Learnaf
Educational Revolution

# CHEMISTRY

## DAILY PRACTICE PAPER

## (DPP) - 1

### ELECTROCHEMISTRY – GALVANIC CELL


Arnav Girvan
CHEMISTRY FACULTY

Learnaf

Educational Revolution

# CHEMISTRY

## DAILY PRACTICE PAPER

## (DPP) - 1

## ELECTROCHEMISTRY – GALVANIC CELL

# DAILY PRACTICE PAPER

DPP - 1

**Unit - Electrochemistry** **Topic - Galvanic Cell**

By - Arnav Girvan
Mob - 9470458687

## Objective Problems

**01.** The standard emf for the cell reaction

$Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$ is 1.10 volt at 25°C. The emf for the cell reaction when 0.1 M $Cu^{2+}$ and 0.1 $Zn^{2+}$ solutions are used at 25°C is

(A) 1.10 volt (B) 0.110 volt
(C) –1.10 volt (D) –0.110 volt

**02.** Three moles of electrons are passed through three solutions in succession containing $AgNO_3$, $CuSO_4$ and $AuCl_3$ respectively. The molar ratio of amounts of cations reduced at cathode will be

(A) 1 : 2 : 3 (B) 2 : 1 : 3
(C) 3 : 2 : 1 (D) 6 : 3 : 2

**03.** The emf of the cell involving the following reaction, $2\ Ag^+ + H_2 \longrightarrow 2\ Ag + 2H^+$ is 0.80 volt. The standard oxidation potential of silver electrode is :-

(A) –0.80 volt (B) 0.80 volt
(C) 0.40 volt (D) –0.40 volt

**04.** For the electrochemical cell -

$M|M^+||X^-|X$, $E°_{M^+/M} = 0.44$ V and $E°_{X/X^-} = 0.33$ V. From this data we can deduce that

(A) $M + X \longrightarrow M^+ + X^-$ is the spontaneous reaction

(B) $M^+ + X^- \longrightarrow M + X$ is the spontaneous reaction

(C) $E_{cell} = 0.77$ V

(D) $E_{cell} = -0.77$ V

**05.** For the net cell reaction of the cell $Zn(s)|Zn^{2+}||Cd^{2+}|Cd(s)$ DG° in kilojules at 25°C is ($E°_{cell} = 0.360$ V)

(A) 112.5 (B) 69.47
(C) –34.73 (D) –69.47

**06.** How many faradays are required to reduce one mol of $MnO_4^-$ to $Mn^{2+}$

(A) 1 (B) 2
(C) 3 (D) 5

**07.** $Cu^+ + e^- \longrightarrow Cu$, $E° = x_1$ volt ;
$Cu^{2+} + 2e^- \longrightarrow Cu$, $E° = x_2$ volt, then for
$Cu^{2+} + e^- \longrightarrow Cu^+$, E° (volt) will be

(A) $x_1 - 2x_2$ (B) $x_1 + 2x_2$
(C) $x_1 - x_2$ (D) $2x_2 - x_1$

**08.** $Zn|Zn^{2+}(C_1)||Zn^{2+}(C_2)|Zn$. for this cell DG is negative if

(A) $C_1 = C_2$ (B) $C_1 > C_2$
(C) $C_2 > C_1$ (D) None

**09.** $Pt\left|\begin{matrix}H_2\\(p_1)\end{matrix}\right|\left|\begin{matrix}H^+\\(1M)\end{matrix}\right|\left|\begin{matrix}H^+\\(1M)\end{matrix}\right|\left|\begin{matrix}H_2\\(p_2)\end{matrix}\right|Pt$

(where $p_1$ and $p_2$ are pressure) cell reaction will be spontaneous if

(A) $p_1 = p_2$ (B) $p_1 > p_2$
(C) $p_2 > p_1$ (D) $p_1 = 1$ atm

**10.** $Pt\left|\begin{matrix}(H_2)\\1\,atm\end{matrix}\right|pH = 2\,||\,pH = 3\left|\begin{matrix}(H_2)\,Pt\\1\,atm\end{matrix}\right.$.

The cell reaction for the given cell is

(A) spontaneous (B) non-spontaneous
(C) equilibrium (D) none is correct

**11.** The cell Pt $(H_2)$ (1 atm) $|H^+$ (pH = ?), $I^-$ (a = 1) AgI(s), Ag has emf, $E_{298\ K} = 0$. The electrode potential for the reaction $AgI + e^- \rightarrow Ag + I^{\ominus}$ is –0.151 volt. Calculate the pH value

(A) 3.37 (B) 5.26
(C) 2.56 (D) 4.62

**12.** Using the information in the preceding problem, calculate the solubility product of AgI in water at 25°C [$E°_{(Ag^+,Ag)} = +0.799$ volt]

(A) $1.97 \times 10^{-17}$ (B) $7.91 \times 10^{-17}$
(C) $1.79 \times 10^{-17}$ (D) $9.17 \times 10^{-17}$

**13.** For $Zn^{2+}$/ Zn, E° = –0.76 V, for $Ag^+$ / Ag, E° = 0.799 V. The correct statement is

(A) the reaction Zn getting reduced Ag getting oxidized is spontaneous

(B) Zn undergoes reduction and Ag is oxidized

(C) Zn undergoes oxidation and $Ag^+$ gets reduced

(D) No suitable answer

**14.** The oxidation potential of a hydrogen electrode at pH = 10 and $P_{H_2} = 1$ is :-

(A) 0.51 V (B) 0.00 V
(C) + 0.59 V (D) 0.059 V

**15.** For the cell

$Pt|H_2(0.4\,atm)|H^+(pH{=}1)||H^+(pH = 2)|H_2\,(0.1\,atm)|Pt$

The measured potential at 25°C is

(A) –0.1 V (B) –0.5
(C) –0.041 (D) none

**16.** Consider the reaction of extraction of gold from its ore

$Au + 2CN^-(aq.) + \frac{1}{4}O_2(g) + \frac{1}{2}H_2O \longrightarrow Au(CN)_2^- + OH^-$

Use the following data to calculate DG° for the reaction $K_f\{Au(CN)_2^-\} = X$

$O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$ ; E° = + 0.41 volt

$Au^{3+} + 3e^- \longrightarrow Au$ ; E° = + 1.5 volt

$Au^{3+} + 2e^- \longrightarrow Au^+$ ; E° = + 1.4 volt

(A) – RT $ln$ X + 1.29 F (B) – RT $ln$ X – 2.11 F

(C) – RT $ln \frac{1}{X}$ + 2.11 F (D) – RT $ln$ X – 1.29 F

**17.** The emf of the cell :

$Ni|Ni^{2+}$ (1.0 M) || $Au^{3+}$(1.0 M) |Au is

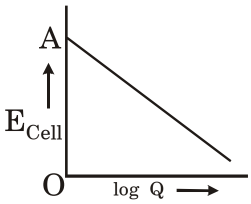
[E° for $Ni^{2+}|Ni$ = –0.25 V; E° for $Au^{3+}|Au$ = 1.5 V]

(A) +1.25 V (B) +1.75 V

(C) –1.25 V (D) –1.75 V

**18.** $Zn(s) + Cu^{2+}(aq.) \rightleftharpoons Cu(s) + Zn^{2+}(aq.)$

Reaction quotient $Q = \frac{[Zn^{2+}]}{[Cu^{2+}]}$, variation $E_{Cell}$ with Q is of the type with OA = 1.10 volt, $E_{Cell}$ = 1.1591 volt when




(A) $[Cu^{+2}]/[Zn^{2+}]$ = 0.01

(B) $[Zn^{2+}]/[Cu^{2+}]$ = 0.01

(C) $[Zn^{2+}]/[Cu^{2+}]$ = 0.1

(D) $[Zn^{2+}]/[Cu^{2+}]$ = 1

**19.** Emf of the cell :

$Pt \cdot H_2$ (1 atm) | $H^+(aq)$ || AgCl | Ag is 0.27 V and 0.26 V at 25°C and 35°C. Heat of reaction occuring inside the cell at 25°C is

(A) –54.8 kJ (B) 26.05 kJ

(C) –26.05 kJ (D) +54.8 kJ

**20.** The oxidation potential of a hydrogen electrode at pH = 10 and $pH_2$ = 1 atm is :

(A) 0.51 V (B) 0.00 V

(C) +0.59 V (D) 0.059 V

**21.** Cell reaction is spontaneous when:

(A) $E^\circ_{red}$ is negative (B) $E^\circ_{red}$ is positive

(C) DG° is negative (D) DG° is positive

**22.** Thy hydrogen electrode is dipped in a solution of pH = 3 at 25°C. The potential of the cell would be (the value of 2.303 RT/F is 0.059 V)

(A) 0.177 V (B) 0.087 V

(C) –0.177 V (D) 0.059 V

**23.** The emf of the cell :

$H_2$(1atm)Pt/$H^+$(pH=?)||$H^+$ (a=1)|$H_2$(1atm) Pt at 25°C is 0.59 V. The pH of the solution is

(A) 1 (B) 4 (C) 7 (D) 10

**24.** E° for the cell :

Zn | $Zn^{2+}$ || $Cu^{2+}$(aq) | Cu is 1.10 V at 25°C. The equilibrium constant for the cell reaction $Zn + Cu^{2+}(aq) \longrightarrow Cu + Zn^{2+}(aq)$ is of the order of

(A) $10^{-37}$ (B) $10^{37}$ (C) $10^{-17}$ (D) $10^{17}$

**25.** The emf of the cell

Zn | $Zn^{+2}$(0.01M) || $Fe^{+2}$(0.001 M) | Fe

at 298 K is 0.2905, then value of $K_{eq}$ for the reaction is

(A) $\frac{0.32}{e^{0.0295}}$ (B) $\frac{0.32}{10^{0.0295}}$

(C) $\frac{0.26}{10^{0.0295}}$ (D) $\frac{0.32}{10^{0.0591}}$

**26.** The emf of the cell in which the following reaction

$Zn(s) + Ni^{2+}(0.1\ M) \longrightarrow Zn^{2+}(1.0\ M) + Ni(s)$

occurs, is found to 0.5105 V at 298 K. The standard emf of the cell is

(A) 0.4810 V (B) 0.5696 V

(C) –0.5105 V (D) 0.5400 V

**27.** Standard reduction potentials at 25°C of $Li^+|Li$, $Ba^{2+}|Ba$, $Na^+|Na$ and $Mg^{2+}|Mg$ are -3.05, -2.9, 2.71 and -2.37 volt respectively. Which one of the following is the strongest oxidising agent?

(A) $Na^+$ (B) $Li^+$ (C) $Ba^{2+}$ (D) $Mg^{2+}$

**28.** The metal that does not displace hydrogen from an acid is:

(A) Al (B) Ca (C) Zn (D) Hg

**29.** Oxidation potential for the following half-cell reactions are :

$Zn \rightleftharpoons Zn^{2+} + 2e^-$; $E^\circ$ = 0.76 V

$Fe \rightleftharpoons Fe^{2+} + 2e^-$; $E^\circ$ = + 0.44 V

The E.M.F. for the cell reaction,

$Fe^{2+}(aq) + Zn \rightarrow Zn^{2+}(aq) + Fe$ will be

(A) -120 V (B) +0.32 V

(C) -0.32 V (D) -1.20 V

**30.** $E°$ for the cell, $Zn|Zn^{2+}(aq)||Cu^{2}(aq)|Cu$ is 1.10 V at 25°C. The equilibrium constant for the reaction,
$Zn(s) + Cu^{2+}(aq) \rightleftharpoons Cu(s) + Zn^{2+}(aq)$
is of the order
(A) $10^{+37}$ (B) $10^{+28}$ (C) $10^{+18}$ (D) $10^{+17}$

**31.** $Cu^{+}(aq)$ is unstable in solution and undergoes simultaneous oxidation and reduction according to the reaction,
$Zn(s) + Cu^{2+}(aq) \rightleftharpoons Cu(s) + Zn^{2+}(aq)$
choose correct $E°$ for above reaction if:
$E°_{Cu^{2+}/Cu} = 0.34$ V and $E°_{Cu^{2+}/Cu^{+}} = 0.15$ V
(A) -0.38 V (B) +0.49 V
(C) +0.38 V (D) -0.19 V

**32.** Given $E^{\circ}_{Fe^{2+}/Fe} = -0.44$ V, $E^{\circ}_{Fe^{3+}/Fe^{2+}} = 0.77$ V. $Fe^{2+}$, $Fe^{3+}$ and Fe are placed together then:
(A) $Fe^{3+}$ increases
(B) $Fe^{3+}$ decreases
(C) $[Fe^{3+}]/[Fe^{2+}]$ remains unchanged
(D) $Fe^{2+}$ decreases

**33.** The e.m.f. of a Daniell cell at 298 K is $E_1$

Zn | $ZnSO_4$ (0.01 M) || $CuSO_4$ (1.0 M) | Cu

When the concentration of $ZnSO_4$ is 1.0 M and that of $CuSO_4$ is 0.01 M, the e.m.f. changed to $E_2$. What is the relationship between $E_1$ and $E_2$?
(A) $E_1 = E_2$ (B) $E_1 = 0 \neq E_2$
(C) $E_1 > E_2$ (D) $E_1 < E_2$

**34.** The standard e.m.f. of a galvanic cell involving cell reaction with n = 2 is found to be 0.295 V at 25°C, The equilibrium constant of the reaction would be
(Given F = 96500 C $mol^{-1}$; $R$ = 8.314 $JK^{-1}$ $mol^{-1}$)
(A) $2.0 \times 10^{11}$ (B) $4.0 \times 10^{12}$
(C) $1.0 \times 10^{2}$ (D) $1.0 \times 10^{10}$

**35.** The values of standard oxidation potentials of following reactions are given below :
$Zn \rightarrow Zn^{2+} + 2e^-$; $E° = 0.762$ V
$Fe \rightarrow Fe^{2+} + 2e^-$; $E° = 0.440$ V
$Cu \rightarrow Cu^{2+} + 2e^-$; $E° = -0.345$ V
$Ag \rightarrow Ag^{+} + e^-$; $E° = -0.800$ V
Which of the following is most easily reduced?
(A) $Fe^{2+}$ (B) $Ag^{+}$
(C) $Zn^{2+}$ (D) $Cu^{2+}$

**36.** In the Daniell cell constructed in the laboratory, the voltage observed was 0.9 V instead of 1.1 V of the standard cell. A possible explanation is:
(A) $[Zn^{2+}] > [Cu^{2+}]$
(B) $[Zn^{2+}] < [Cu^{2+}]$
(C) the Zn electrode has twice the surface of Cu electrode
(D) mol ratio of $Zn^{2+} : Cu^{2+}$ is 2 : 1

**37.** For the cell reaction
$Cu^{2+}(C_1) + Zn(s) \rightleftharpoons Zn^{2+}(c_2) + Cu(s)$
of the electrochemical cell, the change in free energy, at a given temperature is function of:
(A) In $(C_1)$ (B) In $(C_2/C_1)$
(C) In $(C_2)$ (D) In $(C_1 + C_2)$

**38.** If the following half cells have the $E°$ values as:
$Fe^{3+} + e^- \rightarrow Fe^{2+}$; $E = +0.77$ V
and $Fe^{3+} + 2e^- \rightarrow Fe$; $E° = -0.44$ V.
the $E°$ of the half cell $Fe^{3+} + 3e^- \rightarrow Fe$ will be:
(A) 0.605 V (B) 0.33 V
(C) -1.21 V (D) -0.037 V

**39.** A galvanic cell is with standard oxidation electrode potential of zinc = + 0.76 V and of copper = -0.34 V. The e.m.f. with DG° negative is:
(A) +076 V (B) + 0.34 V
(C) -1.1 V (D) + 1.1 V

**40.** The E.M.F. of the cell: $Cu(s) | Cu^{2+}(aq) || Ag^{+}(aq) | Ag(s)$
$E^{\circ}_{Cu/Cu^{2+}} = -0.34$ V; $E^{\circ}_{Ag/Ag^{+}} = -0.80$ V
(A) + 0.46 V (B) - 0.46 V
(C) + 1.14 V (D) - 1.14 V

**41.** The equilibrium constant of the reaction :
$Cu(s) + 2Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$ :
$E^0 = 0.46$V at 298 K is
(A) $2.4 \times 10^{10}$ (B) $2.0 \times 10^{10}$
(C) $4.0 \times 10^{10}$ (D) $4.0 \times 10^{15}$

**42.** An electrochemical cell is set up as:
$Pt(H_2) | HCl(0.1\,M) || CH_3COOH(0.1\,M) | Pt(H_2)$
The e.m.f. of this cell will not be zero, because:
(A) the temperature is constant
(B) pH of 0.1 M HCl and 0.1 M $CH_3COOH$ is not same
(C) acids used in two compartments are different
(D) e.m.f. depends on molarities of acids used

**43.** For the electrochemical cell :
$M | M^{+} || X^{-} | X$, $E^o(M^+/M) = 0.44$ V and $E^o(X/X^-) = 0.33$ V. From this data one can deduce that
(A) $M + X \longrightarrow M^{+} + X^{-}$ is the spontaneous reaction
(B) $M^{+} + X^{-} \longrightarrow M + X$ is the spontaneous reaction
(C) $E_{Cell} = 0.77$ V
(D) $E_{Cell} = -0.77$ V

**44.** Using the data in the preceding problem, calculate the equilibrium constant of the reaction at 25°C.
$Zn + Cu^{++} \longrightarrow Zn^{++} + Cu$, $K = \frac{[Zn^{2+}]}{[Cu^{2+}]}$
(A) $8.314 \times 10^{24}$ (B) $4.831 \times 10^{31}$
(C) $8.314 \times 10^{36}$ (D) $4.831 \times 10^{44}$

**45.** The standard electrode potential (reduction) of $Ag^+|Ag$ is 0.800 V at 25°C. Its electrode potential in a solution containing $10^{-3}$ M ion of $Ag^+$ ions is :-
(A) 0.623 V (B) –0.977 V
(C) 0.892 V (D) 1.246 V

## Subjective Problems

# GALVANIC CELL :

### Representation of Cell diagrams, complete and half cell reactions :

**01.** Make complete cell diagrams of the following cell reactions
(a) $Cd^{2+}$ (aq) + Zn (s) ⟶ $Zn^{2+}$ (aq) + Cd (s)
(b) $2Ag^+$ (aq) + $H_2$ (g) ⟶ $2H^+$ (aq) + 2Ag (s)
(c) $Hg_2Cl_2$ (s) + Cu (s) ⟶ $Cu^{2+}$ (aq) + $2Cl^-$ (aq) + 2Hg (*l*)
(d) $Cr_2O_7^{2-}$ (aq.) + $14H^+$ (aq) + $6Fe^{2+}$ (aq) ⟶ $6Fe^{3+}$ (aq) + $2Cr^{3+}$ (aq) + $7H_2O$ (*l*)

**02.** Write cell reaction of the following cells
(a) Ag | $Ag^+$ (aq) | | $Cu^{2+}$ (aq) | Cu
(b) Pt | $Fe^{2+}$, $Fe^{3+}$ | | $MnO_4^-$, $Mn^{2+}$, $H^+$ | Pt
(c) Pt, $Cl_2$ | $Cl^-$ (aq) | | $Ag^+$ (aq) | Ag
(d) Pt, $H_2$ | $H^+$ (aq) | | $Cd^{2+}$ (aq) | Cd

**03.** Write half cells of each cell with following cell reactions
(a) Zn (s) + $2H^+$ (aq) ⟶ $Zn^{2+}$ (aq) + $H_2$ (g)
(b) $2Fe^{3+}$ (aq) + $Sn^{2+}$ (aq) ⟶ $2Fe^{2+}$ (aq) + $Sn^{4+}$ (aq)
(c) $MnO_4^-$ (aq) + $8H^+$ (aq) + $5Fe^{2+}$ (aq) ⟶ $2Fe^{3+}$ (aq) + $Mn^{2+}$ (aq) + $4H_2O$ (*l*)
(d) Pb (s) + $Br_2$ (*l*) ⟶ $Pb^{2+}$ (aq) + $2Br^-$ (aq)

### Electrode potential and standard electrode potential :

**04.** For the cell reaction
$2Ce^{4+}$ + Co ⟶ $2Ce^{3+}$ + $Co^{2+}$
$E^o_{cell}$ is 1.89 V. If $E^o_{Co^{2+}|Co}$ is – 0.28 V, what is the value of $E^o_{Ce^{4+}|Ce^{3+}}$ ?

**05.** Determine the standard reduction potential for the half reaction :
$Cl_2 + 2e^- \longrightarrow 2Cl^-$, Given :
$Pt^{2+} + 2Cl^- \longrightarrow Pt + Cl_2$ , $E^o_{Cell}$ = – 0.15 V
$Pt^{2+} + 2e^- \longrightarrow Pt$ , E° = 1.20 V

**06.** What is $E^o_{Cell}$ if
$2Cr + 3H_2O + 3OCl^- \longrightarrow 2Cr^{3+} + 3Cl^- + 6OH^-$
$2Cr^{3+} + 3e^- \longrightarrow Cr$ , E° = – 0.74 V
$OCl^- + H_2O + 2e^- \longrightarrow Cl^- + 2OH^-$, E° = 0.94 V

### DG°, $E^o_{Cell}$ and $K_{eq}$ :

**07.** Is 1.0 M $H^+$ solution under $H_2SO_4$ at 1.0 atm capable of oxidising silver metal in the presence of 1.0 M $Ag^+$ ion?
(Given : $E^o_{Ag^+|Ag}$ = 0.80 V, $E^o_{H^+|H_2(Pt)}$ = 0.0 V)

**08.** If for the half cell reactions
$Cu^{2+} + e^- \longrightarrow Cu^+$ E° = 0.15 V
$Cu^{2+} + 2e^- \longrightarrow Cu$ E° = 0.34 V
Calculate E° of the half cell reaction
$Cu^+ + e^- \longrightarrow Cu$, also predict whether $Cu^+$ undergoes disproportionation or not.

**09.** If $E^o_{Fe^{2+}|Fe}$ = – 0.44 V, $E^o_{Fe^{3+}|Fe^{2+}}$ = 0.77 V. Calculate $E^o_{Fe^{3+}|Fe}$ .

**10.** If $E^o_{Cu^+|Cu}$ = 0.52 V, $E^o_{Cu^{2+}|Cu}$ = 0.34 V, what is $E^o_{Cell}$ of the cell reaction Cu + $Cu^{2+}$ ⟶ $2Cu^+$? Is cell reaction spontaneous?

**11.** Calculate the EMF of a Daniel cell when the concentration of $ZnSO_4$ and $CuSO_4$ are 0.001 M and 0.1M respectively. The standard potential of the cell is 1.1V.

**12.** Calculate the equilibrium constant for the reaction Fe + $CuSO_4$ ⇌ $FeSO_4$ + Cu at $25^0$C.
(Given :
$E^0$ ($Fe/Fe^{2+}$) = 0.44V, $E^0$ ($Cu/Cu^{2+}$) = -0.337V.)

**13.** For a cell Mg(s) | $Mg^{2+}$(aq) | | $Ag^+$ (aq) | Ag, Calculate the equilibrium constant at $25^0$C. Also find the maximum work that can be obtained by operating the cell. (Given :
$E^0$ ($Mg^{2+}/Mg$) = -2.37V, $E^0$ ($Ag^+/Ag$) = 0.8 V.)

**14.** The standard reduction potential at $25^0$C for the reduction of water $2H_2O + 2e^- \rightleftharpoons H_2 + 2OH^-$ is -0.8277volt. Calculate the equilibrium constant for the reaction $2H_2O \longrightarrow H_3O^+ + OH^-$ at $25^0$C.

**15.** At $25^0$C the value of K for the equilibrium $Fe^{3+}$ + Ag ⇌ $Fe^{2+}$ + $Ag^+$ is 0.531mol/litre. The standard electrode potential for $Ag^+ + e^- \rightleftharpoons Ag$ is 0.799V. What is the standard potential for $Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$ ?

**16.** The EMF of the cell M | $M^{n+}$ (0.02M) | | $H^+$ (1M) | $H_2$(g) (1 atm), Pt at $25^0$C is 0.81V. Calculate the valency of the metal if the standard oxidation of the metal is 0.76V.

**17.** Equinormal Solutions of two weak acids, HA ($pK_a$ = 3) and HB ($pK_a$ = 5) are each placed in contact with standard hydrogen electrode at $25^0$C. When a cell is constructed by interconnecting them through a salt bridge, find the emf of the cell.

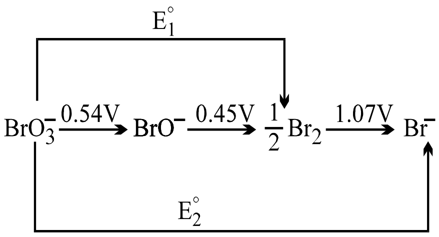
**18.** In two vessels each containing 500ml water, 0.5m mol of aniline ($K_b = 10^{-9}$) and 25mmol of HCl are added separately. Two hydrogen electrodes are constructed using these solutions. Calculate the emf of cell made by connecting them appropriately.

**19.** Calculate $E^0$ and E for the cell Sn | $Sn^{2+}$ (1M) || $Pb^{2+}$ | Pb($10^{-3}$M), $E^0$ ($Sn^{2+}$| Sn) = -0.14V, $E^0$ ($Pb^{2+}$| Pb) = -0.13V. Is cell representation is correct?

**20.** At what concentration of $Cu^{2+}$ in a solution of $CuSO_4$ will the electrode potential be zero at $25^0C$? (Given : $E^0$ (Cu | $Cu^{2+}$) = -0.34 V.)

**21.** A zinc electrode is placed in a 0.1M solution at $25^0C$. Assuming that the salt is 20% dissociated at this dilutions calculate the electrode potential. $E^0$ ($Zn^{2+}$| Zn) = -0.76V.

**22.** From the standard potentials shown in the following diagram, calculate the potentials $E_1^\circ$ and $E_2^\circ$.




**23.** For the reaction, $4Al(s) + 3O_2(g) + 6H_2O + 4OH^- \rightleftharpoons 4\,[Al(OH)_4^-]$ ; $E^\circ_{cell} = 2.73$ V. If $\Delta G_f^\circ(OH^-) = -157$ kJ $mol^{-1}$ and $\Delta G_f^\circ(H_2O) = -237.2$ kJ $mol^{-1}$, determine $\Delta G_f^\circ[Al(OH)_4^-]$.

## Concentration cells :

**24.** Calculate the EMF of the following cell :
Zn | $Zn^{2+}$ (0.01M) || $Zn^{2+}$ (0.1 M) | Zn at 298 K.

**25.** Calculate the EMF of the cell :
Zn – Hg($c_1$M) | $Zn^{2+}$ (aq) | Hg – Zn($c_2$M) at 25°C, if the concentrations of the zinc amalgam are : $c_1$ = 10g per 100g of mercury and $c_2$ = 1g per100 g of mercury.

**26.** Calculate pH using the following cell :
Pt ($H_2$) | $H^+$ (*x* M) || $H^+$ (1 M) | Pt ($H_2$)
1 atm 1 atm
if $E_{cell}$ = 0.2364 V.

**27.** Calculate the EMF of following cells at $25^0C$.

(i) Fe | $Fe^{2+}$ ($a_1$ = 0.3) || $Sn^{2+}$ ($a_2$ = 0.1) | Sn
($E^0$ ($Fe^{2+}$/Fe) = –0.44 V)

(ii) Pt, $H_2$ (2atm) | HCl | $H_2$ (10 atm), Pt.
($E^0$ ($Sn^{2+}$/Sn) = -0.14 V)

**28.** EMF of the cell :
Zn | $ZnSO_4$($a_1$= 0.2) || $ZnSO_4$($a_2$) | Zn
is -0.0088V at $25^0C$. Calculate the value of $a_2$.

Learnaf

Educational Revolution

# ANSWER KEYS OF # Electrochemistry – Galvanic Cell

## Objective Questions # Answers

| Q. | 01. | 02. | 03. | 04. | 05. | 06. | 07. | 08. | 09. | 10. | 11. | 12. | 13. | 14. | 15. |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| A. | A | D | A | B | D | D | D | C | B | B | C | B | C | B | C |
| Q. | 16. | 17. | 18. | 19. | 20. | 21. | 22. | 23. | 24. | 25. | 26. | 27. | 28. | 29. | 30. |
| A. | A | B | B | A | D | C | C | D | B | B | D | D | D | B | A |
| Q. | 31. | 32. | 33. | 34. | 35. | 36. | 37. | 38. | 39. | 40. | 41. | 42. | 43. | 44. | 45. |
| A. | C | A | C | C | B | A | B | D | D | A | D | B | B | C | A |

## Subjective Questions # Answers

**01.** (a) $Zn|Zn^{2+}||Cd^{2+}|Cd$, (b) $Pt, H_2|H^+||Ag^+|Ag$ ,

(c) $Cu|Cu^{2+}||Cl^-|Hg_2Cl_2|Hg$ (d) $Pt|Fe^{2+},Fe^{3+}||Cr^{3+}|Pt$

**02.** (a) $2Ag + Cu^{2+} \longrightarrow 2Ag^+ + Cu$, (b) $+ 5Fe^{2+} + 8H^+ \longrightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

(c) $2Cl^- + 2Ag^+ \longrightarrow 2Ag + Cl_2$, (d) $H_2 + Cd^{2+} \longrightarrow Cd + 2H^+$

**03.**

| | Anode | Cathode |
|---|---|---|
| (a) | $Zn|Zn^{2+}$ | $H^+, H_2|Pt$ |
| (b) | $Pt|Sn^{2+},Sn^{4+}$ | $Fe^{3+}, Fe^{2+}|Pt$ |
| (c) | $Pt|Fe^{2+}, Fe^{3+}$ | $Mn^{2+} | Pt$ |
| (d) | $Pb|Pb^{2+}$ | $Br_2, Br^-|Pt$ |

**04.** 1.61 V **05.** 1.35 V **06.** 1.68 V **07.** – 0.80 V, NO **08.** 0.53 V, disproportionation

**09.** – 0.0367 V **10.** – 0.36 V, not spontaneous **11.** E =1.159V **12.** $K_c = 2.18 \times 10^{26}$

**13.** $K_c = 1.864 \times 10^{107}$, $DG^0 = – 611.8$ kJ **14.** $K_w$» $10^{-14}$ **15.** $E^0 = 0.7826$ V **16.** n = 2

**17.** E = 0.059 **18.** E = 0.395 V **19.** $E^0{}_{cell} = +0.01V$, $E_{cell} = -0.0785V$, correct representation is $Pb|Pb^{2+}(10^{-3}M)||Sn^{2+}(1M)|Sn$ **20.** $[Cu^{2+}] = 2.97 \times 10^{-12}M$ for E =0 **21.** E = -0.81eV

**22.** 0.52 V, 0.61 V **23.** $–1.30 \times 10^3$ kJ $mol^{-1}$ **24.** 0.0295 V **25.** 0.0295 V **26.** pH = 4

**27.** (i) E = 0.286V; (ii) E = -0.0206V **28.** $a_2 = 0.1006$ M

Thank you for your love and support, we hope you are always being happy and get success in your life, we are happy to see you again.

***Regards from Learnaf team***