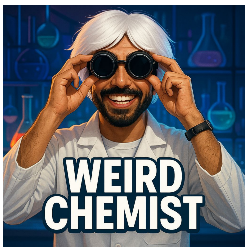

DPP –5 [Nernst Equation -2]

Chapter: Electrochemistry

“You are not preparing for NEET/JEE. You are becoming someone who solves difficult problems calmly.

TYPE–1 : Thermodynamics – ΔG, K_{eq}, E° Relations

1. **For a cell reaction involving a two-electron change, the standard e.m.f. of the cell is found to be 0.295 V at 25° C. The equilibrium constant of the reaction at 25 °C will be :**
   (1) 10 (2) 1×10^{10}
   (3) 1×10^{-10} (4) 29.5×10^{-2}
2. **Standard EMF of a galvanic cell involving $2e^-$ transfer is found to be 0.591 V at 25°C. Calculate equilibrium constant for the cell reaction**
   (1) 20^{10} (2) 10^{20}
   (3) 10^{200} (4) 200^{10}
3. **The equilibrium constant of the reaction :**
   $Cu(s) + 2Ag^+(aq.) \longrightarrow Cu^{+2}(aq.) + 2Ag(s)$
   E° = 0.46 V at 298 K is :–
   (1) 2.0×10^{10} (2) 4.0×10^{10}
   (3) 4.0×10^{15} (4) 2.4×10^{10}
4. **E° for the cell, $Zn|Zn^{2+}(aq)||Cu^{2+}(aq)|Cu$ is 1.10 V at 25°C, the equilibrium constant for the reaction $Zn + Cu^{2+}(aq) \rightleftharpoons Cu + Zn^{2+}(aq)$ is the order of**
   (1) 10^{+18} (2) 10^{+37}
   (3) 10^{-28} (4) 10^{-37}
5. **For a cell involving one electron $E^\circ_{cell} = 0.59$ V at 298 K, the equilibrium constant for the cell reaction is :**
   Given that $\frac{2.303RT}{F} = 0.059$ V at T = 298 K
   [NEET-2019]
   (1) 1.0×10^{2} (2) 1.0×10^{5}
   (3) 1.0×10^{10} (4) 1.0×10^{30}
6. **For a reaction - $A(s) + 2B^+ \rightarrow A^{2+} + 2B(s)$**
   K_c has been found to be 10^{12}. The E°_{eq} is:
   (1) 0.354 V (2) 0.708 V
   (3) 0.0098 V (4) 1.36 V
7. **For given reaction**
   $2B^+_{(aq)} + A_{(s)} \rightleftharpoons 2B_{(s)} + A^{+2}$
   If Keq of reaction is $10^{15.2}$
   and $E^\circ(A^{+2}/A) = 0.34$V

Then calculate $E^\circ(B^+/B)$ will be :–

(1) -0.80 (2) 0.80

(3) 1.27 (4) 1.60

8. **ΔG° of the cell reaction**

$AgCl(s) + \frac{1}{2}H_2(g) \rightarrow Ag(s) + H^+ + Cl^-$ is -21.52 KJ

ΔG° of $2AgCl(s) + H_2(g) \rightarrow 2Ag(s) + 2H^+ + 2Cl^-$ is :

(1) -21.52 kJ (2) -10.76 kJ

(3) -43.04 kJ (4) 43.04 kJ

9. **For the reduction of silver ions with copper metals, the standard cell potential was found to be $+0.46$ V at 25 °C. The value of standard Gibbs energy, ΔG° will be ($F = 96500\ C\ mol^{-1}$)**

(1) -98.0 kJ (2) -89.0 kJ

(3) -89.0 J (4) -44.5 kJ

10. **For the cell reaction**

$2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$

$E^\circ_{cell} = 0.24$ V at 298 K. The standard Gibbs energy ($\Delta_r G^\circ$) of the cell reaction is :

[Given that Faraday constant $F = 96500\ C\ mol^{-1}$]

[NEET-2019]

(1) $-46.32\ kJ\ mol^{-1}$ (2) $-23.16\ kJ\ mol^{-1}$

(3) $46.32\ kJ\ mol^{-1}$ (4) $23.16\ kJ\ mol^{-1}$

11. **Calculate ΔG° for the galvanic cell:**

$2Cr\ (s) + 3Cd^{+2}\ (aq) \longrightarrow 2Cr^{+3}\ (aq) + 3Cd\ (s)$

$E^\circ_{Cr^{+3}/Cr} = -0.74$ V, $E^\circ_{Cd^{+2}/Cd} = -0.40$ V

(1) -196.86 kJ/mol (2) -98.43 kJ/mol

(3) -65.62 kJ/mol (4) -164.05 kJ/mol

12. **For the cell reaction,**

$Cu^{2+}\ (C_1 \cdot aq) + Zn(s) = Zn^{2+}\ (C_2 \cdot aq) + Cu(s)$ of an electrochemical cell, the change in Free Energy (ΔG) at a given temperature is a function of

(1) $\ln (C_2)$ (2) $\ln (C_2/C_1)$

(3) $\ln (C_1)$ (4) $\ln (C_1 + C_2)$

13. **$Zn \mid Zn^{2+}_{(C1)} \parallel Zn^{2+}_{(C2)} \mid Zn\ (s)$. The ΔG is $-$ve if**

(1) $C_1 = C_2$ (2) $C_1 > C_2$

(3) $C_2 > C_1$ (4) Can't predicted

14. **E° for the reaction $Fe + Zn^{2+} \rightarrow Zn + Fe^{2+}$ is -0.35 V. The given cell reaction is :**

(1) feasible (2) not feasible

(3) in equilibrium (4) can't say anything

15. **If the E°_{cell} for a given reaction has a negative value, then which of the following gives the correct relationship for the values of ΔG° and K_{eq} ?**

(1) $\Delta G^\circ > 0;\ K_{eq} > 1$

(2) $\Delta G^\circ < 0;\ K_{eq} > 1$

(3) $\Delta G^\circ < 0;\ K_{eq} < 1$

(4) $\Delta G^\circ > 0;\ K_{eq} < 1$

16. **For a spontaneous reaction the ΔG, equilibrium constant (K) and E°_{cell} will be respectively**

(1) $-$ve, < 1, $-$ve

(2) $-$ve, >1, $-$ve

(3) $-$ve, >1, $+$ve

(4) $+$ve, >1, $-$ve

17. **The EMF of a chemical cell is positive when free energy change of reaction**

(1) > 0

(2) < 0

(3) $= 0$

(4) No relationship of free energy change and e.m.f.

18. **Cell reaction is spontaneous when**

(1) ΔG° is negative

(2) ΔG° is positive

(3) ΔE°_{Red} is positive

(4) ΔE°_{Red} is negative

19. **On the basis of the information available from the reaction $\frac{4}{3}Al + O_2 \rightarrow \frac{2}{3}Al_2O_3$, $\Delta G = -827$ KJ mol^{-1} of O_2, the minimum e.m.f. required to carry out an electrolysis of Al_2O_3 is :–**

(F = 96500 C mol^{-1})

(1) 2.14 V (2) 4.28 V

(3) 6.42 V (4) 8.56 V

20. **Standard free energies of formation (in kJ/mol) at 298 K are -237.2, -394.4 and -8.2 for H_2O (l), CO_2 (g) and pentane (g) respectively. The value of E°_{cell} for the pentane-oxygen fuel cell is**

[AIPMT (Prelims)-2008]

(1) 0.0968 V (2) 1.968 V

(3) 2.0968 V (4) 1.0968 V

21. **Consider the following equations for a cell**

$A + B \rightleftharpoons C + D$; $E^\circ = x$ volt , $K_{eq} = K_1$

$2A + 2B \rightleftharpoons 2C + 2D$; $E^\circ = y$ volt , $K_{eq} = K_2$

Then

(1) $x = y$, $K_1 = K_2$ (2) $x = 2y$, $K_1^2 = K_2$

(3) $x = y$, $K_1^2 = K_2$ (4) $x^2 = y$, $K_1^2 = K_2$

22. **Consider the following equations for a cell reaction:**

$A + B \longrightarrow C + D$; $E^\circ = x$ volt, $\Delta G = \Delta G_1$

$2A + 2B \longrightarrow 2C + 2D$; $E^\circ = y$ volt, $\Delta G = \Delta G_2$

Then,

(1) $x = y$, $\Delta G_1 = \Delta G_2$ (2) $x > y$, $\Delta G_1 > \Delta G_2$

(3) $x = y$, $2\Delta G_2 = \Delta G_1$ (4) $x < y$, $2\Delta G_2 = \Delta G_1$

23. **The standard electrode potentials $E^\circ_{I_2/I^-}$, $E^\circ_{Br_2/Br^-}$ and $E^\circ_{Fe^{2+}/Fe^{+3}}$ are respectively $+0.54$ V, -1.09 V and 0.44 V. On the basis of above data which of the following process is nonspontaneous?**

(1) $Br_2 + 2I^- \longrightarrow 2Br^- + I_2$

(2) $Fe + Br_2 \longrightarrow Fe^{2+} + 2Br^-$

(3) $Fe + I_2 \longrightarrow Fe^{2+} + 2I^-$
(4) $I_2 + 2Br^- \longrightarrow 2I^- + Br_2$

24. **A solution contains Fe^{2+}, Fe^{3+} and I^- ions. This solution was treated with iodine at 35 °C. E° for Fe^{3+}/Fe^{2+} is $+0.77$ V and E° for $I_2/2I^- = 0.536$ V. The favourable redox reaction is**
(1) I^- will be oxidised to I_2
(2) Fe^{2+} will be oxidised to Fe^{3+}
(3) I_2 will be reduced to I^-
(4) There will be no redox reaction

TYPE–2 : Concentration Cell – & Multiple oxidation problem

25. **E value for the cell reaction**
$Cu \mid Cu^{2+}$ (0.001 M) | | Cu^{2+}(0.1 M) | Cu is :–
(1) $-\dfrac{RT}{F}\ln(0.01)$ (2) $-\dfrac{RT}{F}\ln(0.1)$
(3) $+\dfrac{RT}{F}\ln(0.01)$ (4) $\dfrac{RT}{F}\ln(0.1)$

26. **The emf of the cell reaction**
$Ag \mid Ag^+(0.1M) \| Ag^+(1M) \mid Ag$ at 298 K is
(1) 0.0059 V (2) −0.059 V
(3) 5.9 V (4) 0.59 V

27. **What is the potential of the cell containing two hydrogen electrodes as represented below**
Pt, $H_2(g) \mid H^+(10^{-8})M \| H^-(0.001\ M) \mid H_2(g)$,Pt ;
(1) −0.295 V (2) −0.0591 V
(3) 0.295 V (4) 0.0591 V

28. **If hydrogen electrode is dipped in two solutions of pH = 3 and pH = 6 and salt bridge is connected, the emf of resulting cell is**
(1) 0.177 V (2) 0.3 V
(3) 0.052 V (4) 0.104 V

29. **What will be the emf of the given cell?**
Pt | $H_2(P_1)$ | H≡(aq) | $H_2(P_2)$ | Pt
(1) $\dfrac{RT}{F}\ln\dfrac{P_1}{P_2}$ (2) $\dfrac{RT}{2F}\ln\dfrac{P_1}{P_2}$
(3) $\dfrac{RT}{F}\ln\dfrac{P_2}{P_1}$ (4) $\dfrac{RT}{2F}\ln\dfrac{P_2}{P_1}$

30. **An electrochemical cell is set up as :**
Pt, H_2 (1 atm) |HCl (0.1 M) || CH_3COOH (0.1M) | H_2 (1 atm) ; Pt. The e.m.f. of this cell will not be zero, because
(1) Acids used in two compartments are different
(2) e.m.f. depends on molarities of acids used
(3) The temperature is constant
(4) pH of 0.1 M HCl & 0.1 M CH_3COOH is not same

31. **The electrode potentials for**
$Cu^{2+}{}_{(aq)} + e^- \rightarrow Cu_{(aq)}$

and $Cu^{+}{}_{(aq)} + e^{-} \rightarrow Cu_{(s)}$
are $+0.15$ V and $+0.50$ V respectively. The value
of E° $_{Cu^{2+}/Cu}$ will be :–
(1) 0.500 V (2) 0.325 V
(3) 0.650 V (4) 0.150 V

32. **Given :**
(i) $Cu^{2+} + 2e^{-} \rightarrow Cu$, E° = 0.337 V
(ii) $Cu^{2+} + e^{-} \rightarrow Cu^{+}$, E° = 0.153 V
Electrode potential, E° for the reaction, $Cu^{+} + e^{-} \rightarrow Cu$, will be :
(1) 0.38 V (2) 0.52 V
(3) 0.90 V (4) 0.30 V

33. **The standard reduction potentials of Cu^{2+}/Cu and Cu^{2+}/Cu^{+} are 0.337 V and 0.153 V respectively. The standard electrode potential of Cu^{+}/Cu half cell is :**
(1) 0.184 V (2) 0.827 V
(3) 0.521 V (4) 0.490 V

34. **If E° $_{Fe^{3+}/Fe}$ is x_1, E° $_{Fe^{3+}/Fe}$ is x_2; then E° $_{Fe^{3+}/Fe^{+2}}$ will be**
(1) $3x_2 - 2x_1$ (2) $x_2 - x_1$
(3) $x_2 + x_1$ (4) $2x_1 + 3x_2$

35. **Given standard electrode potentials :–**
$Fe^{3+} + 3e^{-} \longrightarrow Fe$; E° = −0.036 volt
$Fe^{2+} + 2e^{-} \longrightarrow Fe$; E° = −0.440 volt
The standard electrode potential E° for
$Fe^{3+} + e^{-} \longrightarrow Fe^{2+}$
(1) −0.476 volt (2) −0.440 volt
(3) 0.440 volt (4) 0.772 volt

36. **For the given reactions**
$Sn^{4+} + 2e^{-} \rightarrow Sn^{2+}$ E° = x V
$Sn \longrightarrow Sn^{2+} + 2e^{-}$ E° = y V
The value of E° for the reaction
$Sn^{4+} + 4e^{-} \longrightarrow Sn$ will be
(1) $\sqrt{\frac{(x-y)}{2}}$ (2) $\frac{(x+y)}{2}$
(3) 2x + 2y (4) $2x + \frac{y}{2}$

37. **Given :**
$Mn^{2+} \longrightarrow MnO_4^{-}$; E° = −1.51 V
$MnO_2 \longrightarrow Mn^{2+}$; E° = 1.23 V
What is the E° value for the reaction?
$MnO_4^{-} \longrightarrow MnO_2$
(1) 1.7 V (2) 2.74 V
(3) 5.09 V (4) 1.1 V

38. **For the disproportionation of copper $2Cu^{+} \rightarrow Cu^{2+} + Cu$, E° is (Given E° for Cu^{2+}/Cu is 0.34 V and E° for Cu^{2+}/Cu^{+} is 0.15 V)**
(1) 0.49 V (2) −0.19 V
(3) −0.38 V (4) −0.38 V

39. **Given below are the half-cell reactions :–**

$Mn^{2+} + 2e^{-} \rightarrow Mn$;   E = −1.18 V
$2(Mn^{3+} + e^{-} \rightarrow Mn^{2+})$; E = +1.51 V
The E° for $3Mn^{2} \rightarrow Mn + 2Mn^{3+}$ will be :
(1) −0.33 V; the reaction will not occur
(2) −0.33 V; the reaction will occur
(3) −2.69 V; the reaction will not occur
(4) −2.69 V; the reaction will occur

TYPE–3 : Special Topics

40. **Standard electrode potentials are**
Fe^{2+}/Fe ; E° = −0.44 volts
Fe^{3+}/Fe^{2+} ; E° = 0.77 volts
Fe^{2+}, Fe^{3+} and Fe blocks are kept together, then
(1) Fe^{3+} increases
(2) Fe^{3+} decreases
(3) Fe^{2+}/Fe^{3+} remains unchanged
(4) Fe^{2+} decreases

41. **Given the data at 25°C,**
$Ag + I^{-} \rightarrow AgI + e^{-}$,   E° = 0.152 V
$Ag \rightarrow Ag^{+} + e^{-}$,   E° = −0.800 V
What is the value of $\log K_{sp}$ for AgI ?
$$\left(2.303\frac{RT}{F} = 0.059V\right)$$
(1) −8.12   (2) +8.612
(3) −37.83   (4) −16.13

42. **$PbO_2 \rightarrow PbO^{-}$; $\Delta G_{298} < 0$**
$SnO_2 \rightarrow SnO^{-}$; $\Delta G_{298} > 0$
Most probable oxidation state of Pb and Sn will be
(1) Pb^{4+}, Sn^{4+}
(2) Pb^{4+}, Sn^{2+}
(3) Pb^{2+}, Sn^{2+}
(4) Pb^{2+}, Sn^{4+}

43. **For the following cell reaction :**
$Cu(s) + 2Ag^{+}(aq) \longrightarrow 2Ag(s) + Cu^{2+}(aq)$
$E°_{cell}$ is 0.46 V at 300 K and 0.50 V at 340 K, then what will be entropy change?
(1) 386 JK^{-1} (2) 276 JK^{-1}
(3) 96.5 JK^{-1} (4) 193 JK^{-1}
Hint: $$\Delta S = \frac{d\Delta G}{dt}$$