

Chemistry

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(Chapter 2) (Electrochemistry) (Intext Questions)
(Class 12)

Question 2.1:

How would you determine the standard electrode potential of the system $\text{Mg}^{2+} | \text{Mg}$?

Answer 2.1:

The standard electrode potential of $\text{Mg}^{2+} | \text{Mg}$ can be measured with respect to the standard hydrogen electrode, represented by $\text{Pt}_{(s)}, \text{H}_{2(g)}(1\text{atm}) | \text{H}^+_{(aq)}(1\text{M})$.

A cell, consisting of $\text{Mg} | \text{MgSO}_4(\text{aq } 1\text{M})$ as the anode and the standard hydrogen electrode as 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.

$$E^\circ = E^\circ_{\text{R}} - E^\circ_{\text{I}}$$

Here, E°_{R} for the standard hydrogen electrode is zero.

$$\therefore E^\circ = 0 - E^\circ_{\text{L}} = -E^\circ_{\text{L}}$$

Question 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: $\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$

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.



This implies that the substances having higher reduction potentials than $+0.77 \text{ V}$ can oxidise ferrous ions to ferric ions. Three substances that can do so are F_2 , Cl_2 , and O_2 .

Question 2.4:

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

Answer 2.4:

For hydrogen electrode, $\text{H}^+ + e^- \rightarrow \frac{1}{2} \text{H}_2$, it is given that $\text{pH} = 10$

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

Now, using Nernst equation:

$$\begin{aligned} E\left(\frac{\text{H}^+}{\frac{1}{2}\text{H}_2}\right) &= E^\circ\left(\frac{\text{H}^+}{\frac{1}{2}\text{H}_2}\right) - \frac{RT}{nF} \ln \frac{1}{[\text{H}^+]} \\ E^\circ\left(\frac{\text{H}^+}{\frac{1}{2}\text{H}_2}\right) &= \frac{0.0591}{1} \log \frac{1}{[\text{H}^+]} = 0 - \frac{0.0591}{1} \log \frac{1}{[10^{-10}]} \\ &= -0.0591 \log 10^{10} = -0.591 \text{ V} \end{aligned}$$

Question 2.5:

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



Given that $E^\circ_{\text{cell}} = 1.05 \text{ V}$

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