

Electrode potential, E_{cell} , Nernt equation and ECS

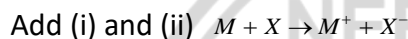
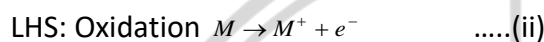
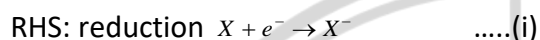
76. (d) $\Delta G = -nFE^\circ$

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

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

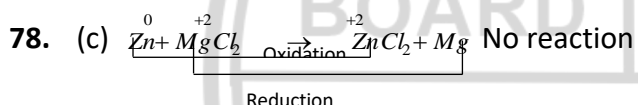
$$\log K = 9.97 = K = 1 \times 10^{10}.$$

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



The cell potential = -0.11 V

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



This type of reaction does not occur because

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

79. (b) In neutral medium Mn^{+7} oxidation state change into +4 oxidation state, hence equivalent

$$\text{weight of } KMnO_4 = \frac{M}{3}.$$

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

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

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

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

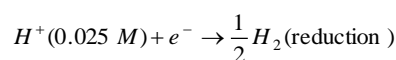


Cathodic reaction : $2H^+ \rightarrow H_2(P_2)$

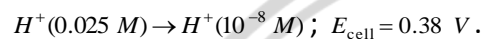
$$E_{cathode} = -\frac{RT}{2F} \ln \frac{P_2}{[H^+]^2} ; E_{anode} = -\frac{RT}{2F} \ln \frac{[H^+]^2}{P_1}$$

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

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



Cell reaction is :

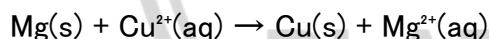


85. (a) $E^\circ \text{ for } Fe / Fe^{2+} = 0.44 V$.

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:



Standard reduction potentials:

- $E^\circ (Mg^{2+}/Mg) = -2.37 V$
- $E^\circ (Cu^{2+}/Cu) = +0.34 V$

Step 1: Identify cathode and anode

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

Step 2: Calculate cell emf

Formula:

$$E^\circ \text{ cell} = E^\circ \text{ cathode} - E^\circ \text{ anode}$$

Substitute values:

$$E^\circ \text{ cell} = (+0.34) - (-2.37)$$



$$E^{\circ}_{\text{cell}} = +0.34 + 2.37$$

$$E^{\circ}_{\text{cell}} = +2.71 \text{ V}$$

89. (c) $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$
 $= 0.34 - (-2.37) = +2.71 \text{ V}$

90. (b) Because fluorine 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.

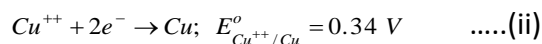
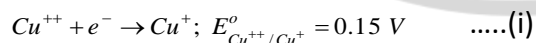
Therefore, the correct option is: (c) Oxygen forms a protective oxide layer

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 ($\text{Cu}^{++} + \text{Cu} \rightarrow 2\text{Cu}^+$) can be obtained by using the following reactions.



Multiplying eq. (i) by 2 we get



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



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



Subtract the eq. (iv) from (iii)



$$\Delta G_3 = -nFE = -1 \times F \times E^{\circ}$$

$$\text{Also } \Delta G_3 = \Delta G_1 - \Delta G_2$$

$$-1FE^{\circ} = (-2F \times 0.15) - (-2F \times 0.34)$$

$$E^{\circ} = -0.38$$

This is the value for the reaction



But the given reaction is just reverse of it

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

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.

