

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

101. What is the potential of a half-cell consisting of zinc electrode in 0.01M ZnSO_4 solution at 25°C ($E^\circ = 0.763\text{V}$)
- (a) 0.8221 V (b) 8.221 V
(c) 0.5282 V (d) 9.232 V
102. The *emf* of a galvanic cell, with electrode potentials of silver = $+0.80\text{ V}$ and that of copper = $+0.34\text{ V}$, is
- (a) -1.1 V (b) $+1.1\text{ V}$
(c) $+0.46\text{ V}$ (d) $+0.76\text{ V}$
103. Copper cannot replace..... from solution
- (a) Fe (b) Au
(c) Hg (d) Ag
104. The strongest reducing agent of the alkali metal is
- (a) Li (b) Na
(c) K (d) Cs
105. Which of the following is the most electropositive element
- (a) Carbon (b) Calcium
(c) Chlorine (d) Potassium
106. The metal that forms a self protecting film of oxide to prevent corrosion, is
- (a) Cu (b) Al
(c) Na (d) Au
107. In a cell that utilises the reaction $\text{Zn}_{(\text{s})} + 2\text{H}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{H}_{2(\text{g})}$ addition of H_2SO_4 to cathode compartment, will
- (a) Increase the E and shift equilibrium to the right
(b) Lower the E and shift equilibrium to the right
(c) Lower the E and shift equilibrium to the left
(d) Increase the E and shift equilibrium to the left
108. For the electrochemical cell, $\text{M}|\text{M}^+||\text{X}^-|\text{X}$, $E^\circ(\text{M}^+|\text{M}) = 0.44\text{ V}$ $E^\circ(\text{X}|\text{X}^-) = 0.33\text{ V}$ From this data, one can deduce that
- (a) $E^\circ_{\text{cell}} = -0.77\text{V}$
(b) $\text{M}^+ + \text{X}^- \rightarrow \text{M} + \text{X}$ is the spontaneous reaction
(c) $\text{M} + \text{X} \rightarrow \text{M}^+ + \text{X}^-$ is the spontaneous reaction
(d) $E^\circ_{\text{cell}} = .77\text{ V}$
109. The standard *e.m.f.* of a cell, involving one electron change is found to be 0.591 V at 25°C . The equilibrium constant of the reaction is



($F = 96,500 \text{ C mol}^{-1}$; $R = 8.314 \text{ JK}^{-1}\text{mol}^{-1}$)

- (a) 1.0×10^{10} (b) 1.0×10^5
(c) 1.0×10^1 (d) 1.0×10^{30}

110. Standard electrode potential of cell $H_2|H^+||Ag^+|Ag$ is

- (a) 0.8 V (b) -0.8 V
(c) -1.2 V (d) 1.2 V

111. A galvanic cell with electrode potential of 'A' = +2.23 V and 'B' = -1.43 V. The value of E°_{cell} is

- (a) 3.66 V (b) 0.80 V
(c) -0.80 V (d) -3.66 V

112. The e.m.f. of a cell whose half cells are given below is $Mg^{2+} + 2e^- \rightarrow Mg(s)$ $E^\circ = -2.37 \text{ V}$ $Cu^{2+} + 2e^- \rightarrow Cu(s)$ $E^\circ = +0.34 \text{ V}$

- (a) +1.36 V (b) +2.71 V
(c) +2.17 V (d) -3.01 V

113. For the cell reaction, $2Ce^{4+} + Co \rightarrow 2Ce^{3+} + Co^{2+}$ E°_{cell} is 1.89 V. If $E^\circ_{Ce^{4+}/Ce^{3+}}$

- (a) -1.64 V (b) +1.64 V
(c) -2.08 V (d) +2.17 V

114. If the ΔG of a cell reaction $AgCl + e^- \rightarrow Ag + Cl^-$ is -21.20 KJ; the standard e.m.f., of cell is

- (a) 0.229 V (b) 0.220 V
(c) -0.220 V (d) -0.110 V

115. The e.m.f. of the cell $Ag|Ag^+(0.1M)||Ag^+(1M)|Ag$ at 298 K is

- (a) 0.0059 V (b) 0.059 V
(c) 5.9 V (d) 0.59 V

116. The e.m.f. of the cell $Zn|Zn^{2+}(0.01M)||Fe^{2+}(0.001M)|Fe$ at 298 K is 0.2905 then the value of equilibrium for the cell reaction is

- (a) $\frac{0.32}{e^{0.0295}}$ (b) $\frac{0.32}{10^{0.0295}}$
(c) $\frac{0.26}{10^{0.0295}}$ (d) $\frac{0.32}{10^{0.0591}}$

117. Aluminium displaces hydrogen from dilute HCl whereas silver does not. The e.m.f. of a cell prepared by combining Al/Al^{3+} and Ag/Ag^+ is 2.46 V. The reduction potential of silver electrode is +0.80 V. The reduction potential of aluminium electrode is

- (a) +1.66 V (b) -3.26V
(c) 3.26 V (d) -1.66 V

118. Consider the following E° values :

$$E^\circ_{Fe^{3+}/Fe^{2+}} = +0.77 \text{ V}$$

$$E^\circ_{Sn^{2+}/Sn} = -0.14 \text{ V}$$



- Under standard conditions the potential for the reaction $\text{Sn}_{(s)} + 2\text{Fe}^{3+}(\text{aq}) \rightarrow 2\text{Fe}^{2+}(\text{aq}) + \text{Sn}^{2+}(\text{aq})$ is
- (a) 0.91 V (b) 1.40 V
(c) 1.68 V (d) 0.63 V
119. $\text{Cr}_2\text{O}_7^{2-} + \text{I}^- \rightarrow \text{I}_2 + \text{Cr}^{3+}$
 $E^\circ_{\text{cell}} = 0.79 \text{ V}$
 $E^\circ_{\text{Cr}_2\text{O}_7^{2-}} = 1.33 \text{ V}$, $E^\circ_{\text{I}_2}$ is
- (a) -0.10 V (b) $+0.18 \text{ V}$
(c) -0.54 V (d) 0.54 V
120. $\text{Zn}(\text{s}) + \text{Cl}_2(1 \text{ atm}) \rightarrow \text{Zn}^{2+} + 2\text{Cl}^-$.
 E°_{cell} of the cell is 2.12 V. To increase E
- (a) $[\text{Zn}^{2+}]$ should be increased
(b) $[\text{Zn}^{2+}]$ should be decreased
(c) $[\text{Cl}^-]$ should be decreased
(d) P_{Cl_2} should be decreased
121. The $E^\circ_{\text{M}^{3+}/\text{M}^{2+}}$ values for Cr, Mn, Fe and Co are -0.41 , $+1.57$, $+0.77$ and $+1.97 \text{ V}$ respectively. For which one of these metals the change in oxidation state from +2 to +3 is easiest
- (a) Fe (b) Mn
(c) Cr (d) Co
122. The rusting of iron takes place as follows
 $2\text{H}^+ + 2\text{e}^- + \frac{1}{2}\text{O}_2 \longrightarrow \text{H}_2\text{O}(\text{l})$;
 $E^\circ = +1.23 \text{ V}$
 $\text{Fe}^{2+} + 2\text{e}^- \longrightarrow \text{Fe}(\text{s})$; $E^\circ = -0.44 \text{ V}$
 Calculate ΔG° for the net process
- (a) -322 kJ mol^{-1}
(b) -161 kJ mol^{-1}
(c) -152 kJ mol^{-1}
(d) -76 kJ mol^{-1}
123. When an acid cell is charged, then
- (a) Voltage of cell increases
(b) Electrolyte of cell dilutes
(c) Resistance of cell increases
(d) None of these
124. The standard electrode potential is measured by
- (a) Electrometer
(b) Voltmeter
(c) Pyrometer
(d) Galvanometer
125. Aluminium displaces hydrogen from acids but copper does not. A galvanic cell prepared by combining Cu/Cu^{2+} and Al/Al^{3+} has an e.m.f. of 2.0 V at 298 K. If the potential of copper electrode is $+0.34 \text{ V}$, that of aluminium is
- (a) $+1.66 \text{ V}$ (b) -1.66 V
(c) $+2.34 \text{ V}$ (d) -2.3 V



126. If the standard electrode potential of Cu^{2+}/Cu electrode is $0.34V$, what is the electrode potential of $0.01M$ concentration of Cu^{2+} ($T = 298K$)
 (a) $0.399 V$ (b) $0.281 V$
 (c) $0.222 V$ (d) $0.176 V$
127. Calculate the electrode potential at $298^{\circ}K$ for $Zn|Zn^{++}$ electrode in which the activity of zinc ions is $0.001 M$ and $E^{\circ}_{Zn/Zn^{++}}$ is -0.74 volts
 (a) 0.38 volts (b) 0.83 volts
 (c) 0.40 volts (d) 0.45 volts
128. Which of the following expression is correct
 (a) $\Delta G^{\circ} = -nFE^{\circ}_{cell}$
 (b) $\Delta G^{\circ} = +nFE^{\circ}_{cell}$
 (c) $\Delta G^{\circ} = -2.303RTnFE^{\circ}_{cell}$
 (d) $\Delta G^{\circ} = -nF \log K_c$
129. For the feasibility of a redox reaction in a cell, the e.m.f. should be
 (a) Positive (b) Fixed
 (c) Zero (d) Negative

