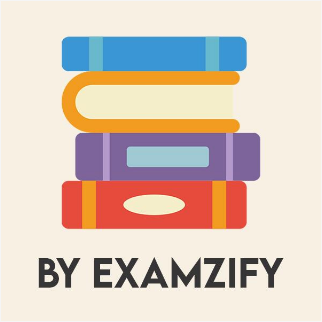
# Electrochemical Practice Exam (Sample)

## Study Guide



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# Questions

1. Why should sulfate ions (SO4^2-) not be capable of oxidizing bromide ions (Br-)?
    - A. SO4^2- has a higher electrode potential
    - B. SO4^2- has a lower electrode potential
    - C. Br- is a stronger oxidizing agent
    - D. SO4^2- does not react with Br-

2. Why is potassium chloride (KCl) not a suitable salt for a salt bridge in the specified cell?
    - A. KCl precipitates under acidic conditions
    - B. KCl does not allow ion movement
    - C. KCl reacts with copper ions
    - D. KCl cannot maintain a constant concentration

3. Increasing which concentration would likely raise the e.m.f of the cell involving H2 and Ag?
    - A. H2(g)
    - B. Ag+(aq)
    - C. H+(aq)
    - D. Cl-(aq)

4. What is the hydrogen ion concentration and hydrogen gas pressure when the potential of the hydrogen electrode is 0.00 V at 298 K?
    - A. H+ ion concentration is 0.01 mol dm-3, H2 pressure is 100 kPa
    - B. H+ ion concentration is 1.00 mol dm-3, H2 pressure is 200 kPa
    - C. H+ ion concentration is 1.00 mol dm-3, H2 pressure is 100 kPa
    - D. H+ ion concentration is 0.10 mol dm-3, H2 pressure is 50 kPa

5. In the context of hydrogen fuel cells, what is one way hydrogen can be utilized?
    - A. To generate heat
    - B. As a fuel in energy cells
    - C. In electrolysis
    - D. To power a combustion engine

6. **What is the standard electrode potential for the positive electrode in the cell configuration?**
   - **A. 2.91V**
   - **B. -3.04V**
   - **C. 0.00V**
   - **D. 3.04V**

7. **How does a salt bridge provide an electrical connection between the two electrodes?**
   - **A. By allowing electrons to flow directly**
   - **B. By allowing ion flow to balance charge**
   - **C. By conducting heat**
   - **D. By insulating against current**

8. **What results from zinc being oxidized to Zn0 in a normal cell reaction?**
   - **A. It can be reduced by applying current**
   - **B. It cannot be reduced by passing current**
   - **C. It produces hydrogen gas**
   - **D. It remains stable**

9. **Which of the following half-reactions represents the reduction of Fe2+ to Fe(s)?**
   - **A. Fe2+ + 2e- → Fe(s)**
   - **B. Fe(s) → Fe2+ + 2e-**
   - **C. Fe3+ + 3e- → Fe(s)**
   - **D. Fe2+ + 2OH- → Fe(OH)2**

10. **What is the standard electrode potential (Eo) for the reduction of ClO3- to Cl-?**
    - **A. +0.17 V**
    - **B. +1.45 V**
    - **C. 0 V**
    - **D. +0.50 V**

# Answers

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

# Explanations

**1. Why should sulfate ions (SO4^2-) not be capable of oxidizing bromide ions (Br-)?**

A. SO4^2- has a higher electrode potential
B. SO4^2- has a lower electrode potential
C. Br- is a stronger oxidizing agent
D. SO4^2- does not react with Br-

The reason sulfate ions (\(SO_4^{2-}\)) should not be capable of oxidizing bromide ions (\(Br^-\)) is primarily due to their electrode potentials. The standard electrode potential values determine the tendency of a species to gain electrons (be reduced) and thus affect its ability to act as an oxidizing agent. A higher electrode potential indicates a greater ability to oxidize other species. In the case of sulfate ions, they have a lower standard electrode potential compared to bromide ions. This means that sulfate is less capable of gaining electrons than bromide, limiting its ability to act as an oxidizer. Since oxidation involves the loss of electrons from one species and the reduction of another, a species with a lower tendency to be reduced (like sulfate) will not effectively oxidize another species (like bromide), which has a higher reduction potential. Therefore, the relationship between the electrode potentials of these ions is a key factor in determining their reactivity and their behavior as oxidizing or reducing agents.

**2. Why is potassium chloride (KCl) not a suitable salt for a salt bridge in the specified cell?**

A. KCl precipitates under acidic conditions
B. KCl does not allow ion movement
C. KCl reacts with copper ions
D. KCl cannot maintain a constant concentration

Potassium chloride (KCl) is typically not a suitable salt for a salt bridge due to its reactivity with copper ions in specific electrochemical cells. In many electrochemical contexts, KCl can dissociate into potassium ($K^+$) and chloride ($Cl^-$) ions. However, copper ions ($Cu^{2+}$) can react with chloride ions to form copper(I) chloride (CuCl), which is a precipitate and can create issues in the electrochemical cell by obstructing the flow of ions. This can lead to an incomplete reaction pathway and disrupt the efficiency and effectiveness of the cell. In contrast, other salts that do not react with metal ions in the cell can maintain the ion flow between the half-cells without causing precipitation or side reactions. Therefore, using a salt that does not interact with the electrodes' metal ions is crucial for the proper functioning of an electrochemical cell.

## 3. Increasing which concentration would likely raise the e.m.f of the cell involving H2 and Ag?

A. H2(g)

B. Ag+(aq)

C. H+(aq)

D. Cl-(aq)

To understand why increasing the concentration of H+(aq) would likely raise the electromotive force (e.m.f) of the cell involving H2 and Ag, it's essential to consider the Nernst equation, which describes how the e.m.f of an electrochemical cell changes with concentration levels of the reactants and products. In the case of a hydrogen electrode, which involves H2 and H+, increasing the concentration of H+ ions shifts the equilibrium toward the reduction of hydrogen ions to hydrogen gas. Since the standard cell potential for the hydrogen half-reaction is set as a reference point (0 volts), any increase in H+ concentration will favor the forward reaction, thus elevating the potential at which this half-cell operates. This, in turn, increases the overall e.m.f of the cell. For the silver electrode, increasing the concentration of Ag+ may influence the cell potential as well, but it does not have the same direct impact on e.m.f as the H+ concentration does in this specific cell. The silver ion reduction is less responsive to changes in concentration compared to the hydrogen half-reaction, particularly when considering standard conditions. In summary, by increasing the concentration of H+(aq), the reaction equilibrium shifts favorably, increasing the cell

## 4. What is the hydrogen ion concentration and hydrogen gas pressure when the potential of the hydrogen electrode is 0.00 V at 298 K?

A. H+ ion concentration is 0.01 mol dm-3, H2 pressure is 100 kPa

B. H+ ion concentration is 1.00 mol dm-3, H2 pressure is 200 kPa

C. H+ ion concentration is 1.00 mol dm-3, H2 pressure is 100 kPa

D. H+ ion concentration is 0.10 mol dm-3, H2 pressure is 50 kPa

In the context of the hydrogen electrode, which serves as a reference for measuring electrode potentials, the standard condition for the hydrogen ion concentration and hydrogen gas pressure is well-defined. When the potential of the hydrogen electrode is set to 0.00 V at 298 K (25 °C), it indicates a standard state of the electrode where typical conditions govern the relationship between the hydrogen ion concentration and hydrogen gas pressure. For a hydrogen electrode, the Nernst equation is used to relate the potential to the concentrations of hydrogen ions and the pressure of hydrogen gas. When the standard potential is 0.00 V, it implies that the activity (or effective concentration) of hydrogen ions $[H^+]$ must be defined such that it adjusts to the standard conditions (usually set at 1.0 M for concentration and 1.0 atm for pressure). However, in this case, we are considering concentrations and pressures that yield a stable environment where the potentials align at this reference point. The accepted standard condition generally describes hydrogen ion concentrations of 1.00 $\text{mol/dm}^3$ and hydrogen gas pressure at 100 kPa for the hydrogen electrode potential at 0.00 V. Thus, in this context, a hydrogen ion concentration of 1.

## 5. In the context of hydrogen fuel cells, what is one way hydrogen can be utilized?

- A. To generate heat
- B. As a fuel in energy cells
- C. In electrolysis
- D. To power a combustion engine

**In the context of hydrogen fuel cells, one way hydrogen can be utilized is as a fuel in energy cells. Hydrogen fuel cells convert the chemical energy of hydrogen directly into electrical energy through an electrochemical reaction, which involves the combination of hydrogen and oxygen. This process results in water and heat as byproducts. Using hydrogen in fuel cells allows for a clean energy source with high efficiency, making it an attractive alternative to traditional fuels. While hydrogen can indeed be combusted in combustion engines, fuel cells represent a distinct method of utilizing hydrogen for producing electricity with minimal emissions, harnessing the energy more efficiently than combustion would allow.**

## 6. What is the standard electrode potential for the positive electrode in the cell configuration?

- A. 2.91V
- B. -3.04V
- C. 0.00V
- D. 3.04V

**The standard electrode potential of the positive electrode in an electrochemical cell indicates the tendency of the electrode to gain electrons, reflecting its ability to function as a cathode in a galvanic (or voltaic) cell. A high positive value, such as 2.91V, signifies a strong oxidizing agent, which effectively pulls electrons towards it. In the context of standard electrode potentials, a positive value indicates a favorable reduction reaction occurring at that electrode compared to the standard hydrogen electrode, which is assigned a potential of 0.00V. This means the electrode with a potential of 2.91V can easily attract electrons from the other half-cell, making it the reaction site where reduction takes place. When evaluating other values, -3.04V represents a scenario where the electrode exhibits a strong tendency to lose electrons rather than gain, indicating a poor oxidizing agent. In contrast, 0.00V is a reference potential which does not indicate a strong tendency for reduction, and 3.04V suggests an inconsistency since electrode potentials are typically referenced against the standard hydrogen electrode and are not usually presented with such high values in practical use. Thus, the recognition of 2.91V as a significant potential aligns well with the**

## 7. How does a salt bridge provide an electrical connection between the two electrodes?

**A. By allowing electrons to flow directly**

**B. By allowing ion flow to balance charge**

**C. By conducting heat**

**D. By insulating against current**

**A salt bridge is an essential component in electrochemical cells, serving to complete the circuit between the two half-cells while maintaining electrical neutrality. It contains an electrolyte solution, typically composed of inert salts, which dissociates into ions. This setup allows the ions in the salt bridge to migrate between the two half-cells. As chemical reactions occur at the electrodes, one electrode generates positive ions (cations) while the other generates negative ions (anions). This leads to a build-up of charge in each half-cell, which would eventually stop the flow of electrons if not balanced. The salt bridge facilitates the movement of ions; cations flow toward the cathode (the electrode where reduction occurs) to neutralize the excess negative charge, while anions flow toward the anode (the electrode where oxidation occurs) to counterbalance the positive charge buildup. This movement of ions is crucial for maintaining the overall charge balance in the electrochemical cell and allows for continuous electron flow through the external circuit. In contrast, the other options do not align with the function of the salt bridge. Specifically, electrons do not flow directly through the bridge; they travel through the external circuit. Heat conduction is not a role of the salt bridge in terms of its primary purpose**

## 8. What results from zinc being oxidized to Zn0 in a normal cell reaction?

**A. It can be reduced by applying current**

**B. It cannot be reduced by passing current**

**C. It produces hydrogen gas**

**D. It remains stable**

**When zinc is oxidized to Zn0, it means that zinc is losing electrons in an electrochemical reaction, transitioning from a more reduced form (e.g., Zn2+) to its elemental metallic state (Zn0). In this context, the process of oxidation is crucial because it indicates that zinc is giving away electrons, resulting in a potential for electron flow in an electrochemical cell. The statement that it cannot be reduced by passing current is correct within the framework of conventional electrochemical reactions. Once zinc has been oxidized to Zn0, it is at a lower energy state compared to its oxidized form. The application of current typically encourages reduction reactions, where electrons are transferred to a species, potentially reversing oxidation. However, since Zn0 is already in its elemental metallic state, it cannot be further reduced under standard conditions. Understanding why the other options do not hold reinforces the rationale. While producing hydrogen gas can be typical in reactions involving specific conditions and species, it is not a direct outcome of zinc being oxidized in standard electrochemical processes. Similarly, stating that zinc can be reduced contradicts the fundamental definition of oxidation, where the substance has already reached a state where it has lost electrons. The notion that it remains stable does not capture the dynamic**

### 9. Which of the following half-reactions represents the reduction of Fe2+ to Fe(s)?

A. Fe2+ + 2e- → Fe(s)

B. Fe(s) → Fe2+ + 2e-

C. Fe3+ + 3e- → Fe(s)

D. Fe2+ + 2OH- → Fe(OH)2

The reduction half-reaction of Fe2+ to Fe(s) involves the gain of electrons by the iron ion. In electrochemistry, a reduction process is defined as the gain of electrons, leading to a decrease in the oxidation state of the species. In the half-reaction presented, Fe2+ receives two electrons (2e-) to transform into solid iron (Fe). The formula highlights that for every Fe2+ ion, two electrons are needed to complete the reduction process. This change illustrates how Fe2+, which is in the +2 oxidation state, is converted to Fe, which is in the zero oxidation state, indicating reduction. This accurately describes the fundamental principles of a reduction reaction. The other choices represent different processes that either depict oxidation, combine iron with hydroxide ions, or involve other oxidation states, and do not indicate the reduction of Fe2+.

### 10. What is the standard electrode potential (Eo) for the reduction of ClO3- to Cl-?

A. +0.17 V

B. +1.45 V

C. 0 V

D. +0.50 V

The standard electrode potential for the reduction of chlorate ions (ClO3-) to chloride ions (Cl-) is +1.45 V, which indicates that this reduction is a thermodynamically favorable process under standard conditions. In electrochemistry, the standard electrode potential, denoted as E°, provides insight into the tendency of a species to gain electrons (be reduced). A positive value signifies that the reduction is likely to occur, as it has a greater tendency to gain electrons compared to the standard hydrogen electrode (which is defined as 0 V). For the reduction of ClO3- to Cl-, the half-reaction can be represented as follows: ClO3- + 6 H+ + 5 e- → Cl- + 3 H2O This reaction occurs in an acidic medium and the involvement of multiple protons and electrons indicates its complexity. The relatively high positive potential of +1.45 V suggests a strong driving force for the reaction, making it effective in various electrochemical applications, including disinfection processes. By understanding the significance of the standard electrode potential, it becomes evident why this particular value reflects the favorable nature of the reduction of chlorate to chloride in electrochemical contexts.