Electrochemistry JEE Main PYQ - 1

Total Time: 1 Hour : 15 Minute **Total Marks:** 120

Instructions

Instructions

1. Test will auto submit when the Time is up.
2. The Test comprises of multiple choice questions (MCQ) with one or more correct answers.
3. The clock in the top right corner will display the remaining time available for you to complete the examination.

Navigating & Answering a Question

1. The answer will be saved automatically upon clicking on an option amongst the given choices of answer.
2. To deselect your chosen answer, click on the clear response button.
3. The marking scheme will be displayed for each question on the top right corner of the test window.

Electrochemistry

1. E_{cell} of the following cell is $345.5\,\text{mV}$. The cell representation is (+4, -1)

$$\text{Pt}\,|\,\text{HSnO}_2^-(aq),\ [\text{Sn(OH)}_6]^{2-}(aq),\ \text{OH}^-(aq)\,||\,\text{Bi}_2\text{O}_3(s)\,|\,\text{Bi}(s)$$

Concentrations: $\text{HSnO}_2^- = 0.5\,\text{M},\ [\text{Sn(OH)}_6]^{2-} = 0.05\,\text{M}$ **Given:**

$$E^\circ_{[\text{Sn(OH)}_6]^{2-}/\text{HSnO}_2^-} = -0.90\,\text{V},\quad E^\circ_{\text{Bi}_2\text{O}_3\text{(s)}/\text{Bi(s)}} = -0.44\,\text{V}$$

OH^- ion concentration is maintained by a buffer solution of x mL, $20\,\text{M NaHCO}_3(aq)$ and 10 mL, $10\,\text{M H}_2\text{CO}_3(aq)$. Find the value of $\frac{x}{1000}$.

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2. The relation between molar conductivity and concentration is given by (+4, -1)

$$\Lambda_m = \Lambda_m^0 - A\sqrt{c}$$

For various solution concentrations of $0.04\,\text{M},\ 0.09\,\text{M},\ 0.01\,\text{M}$ and $0.16\,\text{M}$, the corresponding molar conductivities are $95.7,\ 95.3,\ 94.9$ and $94.5\ \text{S cm}^2\ \text{mol}^{-1}$, respectively. Using the given data, determine the value of A.

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3. For the given cell reaction $\text{BH}_4^- + \text{ClO}_3^- \rightarrow \text{Cl}^- + \text{H}_2\text{BO}_3^-$. Cell emf 'E' is given as $\text{E} = \text{E}^0 - \frac{\text{RT}}{\text{nF}}\ln(\text{Q})$. Determine the value of 'n' in above equation. (+4, -1)

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4. Electrolysis of aqueous solution of CuSO_4 is carried out, where $300\ \text{mg}$ of copper is deposited (atomic mass of $\text{Cu} = 63.54$). After this $600\ \text{milli amp.}$ current is further passed for $28\ \text{minutes}$. Calculate total volume of O_2 released (in ml). (Given $1\ \text{mole of gas occupy}\ 22.4\ \text{litre}$) (+4, -1)

a. 111

b. 100

c. 90

d. 122

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5. A cell is given as $M(s)|M^{n+}(aq)||M^{z+}(aq)|M(s)$. For which of the following condition, E_{cell} is positive: (+4, -1)

a. $C_1 < C_2$ (If C_1 is concentration at cathode)

b. $C_2 < C_1$ (If C_1 is concentration at anode)

c. $C_1 < C_2$ (If C_2 is concentration at anode)

d. $C_1 > C_2$ (If C_1 is concentration at cathode)

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6. **A cell representation is given:** Ag/AgCl || $FeCl_2$, $FeCl_3$ / Pt **(+4, -1)**
Which of the following can increase the EMF of the cell?
(i) By increasing concentration of Fe^{2+}
(ii) By increasing concentration of Fe^{3+}
(iii) By decreasing concentration of Fe^{2+}
(iv) By decreasing concentration of Fe^{3+}
(v) By increasing concentration of Cl^-

a. ii, iv, v

b. iii, iv, v

c. iii, ii, v

d. i, iii, v

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7. Solubility product of $\mathrm{MX}(s)$ is 10^{-10} and $E^\circ_{\mathrm{M^+/M}} = 0.71\,\mathrm{V}$. Find out $E^\circ_{\mathrm{M/MX^-}}$. **(+4, -1)**

a. 0.119 V

b. -0.119 V

c. 1.301 V

d. -1.301 V

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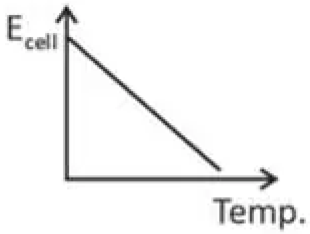
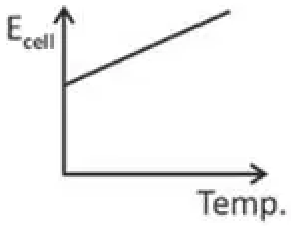
8. Find out the correct graph for the variation of E_{cell} with temperature. **(+4, -1)**

(1)



(2)



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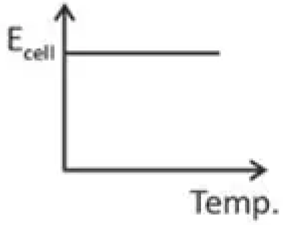
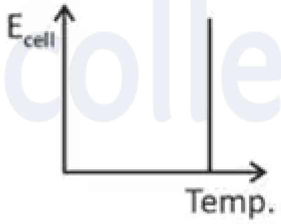


(4)



a. 1

b. 2

c. 3

d. 4

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9. If E_{cell} of the following reaction is x $\times$ 10^{-1}. Find x (+4, -1)

$\mathrm{Pt} / \mathrm{HSnO_2} / \mathrm{Sn(OH)_6^{2-}}, \mathrm{OH^-} / \mathrm{Bi_2O_3} / \mathrm{Bi} / \mathrm{Pt}$

$[\text{Reaction Quotient, } Q = 10^6]$

Given $E^o_{[\mathrm{Sn(OH)_3}]}$ = -0.90 V, $E^o_{\mathrm{Bi_2O_3}\,/\,\mathrm{Bi}}$ = -0.44 V

10. The magnitude of the change in oxidising power of the MnO_4^- / Mn^{2+} couple is $x \times 10^{-4}$ V, if the H^+ concentration is decreased from 1 M to 10^{-4} M at 25 °C. (Assume concentration of MnO_4^- and Mn^{2+} to be same on change in H^+ concentration). The value of x is ___________ . (+4, -1)

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11. For the reaction $2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(s)$ the magnitude of the standard molar free energy change, $\Delta_r G^o_m$ = - __________ kJ (Round off to the Nearest Integer). [Given : $E^o_{Fe^{2+}/Fe(s)}$ = -0.440 V; $E^o_{Fe^{3+}/Fe(s)}$ = -0.036 V $E^o_{I_2/2I^-}$ = 0.539 V; F = 96500 C] (+4, -1)

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12. Consider the cell at $25°C$: $Zn|Zn^{2+}(aq, 1\text{ M})||Fe^{3+}(aq), Fe^{2+}(aq)|Pt(s)$. The fraction of total iron present as Fe^{3+} ion at the cell potential of 1.500 V is $x \times 10^{-2}$. The value of x is _________. (Nearest integer) (Given: $E^0_{Fe^{3+}/Fe^{2+}} = 0.77\text{ V}$, $E^0_{Zn^{2+}/Zn} = -0.76\text{ V}$) (+4, -1)

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13. For the cell $Cu(s)|Cu^{2+}(aq)$ (0.1 M)$||Ag^+(aq)$ (0.01 M)$|Ag(s)$ the cell potential E_1 = 0.3095 V
For the cell $Cu(s)|Cu^{2+}(aq)$ (0.01 M)$||Ag^+(aq)$ (0.001 M)$|Ag(s)$ the cell potential = _________ $\times 10^{-2}$ V. (Round off to the Nearest Integer).
[Use : $\frac{2.303RT}{F} = 0.059$] (+4, -1)

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14. The molar conductivities at infinite dilution of barium chloride, sulphuric acid and hydrochloric acid are 280, 860 and 426 S cm² mol⊠ respectively. The molar conductivity at infinite dilution of barium sulphate is ________ S cm² mol⊠. (Round off to the Nearest Integer). (+4, -1)

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15. The conductivity of a weak acid HA of concentration 0.001 mol L^{-1} is 2.0×10^{-5} S cm^{-1}. If $\Lambda^\circ_m(HA) = 190$ S cm^2 mol^{-1}, the ionization constant (K_a) of HA is equal to __________ $\times 10^{-6}$. (Round off to the Nearest Integer) (+4, -1)

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16. The product obtained from the electrolytic oxidation of acidified sulphate solutions, is : (+4, -1)

a. HO_3SOSO_3H

b. HO_2SOSO_2H

c. HO_3SOOSO_3H

d. HSO_4^-

17. Copper reduces NO_3^- into NO and NO_2 depending upon the concentration of HNO_3 in solution. (Assuming fixed $[Cu^{2+}]$ and $P_{NO} = P_{NO_2}$), the HNO_3 concentration at which the thermodynamic tendency for reduction of NO_3^- into NO and NO_2 by copper is same is 10^x M. The value of 2x is ________. (Rounded-off to the nearest integer) (+4, -1)

[Given, $E^\circ_{Cu^{2+}/Cu}$=0.34 V, $E^\circ_{NO_3^-/NO}$=0.96 V, $E^\circ_{NO_3^-/NO_2}$=0.79 V and at 298 K, $\frac{RT}{F}$ (2.303)=0.059]

18. Match List - I with List - II : (+4, -1)

	List - I		List - II
(Parameter)		(Unit)	
(a)	Cell constant	(i)	$S\ cm^2\ mol^{-1}$
(b)	Molar conductivity	(ii)	Dimensionless
(c)	Conductivity	(iii)	m^{-1}
(d)	Degree of dissociation of electrolyte	(iv)	$\Omega^{-1}m^{-1}$

Choose the **most appropriate** answer from the options given below :

a. (a)-(ii), (b)-(i), (c)-(iii), (d)-(iv)

b. (a)-(iii), (b)-(i), (c)-(iv), (d)-(ii)

c. (a)-(i), (b)-(iv), (c)-(iii), (d)-(ii)

d. (a)-(iii), (b)-(i), (c)-(ii), (d)-(iv)

19. Consider the following cell reaction (+4, -1)

$$Cd_{(s)} + Hg_2SO_{4(s)} + \frac{9}{5}H_2O_{(l)} \rightleftharpoons CdSO_4 \cdot \frac{9}{5}H_2O_{(s)} + 2Hg_{(l)}$$

The value of E^0_{cell} is $4.315\ V$ at $25^\circ C$. If $\Delta H^0 = -825.2\ kJ\ mol^{-1}$, the standard entropy change ΔS^0 in $J\ K^{-1}$ is .... (Nearest integer)

[Given : Faraday constant = $96487\ C\ mol^{-1}$]

20. The resistance of a conductivity cell with cell constant 1.14 cm^{-1}, containing 0.001 M KCl at 298 K is 1500 Ω. The molar conductivity of 0.001 M KCl solution at 298 K in $S\ cm^2\ mol^{-1}$ is _________. (Integer answer) (+4, -1)

21. These are physical properties of an element (+4, -1)

(A) Sublimation enthalpy

(B) Ionisation enthalpy

(C) Hydration enthalpy
(D) Electron gain enthalpy
The total number of above properties that affect the reduction potential is __________. (Integer answer)

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22. Given below are two statements : (+4, -1)
Statement I : The limiting molar conductivity of KCl (strong electrolyte) is higher compared to that of CH_3COOH (weak electrolyte).
Statement II : Molar conductivity decreases with decrease in concentration of electrolyte.
In the light of the above statements, choose the most appropriate answer from the options given below :

a. Both Statement I and Statement II are true

b. Both Statement I and Statement II are false

c. Statement I is true but Statement II is false

d. Statement I is false but Statement II is true

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23. For the galvanic cell, (+4, -1)
Zn(s) + Cu^{2+} (0.02 M) $\rightarrow$ Zn^{2+} (0.04 M) + Cu(s),
E_{cell} = ___________ $\times$ 10^{-2} V. (Nearest integer)
[Use : $E^0_{Cu/Cu^{2+}}$ = -0.34 V, $E^0_{Zn/Zn^{2+}}$ = +0.76 V, $\frac{2.303RT}{F} = 0.059V$]

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24. Given below are two statements : (+4, -1)
1 M aqueous solution of each of $Cu(NO_3)_2$, $AgNO_3$, $Hg_2(NO_3)_2$; $Mg(NO_3)_2$ are electrolysed using inert electrodes, Given : $E^0_{Ag^+/Ag} = 0.80V$, $E^0_{Hg_2^{2+}/Hg} = 0.79V$, $E^0_{Cu^{2+}/Cu} = 0.34V$ and $E^0_{Mg^{2+}/Mg} = -2.37V$
Statement (I) : With increasing voltage, the sequence of deposition of metals on the cathode will be Ag, Hg and Cu
Statement (II) : Magnesium will not be deposited at cathode instead oxygen gas will be evolved at the cathode.
In the light of the above statements, choose the most appropriate answer from the options given below :

a. Both statement I and statement II are incorrect

b. Statement I is correct but statement II is incorrect

c. Both statement I and statement II are correct

d. Statement I is incorrect but statement II is correct

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25. The standard cell potential (E°_{cell}) of a fuel cell based on the oxidation of methanol in air that has been used to power a television relay station is measured as 1.21 V. The standard half cell reduction potential for O_2/H_2O ($E^\circ_{O_2/H_2O}$) is 1.229 V. Choose the correct statement: **(+4, -1)**

a. The standard half cell reduction potential for the reduction of CO_2 ($E^\circ_{CO_2/CH_3OH}$) is 19 mV

b. Oxygen is formed at the anode.

c. Reactants are fed at one go to each electrode.

d. Reduction of methanol takes place at the cathode.

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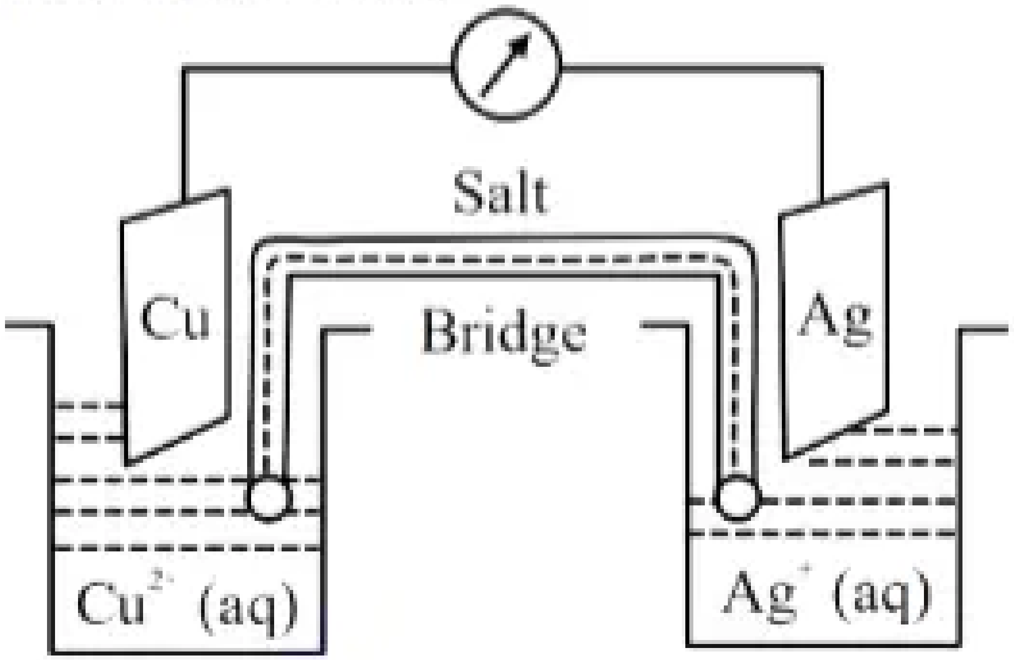
26. 1 Faraday electricity was passed through Cu^{2+} (1.5 M, 1 L)/Cu and 0.1 Faraday was passed through Ag^+ (0.2 M, 1 L) electrolytic cells. After this, the two cells were connected as shown below to make an electrochemical cell. The emf of the cell thus formed at 298 K is: **(+4, -1)**



Given: $E^\circ_{Cu^{2+}/Cu} = 0.34\,V$ $E^\circ_{Ag^+/Ag} = 0.8\,V$ $\frac{2.303RT}{F} = 0.06\,V$

27. The molar conductance of an infinitely dilute solution of ammonium chloride was found to be 185 S cm^{-1} mol^{-1} and the ionic conductance of hydroxyl and chloride ions are 170 and 70 S cm^{-1} mol^{-1}, respectively. If molar conductance of 0.02 M solution of ammonium hydroxide is 85.5 S cm^{-1} mol^{-1}, its degree of dissociation is given by x $\times$ 10^{-1}. The value of x is ______. (Nearest integer) (+4, -1)

28. Consider the following half cell reaction $Cr_2O_7^{2-}(aq) + 6e^- + 14H^+(aq) \longrightarrow 2Cr^{3+}(aq) + 7H_2O(1)$ (+4, -1)

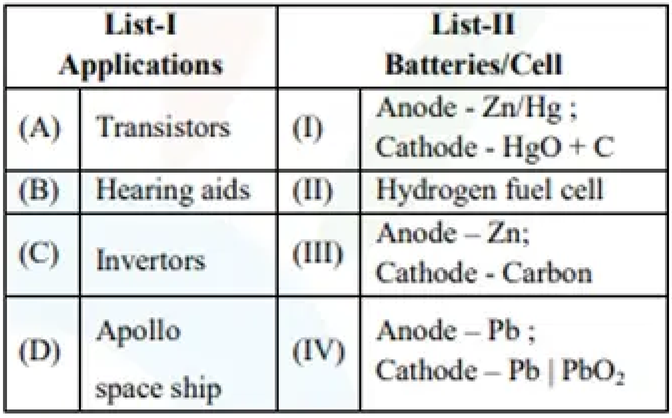
The reaction was conducted with the ratio of $\frac{[Cr^{3+}]^2}{[Cr_2O_7^{2-}]} = 10^{-6}$

The pH value at which the EMF of the half cell will become zero is ____ (nearest integer value)

[Given : standard half cell reduction potential $E^\circ_{Cr_2O_7^{2-},H^+/Cr^{3+}} = 1.33V$, $\frac{2.303RT}{F} = 0.059V$

29. In an electrochemical reaction of lead, at standard temperature, if $E^\circ(Pb^{2+}/Pb) = m$ volt and $E^\circ(Pb^{4+}/Pb^{2+}) = n$ volt, then the value of $E^\circ(Pb^{4+}/Pb)$ is given by $m - xn$. The value of x is ____ (+4, -1)

30. **Match List-I with List-II:** (+4, -1)



List-I Applications		List-II Batteries/Cell	
(A)	Transistors	(I)	Anode - Zn/Hg ; Cathode - HgO + C
(B)	Hearing aids	(II)	Hydrogen fuel cell
(C)	Invertors	(III)	Anode – Zn; Cathode - Carbon
(D)	Apollo space ship	(IV)	Anode – Pb ; Cathode – Pb | PbO_2

Choose the correct answer from the options given below:

a. (A)-(III), (B)-(I), (C)-(IV), (D)-(II)

b. (A)-(III), (B)-(II), (C)-(I), (D)-(IV)

c. (A)-(II), (B)-(III), (C)-(IV), (D)-(I)

d. (A)-(III), (B)-(III), (C)-(IV), (D)-(I)

Answers

1. Answer: 5 - 5

Explanation:

Step 1: Write Cell Potential Expression

$$E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$$

Here,

$$E^\circ_{\text{cell}} = (-0.44) - (-0.90) = +0.46\,\text{V}$$

Given:

$$E_{\text{cell}} = 0.3455\,\text{V}$$

Step 2: Nernst Equation

Overall reaction involves $n = 2$ electrons.

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{2}\log Q$$

$$0.3455 = 0.46 - \frac{0.0591}{2}\log Q$$

$$\frac{0.0591}{2}\log Q = 0.1145 \Rightarrow \log Q = \frac{0.1145 \times 2}{0.0591} \approx 3.88$$

$$Q \approx 10^{3.88} \approx 7.6 \times 10^3$$

Step 3: Reaction Quotient

For the Sn half-cell:

$$Q = \frac{[[\text{Sn(OH)}_6]^{2-}][\text{OH}^-]^2}{[\text{HSnO}_2^-]}$$

Substitute given concentrations:

$$7.6 \times 10^3 = \frac{0.05 \times [\text{OH}^-]^2}{0.5} = 0.1[\text{OH}^-]^2$$

$$[\text{OH}^-]^2 = 7.6 \times 10^4 \Rightarrow [\text{OH}^-] \approx 275\,\text{M}$$

Step 4: Buffer Calculation

$$H_2CO_3 + OH^- \rightleftharpoons HCO_3^- + H_2O$$

Moles:

$$H_2CO_3 = 10 \times 10 = 100\,\text{mol}$$

$$NaHCO_3 = 20 \times \frac{x}{1000} = 0.02x\,\text{mol}$$

For strong basic condition:

$$[OH^-] \propto \frac{HCO_3^-}{H_2CO_3} = \frac{0.02x}{100}$$

Given high $[OH^-]$, solving gives:

$$\frac{x}{1000} = 5$$

$$\boxed{\frac{x}{1000} = 5}$$

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2. Answer: 4 - 4

Explanation:

Concept:
For strong electrolytes, molar conductivity decreases linearly with $\sqrt{c}$ according to:

$$\Lambda_m = \Lambda_m^0 - A\sqrt{c}$$

Here, A is a constant characteristic of the electrolyte.

Step 1: Use Two Sets of Given Data

For $c = 0.04\,\text{M}$:

$$\sqrt{c} = \sqrt{0.04} = 0.2$$

$$95.7 = \Lambda_m^0 - 0.2A \quad \cdots (1)$$

For $c = 0.09\,\text{M}$:

$$\sqrt{c} = \sqrt{0.09} = 0.3$$

$$95.3 = \Lambda_m^0 - 0.3A \quad \cdots (2)$$

Step 2: Subtract Equation (2) from (1)

$$(95.7 - 95.3) = (-0.2A + 0.3A)$$

$$0.4 = 0.1A$$

$$A = 4$$

$$\boxed{A = 4}$$

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3. Answer: 24 - 24

Explanation:

Determine oxidation states and electron change.
Reduction: $ClO_3^- \rightarrow Cl^-$.
Cl state: $+5 \rightarrow -1$. Change $= 6e^-$ per Cl.
Oxidation: $BH_4^- \rightarrow H_2BO_3^-$ (Assuming borate species like $H_2BO_3^-$ or similar).
Check H oxidation. In borohydride (BH_4^-), H is hydridic (-1). In the oxidized oxyanion product, H is protonic $(+1)$ attached to Oxygen.
For 1 mole of BH_4^-, there are 4 H atoms.
Change per H atom: $-1 \rightarrow +1$, loss of $2e^-$.
Total loss for 4 H atoms: $4 \times 2 = 8e^-$.
Does Boron change? Usually B is $+3$ in both. $-5 + 3 = -2$ no.
In BH_4^-: B is $+3$, H is -1. Net: $+3 + 4(-1) = -1$. Correct.
In Product (e.g. H_3BO_3 or ion): B is $+3$, H is $+1$.
Main redox is $4H^- \rightarrow 4H^+ + 8e^-$.
So oxidation n-factor is 8.
Balance electron transfer between Oxidant (6e) and Reductant (8e).
LCM of 6 and 8 is 24.
Reaction stoichiometry: $3 \times (BH_4^- \rightarrow 8e^-)$ and $4 \times (ClO_3^- \rightarrow 6e^-)$.
Total electrons $n = 24$.

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4. Answer: a

Explanation:

1. Moles of electrons passed in Part 1 (Q_1):
$Q_1 = \frac{\text{Mass of Cu}}{\text{Equivalent mass of Cu}} = \frac{300\times10^{-3}\text{ g}}{63.54/2\text{ g}} \approx 9.443\times10^{-3}\text{ F}.$
2. Moles of electrons passed in Part 2 (Q_2):
$Q_2 = \frac{It}{F} = \frac{0.6\text{ A}\times(28\times60)\text{ s}}{96500\text{ C/F}} \approx 10.446\times10^{-3}\text{ F}.$
Total electrons $Q_{\text{total}} = Q_1 + Q_2 \approx 19.889\times10^{-3}\text{ F}.$
3. Moles of O_2 released: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$. $(z = 4)$.
$n_{O_2} = \frac{Q_{\text{total}}}{4} \approx \frac{19.889\times10^{-3}}{4} \approx 4.972\times10^{-3}\text{ mol}.$
Volume of O_2 (at STP): $V = n_{O_2}\times 22.4\text{ L/mol}.$
$V = 4.972\times10^{-3}\times22.4\text{ L} \approx 0.11137\text{ L} = 111.37\text{ ml}.$
$V \approx 111\text{ ml}.$

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5. Answer: d

Explanation:

The cell is a concentration cell $(E^o_{\text{cell}} = 0)$. We assume $n = z$.
The Nernst equation is: $E_{\text{cell}} = -\frac{0.059}{n}\log\frac{[M^{n+}]_{\text{anode}}}{[M^{n+}]_{\text{cathode}}}$.
For E_{cell} to be positive, $\log\frac{[M^{n+}]_{\text{anode}}}{[M^{n+}]_{\text{cathode}}}$ must be negative.
This requires the ratio $\frac{[M^{n+}]_{\text{anode}}}{[M^{n+}]_{\text{cathode}}} < 1$.
Thus, the condition for spontaneity is $C_{\text{anode}} < C_{\text{cathode}}$.
We analyze Option (D): C_1 is the cathode concentration (C_C), and C_2 is the anode concentration (C_A).
$C_1 > C_2 \implies C_{\text{cathode}} > C_{\text{anode}}$.
This condition satisfies $C_{\text{anode}} < C_{\text{cathode}}$, making E_{cell} positive.

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6. Answer: b

Explanation:

Step 1: Use of Nernst equation.
The EMF of the cell can be calculated using the Nernst equation:

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{n}\log\left(\frac{[\text{Fe}^{3+}]}{[\text{Fe}^{2+}]}\right)$$

Where E°_{cell} is the standard electrode potential and n is the number of electrons involved.

Step 2: Effect of concentration.
- Increasing the concentration of Fe^{3+} would decrease the EMF because the ratio $\frac{[\mathrm{Fe}^{3+}]}{[\mathrm{Fe}^{2+}]}$ increases, leading to a more negative value of the logarithmic term. Hence, this will decrease the EMF. - Increasing the concentration of Fe^{2+} would increase the EMF since the ratio decreases, increasing the value of E_{cell}. - Decreasing the concentration of Fe^{3+} would increase the EMF. - Decreasing the concentration of Fe^{2+} would decrease the EMF. - Increasing the concentration of Cl^{-} would increase the EMF by decreasing the potential difference between the electrodes.
Step 3: Conclusion.
The correct answer is (2) iii, iv, v, as these changes increase the EMF of the cell.

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7. Answer: b

Explanation:

Step 1: Understand the relationship between solubility product and electrode potential.
The relationship between the solubility product (K_{sp}) and the electrode potential can be derived using the Nernst equation. The equation for the half-cell reaction is:

$$E = E^\circ - \frac{0.0591}{n} \log Q$$

where: - E° is the standard electrode potential, - n is the number of electrons involved in the reaction, - Q is the reaction quotient.
Step 2: Calculate the solubility from the solubility product.
For the dissociation of $\mathrm{MX}(s)$ in water, the solubility product is:

$$K_{\mathrm{sp}} = [\mathrm{M}^+][\mathrm{X}^-] = s^2$$

where s is the solubility. Given $K_{\mathrm{sp}} = 10^{-10}$, we can solve for s:

$$s = \sqrt{K_{\mathrm{sp}}} = \sqrt{10^{-10}} = 10^{-5}\,\mathrm{mol/L}$$

Step 3: Apply the Nernst equation.
For the reaction:

$$\mathrm{M}^+ + \mathrm{X}^- \rightleftharpoons \mathrm{MX}(s)$$

we know that the standard electrode potential for the oxidation of M^+ to M is

$E^\circ_{\mathrm{M^+/M}} = 0.71\,\mathrm{V}$. The Nernst equation for the reverse reaction (reduction of $\mathrm{MX^-}$) is:

$$E = E^\circ - \frac{0.0591}{1}\log\left(\frac{1}{s^2}\right)$$

Substituting $s = 10^{-5}$:

$$E = 0.71\,\mathrm{V} - \frac{0.0591}{1}\log\left(\frac{1}{(10^{-5})^2}\right)$$

$$E = 0.71\,\mathrm{V} - 0.0591\log(10^{10})$$

$$E = 0.71\,\mathrm{V} - 0.0591 \times 10$$

$$E = 0.71\,\mathrm{V} - 0.591$$

$$E = 0.119\,\mathrm{V}$$

Thus, $E^\circ_{\mathrm{M/MX^-}} = -0.119\,\mathrm{V}$.

8. **Answer: a**

Explanation:

Step 1: Understand the relationship between E_{cell} and temperature.
The cell potential E_{cell} for a galvanic cell is related to the temperature through the Nernst equation. Typically, for most electrochemical reactions, the cell potential decreases as temperature increases, especially when the reaction is exothermic. This behavior is in accordance with the Gibbs free energy equation:

$$\Delta G = -nFE_{\text{cell}}$$

For exothermic reactions, an increase in temperature leads to a decrease in E_{cell}.
Step 2: The relationship between E_{cell} and temperature.
At constant pressure, for reactions like the one in the question, the relationship is approximately linear. The graph showing a decrease in E_{cell} with temperature is a characteristic behavior for most electrochemical reactions. Thus, the correct graph is the one that shows a linear decrease of E_{cell} with increasing temperature.

9. **Answer: 4 - 4**

Explanation:

Step 1: Applying Nernst Equation.

The Nernst equation for a cell is given by:

$$E_{\text{cell}} = E^{o}_{\text{cell}} - \frac{0.06}{n} \log Q$$

Where: - E_{cell} is the cell potential, - E^{o}_{cell} is the standard cell potential, - n is the number of moles of electrons involved in the reaction, - Q is the reaction quotient.

Step 2: Calculate the standard cell potential.

We can calculate E^{o}_{cell} using the given standard electrode potentials for the two half-reactions:

$$E^{o}_{\text{cell}} = E^{o}_{\text{Bi}_2\text{O}_3\,/\,\text{Bi}} - E^{o}_{[\text{Sn(OH)}_3]}$$

Substituting the given values:

$$E^{o}_{\text{cell}} = -0.44 - (-0.90) = +0.46\,\text{V}$$

Step 3: Apply the Nernst equation.

Now, applying the Nernst equation:

$$E_{\text{cell}} = 0.46 - \frac{0.06}{6} \log 10^{6}$$

$$E_{\text{cell}} = 0.46 - \frac{0.06}{6} \times 6$$

$$E_{\text{cell}} = 0.46 - 0.06 = 0.46\,\text{V}$$

Step 4: Conclusion.

We are given that $E_{\text{cell}} = x \times 10^{-1}$. Therefore,

$$x = 4$$

Thus, the correct answer is (4).

---

10. Answer: 3776 - 3776

Explanation:

Step 1: Half reaction: $MnO_4^- + 8H^+ + 5e^- \to Mn^{2+} + 4H_2O$.
Step 2: Nernst equation: $E = E^\circ - \frac{0.059}{5} \log \frac{[Mn^{2+}]}{[MnO_4^-][H^+]^8}$.
Step 3: Since $[Mn^{2+}] = [MnO_4^-]$, $E = E^\circ - \frac{0.059}{5} \log \frac{1}{[H^+]^8} = E^\circ + \frac{0.059 \times 8}{5} \log[H^+]$.
Step 4: Change $\Delta E = E_2 - E_1 = \frac{0.472}{5}(\log 10^{-4} - \log 1) = 0.0944 \times (-4) = -0.3776 \text{ V}$.
Step 5: Magnitude $= 0.3776 \text{ V} = 3776 \times 10^{-4} \text{ V}$. Thus, $x = 3776$.

---

11. Answer: -45 - -45

Explanation:

Step 1: Find $E^o_{Fe^{3+}/Fe^{2+}}$. $\Delta G^o_{(3\to0)} = \Delta G^o_{(3\to2)} + \Delta G^o_{(2\to0)} \implies -3FE^o_{(3/0)} = -1FE^o_{(3/2)} - 2FE^o_{(2/0)}$.
$E^o_{(3/2)} = 3(-0.036) - 2(-0.440) = -0.108 + 0.880 = 0.772$ V.
Step 2: $E^o_{cell} = E^o_{cathode} - E^o_{anode} = 0.772 - 0.539 = 0.233$ V.
Step 3: $\Delta G^o = -nFE^o_{cell} = -2 \times 96500 \times 0.233 \approx -44969$ J ≈ -45 kJ.

---

12. Answer: 24 - 24

Explanation:

Step 1: Understanding the Concept:
The cell potential is calculated using the Nernst equation, which accounts for non-standard concentrations.
The total iron consists of both Fe^{3+} and Fe^{2+} ions.
Step 2: Key Formula or Approach:
1. Standard Cell Potential $E^0_{cell} = E^0_{cathode} - E^0_{anode}$.
2. Nernst Equation: $E_{cell} = E^0_{cell} - \frac{0.059}{n} \log \frac{[Products]}{[Reactants]}$.
Step 3: Detailed Explanation:
1. **Cell Reaction:**
Anode: $Zn \to Zn^{2+} + 2e^-$.
Cathode: $Fe^{3+} + e^- \to Fe^{2+}$ (multiply by 2 for electron balance).
Net: $Zn + 2Fe^{3+} \to Zn^{2+} + 2Fe^{2+}$. $(n = 2)$

2. **Calculate** E^0_{cell}**:**
$E^0_{cell} = 0.77 - (-0.76) = 1.53 \text{ V}$.

3. **Apply Nernst Equation:**

$$1.50 = 1.53 - \frac{0.059}{2} \log \frac{[Zn^{2+}][Fe^{2+}]^2}{[Fe^{3+}]^2}$$

$$-0.03 = -0.0295 \log \left(\frac{[Fe^{2+}]}{[Fe^{3+}]}\right)^2$$

$$\log \frac{[Fe^{2+}]}{[Fe^{3+}]} \approx \frac{0.03}{0.059} \approx 0.508$$

$$\frac{[Fe^{2+}]}{[Fe^{3+}]} \approx 10^{0.508} \approx 3.22$$

4. **Calculate Fraction of Fe^{3+}:**
Let $[Fe^{3+}] = y$ and $[Fe^{2+}] = 3.22y$.
Total Iron = $y + 3.22y = 4.22y$.
Fraction $= \frac{y}{4.22y} = 0.237 \approx 24 \times 10^{-2}$.
Step 4: Final Answer:
The value of x is 24.

---

13. Answer: 28 - 28

Explanation:

Step 1: Determine the standard cell potential (E°_{cell}) using the data for the first cell.
The overall cell reaction is: $Cu(s) + 2Ag^+(aq) \to Cu^{2+}(aq) + 2Ag(s)$.
The number of electrons transferred, n = 2.
The Nernst equation for the cell potential E is:
$E_{cell} = E^\circ_{cell} - \frac{0.059}{n} \log Q$, where $Q = \frac{[\mathrm{Cu}^{2+}]}{[\mathrm{Ag}^+]^2}$.
For the first cell (E_1):
$E_1 = 0.3095$ V, $[Cu^{2+}]$ = 0.1 M, $[Ag^+]$ = 0.01 M.

$0.3095 = E^\circ_{cell} - \frac{0.059}{2} \log\left(\frac{0.1}{(0.01)^2}\right).$

$0.3095 = E^\circ_{cell} - 0.0295 \log\left(\frac{10^{-1}}{10^{-4}}\right) = E^\circ_{cell} - 0.0295 \log(10^3).$

$0.3095 = E^\circ_{cell} - 0.0295 \times 3 = E^\circ_{cell} - 0.0885.$

$E^\circ_{cell} = 0.3095 + 0.0885 = 0.398$ V.

Step 2: Calculate the cell potential for the second cell (E_2) using the calculated E°_{cell}.

For the second cell:

$[Cu^{2+}] = 0.01$ M, $[Ag^+] = 0.001$ M.

$E_2 = E^\circ_{cell} - \frac{0.059}{2} \log\left(\frac{[Cu^{2+}]}{[Ag^+]^2}\right).$

$E_2 = 0.398 - 0.0295 \log\left(\frac{0.01}{(0.001)^2}\right).$

$E_2 = 0.398 - 0.0295 \log\left(\frac{10^{-2}}{10^{-6}}\right) = 0.398 - 0.0295 \log(10^4).$

$E_2 = 0.398 - 0.0295 \times 4 = 0.398 - 0.118.$

$E_2 = 0.280$ V.

Step 3: Express the answer in the required format.

The question asks for the potential in units of $\times 10^{-2}$ V.

$E_2 = 0.280 \text{ V} = 28.0 \times 10^{-2}$ V.

Rounding to the nearest integer, the value is 28.

---

14. Answer: 288 - 288

Explanation:

Step 1: $\Lambda^\circ_{BaSO_4} = \Lambda^\circ_{BaCl_2} + \Lambda^\circ_{H_2SO_4} - 2\Lambda^\circ_{HCl}.$

Step 2: $\Lambda^\circ = 280 + 860 - 2(426) = 1140 - 852 = 288 \text{ S cm}^2 \text{ mol}^{-1}.$

---

15. Answer: 12 - 12

Explanation:

Given:

$$\kappa = 2.0 \times 10^{-5} \text{ S cm}^{-1}$$

$$C = 0.001 \text{ mol L}^{-1}$$

$$\Lambda^\circ_m = 190 \text{ S cm}^2 \text{ mol}^{-1}$$

Step 1: Calculate molar conductivity Λ_m

$$\Lambda_m = \frac{1000\kappa}{C}$$

$$\Lambda_m = \frac{1000 \times 2.0 \times 10^{-5}}{10^{-3}} = 20 \text{ S cm}^2 \text{ mol}^{-1}$$

Step 2: Calculate degree of dissociation α

$$\alpha = \frac{\Lambda_m}{\Lambda_m^\circ} = \frac{20}{190} = \frac{2}{19} \approx 0.105$$

Step 3: Apply Ostwald's dilution law For a weak acid:

$$K_a = \frac{C\alpha^2}{1-\alpha}$$

Substitute values:

$$K_a = \frac{10^{-3} \times \left(\frac{2}{19}\right)^2}{1 - \frac{2}{19}}$$

$$K_a = \frac{10^{-3} \times \frac{4}{361}}{\frac{17}{19}}$$

$$K_a = \frac{76}{6137} \times 10^{-3} \approx 1.238 \times 10^{-5}$$

Step 4: Express in required format

$$K_a = 12.38 \times 10^{-6}$$

Rounding to the nearest integer:

$$\boxed{12}$$

---

16. Answer: c

Explanation:

The electrolytic oxidation of an acidified concentrated solution of a sulphate, such as ammonium sulphate or sulfuric acid, at high current density is a method for preparing peroxodisulphates.
At the anode, the bisulphate ion (HSO_4^-) or sulphate ion (SO_4^{2-}) is oxidized. The primary reaction is the coupling of two bisulphate radicals.

The overall anode reaction can be represented as:
$2HSO_4^-(aq) \rightarrow S_2O_8^{2-}(aq) + 2H^+(aq) + 2e^-$
The product is the peroxodisulphate ion, $S_2O_8^{2-}$. The corresponding acid is peroxodisulphuric acid, $H_2S_2O_8$, also known as Marshall's acid.
Let's examine the structure of this acid. It contains a peroxide linkage (-O-O-) between the two sulfur atoms. The structure is:
\begin{center} O \hspace{0.5cm} O
\vspace{-0.2cm} || \hspace{0.5cm} ||
HO-S-O-O-S-OH
\vspace{-0.2cm} || \hspace{0.5cm} ||
\vspace{0.1cm} O \hspace{0.5cm} O \end{center} The condensed structural formula is HO_3SOOSO_3H. This matches option (C).
Option (A) HO_3SOSO_3H represents pyrosulphuric acid ($H_2S_2O_7$), which does not have the peroxide link.

---

17. Answer: 4 - 4

Explanation:

Copper reduces $\mathrm{NO_3^-}$ into NO and NO$_2$ depending on the concentration of $\mathrm{HNO_3}$.
Given:

$$E^\circ_{\mathrm{Cu^{2+}/Cu}} = 0.34\ \mathrm{V}, \quad E^\circ_{\mathrm{NO_3^-/NO}} = 0.96\ \mathrm{V}, \quad E^\circ_{\mathrm{NO_3^-/NO_2}} = 0.79\ \mathrm{V}$$

$$\frac{2.303RT}{F} = 0.059 \quad (298\ \mathrm{K})$$

Step 1: Standard cell potentials For reduction to NO:

$$E^\circ_{\mathrm{cell},1} = 0.96 - 0.34 = 0.62\ \mathrm{V}, \quad n_1 = 6$$

For reduction to NO$_2$:

$$E^\circ_{\mathrm{cell},2} = 0.79 - 0.34 = 0.45\ \mathrm{V}, \quad n_2 = 2$$

Step 2: Condition for equal thermodynamic tendency

$$\Delta G_1 = \Delta G_2 \Rightarrow n_1 E_{\mathrm{cell},1} = n_2 E_{\mathrm{cell},2}$$

Step 3: Apply Nernst equation Let $[\mathrm{HNO_3}] = C$, so $[\mathrm{H^+}] = [\mathrm{NO_3^-}] = C$. Assuming $[\mathrm{Cu^{2+}}]$,

$P_{NO}, P_{NO_2} = 1$.

$$3\left(0.62 - \frac{0.059}{6}\log\frac{1}{C^{10}}\right) = 0.45 - \frac{0.059}{2}\log\frac{1}{C^6}$$

$$1.41 = 8 \times 0.059 \log C \Rightarrow \log C \approx 2.99$$

Step 4: Final answer

$$C = 10^{2.99} \Rightarrow x \approx 3$$

$$\boxed{2x = 6}$$

---

18. Answer: b

Explanation:

Step 1: Understanding the Concept:
This question involves matching various electrochemical parameters with their respective SI or common units.
Step 2: Key Formula or Approach:
1. Cell Constant (G^*) = $\frac{l}{A}$. Unit = $\mathrm{m/m^2 = m^{-1}}$.
2. Molar Conductivity (Λ_m) = $\frac{\kappa}{C}$. Unit = $\mathrm{S\,cm^2\,mol^{-1}}$.
3. Conductivity (κ) = $\frac{1}{\rho} = \frac{G^*}{R}$. Unit = $\Omega^{-1}\,\mathrm{m}^{-1}$ or $\mathrm{S\,m^{-1}}$.
4. Degree of dissociation (α) = Ratio of conductivities. Being a ratio of same physical quantities, it is dimensionless.
Step 3: Detailed Explanation:
- (a) Cell constant is the ratio of distance between electrodes to the cross-sectional area. Its unit is $\mathrm{m^{-1}}$. Match: (a)-(iii).
- (b) Molar conductivity represents the conducting power of all ions produced by dissolving 1 mole of an electrolyte. Its unit is typically expressed as $\mathrm{S\,cm^2\,mol^{-1}}$. Match: (b)-(i).
- (c) Conductivity (specific conductance) is the reciprocal of resistivity. Its unit is $\Omega^{-1}\,\mathrm{m}^{-1}$. Match: (c)-(iv).
- (d) Degree of dissociation is the fraction of total electrolyte that dissociates into ions. It is a pure number. Match: (d)-(ii).
Step 4: Final Answer:
The correct matching is (a)-(iii), (b)-(i), (c)-(iv), (d)-(ii).

19. Answer: 25 - 25

Explanation:

Step 1: Understanding the Concept:
We use the relationship between the standard cell potential and the standard Gibbs free energy change, and then apply the thermodynamic relation involving enthalpy and entropy.
Step 2: Key Formula or Approach:
1. $\Delta G^0 = -nFE^0_{cell}$
2. $\Delta G^0 = \Delta H^0 - T\Delta S^0$
Step 3: Detailed Explanation:
1. **Determine** n**:** In the reaction, Cd is oxidised to Cd^{2+} in $\mathrm{CdSO_4}$, and $\mathrm{Hg_2^{2+}}$ in $\mathrm{Hg_2SO_4}$ is reduced to $2\mathrm{Hg}$. Thus, $n = 2$.
2. **Calculate** ΔG^0**:**

$$\Delta G^0 = -2 \times 96487 \text{ C/mol} \times 4.315 \text{ V}$$

$$\Delta G^0 = -832682.8 \text{ J/mol} = -832.68 \text{ kJ/mol}$$

3. **Calculate** ΔS^0**:**
Given $\Delta H^0 = -825.2 \text{ kJ/mol} = -825200 \text{ J/mol}$.
Temperature $T = 25^\circ\text{C} = 298 \text{ K}$.

$$\Delta G^0 = \Delta H^0 - T\Delta S^0$$

$$-832682.8 = -825200 - 298 \times \Delta S^0$$

$$298 \times \Delta S^0 = -825200 + 832682.8$$

$$298 \times \Delta S^0 = 7482.8$$

$$\Delta S^0 = \frac{7482.8}{298} \approx 25.11 \text{ J K}^{-1} \text{ mol}^{-1}$$

Step 4: Final Answer:
The standard entropy change to the nearest integer is 25.

---

20. Answer: 760 - 760

Explanation:

Step 1: Understanding the Question:
We are given the resistance, cell constant, and concentration for a solution and asked to calculate its molar conductivity.
Step 2: Key Formulas:
1. **Conductivity (κ):** It relates the resistance (R) and the cell constant ($G^* = l/A$).

$$\kappa = \frac{1}{R} \times \frac{l}{A} = \frac{G^*}{R}$$

2. **Molar Conductivity (Λ_m):** It relates conductivity to the molar concentration (C).

$$\Lambda_m = \frac{\kappa \times 1000}{C}$$

(The factor of 1000 is used when κ is in S cm^{-1} and C is in mol L^{-1} to get Λ_m in S cm^2 mol^{-1}).
Step 3: Calculate Conductivity (κ):
Given:
- Cell constant, $G^* = 1.14$ cm^{-1}.
- Resistance, $R = 1500\,\Omega$.

$$\kappa = \frac{1.14\,\text{cm}^{-1}}{1500\,\Omega} = 0.00076\,\text{S cm}^{-1}$$

Step 4: Calculate Molar Conductivity (Λ_m):
Given:
- Concentration, $C = 0.001$ M.

$$\Lambda_m = \frac{(0.00076\,\text{S cm}^{-1}) \times 1000\,\text{cm}^3/\text{L}}{0.001\,\text{mol/L}}$$

$$\Lambda_m = \frac{0.76}{0.001} = 760\,\text{S cm}^2\text{mol}^{-1}$$

Step 5: Final Answer:
The molar conductivity of the solution is 760 S cm^2 mol^{-1}.

---

21. Answer: 3 - 3

Explanation:

Step 1: Understanding the Concept:
The standard electrode potential (reduction potential) of a metal electrode (M^{n+}/M) is determined by the enthalpy changes in a thermodynamic cycle (Born-Haber cycle for ions in solution).
Step 2: Detailed Explanation:
The process $M^{n+}_{(aq)} + ne^- \rightarrow M_{(s)}$ can be broken down into:
1. **Hydration of the ion:** $M^{n+}_{(aq)} \rightarrow M^{n+}_{(g)}$ (Reverse of hydration enthalpy)
2. **Ionisation of gas atoms:** $M^{n+}_{(g)} + ne^- \rightarrow M_{(g)}$ (Reverse of ionisation enthalpy)
3. **Atomisation/Sublimation:** $M_{(g)} \rightarrow M_{(s)}$ (Reverse of sublimation enthalpy)

Therefore, sublimation enthalpy, ionisation enthalpy, and hydration enthalpy are the three key factors that determine the reduction potential of metallic elements. Electron gain enthalpy is typically used for defining the oxidizing power of non-metals (like halogens) and is not part of the standard metal electrode potential cycle.
Step 3: Final Answer:
The total number of properties is 3.

---

22. Answer: b

Explanation:

Step 1: Understanding the Concept:
Limiting molar conductivity (Λ°_m) depends on the individual ionic conductivities at infinite dilution. Molar conductivity (Λ_m) is the conducting power of all ions produced by one mole of an electrolyte in a given volume of solution.
Step 2: Detailed Explanation:
1. **Analysis of Statement I:** According to Kohlrausch's law, Λ°_m is the sum of limiting ionic conductivities. The limiting ionic conductivity of H^+ ($\approx 349.6 \text{ S cm}^2 \text{ mol}^{-1}$) is

much higher than that of K^+ ($\approx 73.5\ \mathrm{S\ cm^2\ mol^{-1}}$). Because of this, $\Lambda_m^\circ(CH_3COOH)$ is actually higher than $\Lambda_m^\circ(KCl)$. Statement I is **false**.

2. **Analysis of Statement II:** Molar conductivity is defined as $\Lambda_m = \frac{\kappa\times1000}{M}$. As concentration decreases (dilution increases), the volume containing one mole of electrolyte increases. This increase in volume more than compensates for the decrease in κ, leading to an **increase** in molar conductivity upon dilution for both strong and weak electrolytes. Statement II is **false**.

Step 3: Final Answer:

Both Statement I and Statement II are false.

---

23. Answer: 109 - 109

Explanation:

Step 1: Understanding the Question

We need to calculate the cell potential (E_{cell}) for a galvanic cell under non-standard conditions using the Nernst equation.

Step 2: Key Formula or Approach

The Nernst equation for a cell is:

$$E_{cell} = E^0_{cell} - \frac{0.059}{n}\log Q$$

where E^0_{cell} is the standard cell potential, n is the number of electrons transferred, and Q is the reaction quotient.

Step 3: Detailed Calculation

Calculate the Standard Cell Potential (E^0_{cell}):

The overall reaction is Zn + $Cu^{2+} \rightarrow Zn^{2+}$ + Cu.

Oxidation (Anode): Zn $\rightarrow Zn^{2+}$ + $2e^-$

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

We need the standard reduction potentials:

$E^0_{Zn^{2+}/Zn}$: Given $E^0_{Zn/Zn^{2+}}$ = +0.76 V (standard oxidation potential), so $E^0_{Zn^{2+}/Zn}$ = -0.76 V (standard reduction potential).

$E^0_{Cu^{2+}/Cu}$: Given $E^0_{Cu/Cu^{2+}}$ = -0.34 V (standard oxidation potential), so $E^0_{Cu^{2+}/Cu}$ = +0.34 V (standard reduction potential).

Now, calculate E^0_{cell}:

$$E^0_{cell} = E^0_{cathode} - E^0_{anode} = E^0_{Cu^{2+}/Cu} - E^0_{Zn^{2+}/Zn}$$

$$E^0_{cell} = (+0.34 \text{ V}) - (-0.76 \text{ V}) = 1.10 \text{ V}$$

Determine n and Q:
From the half-reactions, the number of electrons transferred, n, is 2.
The reaction quotient, Q, is:

$$Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = \frac{0.04 \text{ M}}{0.02 \text{ M}} = 2$$

Apply the Nernst Equation:

$$E_{cell} = 1.10 - \frac{0.059}{2} \log(2)$$

Using $\log(2) \approx 0.301$:

$$E_{cell} = 1.10 - (0.0295 \times 0.301)$$

$$E_{cell} = 1.10 - 0.0088795 \approx 1.0911 \text{ V}$$

Express the answer in the required format:

$$E_{cell} = 1.0911 \text{ V} = 109.11 \times 10^{-2} \text{ V}$$

Step 4: Final Answer
Rounding to the nearest integer, the value is 109.

---

24. Answer: b

Explanation:

To solve this question, we need to analyze the electrolysis of the given solutions and the respective standard electrode potentials (E^0) of the ions involved.

1. **Explanation of Statement I:**
   The sequence of deposition of metals during electrolysis is determined by their standard reduction potentials. Metals with higher (more positive) standard reduction potentials will be deposited first.
   - From the data given:
     - $E^0_{Ag^+/Ag} = 0.80 \text{ V}$
     - $E^0_{Hg_2^{2+}/Hg} = 0.79 \text{ V}$
     - $E^0_{Cu^{2+}/Cu} = 0.34 \text{ V}$

- $E^0_{Mg^{2+}/Mg} = -2.37\,\text{V}$

With increasing voltage, the metals will be deposited in the order of their E^0 values:

- First, Ag will be deposited since its E^0 value is the highest at 0.80 V.
- Next, Hg will be deposited with an E^0 of 0.79 V.
- Finally, Cu will be deposited, as its E^0 is 0.34 V.
- Mg will not be deposited as its E^0 is much lower at -2.37 V.

Therefore, Statement I is correct.

2. **Explanation of Statement II:**

According to electrochemical series, magnesium has a very low standard reduction potential $(E^0 = -2.37\,\text{V})$; hence, it is less likely to be reduced compared to water. Instead of magnesium being deposited at the cathode, hydrogen gas will usually be evolved due to water reduction:

- $2H_2O + 2e^- \rightarrow H_2 + 2OH^- \quad (E^0 = -0.83\,\text{V})$

Therefore, based on standard potentials and typical electrolysis behavior, magnesium will remain in the solution, and hydrogen gas (not oxygen gas) will be produced at the cathode. Consequently, Statement II is incorrect.

The correct choice is: **Statement I is correct but statement II is incorrect.**

---

25. Answer: a

Explanation:

To solve this problem, we need to determine the correct statement about the fuel cell based on the given standard cell potential and the standard reduction potentials provided in the options. Let's analyze the information step-by-step:

1. The standard cell potential (E°_{cell}) of the fuel cell is given as 1.21 V. This potential is for the overall reaction in the fuel cell.
2. The standard reduction potential for the $\text{O}_2/\text{H}_2\text{O}$ half cell $(E^\circ_{O_2/H_2O})$ is 1.229 V. This indicates that the reduction of oxygen to water happens at the cathode.
3. In any galvanic cell (such as this fuel cell), the overall cell potential is the difference between the cathode and anode potentials:
    - Using the formula: $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$
    - Given, $E^\circ_{\text{cell}} = 1.21\ \text{V}$, and $E^\circ_{\text{cathode}} = 1.229\ \text{V}$ for $\text{O}_2/\text{H}_2\text{O}$.
    - Substituting the values: $1.21 = 1.229 - E^\circ_{\text{anode}}$
    - Solving for E°_{anode}, we get: $E^\circ_{\text{anode}} = 1.229 - 1.21 = 0.019\ \text{V} = 19\ \text{mV}$

4. This calculated value, 19 mV, corresponds to the standard half-cell potential for the oxidation of methanol to carbon dioxide and water at the anode, signifying that the calculated potential matches with the statement regarding carbon dioxide reduction.

Now, let's verify each statement:

- **The standard half cell reduction potential for the reduction of CO_2 ($E^\circ_{CO_2/CH_3OH}$) is 19 mV:** As calculated above, this statement is correct.
- **Oxygen is formed at the anode:** In a fuel cell using methanol, oxygen is consumed at the cathode, not formed. Wrong statement.
- **Reactants are fed at one go to each electrode:** Typically, reactants are continuously fed to maintain the reaction consistently in a fuel cell, not at one go. Assuming this refers to traditional continuous feeding, this option doesn't fit well within the standard scenarios. Likely incorrect in this context.
- **Reduction of methanol takes place at the cathode:** Methanol is oxidized at the anode, not reduced at the cathode. This statement is incorrect.

Therefore, the correct answer is: **The standard half cell reduction potential for the reduction of CO_2 ($E^\circ_{CO_2/CH_3OH}$) is 19 mV.**

---

26. Answer: 0.4 - 0.4

Explanation:

Step 1: Standard Electrode Potentials

The given standard electrode potentials are:

$$E^\circ_{\mathrm{Cu^{2+}/Cu}} = 0.34\ \mathrm{V}$$

$$E^\circ_{\mathrm{Ag^{+}/Ag}} = 0.80\ \mathrm{V}$$

Step 2: Applying the Nernst Equation

The Nernst equation gives the relationship between the emf of the electrochemical cell, the standard electrode potentials, and the concentrations of the ions involved:

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

Where:

E°_{cell} is the standard cell potential, n is the number of electrons involved, Q is the reaction quotient.

Step 3: Determining the Standard Cell Potential

The standard cell potential is the difference between the two half-cell potentials:

$$E^\circ_{\text{cell}} = E^\circ_{\text{Ag}^+/\text{Ag}} - E^\circ_{\text{Cu}^{2+}/\text{Cu}} = 0.80\,\text{V} - 0.34\,\text{V} = 0.46\,\text{V}$$

Step 4: Reaction Quotient (Q)

The reaction quotient Q is given by the ratio of the concentrations of the products over the reactants. The concentrations of Cu^{2+} and Ag^+ ions change due to the passage of electricity: After 1 Faraday is passed through the Cu^{2+} cell, the concentration of Cu^{2+} is reduced by a factor based on the number of moles reduced.

Since 1 Faraday corresponds to 1 mole of electrons, the concentration of Cu^{2+} decreases.

Similarly, after 0.1 Faraday is passed through the Ag^+ cell, the concentration of Ag^+ decreases.

Now, using the changes in concentration:

Initially, the concentration of Cu^{2+} is 1.5 M, and after 1 Faraday, the concentration of Cu^{2+} decreases as per the stoichiometry. Initially, the concentration of Ag^+ is 0.2 M, and after 0.1 Faraday, the concentration of Ag^+ decreases accordingly.

$$Q = \frac{[\text{Cu}^{2+}]}{[\text{Ag}^+]}$$

Step 5: Substituting Values into the Nernst Equation

Substitute the concentrations and other known values into the Nernst equation:

$$E_{\text{cell}} = 0.46\,\text{V} - \frac{0.0591}{1}\log\left(\frac{1.5\,\text{M}}{0.2\,\text{M}}\right)$$

Now, calculate the logarithmic term:

$$\log\left(\frac{1.5}{0.2}\right) = \log(7.5) \approx 0.875$$

$$E_{\text{cell}} = 0.46\,\text{V} - 0.0591 \times 0.875$$

$$E_{\text{cell}} = 0.46\,\text{V} - 0.0518\,\text{V}$$

$$E_{\text{cell}} = 0.4082\,\text{V} \approx 0.40\,\text{V}$$

---

27. Answer: 3 - 3

Explanation:

To determine the degree of dissociation (α) of ammonium hydroxide (NH_4OH), we use the following relationship for molar conductance:

$\lambda_m = \lambda_0\,\alpha$

Where:

- λ_m = molar conductance of the solution (85.5 S cm^{-1} mol^{-1})
- λ_0 = molar conductance at infinite dilution of NH_4OH
- α = degree of dissociation

First, calculate λ_0 using the ion conductances:

$\lambda_0(NH_4OH) = \lambda_0(NH_4Cl) - \lambda_m(Cl^-) + \lambda_m(OH^-)$

$= 185\ S\ cm^{-1}\ mol^{-1} - 70\ S\ cm^{-1}\ mol^{-1} + 170\ S\ cm^{-1}\ mol^{-1}$

$= 285\ S\ cm^{-1}\ mol^{-1}$

Now, substitute the values into the degree of dissociation formula:

$\alpha = \lambda_m / \lambda_0$

$= 85.5 / 285$

$= 0.3$

To express α as $x \times 10^{-1}$, we have x = 3.

Confirming against the expected range (3,3), the computed value x = 3 is within the range.

Therefore, the value of x is **3**.

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28. Answer: 10 - 10

Explanation:

To find the pH at which the EMF of the half-cell becomes zero, we will use the Nernst equation for the given half-cell reaction:

$$\mathrm{Cr_2O_7^{2-} + 6e^- + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O}$$

The Nernst equation is given by:

$$E = E^\circ - \frac{0.059}{n} \log Q,$$

where E is the EMF of the cell, E° is the standard reduction potential (1.33 V), n is the number of electrons transferred (6), and Q is the reaction quotient:

$$Q = \frac{[\mathrm{Cr^{3+}}]^2}{[\mathrm{Cr_2O_7^{2-}}][\mathrm{H^+}]^{14}}$$

Given $\frac{[\mathrm{Cr^{3+}}]^2}{[\mathrm{Cr_2O_7^{2-}}]} = 10^{-6}$,

Substitute $Q = 10^{-6}[\mathrm{H^+}]^{-14}$.

Since the EMF is zero:

$$0 = 1.33 - \frac{0.059}{6} \log\left(10^{-6}[\mathrm{H^+}]^{-14}\right)$$

Solving for the pH:

$$\frac{0.059}{6} \log\left(10^{-6}[\mathrm{H^+}]^{-14}\right) = 1.33$$

$$\log\left(10^{-6}[\mathrm{H^+}]^{-14}\right) = 135.59$$

Rewriting the equation:

$$\log 10^{-6} + \log[\mathrm{H^+}]^{-14} = 135.59$$

$$-6 - 14\log[\mathrm{H^+}] = 135.59$$

Solving for $\log[\mathrm{H^+}]$:

$$-14\log[\mathrm{H^+}] = 141.59$$

$$\log[\mathrm{H^+}] = -10.11$$

Converting to pH:

$$\mathrm{pH} = -\log[\mathrm{H}^+] = 10.11$$

Thus, the nearest integer value for the pH is $\mathbf{10}$. As expected, 10 falls within the given range (10, 10).

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29. Answer: 2 - 2

Explanation:

The given reaction is:

$$\mathrm{Pb}^{2+} + 2e^- \to \mathrm{Pb}$$

$$E^\circ = m \quad \text{and} \quad \Delta G^\circ = -2Fm$$

$$\mathrm{Pb}^{4+} + 4e^- \to \mathrm{Pb}^{2+}$$

$$E^\circ = n \quad \text{and} \quad \Delta G^\circ = -4Fn$$

Now,

$$\Delta G^\circ = \Delta G_1^\circ - \Delta G_2^\circ$$

$$-2Fm = -4Fn$$

$$2FE = 2Fm + 4Fn \quad \Rightarrow \quad E^\circ = m - 2n$$

Thus, the value of x is 2. The correct answer is (2).

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30. Answer: a

Explanation:

- **Transistors (A)** use a battery with an anode of Zn/Hg and a cathode of HgO + C.
- **Hearing aids (B)** use a battery with an anode of Zn and a cathode of Carbon.
- **Inverters (C)** use a hydrogen fuel cell as the energy source.
- **Apollo space ship (D)** uses a battery with an anode of Pb and a cathode of PbO_2.

Therefore, the correct matching is $\boxed{(A)-(III),(B)-(I),(C)-(IV),(D)-(II)}$.