DPP

DAILY PRACTICE PROBLEMS

Class : XIIth
Date :

Solutions

Subject : CHEMISTRY
DPP No. : 10

Topic :- Electro Chemistry

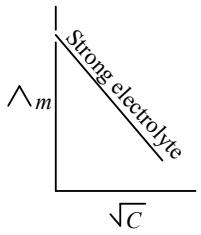
1 **(c)**

Equivalent conductance $= \frac{1000 \times \text{conductance} \times \text{cell constant}}{\text{normality}}$

So, units are, $\Omega^{-1}\,\text{cm}^2\text{equiv}^{-1}$ or $\text{S}\,\text{cm}^2\text{equiv}^{-1}$.

2 **(b)**

For strong electrolytes the plot of molar conductance (Λ_m) $vs.$ $\sqrt{C}$ is linear.



Variation of molar conductance

$(\Lambda_m)$with$\sqrt{C}$ for strong electrolyte.

3 **(b)**

$\Lambda^{\infty}_{eq}(NH_4\,OH) = \Lambda^{\infty}_{eq}\,(NH_4Cl) + \Lambda^{\infty}_{eq}\,(NaOH) - \Lambda^{\infty}_{eq}\,(NaCl)$

$= (149.74 + 248.1 - 126.4)$

$= 271.44\,\Omega^{-1}\,\text{cm}^2\text{eq}^{-1}$

4 **(c)**

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

$E^{\circ}_{Ag^+/Ag} - E^{\circ}_{Cu^+/Cu}$

$= -0.80 - 0.34$

$= +0.46$

5 **(c)**

$E_{cell} = E_{OP_{Ni/Ni^{2+}_2}} + E_{RP_{Au^{3+}/Au}}$

$= E^{\circ}_{OP_{Ni}} - \frac{0.059}{2}\log[Ni^{2+}] + E^{\circ}_{RP_{Au}} + \frac{0.059}{3}\log[Au^{3+}]$

$= 0.25 - \frac{0.059}{2}\log(1.0) + 1.50 + \frac{0.059}{3}\log 1.0 = 1.75\ \text{V}$

6 **(d)**

The metal placed below in electrochemical series does not react with that metal salt solution which metal is placed above in series.

8 **(a)**

Cell representation is done as follows

Anode | Anodic electrolyte || cathodic electrolyte | cathode

(i) Oxidation is loss of electron and it takes place at anode. Reduction is gain of electron and it takes place at cathode.

∴ For cell reaction,

$Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$

Zn is anode and Cu is cathode.

∴ Cell representation is

$Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$

9 **(c)**

$2 \times$ 96500 C electricity is used to liberate

$=$ 22400 mL O_2 at STP

∴ 9.65×1000 C electricity will liberate

$$= \frac{22400 \times 9.65 \times 1000}{2 \times 96500}$$

$= 1120$mL

10 **(d)**

Oxidation half-cell

$ClO_3^- \rightarrow ClO_2^- + 2e^-$; $E°_{cell} = -0.36$ V

Reduction half-cell

$ClO_3^- + 2e^- \rightarrow ClO_2^-$; $E°_{cell} = 0.33$ V

$$E°_{\text{cell}} = 0.33 - 0.36 = -0.03 = \frac{RT}{2F} \ln K$$

or $-0.03 = \frac{0.059}{2} \log K$ or $K = 0.1$

$2ClO_3^- \rightleftharpoons ClO_4^- + ClO_2^-$

$0.1 - 2x \quad x \quad x$

$$\frac{x^2}{(0.1-2x)^2} = 0.1$$

or $x = 1.9 \times 10^{-2}$

11 **(a)**

Cu is placed above Ag in electrochemical series, hence it can replace Ag from its salts solution. Therefore, the reaction occur as follows

$$\overset{0}{Cu} + AgNO_3 \xrightarrow{\text{Oxidation}} \overset{+1}{CuNO_3} + Ag$$

12 **(a)**

$E°$ does not depend on stoichiometry of change.

13 **(d)**

HCl is strong electrolyte and H^+ has highest conducting power due to Grothus conductance.

14 **(d)**

$Fe^{2+} + 2e \longrightarrow Fe; \quad -\Delta G_1 = 2 \times (-0.44) \times F$

$Fe^{3+} + 3e \longrightarrow Fe; \quad -\Delta G_2 = 3 \times (-0.036) \times F$

$Fe^{2+} \longrightarrow Fe^{3+} + e; \Delta G_3 = -0.88 + 0.108 = 0.772$ or $0.772 = 1 \times E° \times F$

$\therefore Fe^{3+} + e \longrightarrow Fe^{2+}; E° = +0.772$ V

Above procedure should be used only when two half reactions on algebraic sum give a third half reaction.

15 **(b)**

Chromium is more electropositive metal than iron. In stainless steel, chromium forms an oxide layer and thus it protects steel from corrosion.

16 **(b)**

$Cu^{2+} + 2KI \longrightarrow CuI_2 + 2K^+;$

$2CuI_2 \longrightarrow Cu_2I_2 + I_2$

17 **(d)**

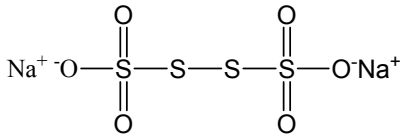
Salt bridge is used to remove or eliminate liquid junction potential arised due to different relative speed of ions of electrolytes at the junction of two electrolytes in an electrochemical cell. Thus, a salt bridge such as KCl is placed in between two electrolytes. A salt used for this purpose should have almost same speeds of its cation and anion.

18 **(b)**

$$w_{Ag} = \frac{E_{Ag} \times Q}{96500} = \frac{108 \times 9.65}{96500} = 1.08 \times 10^{-2} \text{ g} = 10.8 \text{ mg}$$

19 **(d)**

$Na_2S_4O_6$ is



The two S atoms which are linked to each other have 0 oxidation number. The oxidation number of other two S-atoms can be calculated as

$2x + 2 \times 0 + 6 \times -2 = -2$

$2x = 12 - 2 = 10$

$x = +5$

20 **(a)**

Higher the negative value of E°, more is the reducing power.

The order of E° values (negative value) is

$$-2.37 > -0.76 > -0.44$$

(Mg) (Zn) (Fe)

$\therefore$ Mg can reduce both Zn^{2+} and Fe^{2+}. Zn can reduce Fe^{2+}, but not Mg^{2+}. Fe cannot reduce Mg and Zn but can oxidize them.

ANSWER-KEY										
Q.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
A.	**C**	**B**	**B**	**C**	**C**	**D**	**A**	**A**	**C**	**D**
Q.	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
A.	**A**	**A**	**D**	**D**	**B**	**B**	**D**	**B**	**D**	**A**