)_2$; $AgNO_3$; $Hg_2(NO_3)_2$; $Mg(NO_3)_2$ is being electrolysed using inert electrodes. The values of standard electrode potentials (reduction potentials in volts are Ag/Ag^+ = 0.80 V, $2Hg/Hg_2^{++}$ = 0.79 V, Cu/Cu^{++} = + 0.24 V , Mg/Mg^{++} - 2.37 V. With increasing volt age, the sequence of deposition of metals on the cathode will be
   (a) Ag, Hg, Cu   (b) Cu, Hg, Ag
   (c) Ag, Hg, Cu, Mg   (d) Mg, Cu, Hg, Ag

6. When a lead storage battery is discharged
   (a) SO_2 is evolved   (b) Lead sulphate is consumed
   (c) Lead is formed   (d) H_2SO_4 is consumed

7. 58. Cu^+ is not stable and undergoes disproportionation .E^o for Cu^+ disproportionation.....)
   $(E^0_{cu^{+2}/Cu^+} = +0.153V, E^0_{Cu^+/Cu} = 0.53V)$
   (a) + 0.683 V   (b) - 0. 367   (c) + 0. 3415 V   (d) + 0.367

8. The minimum equivalent conductance in fused state is shown by -
   (a) $MgCl_2$ (b) $BeCl_2$ (c) $CrCl_2$ (d) $SrCl_2$

9. In which of the following electrochemical cell overall cell reaction is :
   $HgO (s) + H_2 (g) \longrightarrow Hg (\ell) + H_2O (\ell)$
   (a) Pt | H_2 (g) | H^+ (aq) | HgO (s) | Hg (ℓ) | Pt
   (b) Pt | H_2 (g) | NaOH (aq) | HgO (s) | Hg (ℓ) |Pt
   (c) Pt | H_2 (g)|H^+||NaOH (aq)|HgO (s) | Hg (ℓ) | Pt
   (d) Pt | H_2 (g) | H^+ || HgO (s) | Hg (ℓ) | Pt

10. What would be the product of electrolysis if molten ICl_3 is electrolysed ?
    (a) I_2 is produced at cathode and Cl_2 is produced at anode
    (b) Cl_2 is produced at cathode and I_2 is produced at anode
    (c) Both I_2 and Cl_2 are liberated at both electrodes
    (d) ICl_2 is produced at cathode and ICl_4 is produced at anode

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11. Molar conductivity is defined as (G is the conductance)
(a) κ/V (b) κV (c) $G\,\ell^2$ (d) $R\ell^2$

12. At 298 K , the conductivity of a saturated solution of AgCl in water is 2.6×10^{-6} ohm^{-1}cm^{-1}. Given , λ_m^∞ (Ag$^+$) = 63 ohm^{-1} cm^{2}mol^{-1} & λ_m^∞ (Cl$^-$) = 67 ohm^{-1} cm^{2}mol^{-1}
Therefore solubility product of AgCl is
(a) 2×10^{-5} (b) 4×10^{-10}
(c) 4×10^{-16} (d) 2×10^{-8}

13. Which of the following would occur when lead storage cell is charged ?
(a) Sulphuric acid is consumed   (b) Sulphuric acid is formed
(c) Leadsulphate is formed   (d) Lead is consumed

14. During electrolysis, the gm of product produced (m) and when plotted against the electric charge Q$'$ in faradays which of the following plot is correct ?
(a)

(b)

(c)

(d)

15. The standard potentials at 298 K for the following half reactions are as given
$Zn^{2+} + 2e \rightleftharpoons Zn$ E° = –0.762 V

$2H^+ + 2e \rightarrow H_2 (g)$ E° = 0.000 V

$Cr^{3+} + 3e \rightleftharpoons Cr$ E° = – 0.740 V

$Fe^{3+} + 2e \rightarrow Fe^{2+}$ E° = 0.772 V

Which of the following is the strongest reducing agent ?
(a) Zn (s)   (b) Cr   (c) H_2 (g)   (d) Fe^{2+} (aq)

16. The thermodynamic efficiency of cell is given by –

(a) $\Delta H/\Delta G$ (b) $nFE/\Delta G$ (c) $nFE/\Delta H$ (d) Nfe^0

17. The equivalent conductivity of 0.1 M weak acid is 100 times less than that at infinite dilution. The degree of dissociation of weak electrolyte at 0.1 M is –
(a) 100   (b) 10   (c) 0.01   (d) 0.001

18. For the following cell with hydrogen electrodes at two different pressure p_1 and p

$$Pt\,|\,H_2(g)\,|\,H^+(aq)\,|\,H_2\,(g)\,|\,Pt$$
$$\quad p_1 \qquad 1M \qquad p_2$$
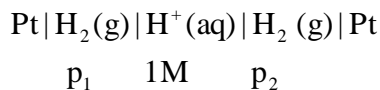
emf is given by : $\quad Pt\,|\,H_2(g)\,|\,H^+(aq)\,|\,H_2\,(g)\,|\,Pt$
$$\qquad\qquad\qquad p_1 \qquad 1M \qquad p_2$$

(a) $\dfrac{RT}{F}\log_e\dfrac{p_1}{p_2}$ (b) $\dfrac{RT}{2F}\log_e\dfrac{p_1}{p_2}$

(c) $\dfrac{RT}{F}\log e\dfrac{p_2}{p_1}$ (d) $\dfrac{RT}{2F}\log_e\dfrac{p_2}{p_1}$

19. Conductance (with unit Siemens S) is directly proportional to area of the electrode plates and the concentration of the solution in the cell and is inversely proportional to the separation between the electrode plates. Then the unit of the constant of proportionality is
(a) $Smmol^{-1}$ (b) Sm^2mol^{-1}
(c) $S^{-2}m^2mol$ (d) $S^2m^2mol^{-1}$.

20. The cell
$Zn\,|\,Zn^{2+}(1M)\,\|\,Cu^{2+}(1M)\,|\,Cu:(E^0_{cell}=1.10V)$ was allowed to completely discharge at 298 K. The relative concentration of Zn^{2+} to $Cu^{2+}\left(\dfrac{[Zn^{2+}]}{[Cu^{2+}]}\right)$ is :

(Take $\dfrac{1.1}{0.059}=18.65$)

(a) $10^{37.3}$ (b) 9.65×10^4 (c) Antilog (24.08)   (d) 37.3

21. Given : $E^0_{Fe^{3+}/Fe=-0.036}$ V, $E^0_{Fe^{2+}/Fe}=-0.439$ V   The value of standard electrode potential for the change, $+\,e^- \longrightarrow$ will be :
(a) 0.385V   (b) 0.770V   (c) – 0.270V   (d) – 0.072V

22. The Gibbs energy foar the decomposition of Al_2O_3 at 500ºC is as follows :
$\dfrac{2}{3}Al_2O_3 \rightarrow \dfrac{4}{3}Al + O_2; \Delta_rG = +966\ kJmol^{-1}$. The potential difference needed for electrolytic reduction of Al_2O_3 at 500^0C is at least :
(a) 4.5 V   (b) 3.0 V   (c) 2.5 V   (d) 5.0 V

23. The standard reduction potentials for Zn^{2+}/Zn, Ni^{2+}/Ni and Fe^{2+}/Fe are –0.76, – 0.23 and –0.44 V respectively. The reaction $X + Y^{2+} \rightarrow X^{2+} + Y$ will be spontaneous, when :
(a) X = Ni, Y = Fe   (b) X = Ni, Y = Zn
(c) X= Fe, Y = Zn   (d) X= Zn, Y = Ni

24. Given :
$E^0_{Cr^{3+}/Cr}=-0.74V;\ E^0_{MnO_4^-/Mn^{2+}}=1.51$ V

$E^0_{Cr_2O_7^{2-}/Cr^{3+}}=1.33$ V; $E^0_{Cl/Cl^-}=1.36V$ Based on the

data given above, strongest oxidising agent will be :
(a) Cl   (b) Cr^{3+} (c) Mn^{2+} (d) MnO_4^-

25. E^0 for $F_2 + 2e^- \rightarrow 2F^-$ is 2.8 V, E^0 for $\frac{1}{2}F_2 + e^- \rightarrow F^-$ is

    (a) 2.8 V   (b) 1.4 V   (c) – 2.8 V   (d) – 1.4 V

26. The emf of the cell,

   $Ni \mid Ni^{2+} (1.0\,M) \parallel Ag^+ (1.0\,M) \mid Ag\ [E^0$ for $Ni^{2+}/Ni = -0.25$

   volt, E° for $Ag^+/Ag = 0.80$ volt] is given by
   (a) –0.25 + 0.80 = 0.55 volt
   (b) –0.25 – (+0.80) = –1.05 volt
   (c) 0 + 0.80 – (–0.25) = + 1.05 volt
   (d) –0.80 – (–0.25) = – 0.55 vol

27. The electrode potential of electrode

   $M(s) \rightarrow M^{n+}(aq)(2M) + ne^-$. at 298 K is E_1. When temperature is doubled and concentration is made half, then the electrode potential becomes E_2. Which of the following represents the correct relationship between E_1 and E_2 ?

   (a) $E_1 > E_2$ (b) $E_1 < E_2$

   (c) $E_1 = E_2$ (d) Cann't be predicted

28. Using the standard potential values given below, decide which of the statements I, II, III, IV are correct.
   Choose the right answer from (a),(b),(c) and (d)

   $Fe^{2+} + 2e^- = Fe,$ $E^0 = -0.44\,V$

   $Cu^{2+} + 2e^- = Cu,$ $E^0 = +0.34\,V$

   $Ag^+ + e^- = Ag,$ $E^0 = +0.80\,V$

   I. Copper can
   displace iron from $FeSO_4$ solution

   II. Iron can displace copper from $CuSO_4$ solution

   III. Silver can displace Cu from $CuSO_4$ solution

   IV. Iron can displace silver from $AgNO_3$ solution

   (a) I and II   (b) II and III
   (c) II and IV   (d) I and IV

29. Four moles of electrons were transferred from anode to cathode in an experiment on electrolysis of water. The total volume of the two gases (dry and at STP) produced will be approximately (in litres)
   (a) 22.4   (b) 44.8   (c) 67.2   (d) 89.4

30. A cell $Cu \mid Cu^{++} \parallel Ag^+ \mid Ag$ initially contains 2M Ag^+ and 2M Cu^{++} ions in 1L electrolyte. The change in cell potential after the passage of 10 amp current for 4825 sec is :
   (a) – 0.00738 V   (b) – 1.00738 V
   (c) – 0.0038 V   (d) None

1. (a)

$\kappa = G \times \text{cell constant} \quad = \dfrac{\text{Cell constant}}{R} = \dfrac{1.15}{250}$

$\Lambda_{eq} = \dfrac{\kappa \times 1000}{\text{Normality}} = \dfrac{1.15}{250} \times \dfrac{1000}{1} = 4.6\,\Omega^{-1}cm^2 eq^{-1}.$

2. (a)

Adding the first two reactions, we get the third equation and using the free energy concept, we have

$\Delta G_1^0 + \Delta G_2^0 = \Delta G_3^0$; $-n_1 E_1^0 F - n_2 E_2^0 F = -n_3 E_3^0 F$

(n = number of electron involved)

$E_3^0 = \dfrac{n_1 E_1^0 + n_2 E_2^0}{n_3} = \dfrac{1 \times (-0.424) + 2 \times (-0.900)}{3} = -0.741\,V$

3. (c)

Terepthalic acid is a para dicarboxylic acid

4. (a)

Rest all show less tendency to donate electron pair due to resonance.

5. (a)

Greater the value of standard reduction potential, greater will be it's tendency to undergo reduction . So the sequence of deposition of metals on cathdoe will be Ag, Hg, Cu. Here magnesium will not be deposited because its' standard reduction potential is negative . So it is strong tandency to undergo oxidation. Therefore, on electrolysis of $Mg(NO_3)_2$ solution, H_2 gas will be evolved at cathode.

6. (d)

At anode :

$Pb + H_2SO_4SO_4 \rightarrow PbSO_4 + 2H_2O$

At cathode:

$PbO_2 + H_2SO_4 + 2H^+ + 2e^- \rightarrow PbSO_4 + 2H_2O$

Net dishcharging equation

So, H_2SO_4 is consumed during discharging of lead storage battery

7. (d)

$2Cu^+ \rightarrow Cu^{+2} + Cu$

$E^0{}_{cell} = E^0{}_{cu^+/Cu^{+2}} + E^0{}_{Cu^+/Cu}$

$= -0.153 + 0.53 = +0.367V$

8. (b)

$BeCl_2$ has higher covalent character hence it ionizes in least extent in the fused state

9. (b)

Pt | H_2 (g) | NaOH (aq) | HgO (s) | Hg (ℓ) | Pt

LHE reaction : H_2 (g) + $2OH^- \longrightarrow 2H_2O$ (ℓ) + 2e

RHE reaction : HgO (s) + H_2O (ℓ) + 2e $\longrightarrow$ Hg (ℓ) + $2OH^-$

Net cell reaction : HgO (s) + H_2 (g) $\longrightarrow$ Hg (ℓ) + H_2O

10. (c)

In the molten state ICl_3 ionises as follows

$2ICl_3 \rightleftharpoons ICl_2^+ + ICl_4^-$

Hence both I_2 and Cl_2 are produced at both the electrodes.

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11. (c)

Molar conductivity is defined as $\pi_m = G \, \ell^2$

12. (b)

$$s = \frac{10^3 \times \kappa}{\lambda_m^\infty (\text{AgCl})} = \frac{10^3 \times 2.6 \times 10^{-6}}{(63+67)} = 2 \times 10^{-5} \, (M)$$

$\therefore \, K_{sp} = (2 \times 10^{-5})^2 = 4 \times 10^{-10}$.

13. (b)

The charging reaction in the lead storage battery is

$2PbSO_4 + 2H_2O \longrightarrow Pb + PbO_2 + 2H_2SO_4$.

14. (b)

$m = \dfrac{EQ}{F}$; $\dfrac{Q}{F}$ represent the electric charge Q'

$\therefore \qquad m = EQ'$

15. (a)

More negative the standard potential, least the reduction tendency of the ion. The corresponding atom has largest oxidation tendency and thus is a strong reducing agent. Zn is the strongest reducing agent.

16. (c)

Thermodynamic efficiency = $\dfrac{nFE}{\Delta H}$, i.e., the part of enthalpy change during the course of redox reaction, given out as electrical energy.

17. (c)

$$\Lambda_v = \frac{\Lambda^0}{100}$$

$$\therefore \quad \alpha = \frac{\Lambda_v}{\Lambda^0} = \frac{\Lambda^0}{100\Lambda^0} = 0.01$$

18. (b)

The E^0 of cell will be zero

19. (b)

$$K = \frac{1}{\rho} = \frac{1}{R}\frac{\ell}{A}$$

20. (a)

$$0 = +1.1 - \frac{0.0591}{2}\log\left[\frac{Zn^{2+}}{Cu^{2+}}\right]; \log\left[\frac{Zn^{2+}}{Cu^{2+}}\right] = 37.3.; \left[\frac{Zn^{2+}}{Cu^{2+}}\right] = 10^{37.3}$$

21. (b)

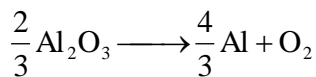$Fe^{3+} + 3e^- \longrightarrow Fe \qquad \Delta G_1 = -3 \times F \times E^0_{Fe^{3+}/Fe} \quad Fe^{2+} \longrightarrow Fe \qquad\qquad \Delta G_2 = -2 \times F \times E^0_{Fe^{2+}/Fe}$

$Fe^{3+} + e^- \longrightarrow Fe^{2+} \;\; \Delta G = \Delta G_1 - \Delta G_2 \quad \Delta G = 3 \times 0.036F - 2 \times 0.439 \times F = -1 \times E^0 \,(Fe^{3+}/Fe^{+2}) \times F$

$E^0 \,(Fe^{3+}/Fe^{+2}) \;\; = 2 \times 0.439 - 3 \times 0.036$

$= 0.878 - 0.108$

$= 0.770 \, V$

22. (c)

$$\frac{2}{3}Al_2O_3 \longrightarrow \frac{4}{3}Al + O_2$$

$$\Delta G = +966\,kJ\,mol^{-1} = 966\times10^3\,J\,mol^{-1}$$

$$\Delta G = -nFE_{cell} \quad 966\times10^3 = -4\times96500\times E_{cell} \quad E_{cell} = 2.5\,V$$

23. (d)

$X + Y^{2+} \longrightarrow X^{2+} + Y$ For reaction to be spontaneous E° must be positive.

$$E^0_{Zn/Zn+2} + E^0_{Ni2+/Ni} = 0.76 + (-0.23) = +0.53 (positive)$$

24. (d)

Higher the SRP, better is oxidising agent

Hence MnO_4^- is strongest oxidising agent.

25. (a)

E^0 is intensive property and it does not depend on mass of F_2 taking part.

26. (c)

$$E_{cell} = E^0_{Ni/N2+} + = 0.25 + 0.80 = 1.05\,Volt$$

27. (b)

$$E_1 = E^0 - \frac{RT}{nF}\ln 2$$

$$E_2 = E^0 - \frac{R\times 2T}{nF}\ln 1 = E^0$$

$$\therefore E_2 > E_1$$

28. (c)

Lower standard reduction potential related metal ions can displace higher standard reduction potential related metal ions.

29. (c)

At Cathode : $2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$]×2 At Anode : $2H_2O \longrightarrow O_2 + 4H^+ + 4e^-$ $H_2 = 2$ mole

$O_2 = 1$ mole

Total volume = 3×22.4 **= 67.2 L.**

30. (a)

$Q = 10 \times 4825 = 48250\,C$

no. of faraday = $\frac{48250}{96500} = 0.5$

$$Ag + \frac{1}{2}Cu^{++} \longrightarrow Ag^+$$

$$+ \frac{1}{2}Cu$$
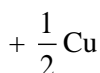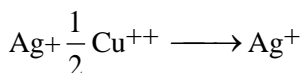
2.0

2.00

20.250

$$E_{cell} = E^0_{cell} \frac{0.0591}{1}\log\frac{[Ag^+]}{[Cu^{++}]^{1/2}}$$

$$E_1 = E^0_{Cell} - \frac{0.0591}{1} \log \log \frac{2.00}{(2.00)^{1/2}}$$

$$E_2 = E^0_{cell} - \frac{0.0591}{1} \log \frac{2.50}{(1.75)^{1/2}}$$

$$\Delta E = E_2 - E_1 = \frac{0.0591}{1} \left[\log \sqrt{2} - \log \frac{2.50}{\sqrt{1.75}} \right] = \frac{0.0591}{1} [\log 1.41 - \log 1.88]$$

$$= \frac{0.0591}{1} [0.1492 - 0.2742] = -\frac{0.0591}{1} \times 0.125 = -0.00738 \, V.$$