

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.

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

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:

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

Question 2.5:

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



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

Chemistry Chapter 12 Solutions Answers

Darrell D. Ebbing



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1. Understanding the eBook Chemistry Chapter 12 Solutions Answers
 - The Rise of Digital Reading Chemistry Chapter 12 Solutions Answers
 - Advantages of eBooks Over Traditional Books
2. Identifying Chemistry Chapter 12 Solutions Answers
 - 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 Chemistry Chapter 12 Solutions Answers
 - User-Friendly Interface
4. Exploring eBook Recommendations from Chemistry Chapter 12 Solutions Answers
 - Personalized Recommendations
 - Chemistry Chapter 12 Solutions Answers User Reviews and Ratings
 - Chemistry Chapter 12 Solutions Answers and Bestseller Lists
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 - Adjustable Fonts and Text Sizes of Chemistry Chapter 12 Solutions Answers
 - Highlighting and Note-Taking Chemistry Chapter 12 Solutions Answers
 - Interactive Elements Chemistry Chapter 12 Solutions Answers
- 8. Staying Engaged with Chemistry Chapter 12 Solutions Answers
 - Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Chemistry Chapter 12 Solutions Answers
- 9. Balancing eBooks and Physical Books Chemistry Chapter 12 Solutions Answers
 - Benefits of a Digital Library
 - Creating a Diverse Reading Collection Chemistry Chapter 12 Solutions Answers
- 10. Overcoming Reading Challenges
 - Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
- 11. Cultivating a Reading Routine Chemistry Chapter 12 Solutions Answers
 - Setting Reading Goals Chemistry Chapter 12 Solutions Answers
 - Carving Out Dedicated Reading Time
- 12. Sourcing Reliable Information of Chemistry Chapter 12 Solutions Answers
 - Fact-Checking eBook Content of Chemistry Chapter 12 Solutions Answers
 - 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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