Q1. $Mg(s) + 2Ag^{2+}(0.0001M) \rightarrow Mg^{2+}(0.13M) + 2Ag(s)$

$E^\circ_{cell} = 3.17V$ [Galvanic cell] Find E_{cell}

Sol^n:- Here, $n = 2$

$E^\circ_{cell} = 3.17V$

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{2} \log \frac{[Mg^{2+}]}{[Ag^{2+}]^2} \quad [\because 1\,Mg^{2+}] \; [\because 2\,Ag^{2+}]$$

$$= 3.17 - \frac{0.0591}{2} \cdot \log\left(\frac{0.13}{(0.0001)^2}\right)$$

$$= 3.17 - 0.02955 \times 7.114$$

$$= 3.17 - 0.21$$

$$\Rightarrow E_{cell} = 2.96\,V$$

Q2. Calculate Equilibrium constant (K) for Daniel cell

$E^\circ_{cell} = 1.1V$ ~~Find E_{cell}~~

Sol^n:- Anode : $Zn \rightarrow Zn^{2+} + 2e^-$

Cathode : $Cu^{2+} + 2e^- \rightarrow Cu$

$Zn(s) + Cu^{2+} \rightarrow Zn^{2+} + Cu(s)$

Here, $n = 2$, $E^\circ_{cell} = 1.1V$

At EQ^m, $E_{cell} = 0$

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{n} \log K$$

$$\Rightarrow 0 = 1.1 - \frac{0.0591}{2} \log K$$

$$\Rightarrow 1.1 = \frac{0.0591}{2} \log K$$

$$\Rightarrow K = \log^{-1}\left(\frac{2.2}{0.0591}\right)$$

$$\Rightarrow K = 1.679 \times 10^{37}$$

(K → unitless)

* $\log_{10}(1000) = y$ (Logarithmic form)

(anti-log) ⇓ $10^y = 1000$ (exponential form)

$\frac{[Product]}{[Reactant]} = \frac{[M]}{[M]} = M^\circ$

$[M^\circ L^\circ T^\circ]$

Q3. Calculate E_{cell} for following cell reactions :-

i) $Mg(s)/Mg^{2+}(0.001M)\,||\,Cu^{2+}(0.0001M)/Cu(s)$ $\quad E^\circ_{cell} = 2.7V$

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

Sol^n :- $E_{cell} = E^\circ_{cell} - \frac{0.0591}{2} \log \frac{[Cu^{2+}]}{[Mg^{2+}]}$

$= 2.7 - 0.02955 \log\left(\frac{0.0001}{0.001}\right)$

$= 2.7 - 0.02955 \times (-1)$

$E_{cell} = 2.729\,V$

ii) $Fe(s)/Fe^{2+}(0.001M)\,||\,H^+(1M)/H_2(g)$ $\quad E^\circ_{cell} = 0.44V$

$Fe + 2H^+ \rightarrow Fe^{2+} + H_2$

Sol^n :- $E_{cell} = E^\circ_{cell} - \frac{0.0591}{2} \log \frac{[H^+]^2}{[Fe^{2+}]}$

$= 0.44 - 0.02955 \log\left(\frac{1 \times 1}{0.001}\right)$

$= 0.44 - 0.02955 \times 3$

$= 0.44 - 0.09$

$E_{cell} = 0.35\,V$

iii) $Sn(s)/Sn^{2+}(0.05M)\,||\,H^+(0.02M)/H_2(g)$ $\quad E^\circ_{cell} = 0.14V$

$Sn + 2H^+ \rightarrow Sn^{2+} + H_2$

Sol^n :- $E_{cell} = E^\circ_{cell} - \frac{0.0591}{2} \log \frac{[H^+]^2}{[Sn^{2+}]}$

$= 0.14 - 0.02955 \log\left(\frac{0.02 \times 0.02}{0.05}\right)$

$= 0.14 - 0.03 \times (-2.1)$

$= 0.14 + 0.063$

$E_{cell} = 0.203\,V$

iv) ~~Pt(s)~~ $Br_2(l)/Br^-(0.01M)\,||\,H^+(0.03M)/H_2(g)$ $E^\circ_{cell} = -0.09V$

Sol^n:- $E_{cell} = E^\circ_{cell} - \frac{0.0591}{2}\log\frac{[H^+]^2}{[Br^-]^2}$

$Br_2 + 2H^+ \rightarrow 2Br^- + H_2$

$= -0.09 - 0.03\log\left(\frac{0.03\times0.03}{0.01\times0.01}\right)$

$= -0.09 - 0.03 \times 0.954$

$= -0.09 - 0.0029$

$E_{cell} = -0.0929V$

Q4. Calculate the electrode potential for the following:-

i) $Fe^{2+}(0.01M)/Fe(s)$ $E^\circ_{Fe^{2+}/Fe} = -0.45V$

Sol^n:- $Fe^{2+} + 2e^- \longrightarrow Fe(s)$ ⇐ [Reduction potential]

$\therefore E_{Fe^{2+}/Fe} = E^\circ_{Fe^{2+}/Fe} - \frac{0.0591}{2}\log\frac{[Fe]}{[Fe^{2+}]}$

$= -0.45 - \frac{0.0591}{2}\log\left(\frac{1}{0.01}\right)$

$= -0.45 - \frac{0.0591}{2} \times 2$

$\Rightarrow E_{Fe^{2+}/Fe} = -0.509V$ [$E_{Fe^{2+}/Fe}$ → Reduction potential]

* For reduction potential, −ve : $-\frac{0.0591}{n}\log Q$

For oxidation potential, +ve : $+\frac{0.0591}{n}\log Q$

* Conc of Product/Reactant (solid/liquid/gas)

[Product]/[Reactant] is taken as 1

(ii) $Cr(s)/Cr^{3+}(0.1M)$ $E^\circ_{Cr^{3+}/Cr} = -0.75V$

Sol^n:- $Cr^{3+} + 3e^- \longrightarrow Cr(s)$ [Reduction potential]

$$\therefore E_{Cr^{3+}/Cr} = E^\circ_{Cr^{3+}/Cr} - \frac{0.0591}{3}\log\frac{[Cr]}{[Cr^{3+}]}$$

$$= -0.75 - \frac{0.0591}{3}\log\left(\frac{1}{0.1}\right)$$

$$= -0.75 - 0.0197 \times (1)$$

$$E_{Cr^{3+}/Cr} = -0.769V$$

(*) Qs. Find EMF of Cell :-

(Anode) (Cathode)

(i) $Cr/Cr^{3+}(0.1M) \,||\, Fe^{2+}(0.01M)/Fe$ $E^\circ_{Cr^{3+}/Cr} = -0.75V$, $E^\circ_{Fe^{2+}/Fe} = -0.45V$

Sol^n:- $Cr + Fe^{2+} \longrightarrow Cr^{3+} + Fe$

By default, we always take Reduction Potential

$$\therefore E^\circ_{cell} = E^\circ_{Cathode} - E^\circ_{Anode} \quad (\text{or}) \quad E^\circ_{Right} - E^\circ_{Left}$$

$$= -0.45 - (-0.75) \qquad (E^\circ_{Fe^{2+}/Fe}) \quad (E^\circ_{Cr^{3+}/Cr})$$

$$\Rightarrow E^\circ_{cell} = 0.3V$$

Anode : $Cr \longrightarrow Cr^{3+} + 3e^-$ $\times 2 \Rightarrow 2Cr \longrightarrow 2Cr^{3+} + 6e^-$

Cathode : $Fe^{2+} + 2e^- \longrightarrow Fe$ $\times 3 \Rightarrow 3Fe^{2+} + 6e^- \longrightarrow 3Fe$

$\therefore n = 6$ $\qquad 2Cr + 3Fe^{2+} \longrightarrow 2Cr^{3+} + 3Fe$

$$\therefore E_{cell} = E^\circ_{cell} - \frac{0.0591}{n}\log\frac{[Cr^{3+}]^2}{[Fe^{2+}]^3}$$

$$= 0.3 - \frac{0.0591}{6}\log\left(\frac{0.1 \times 0.1}{0.01 \times 0.01 \times 0.01}\right)$$

$$= 0.3 - 0.0394$$

$$= 0.3 - 0.04$$

$$\Rightarrow E_{cell} = 0.26$$

(ii) $Zn/Zn^{2+}(0.1M) \,||\, Sn^{2+}(0.001M)/Sn$ $E^\circ_{Zn^{2+}/Zn} = -0.76V$, $E^\circ_{Sn^{2+}/Sn} = -0.14V$

Sol^n:- $Zn + Sn^{2+} \longrightarrow Zn^{2+} + Sn$ $[n=2]$

$\therefore E^\circ_{cell} = E^\circ_{Sn^{2+}/Sn} - E^\circ_{Zn^{2+}/Zn}$
$= -0.14 - (-0.76)$
$= 0.62V$

$E_{cell} = E^\circ_{cell} - \frac{0.0591}{n} \log \frac{[Zn^{2+}]}{[Sn^{2+}]}$
$= 0.62 - \frac{0.06}{2} \log\left(\frac{0.1}{0.001}\right)$
$= 0.62 - 0.03 \times 2$
$E_{cell} = 0.56V$

Q6) $Pt/H_2(g, 1\ bar)/H^+(aq, 1M) \,||\, M^{4+}(aq)/M^{2+}(aq)/Pt$

Sol^n:- $E^\circ_{M^{4+}/M^{2+}} = 0.151V$, $E^\circ_{H^+/\frac{1}{2}H_2} = 0V$, $E_{cell} = 0.092V$
Find molar concentration of M

Sol^n:- A/Q, $\frac{[M^{4+}]}{[M^{4+}]} = 10^x$ where $x \rightarrow$ molar concentration

Anode : $H_2 \longrightarrow 2H^+ + 2e^-$
Cathode : $M^{4+} + 2e^- \longrightarrow M^{2+}$
$H_2 + M^{4+} \longrightarrow 2H^+ + M^{2+}$
(x) (1M)

$n = 2$
$\therefore E^\circ_{cell} = E^\circ_{M^{4+}/M^{2+}} - E^\circ_{H^+/\frac{1}{2}H_2}$
$= 0.151 - 0$
$\Rightarrow E^\circ_{cell} = 0.151V$

$\therefore E_{cell} = E^\circ_{cell} - \frac{0.0591}{2} \log\left(\frac{[H^+]^2[M^{2+}]}{[H_2][M^{4+}]}\right)$
$\Rightarrow 0.092 = 0.151 - \frac{0.06}{x} \log\left(\frac{(1)^2 \times 10^x}{1}\right)$

$\Rightarrow -0.059 = -0.03 \cdot x$
$\Rightarrow x = \frac{0.059}{0.03} = \frac{0.06}{0.03} \Rightarrow \boxed{x = 2M}$