Electrochemistry Practice Test (Sample)

Study Guide

Everything you need from our exam experts!

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Introduction

Preparing for a certification exam can feel overwhelming, but with the right tools, it becomes an opportunity to build confidence, sharpen your skills, and move one step closer to your goals. At Examzify, we believe that effective exam preparation isn't just about memorization, it's about understanding the material, identifying knowledge gaps, and building the test-taking strategies that lead to success.

This guide was designed to help you do exactly that.

Whether you're preparing for a licensing exam, professional certification, or entry-level qualification, this book offers structured practice to reinforce key concepts. You'll find a wide range of multiple-choice questions, each followed by clear explanations to help you understand not just the right answer, but why it's correct.

The content in this guide is based on real-world exam objectives and aligned with the types of questions and topics commonly found on official tests. It's ideal for learners who want to:

- **Practice answering questions under realistic conditions,**
- **Improve accuracy and speed,**
- **Review explanations to strengthen weak areas, and**
- **Approach the exam with greater confidence.**

We recommend using this book not as a stand-alone study tool, but alongside other resources like flashcards, textbooks, or hands-on training. For best results, we recommend working through each question, reflecting on the explanation provided, and revisiting the topics that challenge you most.

Remember: successful test preparation isn't about getting every question right the first time, it's about learning from your mistakes and improving over time. Stay focused, trust the process, and know that every page you turn brings you closer to success.

Let's begin.

How to Use This Guide

This guide is designed to help you study more effectively and approach your exam with confidence. Whether you're reviewing for the first time or doing a final refresh, here’s how to get the most out of your Examzify study guide:

1. Start with a Diagnostic Review

Skim through the questions to get a sense of what you know and what you need to focus on. Your goal is to identify knowledge gaps early.

2. Study in Short, Focused Sessions

Break your study time into manageable blocks (e.g. 30 – 45 minutes). Review a handful of questions, reflect on the explanations.

3. Learn from the Explanations

After answering a question, always read the explanation, even if you got it right. It reinforces key points, corrects misunderstandings, and teaches subtle distinctions between similar answers.

4. Track Your Progress

Use bookmarks or notes (if reading digitally) to mark difficult questions. Revisit these regularly and track improvements over time.

5. Simulate the Real Exam

Once you're comfortable, try taking a full set of questions without pausing. Set a timer and simulate test-day conditions to build confidence and time management skills.

6. Repeat and Review

Don’t just study once, repetition builds retention. Re-attempt questions after a few days and revisit explanations to reinforce learning. Pair this guide with other Examzify tools like flashcards, and digital practice tests to strengthen your preparation across formats.

There’s no single right way to study, but consistent, thoughtful effort always wins. Use this guide flexibly, adapt the tips above to fit your pace and learning style. You've got this!

Questions

1. **What does the reaction quotient Q represent in the Nernst equation?**
   - **A. Temperature**
   - **B. Pressure**
   - **C. The ratio of product activities to reactant activities**
   - **D. The number of electrons transferred**

2. **If a redox reaction is non-spontaneous, what can be said about K?**
   - **A. $K > 1$**
   - **B. $K = 1$**
   - **C. $K < 1$**
   - **D. Cannot determine**

3. **In the reaction $2\ Fe^{2+}(aq) + H_2O_2(aq) \rightarrow 2\ Fe^{3+}(aq) + 2\ OH^-(aq)$, which species is the oxidizing agent?**
   - **A. Fe^{2+}**
   - **B. H_2O_2**
   - **C. Fe^{3+}**
   - **D. OH^-**

4. **In the reaction $Cl_2(g) + S_2O_3^{2-}(aq) \rightarrow Cl^-(aq) + SO_4^{2-}(aq)$ in acidic solution, which species is the reducing agent?**
   - **A. Cl_2**
   - **B. Cl^-**
   - **C. $S_2O_3^{2-}$**
   - **D. SO_4^{2-}**

5. **For a redox reaction that is non-spontaneous, which of the following is true under standard conditions?**
   - **A. $E^\circ > 0$**
   - **B. $\Delta G^\circ < 0$**
   - **C. $K > 1$**
   - **D. $E^\circ < 0$; $\Delta G^\circ > 0$; $K < 1$**

6. In each half-cell of an electrochemical cell, which process occurs at the anode and which at the cathode?

A. Oxidation occurs at the cathode; reduction occurs at the anode

B. Oxidation occurs at both electrodes

C. Reduction occurs at both electrodes

D. Oxidation occurs at the anode; reduction occurs at the cathode

7. In standard cell notation, which side corresponds to the anode?

A. The left side

B. The right side

C. The middle

D. The salt bridge

8. Balancing FeO(aq) + V2O5(aq) → Fe2O3(aq) + VO(aq) in acidic solution, how many electrons are transferred?

A. 4

B. 6

C. 8

D. 2

9. Which redox pair cannot be used to make a voltaic cell?

A. 2 Ga3+ + 3 Fe ⟶ 2 Ga + 3 Fe2+

B. Mn2+ + 2e- ⟶ Mn

C. Fe2+ + 2e- ⟶ Fe

D. Sn2+ + 2e- ⟶ Sn

10. How does temperature affect the Nernst equation?

A. Temperature changes the factor RT/nF.

B. The simple 0.05916/n log form is specific to 25°C.

C. It changes the standard potential E°.

D. At other temperatures, E = E° − (RT/nF) ln Q.

Answers

1. **C**
2. **C**
3. **B**
4. **C**
5. **D**
6. **D**
7. **A**
8. **B**
9. **A**
10. **D**

Explanations

1. What does the reaction quotient Q represent in the Nernst equation?

A. Temperature

B. Pressure

C. The ratio of product activities to reactant activities

D. The number of electrons transferred

The reaction quotient Q in the Nernst equation expresses the current state of the reaction by comparing how much product activity there is to how much reactant activity there is. It isn’t about temperature or pressure itself; those factors enter the equation separately (through E° and the RT/nF term), while Q captures the actual composition of the system at that moment. For a balanced reaction, Q is the product of the activities of the products, each raised to its stoichiometric power, divided by the product of the activities of the reactants, each raised to its stoichiometric power. In dilute solutions, activities are well approximated by concentrations; for gases, by partial pressures; for pure solids/liquids, activities are about 1. The Nernst equation uses this Q via $E = E^\circ - (RT/nF) \ln Q$, so as the system moves toward more products or more reactants, Q changes and the cell potential shifts accordingly. When Q equals the equilibrium quotient K, the potential equals the standard potential E°. The key point is that Q is the ratio of product activities to reactant activities, not temperature, pressure, or the number of electrons transferred.

2. If a redox reaction is non-spontaneous, what can be said about K?

A. $K > 1$

B. $K = 1$

C. $K < 1$

D. Cannot determine

Spontaneity of a reaction is linked to Gibbs free energy and to the equilibrium constant. If a redox reaction is non-spontaneous under standard conditions, its standard Gibbs energy change ΔG° is positive. The relationship $\Delta G^\circ = -RT \ln K$ connects this to the equilibrium constant K. With $\Delta G^\circ > 0$, the natural log term must be negative, so $\ln K < 0$ and thus $K < 1$. This means the reactants are favored at equilibrium, aligning with the non-spontaneous nature of the process. Remember, if $K > 1$, products would be favored and the reaction would be spontaneous; if $K = 1$, the system sits at equilibrium with neither side preferred.

3. In the reaction 2 Fe2+(aq) + H2O2(aq) → 2 Fe3+(aq) + 2 OH−(aq), which species is the oxidizing agent?

A. Fe2+

B. H2O2

C. Fe3+

D. OH−

In redox terms, the oxidizing agent is the species that gets reduced. Here, Fe2+ is oxidized to Fe3+ (it loses an electron), so Fe2+ acts as the reducing agent. Hydrogen peroxide, on the other hand, is reduced to hydroxide ions (oxygen goes from -1 in H2O2 to -2 in OH−, gaining electrons). Because H2O2 accepts electrons, it is the oxidizing agent. The two electrons released by the two Fe2+ ions balance the two electrons gained by H2O2, forming 2 OH−. The iron(III) and hydroxide aren’t the electron acceptors, so the oxidizing agent is hydrogen peroxide.

4. In the reaction Cl2(g) + S2O3^2−(aq) → Cl−(aq) + SO4^2−(aq) in acidic solution, which species is the reducing agent?

A. Cl2

B. Cl−

C. S2O3^2−

D. SO4^2−

In a redox reaction, the reducing agent is the species that donates electrons and is oxidized. Here, thiosulfate ion is transformed into sulfate, so its oxidation state increases (sulfur goes from a lower average oxidation state toward +6 in sulfate). That means thiosulfate loses electrons to Cl2, which is reduced to Cl−. Since thiosulfate supplies electrons, it is the reducing agent.

5. For a redox reaction that is non-spontaneous, which of the following is true under standard conditions?

A. E° > 0

B. ΔG° < 0

C. K > 1

D. E° < 0; ΔG° > 0; K < 1

The key idea is how standard electrochemical parameters reflect spontaneity. For a redox reaction under standard conditions, the sign of the standard cell potential tells you whether the written direction is spontaneous: a positive E°cell means the reaction tends to proceed on its own, while a negative E°cell means it does not and would go in reverse. If the reaction is non-spontaneous in the written direction, E°cell must be negative. That directly ties to Gibbs energy through ΔG° = −nF E°cell: with E°cell negative, ΔG° becomes positive, meaning the process requires energy input and is not spontaneous under standard conditions. Relating to the equilibrium constant, ΔG° = −RT ln K, a positive ΔG° implies ln K is negative, so K < 1. That indicates the equilibrium lies toward the reactants, consistent with non-spontaneity in the forward direction. In short, non-spontaneous in standard conditions corresponds to E° < 0, ΔG° > 0, and K < 1.

6. In each half-cell of an electrochemical cell, which process occurs at the anode and which at the cathode?

A. Oxidation occurs at the cathode; reduction occurs at the anode

B. Oxidation occurs at both electrodes

C. Reduction occurs at both electrodes

D. Oxidation occurs at the anode; reduction occurs at the cathode

Oxidation occurs at the anode and reduction occurs at the cathode. In every half-cell, the electrode where electrons are produced is the anode, so the species there loses electrons. Those electrons travel through the external circuit to the other electrode, where they are consumed by a reduction reaction. This flow direction holds for galvanic cells (where electrons flow spontaneously from anode to cathode) and electrolytic cells (where an external source drives the same electron flow). For example, in a zinc–copper cell, zinc at the anode is oxidized to Zn2+ while electrons leave the electrode; at the cathode, Cu2+ ions gain electrons and are reduced to copper metal. Oxidation state rises at the anode and falls at the cathode.

7. In standard cell notation, which side corresponds to the anode?

A. The left side

B. The right side

C. The middle

D. The salt bridge

The left side is the anode. In standard cell notation, oxidation occurs at the anode and electrons flow from that electrode to the cathode through the external circuit, so the electrode where oxidation happens is written on the left and the reduction occurs on the right. The vertical bars separate phases, and the middle region represents the salt bridge or passage between solutions, not an electrode. For example, in a zinc-cathode cell written as Zn | Zn2+ (aq) || Cu2+ (aq) | Cu, the zinc electrode on the left is the anode.

8. Balancing FeO(aq) + V2O5(aq) → Fe2O3(aq) + VO(aq) in acidic solution, how many electrons are transferred?

A. 4

B. 6

C. 8

D. 2

You count electrons by tracking how oxidation states change for the elements that actually undergo oxidation and reduction, then tie those changes to the balanced stoichiometry in the acidic medium. In FeO, the iron is in the +2 state. In Fe2O3, iron is +3. So iron is oxidized, losing one electron per iron atom that ends up in Fe2O3. If two iron centers participate (as implied by the balanced balance with Fe2O3), that accounts for two electrons lost. Vanadium starts as +5 in V2O5. In VO, vanadium is in the +4 state, so each vanadium atom is reduced by one electron. Each V2O5 molecule has two vanadium atoms; if two V2O5 units participate, that's four vanadium atoms reduced, consuming four electrons. Total electrons moved in the redox process = 2 electrons lost by iron + 4 electrons gained by vanadium = 6 electrons. Thus six electrons are transferred.

9. Which redox pair cannot be used to make a voltaic cell?

A. 2 Ga3+ + 3 Fe ⟶ 2 Ga + 3 Fe2+

B. Mn2+ + 2e- ⟶ Mn

C. Fe2+ + 2e- ⟶ Fe

D. Sn2+ + 2e- ⟶ Sn

The key idea is that a voltaic cell must use two half-reactions that transfer the same number of electrons and together produce a positive cell potential (a spontaneous reaction). In the pair that combines Ga3+/Ga reduction with Fe/Fe2+ oxidation, the Ga3+/Ga couple consumes 3 electrons per Ga atom, while the Fe/Fe2+ couple releases 2 electrons per Fe atom. To balance electrons, you'd multiply the half-reactions appropriately (3 Fe oxidized to Fe2+ and 2 Ga3+ reduced to Ga), giving an overall cell reaction. However, when you look at the standard reduction potentials, Ga3+/Ga has a reduction potential around -0.56 V and Fe2+/Fe around -0.44 V. Using Ga3+/Ga as the cathode and Fe2+/Fe as the anode yields a cell potential E°cell ≈ (-0.56) - (-0.44) ≈ -0.11 V, which is negative. That means this arrangement would not proceed spontaneously as a galvanic (voltaic) cell. The other reductions listed can be paired with an oxidation that yields a positive overall potential, so they can form voltaic cells with an appropriate partner. For example, Mn2+/Mn has a very negative reduction potential and can pair with an electrode couple with a higher (less negative) reduction potential to give a positive E°cell, and similarly for Fe and Sn reductions with suitable oxidants. Thus, the redox pair involving Ga3+/Ga with Fe/Fe2+ is the one that cannot function as a spontaneous voltaic cell in the orientation shown.

10. How does temperature affect the Nernst equation?

A. Temperature changes the factor RT/nF.

B. The simple 0.05916/n log form is specific to 25°C.

C. It changes the standard potential E°.

D. At other temperatures, $E = E^\circ - (RT/nF)\ln Q$.

Temperature shows up in the Nernst equation through the RT/nF factor, which sets how strongly the reaction quotient Q influences the electrode potential. The general form is $E = E^\circ - (RT/nF)\ln Q$. The familiar 0.05916/n log Q form comes from evaluating RT/F at 25°C and converting natural logarithm to base-10. So when temperature changes, you can’t use the 0.05916 constant anymore—you must use the actual RT/nF value (and the natural logarithm form, or convert with the correct temperature). While E° is defined at a reference temperature, if you truly vary temperature, E° may also shift; but the primary way temperature affects the equation is through the RT/nF factor, making the general expression $E = E^\circ - (RT/nF)\ln Q$ the correct form for other temperatures.

Next Steps

Congratulations on reaching the final section of this guide. You've taken a meaningful step toward passing your certification exam and advancing your career.

As you continue preparing, remember that consistent practice, review, and self-reflection are key to success. Make time to revisit difficult topics, simulate exam conditions, and track your progress along the way.

If you need help, have suggestions, or want to share feedback, we’d love to hear from you. Reach out to our team at hello@examzify.com.

Or visit your dedicated course page for more study tools and resources:

https://electrochemistry.examzify.com

We wish you the very best on your exam journey. You've got this!