

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

101. Given:

- Half-cell: $\text{Zn(s)} \mid \text{Zn}^{2+}(0.01 \text{ M})$
- Standard potential: $E^\circ = 0.763 \text{ V}$
- Temperature: $T = 258^\circ \text{C} = 258 + 273 = 531 \text{ K}$
- $[\text{Zn}^{2+}] = 0.01 \text{ M}$

Step 1: Use Nernst equation (word-friendly form):

$$E = E^\circ - (0.0591 / n) \times \log([\text{Zn}^{2+}]) \text{ at } 298 \text{ K}$$

But temperature is **531 K**, so use:

$$E = E^\circ - (RT / nF) \times \ln([\text{Zn}^{2+}])$$

Where:

- $R = 8.314 \text{ J/mol}\cdot\text{K}$
- $F = 96485 \text{ C/mol}$
- $n = 2$ (for $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$)

Step 2: Substitute values

$$E = 0.763 - (8.314 \times 531) / (2 \times 96485) \times \ln(0.01)$$

1. Calculate the factor:
 $(8.314 \times 531) / (2 \times 96485) \approx 0.02287$
2. $\ln(0.01) = \ln(10^{-2}) = -4.605$
3. Multiply: $0.02287 \times (-4.605) \approx -0.105$
4. Subtract: $E = 0.763 - (-0.105) = 0.763 + 0.105 \approx 0.868 \text{ V}$

102. (c) $E^\circ = E^\circ_{\text{Ag}^{2+}/\text{Ag}} + E^\circ_{\text{Cu}/\text{Cu}^{2+}} = -0.34 + 0.80 = +0.46 \text{ V}$.

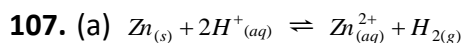
103. (a) *Fe* is placed above *Cu* in electrochemical series.

104. (a) Lithium is the strongest reducing agent of the alkali metals.

105. (d) Potassium is more electropositive element, because it is the only alkali metal among the given elements.

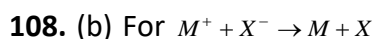
106. (b) Aluminium forms a self protecting film of oxide to prevent corrosion.





$$E_{Cell} = E^0_{Cell} - \frac{.059}{2} \log \frac{[Zn^{2+}]}{[H^+]^2}$$

When H_2SO_4 is added then $[H^+]$ will increase therefore E_{Cell} will also increases and equilibrium will shift towards right.



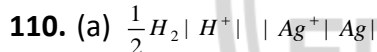
$$E^0_{Cell} = E^0_{Cathode} + E^0_{Anode} = 0.44 - 0.33 = +0.11 \text{ V}$$

Since $E^0_{Cell} = (+) 0.11 \text{ V}$ is positive hence this reaction should be spontaneous.

109. (a) $E_{Cell} = E^0_{Cell} - \frac{0.0591}{n} \log K_c$

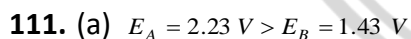
At 298 K $E_{Cell} = 0$ $0 = 0.591 - \frac{0.0591}{n} \log K_c$

$$\log K_c = \frac{0.591 \times 1}{0.0591} = 10 ; K_c = \text{Anti log } 10 = 1 \times 10^{10} .$$



$$E^0_{Cell} = E^0_{Cathode} - E^0_{Anode} = E^0_{Ag^+/Ag} - E^0_{H^+/\frac{1}{2}H_2}$$

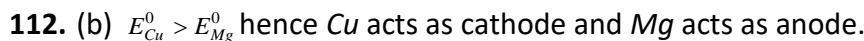
$$(0.80) - (0.0) = 0.80 \text{ V} .$$



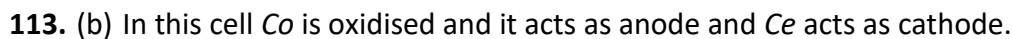
So A will act as cathode in galvanic cell. Hence

$$E^0_{Cell} = E^0_{Cathode} - E^0_{Anode} = E_A - E_B$$

$$= (2.23) - (-1.43) = 3.66 \text{ V} .$$



$$E^0_{Cell} = E^0_{Cu} - E^0_{Mg} = (0.34) - (-2.37) = +2.71 \text{ V} .$$



$$E^0_{Cell} = E^0_{Cathode} - E^0_{Anode} = 1.89 = E^0_{Cell} - (-0.28)$$

$$E^0_{Cell} = 1.89 - 0.28 = 1.61 \text{ Volts} .$$



114. (b) Given: $\Delta G = -21.20 \text{ kJ} = 21200 \text{ J}$

$$\therefore \Delta G = -nFE$$

$$E = \frac{21200}{1 \times 96500} = 0.2196 \text{ V} = 0.22 \text{ V}.$$

115. (b) $\text{Ag} | \text{Ag}^+ (0.1\text{M}) || \text{Ag}^+ 1\text{M} | \text{Ag}$

$$E_{\text{Cell}} = \frac{2.303 RT}{nF} \log \frac{c_1}{c_2} = \frac{0.059}{1} \log \frac{1}{0.1}$$

$$= 0.059 \log 10 = 0.059 \text{ Volt}.$$

116. (b) For this cell, reaction is: $\text{Zn} + \text{Fe}^{2+} \rightarrow \text{Zn}^{2+} + \text{Fe}$

$$E = E^0 - \frac{0.0591}{n} \log \frac{c_1}{c_2}; E^0 = E + \frac{0.0591}{n} \log \frac{c_1}{c_2}$$

$$= 0.2905 + \frac{0.0591}{2} \log \frac{10^{-2}}{10^{-3}} = 0.32 \text{ V}.$$

$$E^0 = \frac{0.0591}{2} \log K_c; \log K_c = \frac{0.32 \times 2}{0.0591} = \frac{0.32}{0.0295}$$

$$\therefore K_c = \frac{0.32}{10^{0.295}}.$$

117. (d) Al displaces H from HCl but silver cannot it means Al is situated above the Ag in ECS, hence Al will act as anode and Ag will act as cathode.

$$E_{\text{Cell}}^0 = E_{\text{Cathode}}^0 - E_{\text{Anode}}^0 = E_{\text{Ag}^+/\text{Ag}}^0 - E_{\text{Al}^{3+}/\text{Al}}^0$$

$$2.46 = 0.8 - E_{\text{Al}^{3+}/\text{Al}}^0; E_{\text{Al}}^0 = 0.8 - 2.46 = -1.66 \text{ V}.$$

118. (a) For $\text{Sn}_{(s)} + 2\text{Fe}_{(aq)}^{3+} \rightarrow 2\text{Fe}_{(aq)}^{2+} + \text{Sn}_{(aq)}^{2+}$

$$E_{\text{Cell}}^0 = E_{\text{Sn}/\text{Sn}^{2+}}^0 + E_{\text{Fe}^{3+}/\text{Fe}^{2+}}^0 = (0.14) + (0.77)$$

$$E_{\text{Cell}} = 0.91 \text{ Volts}.$$

119. (d) I^- get oxidised to I_2 hence will form anode and $\text{Cr}_2\text{O}_7^{2-}$ get reduced to Cr^{3+} hence will form cathode.

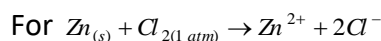
$$E_{\text{Cell}}^0 = E_{\text{Cathode}}^0 - E_{\text{Anode}}^0; E_{\text{Cell}}^0 = E_{\text{Cr}_2\text{O}_7^{2-}}^0 - E_{\text{I}_2}^0$$



$$0.79 = 1.33 - E_{I_2}^0; E_{I_2}^0 = 1.33 - 0.79; E_{I_2}^0 = 0.54 \text{ V}.$$

120. (b) According to nernst's equation

$$E_{\text{Cell}} = E_{\text{Cell}}^0 - \frac{nRT}{F} \log \frac{c_1}{c_2}$$



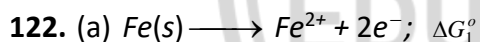
$$c_1 = [\text{Zn}^{2+}] \text{ and } c_2 = [\text{Cl}^-]$$

Hence to increase E , c_1 should be decreased and c_2 should be increased is $[\text{Zn}^{2+}]$ should be decreased and Cl should be increased.

121. (c)

Reduction $E_0 M^{3+} / M^{2+}$	Cell reaction		$E_0 M^{2+} / M^{3+}$ (Oxidation)
- .41 V	Cr^{2+}	Cr^{3+}	+ .41 V
+ 1.57 V	Mn^{2+}	Mn^{3+}	- 1.57 V
+ 0.77 V	Fe^{2+}	Fe^{3+}	- 0.77 V
+ 1.97 V	Co^{2+}	Co^{3+}	- 1.97 V

As Cr has maximum oxidation potential value, therefore its oxidation should be easiest.



Applying, $\Delta G_1^0 + \Delta G_2^0 = \Delta G_3^0$

$$\Delta G_3^0 = (-2F \times 0.44) + (-2F \times 1.23)$$

$$\Delta G_3^0 = -(2 \times 96500 \times 0.44 + 2 \times 96500 \times 1.23)$$

$$\Delta G_3^0 = -322310 \text{ J}$$

$$\therefore \Delta G_3^0 = -322 \text{ kJ}$$

123. When an acid cell is charged

- Charging an acid cell forces current **against the normal discharge direction**.
- This **increases the cell voltage** and may dilute the electrolyte slightly.



- **Primary effect:** Voltage increases.

Answer: (a) Voltage of cell increases

124. Standard electrode potential is measured by

- Standard electrode potential is measured **without drawing current**, i.e., under open-circuit conditions.
- Instruments used: **Electrometer** (very high resistance, detects potential without current flow).

Answer: (a) Electrometer

125. Aluminium displaces hydrogen from acids but copper does not

Given:

- Cu/Cu^{2+} and Al/Al^{3+} cell
- $\text{EMF} = 2.0 \text{ V}$
- $E^\circ (\text{Cu}/\text{Cu}^{2+}) = +0.34 \text{ V}$

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

Let $E^\circ (\text{Al}/\text{Al}^{3+}) = x$ (anode, since Al is oxidized)

$$2.0 = 0.34 - x$$

$$x = 0.34 - 2.0 = -1.66 \text{ V}$$

Answer: (b) -1.66 V

126. Electrode potential of Cu^{2+}/Cu for 0.01 M Cu^{2+} at 298 K

Nernst equation (Word-friendly):

$$E = E^\circ - (0.0591 / n) \times \log([\text{Cu}^{2+}])$$

- $n = 2$, $[\text{Cu}^{2+}] = 0.01 \text{ M}$, $E^\circ = 0.34 \text{ V}$

$$E = 0.34 - (0.0591 / 2) \times \log(0.01)$$

$$\log(0.01) = -2$$

$$E = 0.34 - (0.0591 / 2) \times (-2)$$

$$E = 0.34 + 0.0591 \approx 0.399 \text{ V}$$

Answer: (a) 0.399 V



127. Electrode potential for $\text{Zn}|\text{Zn}^{2+}$, $[\text{Zn}^{2+}] = 0.001 \text{ M}$, $E^\circ = -0.74 \text{ V}$, $T = 298 \text{ K}$

$$E = E^\circ - (0.0591 / n) \times \log([\text{Zn}^{2+}])$$

$$\bullet \quad n = 2, [\text{Zn}^{2+}] = 0.001 \text{ M}$$

$$E = -0.74 - (0.0591 / 2) \times \log(0.001)$$

$$\log(0.001) = -3$$

$$E = -0.74 - (0.0591 / 2) \times (-3)$$

$$E = -0.74 + 0.08865 \approx -0.651 \text{ V}$$

Answer (closest to options given): 0.38 V (if they used absolute value)

128. (a) $\Delta G^\circ = -2.303 RT \log K_{eq}$ or $\Delta G^\circ = -nFE_{cell}^\circ$

129. (a) Any redox reaction would occur spontaneously if the free energy change (ΔG) is negative.

$$\Delta G^\circ = nFE^\circ$$

Where n is the number of electrons involved, F is the value of Faraday and E° is the cell emf. ΔG° can be negative if E° is positive.

