Chemistry

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(Chapter 2) (Electrochemistry) (Intext Questions)

(Class 12)

Ouestion 2.1:

How would you determine the standard electrode potential of the system Mg2+ | Mg?

Answer 2.1:

The standard electrode potential of Mg2+ | Mg can be measured with respect to the standard hydrogen electrode, represented by Pt(s), H2(s) (1atm) | H(an) (1M).

A cell, consisting of Mg | MgSO4 (aq 1M) as the anode and the standard hydrogen electrodeas the cathode, is set up.

Then, the emf of the cell is measured and this measured emf is the standard electrode potential of the magnesium electrode.

$$E^{\circ} = E_{R}^{\circ} - E_{I}^{\circ}$$

Here, E^o_B for the standard hydrogen electrode is zero.

$$\therefore E^{o} = 0 - E^{o}_{L} = -E^{o}_{L}$$

Opestion 2.2:

Can you store copper sulphate solutions in a zinc pot?

Answer 2.2:

Zinc is more reactive than copper. Therefore, zinc can displace copper from its salt solution.

If copper sulphate solution is stored in a zinc pot, then zinc will displace copper from the Copper sulphate solution: Zn + CuSO₄ → ZnSO₄ + Cu

Hence, copper sulphate solution cannot be stored in a zinc pot.

Question 2.3:

Consult the table of standard electrode potentials and suggest three substances that can oxidise ferrous ions under suitable conditions.

Answer 2.3:

Substances that are stronger oxidizing agents than ferrous ions can oxidise ferrous ions.

$$Fe^{2+} \rightarrow Fe^{3+} + e^{-}; E^{o} = -0.77 \text{ V}$$

This implies that the substances having higher reduction potentials than +0.77 V can oxidise ferrous ions to ferric ions. Three substances that can do so are F2, Cl2, and O2.

Ouestion 2.4:

Calculate the potential of hydrogen electrode in contact with a solution whose pH is 10. Answer 2.4:

For hydrogen electrode, $H^+ + e^- \rightarrow \frac{1}{2}H_2$, it is given that pH = 10

$$..[H^+] = 10^{-10} \text{ M}$$

Now, using Nernst equation:

$$H_{\left(\frac{H^{+}}{2H_{2}}\right)} = E_{\left(\frac{H^{+}}{2H_{2}}\right)}^{0} - \frac{RT}{nF} \ln \frac{1}{[H^{+}]}$$

$$E_{(\frac{16^{+}}{1/2102})}^{0} - \frac{0.0591}{1} \log \frac{1}{[H^{+}]} = 0 - \frac{0.0591}{1} \log \frac{1}{[10^{-10}]}$$
= - 0.0591 log 10¹⁰ = - 0.591 V

Question 2.5:

Calculate the emf of the cell in which the following reaction takes place:

Ni (s) + $2Ag^{+}(0.002 \text{ M}) \rightarrow Ni^{2+}(0.160 \text{ M}) + 2Ag$ (s) Given that $E_{cell}^{o} = 1.05 \text{ V}$

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