Electrode potential, E$_{Cell}$, Nernt equation and ECS

76. (d) $\Delta G = -nFE^o$

$\Delta G = -2.303\, RT \log K$; $nFE^o = 2.303\, RT \log K$

$$\log K = \frac{nFE^o}{2.303\, RT} = \frac{2 \times 96500 \times 0.295}{2.303 \times 8.314 \times 298}$$

$\log K = 9.97 = K = 1 \times 10^{10}$.

77. (b) For the given cell $M | M^+ || X^- | X$, the cell reaction is derived as follows:

RHS: reduction $X + e^- \rightarrow X^-$ .....(i)

LHS: Oxidation $M \rightarrow M^+ + e^-$ .....(ii)

Add (i) and (ii) $M + X \rightarrow M^+ + X^-$

The cell potential = $-0.11\, V$

Since $E_{cell} = -$ ve, the cell reaction derived above is not spontaneous. In fact, the reverse reaction will occur spontaneously.

78. (c) $\overset{0}{Zn} + \overset{+2}{Mg}Cl_2 \rightarrow \overset{+2}{Zn}Cl_2 + Mg$ No reaction

Oxidation

Reduction

This type of reaction does not occur because

$Mg^{2+} E^o = -2.37\, V$ while $Zn^{2+} E^o = -0.76\, V$.

79. (b) In neutral medium Mn^{+7} oxidation state change into +4 oxidation state, hence equivalent weight of $KMnO_4 = \frac{M}{3}$.

80. (a) Increase in the concentration of Ag^+ ion increase the voltage of the cell.

81. (a) $E_{cell} = E^o_{cell} + \frac{0.059}{2} \log \frac{(Ag^+)}{(Sn^{2+})}$.

82. (b) The *K.E.* of proton is 1 *KeV.*

83. (b) Anodic reaction : $H_2(P_1) \rightarrow 2H^+$

Cathodic reaction : $2H^+ \to H_2(P_2)$

$E_{cathode} = -\frac{RT}{2F}\ln\frac{P_2}{[H^+]^2}$; $E_{\text{anode}} = -\frac{RT}{2F}\ln\frac{[H^+]^2}{P_1}$

$E_{\inf} = E_{anode} + E_{cathode} \quad = -\frac{RT}{2F}\ln\frac{(H^+)^2}{P_1} - \frac{RT}{2F}\ln\frac{P_2}{(H^+)^2} \quad = -\frac{RT}{2F}\ln\frac{P_2}{P_1} = \frac{RT}{2F}\ln\frac{P_1}{P_2}$.

84. (c) $\frac{1}{2}H_2 \to H^+(10^{-8}M) + e^-(\text{oxidation })$

$H^+(0.025\ M) + e^- \to \frac{1}{2}H_2(\text{reduction })$

Cell reaction is :

$H^+(0.025\ M) \to H^+(10^{-8}\ M)$; $E_{\text{cell}} = 0.38\ V$.

85. (a) E^o *for* $Fe/Fe^{2+} = 0.44$ V .

86. (c) (Reduction potential of cathode) – (reduction potential of anode).

87. (a) The correct decreasing electrode potential order is : K, Ba, Ca, Mg.

88. **Given reaction:**

Mg(s) + Cu^{2+}(aq) ⟶ Cu(s) + Mg^{2+}(aq)

Standard reduction potentials:

- E° (Mg^{2+}/Mg) = −2.37 V
- E° (Cu^{2+}/Cu) = +0.34 V

Step 1: Identify cathode and anode

- Cu^{2+} ⟶ Cu (reduction) happens at the **cathode** (higher E°).
- Mg ⟶ Mg^{2+} + $2e^-$ (oxidation) happens at the **anode** (lower E°).

Step 2: Calculate cell emf

Formula:

E° cell = E° cathode − E° anode

Substitute values:

E° cell = (+0.34) − (−2.37)

E° cell = +0.34 + 2.37

E° cell = **+2.71 V**

89. (c) $E^o_{cell} = E^o_{cathode} - E^o_{anode}$

$= 0.34 - (-2.37) = +2.71\ V$.

90. (b) Because flourine is most powerful reducing agent than other halogens.

91. **Word–friendly explanation:**

Although **aluminium is more reactive** than iron, it does **not corrode easily** because:

Aluminium reacts quickly with oxygen in air to form a **thin, tough, and adherent layer of aluminium oxide (Al_2O_3)** on its surface.

This oxide layer is **impermeable** and **protects the underlying metal** from further oxidation or corrosion.

Therefore, the correct option is: **(c) Oxygen forms a protective oxide layer**

92. (c) Aluminium forms a protective oxide layer but iron does not.

93. (d) The reduction potential of *Zn* is very higher than *Cu*.

94. (a) The *pH* of 0.1 *M* HCl and 0.1 *M* acetic acid is not the same, because HCl is a strong acid so its *pH* is less and CH_3COOH is a weak acid, so its *pH* is more.

95. (d) The required reaction $(Cu^{++} + Cu \rightarrow 2Cu^+)$ can be obtained by using the following reactions.

$Cu^{++} + e^- \rightarrow Cu^+;\ E^o_{Cu^{++}/Cu^+} = 0.15\ V$ .....(i)

$Cu^{++} + 2e^- \rightarrow Cu;\ E^o_{Cu^{++}/Cu} = 0.34\ V$ .....(ii)

Multiplying eq. (i) by 2 we get

$2Cu^{++} + 2e^- \rightarrow 2Cu^+$ .....(iii)

$\Delta G_1 = -nFE = -2 \times F \times 0.15$

$Cu^{++} + 2e^- \rightarrow Cu$ .....(iv)

$\Delta G_2 = -nFE = -2 \times F \times 0.34$

Subtract the *eq*. (iv) from (iii)

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

$\Delta G_3 = -nFE = -1 \times F \times E^o$

Also $\Delta G_3 = \Delta G_1 - \Delta G_2$

$-1FE^o = (-2F \times 0.15) - (-2F \times 0.34)$

$E^o = -0.38$

This is the value for the reaction

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

But the given reaction is just reverse of it

$\therefore E_{\text{cell}}$ for given reaction = + 0.38*V*.

96. A **salt bridge** completes the electrical circuit by allowing **ion flow** between the two half–cells.

- If the salt bridge is **removed**, **no ions can flow**, so **the circuit is broken**.
- As a result, **no current flows**, and the **measured voltage (EMF) drops to zero**.

Correct option: **(d) Drops to zero**

97. (d) It connect two solutions and complete the circuit.

98. (a) Greater the oxidation potential, greater is the reactivity.

99. (b) Electrochemical series compare the relative reactivity of metals.

100. (d) Fuel cells are more efficient, free from pollution and they run till reactants are active.