

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:

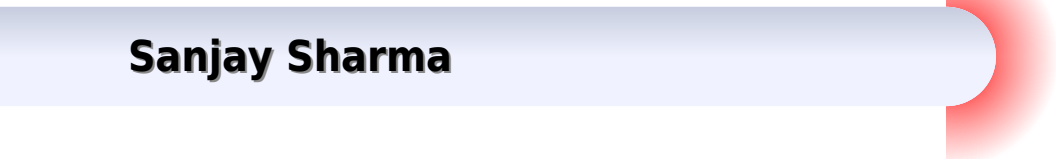
Calculate the *emf* of the cell in which the following reaction takes place:



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

Ncert Chemistry 12 Intext Answer Chapter 2

Sanjay Sharma



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TABLE OF

CONTENT The Solid State Solutions Electrochemistry Chemical Kinetics Surface Chemistry Elements ke Isolation ke General Principles evmProcesses The p Block Elements The d and f Block Elements Coordination Compounds Haloalkanes and Haloarenes Alcohols Phenols and Ethers Aldehydes Ketones va Carboxylic Acids Amines Biomolecules Polymers Chemistry in Everyday Life **NCERT Solutions Chemistry 12th** Geeta Rastogi,2014-01-01 **Errorless New Syllabus**

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Table of Contents Ncert Chemistry 12 Intext Answer Chapter 2

1. Understanding the eBook Ncert Chemistry 12 Intext Answer Chapter 2
 - The Rise of Digital Reading Ncert Chemistry 12 Intext Answer Chapter 2
 - Advantages of eBooks Over Traditional Books
2. Identifying Ncert Chemistry 12 Intext Answer Chapter 2
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Ncert Chemistry 12 Intext Answer Chapter 2
 - User-Friendly Interface
4. Exploring eBook Recommendations from Ncert Chemistry 12 Intext Answer Chapter 2
 - Personalized Recommendations
 - Ncert Chemistry 12 Intext Answer Chapter 2 User Reviews and Ratings
 - Ncert Chemistry 12 Intext Answer Chapter 2 and Bestseller Lists
5. Accessing Ncert Chemistry 12 Intext Answer Chapter 2 Free and Paid eBooks
 - Ncert Chemistry 12 Intext Answer Chapter 2 Public Domain eBooks
 - Ncert Chemistry 12 Intext Answer Chapter 2 eBook Subscription Services
 - Ncert Chemistry 12 Intext Answer Chapter 2 Budget-Friendly Options
6. Navigating Ncert Chemistry 12 Intext Answer Chapter 2 eBook Formats

- ePub, PDF, MOBI, and More
 - Ncert Chemistry 12 Intext Answer Chapter 2 Compatibility with Devices
 - Ncert Chemistry 12 Intext Answer Chapter 2 Enhanced eBook Features
7. Enhancing Your Reading Experience
 - Adjustable Fonts and Text Sizes of Ncert Chemistry 12 Intext Answer Chapter 2
 - Highlighting and Note-Taking Ncert Chemistry 12 Intext Answer Chapter 2
 - Interactive Elements Ncert Chemistry 12 Intext Answer Chapter 2
 8. Staying Engaged with Ncert Chemistry 12 Intext Answer Chapter 2
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Ncert Chemistry 12 Intext Answer Chapter 2
 9. Balancing eBooks and Physical Books Ncert Chemistry 12 Intext Answer Chapter 2
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Ncert Chemistry 12 Intext Answer Chapter 2
 10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
 11. Cultivating a Reading Routine Ncert Chemistry 12 Intext Answer Chapter 2
 - Setting Reading Goals Ncert Chemistry 12 Intext Answer Chapter 2
 - Carving Out Dedicated Reading Time
 12. Sourcing Reliable Information of Ncert Chemistry 12 Intext Answer Chapter 2
 - Fact-Checking eBook Content of Ncert Chemistry 12 Intext Answer Chapter 2
 - Distinguishing Credible Sources
 13. Promoting Lifelong Learning
 - Utilizing eBooks for Skill Development
 - Exploring Educational eBooks
 14. Embracing eBook Trends
 - Integration of Multimedia Elements
 - Interactive and Gamified eBooks

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