**Electrochemistry Practice Problems**

**Choose one of the following for questions 1–4.**

1. There is no change in the voltage.
2. The voltage becomes zero.
3. The voltage increases.
4. The voltage decreases, but stays positive.
5. The voltage becomes negative.

The following reaction takes place in a voltaic cell.

Zn(s) + Cu2+(1 M) → Cu(s) + Zn2+(1 M)

The cell has a voltage that is measured and found to be +1.10 V.

1. What happens to the voltage when a saturated ZnSO4 solution is added to the zinc compartment of the cell?
2. What happens to the cell voltage when the copper electrode is made smaller?
3. What happens to the cell voltage when the salt bridge is filled with deionized water instead of 1 M KNO3?
4. What happens to the cell voltage after the cell has operated for 10 minutes?
5. MnO4–(aq) + H+(aq) + C2O42–(aq) → Mn2+(aq) + H2O(l) + CO2(g)

What is the coefficient of H+ when the above reaction is balanced?

* 1. 16
	2. 2
	3. 8
	4. 5
	5. 32
1. S2O32– + OH– → SO42– + H2O + e–

After the above half-reaction is balanced, which of the following are the respective coefficients of OH– and SO42–in the balanced half-reaction?

* 1. 8 and 3
	2. 6 and 2
	3. 10 and 2
	4. 5 and 2
	5. 5 and 1
1. How many moles of Pt may be deposited on the cathode when 0.80 F of electricity is passed through a 1.0 M solution of Pt4+?
	1. 1.0 mol
	2. 0.60 mol
	3. 0.20 mol
	4. 0.80 mol
	5. 0.40 mol
2. All of the following may serve as reducing agents, EXCEPT:
	1. Mg
	2. Cs
	3. Fe2+
	4. MnO4–
	5. Br–
3. Cr2O72–– + 14 H+ +3 S2– → 2 Cr3+ + 3 S + 7 H2O

For the above reaction, pick the true statement from the following.

* 1. The S2– is reduced by Cr2O72–.
	2. The oxidation number of chromium changes from +7 to +3.
	3. The oxidation number of sulfur remains –2.
	4. The S2– is oxidized by Cr2O72–.
	5. The H+ oxidizes the S.
1. H+ + NO3– + e– → NO + H2O

What is the coefficient for water arising when the above half-reaction is balanced?

* 1. 3
	2. 4
	3. 2
	4. 1
	5. 6
1. Co2+ + 2 e– → Co *E*° = –0.28 V

Cd2+ + 2 e– → Cd *E*° = –0.40 V

Given the above standard reduction potentials, estimate the approximate value of the equilibrium constant for the following reaction:

Cd + Co2+ → Cd2+ + Co

* 1. 10–4
	2. 10–2
	3. 104
	4. 1016
	5. 102
1. When a basic solution of KMnO4 is added to an SnCl2 solution, a brown precipitate of MnO2 forms and Sn4+ remains in solution. When the same basic solution of KMnO4 is added to an NaF solution, no reaction occurs. Which of the substances involved in these reactions serves as the best reducing agent?
	1. SnCl2
	2. KMnO4
	3. NaF
	4. MnO2
	5. Sn4+
2. A sample of silver is to be purified by electrorefining. This will separate the silver from an impurity of gold. The impure silver is made into an electrode. Which of the following is the best way to set up the electrolytic cell?
	1. an impure silver cathode and an inert anode
	2. an impure silver cathode and a pure gold anode
	3. a pure silver cathode with an impure silver anode
	4. a pure gold cathode with an impure silver anode
	5. an impure silver cathode with a pure silver anode
3. 2 MnO4– + 16 H+ + 5 S2– → 2 Mn2+ + 5 S + 8 H2O

The reducing agent in the above reaction is which of the following?

* 1. MnO4–
	2. H+
	3. S
	4. S2–
	5. Mn2+
1. 2 Fe3+ + Zn → Zn2+ + 2 Fe2+

The reaction shown above was used in an electrolytic cell. The voltage measured for the cell was not equal to the calculated *E* ° for the cell. This discrepancy could be caused by which of the following?

* 1. The anion in the anode compartment was chloride, instead of nitrate as in the cathode compartment.
	2. One or more of the ion concentrations was not 1 M.
	3. Both of the solutions were at 25°C instead of 0°C.
	4. The solution in the salt bridge was Na2SO4 instead of KNO3.
	5. The anode and cathode were different sizes.
1. How many grams of mercury could be produced by electrolyzing a 1.0 M Hg(NO3)2solution with a current of 2.00 A for 3.00 h?
	1. 22.4 g
	2. 201 g
	3. 11.2 g
	4. 44.8 g
	5. 6.00 g
2. An electrolysis cell was constructed with two platinum electrodes in a 1.00 M aqueous solution of KCl. An odorless gas evolves from one electrode, and a gas with a distinctive odor evolves from the other electrode. Choose the correct statement from the following list.
	1. The gas with the distinctive odor was evolved at the anode.
	2. The odorless gas was oxygen.
	3. The gas with the distinctive odor was evolved at the negative electrode.
	4. The odorless gas was evolved at the positive electrode.
	5. The odorless gas was evolved at the anode.
3. H2O2(aq) + KIO4(aq) → KIO3(aq) + O2(g) + H2O(l)

Choose the true statement from the following list.

* 1. The iodine oxidation state is reduced from +8 to +6.
	2. This is not an oxidation–reduction reaction.
	3. H2O2 behaves as a reducing agent.
	4. Hydrogen is reduced from +2 to +1.
	5. H2O2 behaves as an oxidizing agent.

**Questions 19 and 20 are concerned with the following half-reaction in an electrolytic cell:**

2 BrO3– + 12 H+ + 10 e– → Br2 + 6 H2O

1. Choose the correct statement from the following list.
	1. The BrO3– undergoes oxidation at the anode.
	2. Br goes from a –1 oxidation to a 0 oxidation state.
	3. Br2 is oxidized at the anode.
	4. H+ is a catalyst.
	5. The BrO3– undergoes reduction at the cathode.
2. If a current of 5.0 A is passed through the electrolytic cell for 0.50 h, how should you calculate the number of grams of Br2 that will form?
	1. (5.0)(0.50)(3600)(159.8)/(10)
	2. (5.0)(0.50)(3600)(159.8)/(96500)(10)
	3. (5.0)(0.50)(60)(159.8)/(96500)(10)
	4. (5.0)(0.50)(3600)(79.9)/(96500)(10)
	5. (5.0)(0.50)(159.8)/(96500)(10)
3. 2 M(s) + 3 Zn2+(aq) → 2 M3+(aq) + 3 Zn2+(aq) *E*° = 0.90 V

Zn2+(aq) + 2e– → Zn(s) *E*° = –0.76 V

Using the above information, determine the standard reduction potential for the following reaction:

M3+(aq) + 3e–→ M(s)

* 1. 0.90 V
	2. +1.66 V
	3. 0.00 V
	4. –0.62 V
	5. –1.66V

**Answers and Explanations**

1. **D**—The addition of zinc ion, from the ZnSO4, increases the zinc concentration. This increases the numerator in the logarithm part of the Nernst equation. This is a negative term, so the cell voltage will decrease.
2. **A**—The size of the electrode is not important.
3. **B**—The salt bridge serves as an ion source to maintain charge neutrality. Deionized water would not be an ion source, so the cell could not operate.
4. **D**—As the cell operates, the copper ion concentration would decrease and the zinc ion concentration would increase. Both of these changes would make the logarithm term in the Nernst equation more negative. This would decrease the voltage.
5. **A**—The balanced equation is:

2 MnO4–(aq) + 16 H+(aq) + 5 C2O42–(aq) → 2 Mn2+ + 8 H2O(I) + 10 CO2(g)

1. **C**—The balanced equation is:

S2O32– + 10 OH– → 2 SO42– + 5 H2O + 8 e–

1. **C**—It takes 4 mol of electrons (4 F) to change the platinum ions to platinum metal. The calculation would be: (0.80 F)(1 mol Pt/4 F) = 0.20 mol Pt
2. **D**—For a substance to serve as a reducing agent, it must be capable of being oxidized. The manganese, in the MnO4–, is already in its highest oxidation state, so it could not be oxidized. All other answers contain a substance that may be oxidized.
3. **D**—The dichromate ion oxidizes the sulfide ion to elemental sulfur, as the sulfide ion reduces the dichromate ion to the chromium(III) ion. Chromium goes from +6 to +3,while sulfur goes from –2 to 0. The hydrogen remains at +l, so it is neither oxidized nor reduced.
4. **C**—The balanced chemical equation is:

4 H+ + NO3– + 3 e– → NO + 2 H2O

1. **C**—Using the equation:



You should realize that log *K* = 4 gives a *K* = 104. This will give a *K* of about 104(actually, *K* = 1.1 × 104).

1. **A**—The Sn2+, from SnCl2, reduces the manganese from +7 to +4. This makes SnCl2 a reducing agent. The tin is oxidized to Sn4+, so KMnO4 is an oxidizing agent. NaF did nothing, so it behaves as neither an oxidizing nor as a reducing agent.
2. **C**—The impure silver must be oxidized so it will go into solution. Oxidation occurs at the anode. Reduction is required to convert the silver ions to pure silver. Reduction occurs at the cathode. The cathode must be pure silver, otherwise it could be contaminated with the cathode material.
3. **D**—The MnO4– oxidizes the sulfide ion to elemental sulfur, while the sulfide ion reduces the permanganate ion to the manganese(II) ion.
4. **B**—If the voltage was not equal to *E* °, then the cell was not standard. Standard cells have 1 M concentrations, and operate at 25°C with a partial pressure of each gas equal to 1 atm. No gases are involved in this reaction, so the cell must be operating at a different temperature or a different concentration (or both).
5. **A**—

You can estimate the answer by replacing 96,500 with 100,000 and 200.6 with 200.

1. **A**—The gases produced are hydrogen (at the cathode) and chlorine (at the anode). Hydrogen is odorless, while chlorine has a distinctive odor.
2. **C**—The KIO4 oxidizes the H2O2. Thus, H2O2 is the reducing agent. The iodine is reduced from +7 to +5, while the oxygen in the H2O2 is oxidized from –1 to 0 (O2).
3. **E**—The bromate ion, BrO3–, is gaining electrons, so it is being reduced. Reduction always occurs at the cathode.
4. **B**—Recall that 5.0 amp is 5.0 C/s. The calculation would be:



1. **E**—The half-reactions giving the overall reaction must be:



Thus, –0.76 + ? = 0.90, giving ? = 1.66 V. The half-reaction under consideration is the reverse of the one used in this combination, so the sign of the calculated voltage must be reversed. Do not make the mistake of multiplying the voltages when the half-reactions were multiplied to equalize the electrons.