Electrochemistry JEE Main PYQ - 2

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. For a $\mathrm{Mg}|\mathrm{Mg}^{2+}(aq)||\mathrm{Ag}^{+}(aq)|\mathrm{Ag}$, the correct Nernst Equation is: (+4, -1)

   a. $E_{\text{cell}} = E^{\circ}_{\text{cell}} + \frac{RT}{2F} \ln\left(\frac{[\mathrm{Ag}^+]}{[\mathrm{Mg}^{2+}]}\right)$

   b. $E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{2F} \ln\left(\frac{[\mathrm{Ag}^+]}{[\mathrm{Mg}^{2+}]}\right)$

   c. $E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{2F} \ln\left(\frac{[\mathrm{Mg}^{2+}]}{[\mathrm{Ag}^+]}\right)$

   d. $E_{\text{cell}} = E^{\circ}_{\text{cell}} + \frac{RT}{2F} \ln[\mathrm{Ag}^+]^2$

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2. Which of the following electrolyte can be used to obtain $\mathrm{H_2S_2O_8}$ by the process of electrolysis? (+4, -1)

   a. Dilute solution of sodium sulphate

   b. Dilute solution of sulphuric acid

   c. Concentrated solution of sulphuric acid

   d. Acidified dilute solution of sodium sulphate

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3. Electrolysis of 600 mL aqueous solution of NaCl for 5 min changes the pH of the solution to 12. The current in Amperes used for the given electrolysis is ____________________________. (Nearest integer). (+4, -1)

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4. For the given cell: (+4, -1)

$$Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$$

The standard cell potential of the above reaction is given. The standard reduction potentials are given as:

$$Ag^{+} + e^{-} \rightarrow Ag \quad E^{\circ} = x\,V$$

$$Fe^{2+} + 2e^{-} \rightarrow Fe \quad E^{\circ} = y\,V$$

$$Fe^{3+} + 3e^{-} \rightarrow Fe \quad E^{\circ} = z\,V$$

The correct answer is:

a. $x + y - z$

b. $x + 2y - 3z$

c. $y - 2x$

d. $x + 2y$

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5. **Given below are two statements:** **(+4, -1)**

Statement (I): *Corrosion is an electrochemical phenomenon in which pure metal acts as an anode and impure metal as a cathode.*

Statement (II): *The rate of corrosion is more in alkaline medium than in acidic medium.*

In the light of the above statements, choose the correct answer from the options given below:

a. Both Statement I and Statement II are false

b. Statement I is true but Statement II is false

c. Both Statement I and Statement II are true

d. Statement I is false but Statement II is true

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6. The amount of electricity in Coulomb required for the oxidation of 1 mol of H_2O to O_2 is **(+4, -1)**

$$\_ \times 10^5 \text{ C.}$$

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7. The strongest reducing agent amont the following is: **(+4, -1)**

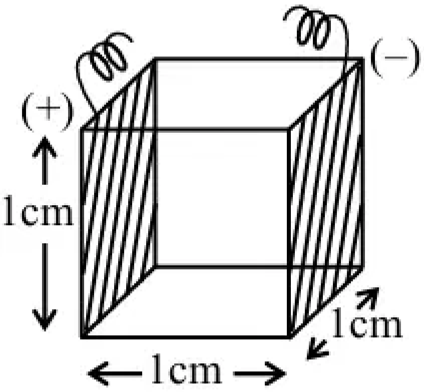
a. NH_3

b. SbH_3

c. BiH_3

d. PH_3

8. A conductivity cell with two electrodes (dark side) are half filled with infinitely dilute aqueous solution of a weak electrolyte. If volume is doubled by adding more water at constant temperature, the molar conductivity of the cell will – **(+4, -1)**



a. increase sharply

b. remain same or can not be measured accurately

c. decrease sharply

d. depend upon type of electrolyte

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9. How can an electrochemical cell be converted into an electrolytic cell ? **(+4, -1)**

a. Applying an external opposite potential greater than E°_{cell}.

b. Reversing the flow of ions in salt bridge.

c. Applying an external opposite potential lower than E°_{cell}.

d. Exchanging the electrodes at anode and cathode.

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10. For the electrochemical cell **(+4, -1)**

M|M^{2+}||X^{2-}|X

If $E^0_{(M^{2+}/M)} = 0.46$ V and $E^0_{(X/X^{2-})} = 0.34$ V.

Which of the following is correct?

a. $E_{cell} = -0.80$ V

b. M + X $\rightarrow$ M^{2+} + X^{2-} is a spontaneous reaction

c. M^{2+} + X^{2-} $\rightarrow$ M + X is a spontaneous reaction

d. $E_{cell} = 0.80$ V}

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11. The quantity of silver deposited when one coulomb charge is passed through $AgNO_3$ solution: **(+4, -1)**

a. 0.1 g atom of silver

b. 1 chemical equivalent of silver

c. 1 g of silver

d. 1 electrochemical equivalent of silver

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12. For a strong electrolyte, a plot of molar conductivity against $(\text{concentration})^{1/2}$ is a straight line, with a negative slope, the correct unit for the slope is: **(+4, -1)**

a. S cm^2 mol$^{-3/2}$ L$^{1/2}$

b. S cm^2 mol^{-1} L$^{1/2}$

c. S cm^2 mol$^{-3/2}$ L

d. S cm^2 mol$^{-3/2}$ L$^{-1/2}$

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13. Fuel cell, using hydrogen and oxygen as fuels, **(+4, -1)**
A. has been used in spaceship
B. has as efficiency of 40% to produce electricity
C. uses aluminium as catalysts
D. is eco-friendly
E. is actually a type of Galvanic cell only

a. A,B,C only

b. A,B,D only

c. A,B,D,E only

d. A,D,E only

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14. The reaction; $\frac{1}{2}H_{2(g)} + AgCl_{(s)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)} + Ag_{(s)}$ occurs in which of the following galvanic cell: **(+4, -1)**

a. Pt | $H_{2(g)}$ | $HCl_{(soln.)}$ | $AgCl_{(s)}$ | Ag

b. Pt | $H_{2(g)}$ | $HCl_{(soln.)}$ | $AgNO_{3(aq)}$ | Ag

c. Pt | $H_{2(g)}$ | $KCl_{(soln.)}$ | $AgCl_{(s)}$ | Ag

d. Ag | $AgCl_{(s)}$ | $KCl_{(soln.)}$ | $AgNO_{3(aq)}$ | Ag

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15. The emf of cell $Tl|^{Tl^+}_{(0.001M)}|^{Cu^{2+}}_{(0.01M)}$ Cu is 0.83 V at 298 K. It could be increased by: **(+4, -1)**

a. increasing concentration of Tl^+ ions

b. increasing concentration of both Tl^+ and Cu^{2+} ions

c. decreasing concentration of both Tl^+ and Cu^{2+} ions

d. increasing concentration of Cu^{2+} ions

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16. The standard reduction potentials at 298 K for the following half cells are given below: **(+4, -1)**

$$\mathrm{Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O},\ E^\circ = 1.33\,\mathrm{V}$$

$$\mathrm{Fe^{3+}(aq) + 3e^- \rightarrow Fe},\ E^\circ = -0.04\,\mathrm{V}$$

$$\mathrm{Ni^{2+}(aq) + 2e^- \rightarrow Ni},\ E^\circ = -0.25\,\mathrm{V}$$

$$\mathrm{Ag^+(aq) + e^- \rightarrow Ag},\ E^\circ = 0.80\,\mathrm{V}$$

$$Au^{3+}(aq) + 3e^- \rightarrow Au, \; E^\circ = 1.40\,V$$

Consider the given electrochemical reactions, the number of metal(s) which will be oxidized by $Cr_2O_7^{2-}$ in aqueous solution is ____.

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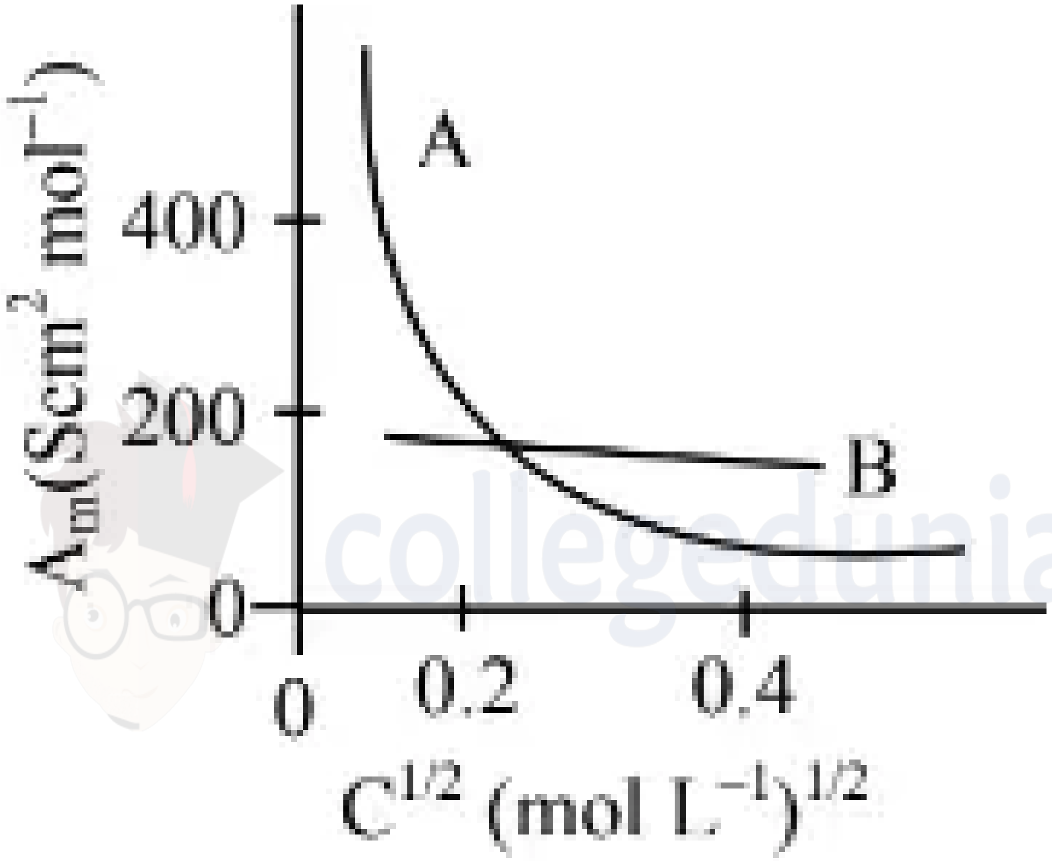
17. The molar conductivity for electrolytes A and B are plotted against $C^{1/2}$ as shown below. Electrolytes A and B respectively are : **(+4, -1)**



a.

A	B
Weak electrolyte	Weak electrolyte

b.

A	B
Strong electrolyte	strong electrolyte

c.

A	B
Weak electrolyte	strong electrolyte

d.

A	B
Strong electrolyte	Weak electrolyte

---

18. molar ionic conductivities of a divalent cation and anion are $57\,\mathrm{S\ cm^2 mol^{-1}}$ and $73\,\mathrm{S\ cm^2 mol^{-1}}$, respectively. The molar conductivity of the solution of an electrolyte with the above cation and anion will be: **(+4, -1)**

a. $65\,\mathrm{S\ cm^2 mol^{-1}}$

b. $130\,\mathrm{S\ cm^2 mol^{-1}}$

c. $187\,\mathrm{S\ cm^2 mol^{-1}}$

d. $260\,\mathrm{S\ cm^2 mol^{-1}}$

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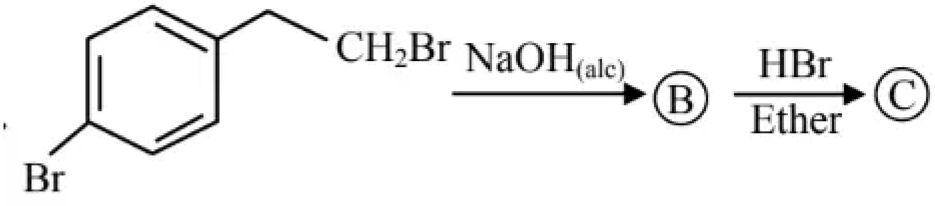
19. The reaction at cathode in the cells commonly used in clocks involves. **(+4, -1)**

a. reduction of Mn from +4 to +3

b. oxidation of Mn from +3 to +4

c. reduction of Mn from + 7 to +2

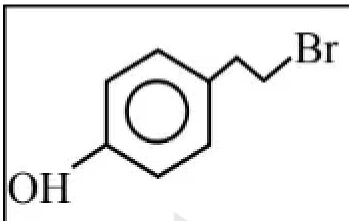
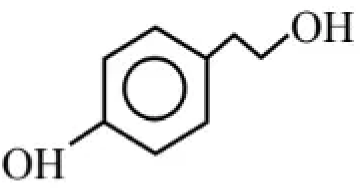
d. oxidation of Mn from + 2 to +7



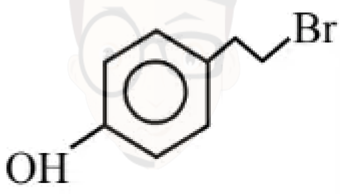
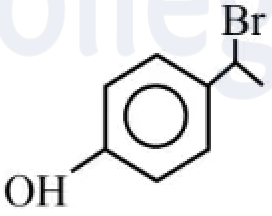
Identify Ⓑ and Ⓒ and how are Ⓐ and Ⓒ related ?

20. **(B)** **(C)** (+4, -1)

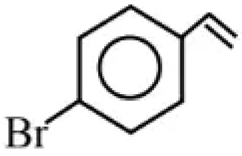
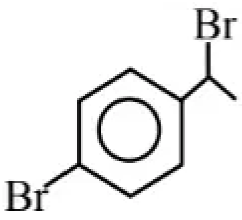
a.

		functional group isomers

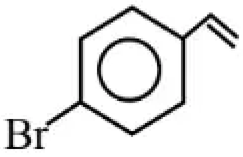
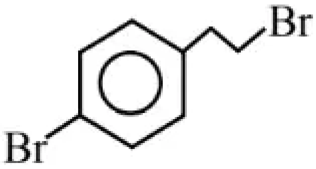
b.

		Derivative

c.

		position isomers

d.

		chain isomers

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21. What pressure (bar) of H_2 would be required to make the emf of a hydrogen electrode zero in pure water at $25°C$? **(+4, -1)**

a. 10^{-14}

b. 10^{-7}

c. 1

d. 0.5

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22. The values of conductivity of some materials at $298.15\,K$ in $S\,m^{-1}$ are: **(+4, -1)**
2.1×10^3, 1.0×10^{-16}, 1.2×10^1, 3.91, 1.5×10^{-2}, 1×10^{-7}, 1.0×10^3.
The number of conductors among the materials is ______.

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23. Which out of the following is a correct equation to show change in molar conductivity with respect to concentration for a weak electrolyte, if the symbols carry their usual meaning **(+4, -1)**

a. $\Lambda_m^2 C - K_a\Lambda_m + K_a\Lambda_m^{\circ 2} = 0$

b. $\Lambda_m - \Lambda_m^\circ + AC\frac{1}{2} = 0$

c. $\Lambda_m - \Lambda_m^\circ - AC\frac{1}{2} = 0$

d. $\Lambda_m^2 C + K_a\Lambda_m^{\circ 2} - K_a\Lambda_m^\circ = 0$

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

LIST-I (Cell)	LIST-II (Use/Property/Reaction)
A. Leclanche cell	IV. Reaction at anode Zn → Zn^{2+} + $2e^-$
B. Ni-Cd cell	III. Rechargeable
C. Fuel cell	I. Converts energy of combustion into electrical energy
D. Mercury cell	II. Does not involve any ion in solution and is used in hearing aids

Choose the correct answer from the options given below:

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

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

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

d. A-II, B-III, C-IV, D-I

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25. One Faraday of electricity liberates $x \times 10^{-1}$gram atom of copper from copper sulphate, x is ______. **(+4, -1)**

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26. Identify the factor from the following that does not affect electrolytic conductance of a solution. **(+4, -1)**

a. The nature of the electrolyte added

b. The nature of the electrode used

c. Concentration of the electrolyte

d. The nature of solvent used

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27. The metals that are employed in the battery industries are **(+4, -1)**

A. Fe
B. Mn
C. Ni
D. Cr
E. Cd

Choose the correct answer from the options given below:

a. B, C, and E only

b. A, B, C, D, and E

c. A, B, C, and D only

d. B, D, and E only

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28. The potential for the given half cell at 298 K is **(+4, -1)**

$$(-)\ldots\times 10^{-2}\,\mathrm{V}.$$

$$2\mathrm{H}^{+}(\mathrm{aq})+2e^{-}\rightarrow \mathrm{H}_2(\mathrm{g})$$

$$[\mathrm{H}^{+}]=1M,\ P_{\mathrm{H}_2}=2\,\mathrm{atm}$$

(Given: $2.303\frac{RT}{F}=0.06\,\mathrm{V}, \log 2=0.3$)

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29. The mass of silver (Molar mass of Ag : 108 gmol^{-1}) displaced by a quantity of electricity which displaces 5600 mL of O_2 at S.T.P. will be _____ g. **(+4, -1)**

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30. The mass of zinc produced by the electrolysis of zinc sulfate solution with a steady current of 0.015 A for 15 minutes is _____ $\times 10^{-4}$ g. **(+4, -1)**
(Atomic mass of zinc = 65.4 amu)

Answers

1. Answer: c

Explanation:

To determine the correct Nernst equation for the given electrochemical cell, Mg | Mg^{2+} (aq) || Ag^+ (aq) | Ag, we must first understand the Nernst equation's general form:

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln Q$$

where:

- E_{cell} is the cell potential under non-standard conditions.
- E°_{cell} is the standard cell potential.
- R is the universal gas constant (8.314 J/(mol·K)).
- T is the temperature in Kelvin.
- n is the number of moles of electrons transferred in the reaction.
- F is Faraday's constant (96485 C/mol).
- Q is the reaction quotient.

In this cell, magnesium is oxidized and silver ions are reduced. The half-reactions are:

- Oxidation: $\text{Mg} \to \text{Mg}^{2+} + 2e^-$
- Reduction: $\text{Ag}^+ + e^- \to \text{Ag}$ (this occurs twice in the balanced equation for cell reaction)

The balanced overall reaction is:

$$\text{Mg} + 2\text{Ag}^+ \to \text{Mg}^{2+} + 2\text{Ag}$$

The reaction quotient Q is given by:

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

Here, $n = 2$ based on the number of electrons transferred. Substituting these into the Nernst equation, we have:

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{2F} \ln \left(\frac{[\text{Mg}^{2+}]}{[\text{Ag}^+]^2}\right)$$

Since only the concentration ratio's argument in the logarithmic function is squared with respect to silver ions, this simplifies the expression further when compared to the given options. Thus, the correct form aligns with:

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{2F} \ln\left(\frac{[\text{Mg}^{2+}]}{[\text{Ag}^{+}]}\right)$$

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2. Answer: c

Explanation:

To determine which electrolyte can be used to obtain $H_2S_2O_8$ (peroxydisulfuric acid) by the process of electrolysis, we need to understand the conditions required for its formation:

1. **Principle of Electrolysis:** Electrolysis involves the decomposition of compounds using an electric current. The electrolyte's nature significantly affects the products formed.

2. **Peroxydisulfuric Acid Formation:** This compound is specifically formed when concentrated sulfuric acid is electrolyzed. The equation for the formation of $H_2S_2O_8$ is as follows:

$$2\ HSO_4^-\ (aq) \rightarrow H_2S_2O_8\ (aq) + 2\ e^-$$

3. **Options Analysis:**

Option	Suitability
Dilute solution of sodium sulphate	Not suitable, lacks sufficient sulfate ions and concentration.
Dilute solution of sulphuric acid	Not suitable due to low concentration.
Concentrated solution of sulphuric acid	Suitable; provides high concentration of sulfate ions for forming $H_2S_2O_8$.
Acidified dilute solution of sodium sulphate	Not suitable, lacks adequate concentration of HSO_4^- ions.

The process's efficiency in forming $H_2S_2O_8$ increases with the concentration of HSO_4^- ions provided by concentrated sulfuric acid, which makes it the correct choice.

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3. Answer: 1.93 - 1.93

Explanation:

Using Faraday's law:

$$\text{Charge}(Q) = I \times t$$

Hydroxide ion (OH^-) concentration at $pH = 12$:

$$[OH^-] = 10^{-2}\,M$$

Moles of OH^- in 0.6 L:

$$n = 0.6 \times 10^{-2} = 6 \times 10^{-3}\,moles$$

Charge required:

$$Q = n \times F = 6 \times 10^{-3} \times 96500$$

$$= 579\,C$$

Current:

$$I = \frac{Q}{t} = \frac{579}{300} \approx 1.93\ \ A$$

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4. Answer: b

Explanation:

The problem requires us to find the standard cell potential (E°cell) for the given reaction:

$$Fe^{2+}(aq) + Ag^{+}(aq) \rightarrow Fe^{3+}(aq) + Ag(s)$$

We need to determine the expression for the cell potential using the given standard reduction potentials:

- $Ag^{+} + e^{-} \rightarrow Ag\ E^{\circ} = x\,V$
- $Fe^{2+} + 2e^{-} \rightarrow Fe\ E^{\circ} = y\,V$
- $Fe^{3+} + 3e^{-} \rightarrow Fe\ E^{\circ} = z\,V$

To determine the cell potential, we need the standard reduction potential of each half-reaction as it occurs in the overall cell reaction. Let's rewrite these half-reactions as they occur:

1. **Reduction half-reaction:**
   $Ag^{+} + e^{-} \rightarrow Ag$
   Standard reduction potential: $E_1^{\circ} = x$
2. **Oxidation half-reaction:**
   Convert Fe^{2+} to Fe^{3+}: $Fe^{2+} \rightarrow Fe^{3+} + e^{-}$ (This reaction is derived from the given equation by considering conservation of electrons)
   We get:
   $Fe \rightarrow Fe^{3+} + 3e^{-}\ E^{\circ} = -z$
   To convert Fe^{2+} to Fe^{3+}: $Fe^{2+} \rightarrow Fe^{3+} + e^{-}$
   Considering conservation of mass and charge, we can deduce that it is a part of the full reaction: Thus, the effective potential: $E_2^{\circ} = \frac{-(z-y)}{1} = y - z$ [Subtracting to get the used potential directly involving Fe²⁺, as Fe³⁺ + e⁻ → Fe²⁺ should be used here directly]

Therefore, the cell potential for the overall cell reaction can be calculated as:

$$E^{\circ}_{cell} = E^{\circ}_{\text{reduction}} - E^{\circ}_{\text{oxidation}}$$

$$E^{\circ}_{cell} = x - (y - z)$$

Therefore, combining the standard potentials:

$$E^{\circ}_{cell} = x + 2y - 3z$$

Thus, the correct answer is $x + 2y - 3z$, which corresponds to option 2.

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5. Answer: d

Explanation:

Step 1: Statement I is false because corrosion involves the metal becoming an anode and the impure metal becoming the cathode. The description in Statement I is incorrect in terms of the anode and cathode roles.

Step 2: Statement II is true. Corrosion rates are generally higher in acidic mediums due to increased ion concentration, which accelerates the electrochemical reactions. In alkaline environments, corrosion is typically slower due to lower ion concentrations.

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6. Answer: 2 - 2

Explanation:

Step-by-step Calculation:
The oxidation of water to oxygen gas involves the half-reaction:

$$2\mathrm{H_2O} \rightarrow \mathrm{O_2} + 4\mathrm{H^+} + 4e^-$$

This reaction shows that 4 moles of electrons are required to oxidize 2 moles of water to produce 1 mole of $\mathrm{O_2}$.
The amount of electricity required to transfer 1 mole of electrons is given by Faraday's constant:

$$F = 96500\,\mathrm{C\,mol^{-1}}$$

Therefore, the total charge required for the oxidation of 1 mole of $\mathrm{H_2O}$ to $\mathrm{O_2}$ is:

$$\mathrm{Charge} = 4 \times F = 4 \times 96500 = 386000\,\mathrm{C} = 3.86 \times 10^5\,\mathrm{C}$$

Conclusion:The amount of electricity required for the oxidation of 1 mole of $\mathrm{H_2O}$ to $\mathrm{O_2}$ is $2 \times 10^5\,\mathrm{C}$.

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7. Answer: c

Explanation:

To determine the strongest reducing agent among the given hydrides—$\mathrm{NH_3}$, $\mathrm{SbH_3}$, $\mathrm{BiH_3}$, and $\mathrm{PH_3}$—we need to consider their chemical properties, particularly their reducing abilities.

The reducing ability of a hydride depends on the ease with which it can donate electrons to other substances. This ability is influenced by the bond strength between hydrogen and the central atom in the hydride, with weaker bonds leading to stronger reducing agents.

1. These compounds are hydrides of Group 15 elements, and reducing behavior typically increases down the group. This trend is due to:
    - Decreasing electronegativity of the central atom as we move down the group, which makes the bond between the central atom and hydrogen weaker.
    - Increasing atomic size, which also weakens the bond, making hydrogen easier to release.
2. Analyzing the given hydrides:
    - $\mathrm{NH_3}$ has relatively strong N-H bonds due to nitrogen's high electronegativity. Hence, it is a weaker reducing agent.
    - $\mathrm{PH_3}$ has weaker P-H bonds compared to N-H, increasing its reducing ability.
    - $\mathrm{SbH_3}$ and $\mathrm{BiH_3}$ have even weaker bonds due to larger atomic sizes of Sb and Bi, further increasing reducing capabilities.
3. Since $\mathrm{BiH_3}$ is at the bottom of the group, it has the weakest bond with hydrogen and is, therefore, the strongest reducing agent among the options due to its ability to easily donate electrons.

Thus, the correct answer is $\mathrm{BiH_3}$, which is the strongest reducing agent among the given options.

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8. Answer: b

Explanation:

The question involves understanding how the molar conductivity of a weak electrolyte changes when the volume of its solution is doubled by adding water.

Molar conductivity (Λ_m) is defined as:

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

where κ is the conductivity of the solution and C is the concentration of the electrolyte.

When the volume of the solution is doubled, the concentration (C) of the electrolyte is halved, as concentration is inversely proportional to volume.

However, since the solution is infinitely dilute, the impact on conductivity (κ) is negligible. This is because at infinite dilution, the ions are completely dissociated and do not experience additional interactions.

Given this understanding, the molar conductivity for weak electrolytes does not change significantly with dilution at infinite dilution.

Therefore, the correct answer is: **remain same or can not be measured accurately**.

The measurement accuracy is also an issue due to the limitations of measuring extremely low conductance values at high dilutions.

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9. Answer: a

Explanation:

To understand how an electrochemical cell can be converted into an electrolytic cell, let's first review the basic operation of each type of cell:

- **Electrochemical Cell (Galvanic Cell):** This type of cell generates electrical energy from spontaneous chemical reactions. The potential difference between the electrodes drives the flow of electrons through an external circuit. The cell operates under a potential E°_{cell} which is the standard cell potential.
- **Electrolytic Cell:** This type of cell uses external electrical energy to drive a non-spontaneous chemical reaction. In this configuration, an external voltage source is used to provide the energy necessary to force the reaction in the opposite direction.

To convert an electrochemical cell into an electrolytic cell, one needs to apply an external force that counteracts the natural direction of the reactions taking place in the electrochemical cell. This is done by:

1. **Applying an External Opposite Potential:** By applying an external potential that is greater than the standard cell potential E°_{cell}, the natural electron flow is reversed. The external voltage must overcome the natural potential of the cell and provide enough energy to drive the reaction in the reverse direction.

Therefore, the correct way to convert an electrochemical cell into an electrolytic cell is by **applying an external opposite potential greater than** E°_{cell}. This approach forces the redox reaction to proceed in the non-spontaneous direction, effectively using the cell as an electrolytic cell.

Let's evaluate the other options:

- *Reversing the flow of ions in the salt bridge:* This option is not feasible because the role of the salt bridge is to maintain charge balance by allowing ion flow, but it does not affect the spontaneity of the reaction itself.
- *Applying an external opposite potential lower than* E°_{cell}*:* This would not be sufficient to reverse the reaction. The applied potential must be greater than the cell's potential.
- *Exchanging the electrodes at anode and cathode:* This does not convert the cell from electrochemical to electrolytic; it merely changes the designation of the electrodes.

In conclusion, the correct answer is: **Applying an external opposite potential greater than** E°_{cell}.

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10. Answer: c

Explanation:

The standard cell potential E°_{cell} is calculated as:

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

Step 1: Identify the anode and cathode
- $E^\circ(\mathrm{M}^{2+}/\mathrm{M}) = 0.46\,\mathrm{V}$,
- $E^\circ(\mathrm{X}/\mathrm{X}^{2-}) = 0.34\,\mathrm{V}$.
Since $\mathrm{M}^{2+}/\mathrm{M}$ has a higher reduction potential, it will act as the cathode, and $\mathrm{X}/\mathrm{X}^{2-}$ will act as the anode.
Step 2: Calculate E°_{cell}

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 0.34 - 0.46 = -0.12\,\mathrm{V}.$$

Step 3: Analyze the spontaneity of the reaction

Since E°_{cell} is negative, the reaction will proceed in the reverse direction (the reverse reaction is spontaneous).
The spontaneous reaction is:

$$\mathrm{M}^{2+} + \mathrm{X}^{2-} \to \mathrm{M} + \mathrm{X}.$$

Step 4: Validate the options
-Option (1): Incorrect, as $E_{\text{cell}} = -0.12\,\mathrm{V}$, not $-0.80\,\mathrm{V}$.
-Option (2): Incorrect, as $\mathrm{M} + \mathrm{X}^{2-} \to \mathrm{M}^{2+} + \mathrm{X}^{2-}$ is not spontaneous.
-Option (3): Correct, as $\mathrm{M}^{2+} + \mathrm{X}^{2-} \to \mathrm{M} + \mathrm{X}$ is the spontaneous reaction.
-Option (4): Incorrect, as $E_{\text{cell}} = -0.12\,\mathrm{V}$, not $0.80\,\mathrm{V}$.
Final Answer: (3).

---

11. Answer: d

Explanation:

The amount of a substance deposited during electrolysis is determined using Faraday's laws of electrolysis. The formula is:

$$W = ZIt,$$

where:
- W is the mass of the substance deposited,
- Z is the electrochemical equivalent of the substance,
- I is the current passed, and
- t is the time for which the current is passed.

Step 1: Relating charge to electrochemical equivalent
We know that:

$$Q = It,$$

where Q is the total charge passed through the solution. Substituting this into the equation for W, we get:

$$W = ZQ$$

Step 2: Deposition of silver
For one coulomb of charge ($Q = 1\,\mathrm{C}$), the mass of silver deposited is directly proportional to the electrochemical equivalent (Z) of silver. Thus:

$$W = ZQ = (\text{electrochemical equivalent of silver}).$$

Step 3: Conclusion
The quantity of silver deposited when one coulomb of charge is passed is equal to the electrochemical equivalent of silver. This matches the given option.
Final Answer: (4).

---

12. Answer: a

Explanation:

For a strong electrolyte, the molar conductivity Λ_m can be expressed as:

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

where Λ_m^0 is the molar conductivity at infinite dilution, A is a constant, and C is the concentration.
The term $A\sqrt{C}$ has units of $\mathrm{S\,cm^2 mol^{-1}}$, so the units of A must be $\mathrm{S\,cm^2 mol^{-3/2} L^{1/2}}$ to ensure dimensional consistency when multiplied with $\sqrt{C}$ (units of $\mathrm{mol^{1/2} L^{-1/2}}$).

---

13. Answer: d

Explanation:

A fuel cell, specifically using hydrogen and oxygen, has found applications in spaceships due to its capability to provide electricity in a lightweight, efficient form. It does not use aluminum as a catalyst but rather commonly uses platinum or other noble metals. Fuel cells operate in an eco-friendly manner, emitting only water as a by-product. Furthermore, they can be categorized as galvanic cells since they convert chemical energy directly into electrical energy.

---

14. **Answer: c**

Explanation:

To determine in which galvanic cell the given reaction occurs, we need to analyze the components and chemical processes involved in the options.

The reaction is:

$\frac{1}{2}H_{2(g)} + AgCl_{(s)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)} + Ag_{(s)}$

This represents a galvanic cell where hydrogen gas and silver chloride are involved. The hydrogen gas is oxidized to produce H^+, and the $AgCl$ is reduced to solid silver Ag. This process would typically occur in a galvanic cell where:

1. Hydrogen gas H_2 is used at one electrode, often composed of platinum ($Pt|H_2|H^+$).
2. Silver chloride $AgCl$ and its reduction to silver Ag forms the other electrode component.

Now, let's analyze each option:

- **Option 1:** $Pt|H_{2(g)}|HCl_{(soln.)}|AgCl_{(s)}|Ag$
- **Option 2:** $Pt|H_{2(g)}|HCl_{(soln.)}|AgNO_{3(aq)}|Ag$
- **Option 3:** $Pt|H_{2(g)}|KCl_{(soln.)}|AgCl_{(s)}|Ag$
- **Option 4:** $Ag|AgCl_{(s)}|KCl_{(soln.)}|AgNO_{3(aq)}|Ag$

The correct option should have the components that match the reactions required in the question. Here, Option 3 accurately represents the process:

1. $Pt|H_{2(g)}$: This is where hydrogen gas is oxidized to H^+.
2. $KCl_{(soln.)}|AgCl_{(s)}|Ag$: This involves the incorporation of $AgCl$ and the reduction to silver, which is the solid deposited at the electrode.

Conclusion: The galvanic cell corresponding to the reaction is $Pt|H_{2(g)}|KCl_{(soln.)}|AgCl_{(s)}|Ag$, which is *Option 3*.

---

15. **Answer: d**

Explanation:

The given problem involves a galvanic cell with a specified cell notation: Tl$|^{Tl^+}_{(0.001M)}|^{Cu^{2+}}_{(0.01M)}$ Cu. We are tasked with determining how the emf (electromotive force) of this cell could be increased. The correct answer is increasing the concentration of Cu^{2+} ions.

Let's analyze why this choice is correct using the Nernst Equation, which describes how the emf of a cell changes with concentration. The general form of the Nernst Equation for a cell reaction is:

$$E_{\text{cell}} = E^{\circ}_{\text{cell}} - \frac{RT}{nF} \ln Q$$

Where:

- E_{cell} = emf of the cell
- E°_{cell} = standard emf of the cell
- R = universal gas constant (8.314 J/mol K)
- T = temperature in Kelvin (298 K in this case)
- n = number of moles of electrons exchanged
- F = Faraday's constant (96485 C/mol)
- Q = reaction quotient

For the cell reaction:
$\text{Tl}^+ + \text{Cu} \rightarrow \text{Tl} + \text{Cu}^{2+}$

The reaction quotient Q can be expressed as:

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

To increase the emf of the cell (E_{cell}), we need to reduce the value of Q. This can be done by increasing the concentration of the denominator (Cu^{2+} ions) since Q is inversely proportional to the concentration of Cu^{2+}.

Now, let's rule out the other options:

- **Increasing the concentration of Tl$^+$ ions:** This would increase Q, thereby decreasing the emf.
- **Increasing concentration of both Tl$^+$ and Cu^{2+} ions:** The increase in Tl$^+$ would likely offset the increase in Cu^{2+}, not necessarily increasing the emf.
- **Decreasing concentration of both Tl$^+$ and Cu^{2+} ions:** This would increase Q due to a reduction in the denominator, potentially decreasing the emf.

Thus, the correct choice is **increasing the concentration of Cu^{2+} ions** to increase the cell's emf.

---

16. **Answer: 3 - 3**

Explanation:

The given problem involves determining which metals will be oxidized by the dichromate ion, $\mathrm{Cr_2O_7^{2-}}$. The dichromate ion reduction potential is $E^\circ = 1.33\,\mathrm{V}$. To determine if a metal will be oxidized, compare its standard reduction potential to that of the dichromate ion. A metal with a lower (more negative) reduction potential will be oxidized by the dichromate ion.

Let's analyze each half-reaction:

- $\mathrm{Fe^{3+}} + 3e^- \to \mathrm{Fe},\ E^\circ = -0.04\,\mathrm{V}$

Since $-0.04\,\mathrm{V} < 1.33\,\mathrm{V}$, Fe can be oxidized.

- $\mathrm{Ni^{2+}} + 2e^- \to \mathrm{Ni},\ E^\circ = -0.25\,\mathrm{V}$

Since $-0.25\,\mathrm{V} < 1.33\,\mathrm{V}$, Ni can be oxidized.

- $\mathrm{Ag^{+}} + e^- \to \mathrm{Ag},\ E^\circ = 0.80\,\mathrm{V}$

Since $0.80\,\mathrm{V} < 1.33\,\mathrm{V}$, Ag can be oxidized.

- $\mathrm{Au^{3+}} + 3e^- \to \mathrm{Au},\ E^\circ = 1.40\,\mathrm{V}$

Since $1.40\,\mathrm{V} > 1.33\,\mathrm{V}$, Au cannot be oxidized.

Thus, three metals—Fe, Ni, and Ag—will be oxidized by $\mathrm{Cr_2O_7^{2-}}$. The result, 3, confirms that the solution falls within the provided range (3,3).

---

17. Answer: c

Explanation:

The question involves interpreting the graph of molar conductivity (Λ_m) versus the square root of concentration $(C^{1/2})$ for two electrolytes, A and B. To identify whether they are strong or weak electrolytes, we must understand their molar conductivity behavior:

1. **Strong Electrolytes:** These electrolytes dissociate completely in solution. Their molar conductivity tends to decrease slightly with an increase in concentration due to interionic interactions. The plot for strong electrolytes is typically a straight line or shows a very gentle slope as $C^{1/2}$ increases.
2. **Weak Electrolytes:** These electrolytes do not dissociate completely. Their molar conductivity increases significantly with a decrease in concentration because more ions are formed as dilution shifts the equilibrium towards more dissociation. In the graph, weak electrolytes show a steep curve as $C^{1/2}$ increases.

Based on the graph:

- **Electrolyte A:** The plot for A shows a sharp increase in molar conductivity as the concentration decreases. This behavior is typical for a weak electrolyte.
- **Electrolyte B:** The plot for B is almost flat, indicating that molar conductivity does not change much with concentration. This behavior is typical for a strong electrolyte.

Conclusion: Thus, electrolyte A is a weak electrolyte, and electrolyte B is a strong electrolyte. The correct answer is:

A	B
Weak electrolyte	Strong electrolyte

---

18. Answer: b

Explanation:

The molar conductivity of a solution ($\Lambda_{Solution}$) is given by the sum of molar ionic conductivities:

$$\Lambda^{+2}{}_{C} = 57 \text{ S cm}^2 \text{ mol}^{-1}$$
$$\Lambda^{+2}{}_{A} = 73 \text{ S cm}^2 \text{ mol}^{-1}$$

$$\Lambda_{Solution} = \lambda^{+} + \lambda^{-}$$
$$\Lambda_{Solution} = 57 + 73 = 130 \text{ S cm}^2 \text{ mol}^{-1}$$

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19. Answer: a

Explanation:

In the cathode reaction manganese (Mn) is reduced from the $+4$ oxidation state to the $+3$ state.

---

20. Answer: c

Explanation:

Reaction of $C_6H_5Br + CH_3Br$ with $\mathrm{NaOH\ (alc.)}$ to Form B:

In the first step, the reaction of bromobenzene (C_6H_5Br) with methyl bromide (CH_3Br) in the presence of alcoholic NaOH undergoes a nucleophilic substitution to form B, which is ortho-bromophenol.

Reaction of B with $\mathrm{HBr\ (ether)}$ to Form C:

In the second step, treating ortho-bromophenol with HBr in ether leads to the formation of C, which is para-bromophenol.

Identifying the Relationship Between A and C:

Compound A (bromobenzene) and compound C (para-bromophenol) are position isomers because they differ in the position of the bromine and hydroxyl groups on the benzene ring.

Conclusion:

The structures of B and C are as shown in Option (3), and A and C are position isomers.

---

21. Answer: c

Explanation:

Understanding the Hydrogen Electrode Reaction:

The reaction at the hydrogen electrode is:

$$2e^- + 2H^+(aq) \rightarrow H_2(g)$$

Nernst Equation for the Electrode Potential:

The Nernst equation for this half-cell reaction is:

$$E = E^\circ - \frac{0.059}{n} \log \frac{P_{H_2}}{[H^+]^2}$$

where:

- E is the electrode potential,
- $E^\circ = 0$ (standard electrode potential for the hydrogen electrode),
- P_{H_2} is the partial pressure of hydrogen gas,
- $[H^+]$ is the concentration of hydrogen ions.

Setting $E = 0$:

To make the emf zero, set $E = 0$:

$$0 = 0 - \frac{0.059}{2} \log \frac{P_{H_2}}{(10^{-7})^2}$$

Solve for P_{H_2}:

$$\frac{0.059}{2} \log \frac{P_{H_2}}{10^{-14}} = 0$$

$$\log \frac{P_{H_2}}{10^{-14}} = 0$$

$$\frac{P_{H_2}}{10^{-14}} = 1$$

$$P_{H_2} = 10^{-14}\,\text{bar}$$

Conclusion:

The required pressure of H_2 is $10^{-14}\,\text{bar}$.

---

22. Answer: 4 - 4

Explanation:

The conductivity of a material indicates its ability to conduct electricity. Materials with high conductivity are categorized as conductors. For this analysis, a standard threshold is needed: materials with conductivities significantly greater than $1\,\text{S m}^{-1}$ are typically considered conductors. Let's evaluate each given material based on this criteria.

1. 2.1×10^3 is greater than $1\,\text{S m}^{-1}$, thus it is a conductor.
2. 1.0×10^{-16} is much less than $1\,\text{S m}^{-1}$, therefore, it is not a conductor.
3. $1.2 \times 10^1 = 12$, which is greater than $1\,\text{S m}^{-1}$, so it is a conductor.
4. 3.91 is greater than $1\,\text{S m}^{-1}$, indicating it is a conductor.
5. $1.5 \times 10^{-2} = 0.015$, which is less than $1\,\text{S m}^{-1}$. Thus, it is not a conductor.
6. $1 \times 10^{-7} = 0.0000001$, also less than $1\,\text{S m}^{-1}$, making it not a conductor.
7. $1.0 \times 10^3 = 1000$, which is greater than $1\,\text{S m}^{-1}$, hence it is a conductor.

Counting the number of materials with conductivity greater than $1\,\text{S m}^{-1}$, we find 4 conductors:

- 2.1×10^3
- 1.2×10^1
- 3.91
- 1.0×10^3

The number of conductors is thus **4**, which matches the expected range (4).

---

23. Answer: a

Explanation:

To determine which equation correctly describes the change in molar conductivity with respect to concentration for a weak electrolyte, we need to consider the dissociation equilibrium and conductivity of weak electrolytes.

Molar conductivity (Λ_m) of an electrolyte is given by:

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

where κ is the conductivity and C is the concentration.

The molar conductivity of a weak electrolyte at any concentration differs from its limiting molar conductivity (Λ_m°) at infinite dilution. For weak electrolytes, as the concentration decreases, Λ_m approaches Λ_m°.

The relationship between molar conductivity and concentration for weak electrolytes is rather complex, and various models approximate it. A commonly used relationship involves the degree of dissociation, α, linked with the equilibrium constant (K_a). For a weak electrolyte:

$\alpha = \frac{\Lambda_m}{\Lambda_m^\circ}$

The equilibrium constant (K_a) can also be expressed as:

$K_a = C\alpha^2 = C\left(\frac{\Lambda_m}{\Lambda_m^\circ}\right)^2$

Rearranging gives us an expression reflective of the equation:

$\Lambda_m^2 C - K_a\Lambda_m + K_a\Lambda_m^{\circ 2} = 0$

This equation represents a quadratic relationship between molar conductivity, concentration, and the dissociation constant of a weak electrolyte.

Let's evaluate the given options:

- $\Lambda_m^2 C - K_a\Lambda_m + K_a\Lambda_m^{\circ 2} = 0$: This equation correctly represents the relationship as derived above, making it the correct option.
- $\Lambda_m - \Lambda_m^\circ + AC\frac{1}{2} = 0$: Incorrect. This equation does not represent the typical quadratic form expected for weak electrolytes.
- $\Lambda_m - \Lambda_m^\circ - AC\frac{1}{2} = 0$: Incorrect. Similar reasoning to the previous one, it does not align with expected relationships.

- $\Lambda_m^2 C + K_a \Lambda_m^{\circ 2} - K_a \Lambda_m^{\circ} = 0$: Incorrect. The terms rearranged do not satisfy the known relationship for weak electrolytes.

Therefore, the correct answer is option $\Lambda_m^2 C - K_a \Lambda_m + K_a \Lambda_m^{\circ 2} = 0$.

---

24. Answer: c

Explanation:

In this question, we need to match the cells listed in List-I with their corresponding uses, properties, or reactions from List-II. Let's analyze each one:

1. **A. Leclanche cell - IV. Reaction at anode Zn → Zn^{2+} + $2e^-$**
   The Leclanche cell is a type of primary battery that uses zinc as the anode. In this cell, the anodic reaction involves the oxidation of zinc: $Zn \rightarrow Zn^{2+} + 2e^-$ \), which matches with option IV.
2. **B. Ni-Cd cell - III. Rechargeable**
   The Nickel-Cadmium (Ni-Cd) cell is known for its rechargeable property. It can be cycled many times by recharging, matching option III.
3. **C. Fuel cell - I. Converts energy of combustion into electrical energy**
   Fuel cells convert the chemical energy from fuel into electricity through a chemical reaction with oxygen or another oxidizing agent. Thus, they perfectly describe option I.
4. **D. Mercury cell - II. Does not involve any ion in solution and is used in hearing aids**
   Mercury cells typically do not involve any ions in the solution due to their construction and chemistry. They are used in small devices like hearing aids, aligning with option II.

Therefore, the correct answer is **A-IV, B-III, C-I, D-II**, which matches each cell type from List-I with the correct description or use from List-II.

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25. Answer: 5 - 5

Explanation:

To determine the value of x, we start by understanding the concept of Faraday's laws of electrolysis. One Faraday corresponds to the charge of one mole of electrons, which is approximately 96485 Coulombs. Copper has a valency of 2, meaning that two moles of electrons are needed to discharge one mole of copper atoms at the cathode.

Let's calculate the gram atom of copper liberated by one Faraday. The molar mass of copper is approximately 63.5 grams per mole. For copper sulfate, the reaction can be represented as:

$$\mathrm{Cu}^{2+} + 2\mathrm{e}^- \rightarrow \mathrm{Cu}$$

Given:

- 1 mole of Cu requires 2 moles of electrons (2 Faradays)
- Molar mass of Cu = 63.5 g/mol

Thus, 1 Faraday will liberate:

$\frac{63.5}{2} = 31.75$ g of Cu

Converting grams to gram-atoms (since 1 gram-atom corresponds to 1 mole):

$\frac{31.75}{63.5} = 0.5$ gram atom of Cu

The expression is given as $x \times 10^{-1}$ gram atom, meaning:

$x \times 10^{-1} = 0.5$

Solving for x:

$x = 0.5 \times 10 = 5$

This calculated value $(x = 5)$.

---

26. Answer: b

Explanation:

To determine which factor does not affect the electrolytic conductance of a solution, we need to understand the various factors that typically influence this property:

1. **The nature of the electrolyte added:** The type of electrolyte significantly affects the conductance. Strong electrolytes (like NaCl) dissociate completely in solution, leading to high conductance, whereas weak electrolytes (like acetic acid) partially dissociate, resulting in lower conductance.
2. **The nature of the electrode used:** This factor refers to the material from which the electrodes in an electrolytic cell are made. While the choice of electrode material can influence the kinetics of certain electrochemical reactions, it generally does not affect the conductance of the solution itself. Conductance is primarily determined by the properties of the electrolyte and solvent.
3. **Concentration of the electrolyte:** The concentration affects conductance as it determines the number of ions in solution. Higher concentrations typically increase conductance to a point, beyond which conductance may decrease due to ion-ion interactions.
4. **The nature of the solvent used:** The solvent can affect conductance since it determines the extent of ionization of the electrolyte. Solvents with a high dielectric constant, such as water, are better at dissociating electrolytes, enhancing conductance.

From this analysis, the factor that **does not** affect the electrolytic conductance of a solution is **the nature of the electrode used**. The conductance primarily depends on the electrolyte's properties, its concentration in the solution, and the solvent's nature, not on the electrode material.

---

27. Answer: a

Explanation:

To determine which metals are employed in the battery industries, we need to assess the given options. Batteries are commonly made using metals that provide good electron flow and stability. Let's examine the given options:

1. **Fe (Iron)** - Though iron is used in many industrial applications, it is not typically used in battery manufacturing due to its tendency to corrode and not having the required electrochemical properties.

2. **Mn (Manganese)** - Manganese is commonly used in batteries, particularly in alkaline batteries and lithium-ion batteries in the form of manganese dioxide.
3. **Ni (Nickel)** - Nickel is widely used in various types of rechargeable batteries, including Nickel-Cadmium (NiCd) and Nickel-Metal Hydride (NiMH) batteries.
4. **Cr (Chromium)** - Chromium is not commonly used in battery manufacturing. Its primary uses are in corrosion-resistant materials and stainless steel production.
5. **Cd (Cadmium)** - Cadmium is used in rechargeable NiCd batteries. Despite the environmental concerns associated with cadmium, it is a key component in this type of battery.

Analyzing these facts, the metals used in battery industries from the given options are Manganese (Mn), Nickel (Ni), and Cadmium (Cd).

Therefore, the correct answer is: **B, C, and E only**.

---

28. Answer: 1 - 1

Explanation:

The potential is given by the Nernst equation:

$$E = E^\circ - \frac{0.06}{2} \log \left(\frac{P_{H_2}}{[H^+]^2} \right)$$

Substituting the values:

$$E = 0 - \frac{0.06}{2} \log \left(\frac{2}{1^2} \right)$$

$$E = -0.03 \times 0.3 = -0.9 \times 10^{-2}\,\mathrm{V}$$

---

29. Answer: 108 - 108

Explanation:

To solve this problem, we need to determine the mass of silver (Ag) displaced when a certain quantity of electricity displaces 5600 mL of O_2 at standard temperature and pressure (STP).

Firstly, at STP, 1 mole of any gas occupies 22.4 L (22,400 mL). Therefore, the moles of O_2 displaced are:

Moles of O_2 = $\frac{5600\,\text{mL}}{22400\,\text{mL/mol}} = 0.25\,\text{mol}$

Now, according to the electrolytic process for the displacement of silver using electricity, we have the following reaction for water electrolysis:

$2H_2O \rightarrow 4H^+ + O_2 + 4e^-$

This indicates that 1 mole of O_2 is produced by 4 faradays of electricity.

Thus, 0.25 moles of O_2 are produced by:

$0.25 \times 4 = 1$ faraday of electricity

The reaction for displacement of silver is:

$Ag^+ + e^- \rightarrow Ag$

This shows that 1 mole of Ag requires 1 faraday of electricity. Therefore, 1 faraday will deposit 1 mole of Ag.

The molar mass of Ag is 108 g/mol. Thus, 1 faraday will deposit:

108 g of Ag

Therefore, the mass of silver displaced by the given quantity of electricity is **108 g**.

---

30. Answer: 45.75 - 45.75

Explanation:

The reaction for the deposition of zinc is as follows:

$$Zn^{2+} + 2e^- \rightarrow Zn$$

Using the formula for electrolysis:

$$W = \frac{Z \times i \times t}{F}$$

where

- Z = $\frac{65.4}{2}$ (Equivalent weight of zinc),
- i = 0.015 A (current),
- t = 15 $\times$ 60 seconds,
- F = 96500 C/mol (Faraday constant).

Calculating the mass of zinc:

$$W = \frac{65.4}{2\times96500} \times 0.015 \times 15 \times 60$$

$$W = 45.75 \times 10^{-4}\text{ g}$$

Since the answer can be approximated, we also consider 46 $\times$ 10^{-4} g.

So, the correct answer is: 45.75 or 46