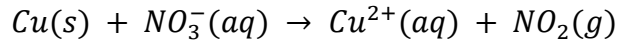


19: Electron Transfer Reactions

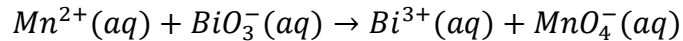
19.2: Oxidation-Reductions Reactions

Balancing Redox Reactions in Acidic Solution

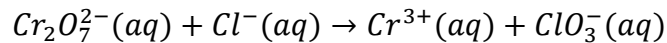
19.2.1. Balance the reaction in an acidic solution:



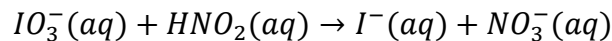
19.2.2. Balance the reaction in an acidic solution:



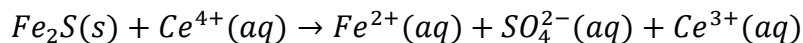
19.2.3. Balance the reaction in an acidic solution:



19.2.4. Balance the reaction in an acidic solution:

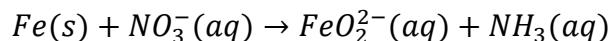


19.2.5. Balance the reaction in an acidic solution:

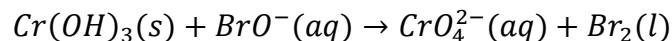


Balancing Redox Reactions in Basic Solutions

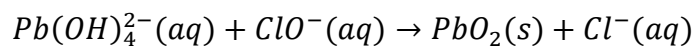
19.2.6. Balance the reactions in the basic solution:



19.2.7. Balance the reactions in the basic solution:

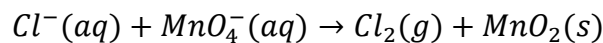


19.2.8. Balance the reactions in the basic solution:

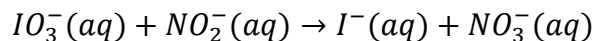


19: Electron Transfer Reactions

19.2.9. Balance the reactions in the basic solution:



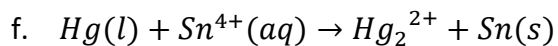
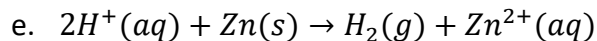
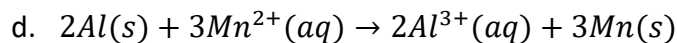
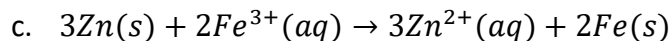
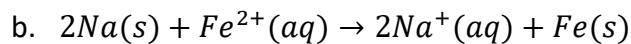
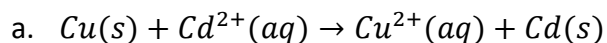
19.2.10. Balance the reactions in the basic solution:



19.3: Electrochemical Cells

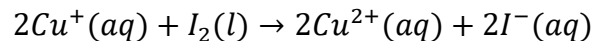
Electrochemical Cell Notation

19.3.1. Write the following reactions in standard electrochemical cell notation at standard state conditions.



Voltaic Cell

19.3.2. What is the half-reaction at the anode in the voltaic reaction?



19.3.3. What is the reaction occurring at the cathode from Question 19.3.2?

19: Electron Transfer Reactions

19.3.4. Which one can occur at the cathode of an electrochemical cell?

- a. $NO \rightarrow NO_3^-$
- b. $Cr_2O_7^{2-} \rightarrow Cr^{7+}$
- c. $I_2 \rightarrow I^-$
- d. none of the above

19.3.5. Which one can occur at the anode of an electrochemical cell?

- a. $Cr_2O_7^{2-} \rightarrow Cr^{7+}$
- b. $Fe^{2+} \rightarrow Fe$
- c. $I_2 \rightarrow I^-$
- d. none of the above

19.3.6. Using the information given, determine which reaction occurs at the anode.

- i. $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$
- ii. $ClO_3^-(aq) + 6H^+(aq) + 6e^- \rightarrow Cl^-(aq) + 3H_2O(l)$
- a. a
- b. b
- c. neither
- d. both

19.3.7. Using the same information in Question 19.3.6, which electrode is consumed?

- a. anode
- b. cathode
- c. neither
- d. both

19.3.8. Using the same information in Question 19.3.6, which electrode is positive?

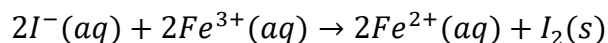
- a. anode
- b. cathode
- c. both
- d. neither

19.4: Electrochemical Cell Fundamentals

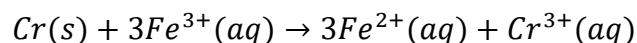
Figure 19.1: Use the following information to answer Questions 19.4.1-19.4.5

Half-reaction	E_{red}^0
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	0.54 V
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	-0.44 V
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	0.77 V
$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$	-0.74 V

19.4.1. Determine the standard potential (V) for the cell reaction.

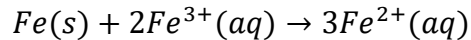


19.4.2. Determine the standard potential (V) for the cell reaction.

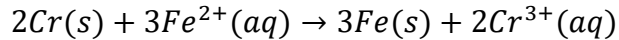


19: Electron Transfer Reactions

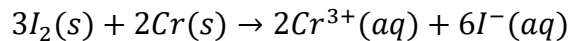
19.4.3. Determine the standard potential (V) for the cell reaction.



19.4.4. Determine the standard potential (V) for the cell reaction.

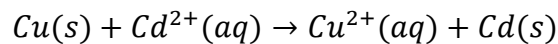


19.4.5. Determine the standard potential (V) for the cell reaction.

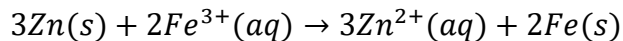


19.5: Standard Electrochemical Potentials

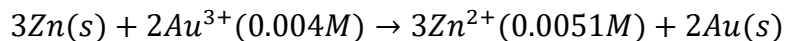
19.5.1. What is the E°_{cell} ?



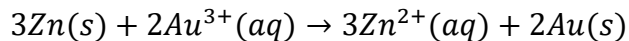
19.5.2. What is the expression for Q, in the equation:



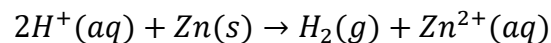
19.5.3. What is the E_{cell} ?



19.5.4. E°_{cell} is +2.28V, what is ΔG ?



19.5.5. E°_{cell} is -0.763, what is the ΔG ?



19: Electron Transfer Reactions

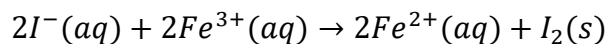
19.6: Electrochemistry and Thermodynamics

Figure 19.2: Use the following table to answer questions 19.6.1-19.6.5

Half-reaction	$\epsilon_{\text{red}}^\circ$
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	0.54 V
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	-0.44 V
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	0.77 V
$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$	-0.74 V

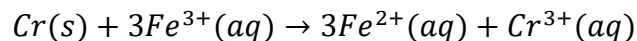
19.6.1. Using the given information, determine the ΔG^0 in J for the following cell reaction.

$F=96500\text{J/V mol}$



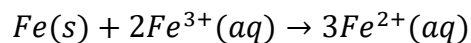
19.6.2. Using the given information, determine the ΔG^0 in J for the following cell reaction.

$F=96500\text{J/V mol}$



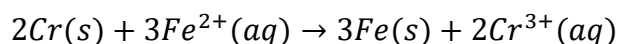
19.6.3. Using the given information, determine the ΔG^0 in J for the following cell reaction.

$F=96500\text{J/V mol}$



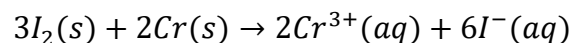
19.6.4. Using the given information, determine the ΔG^0 in J for the following cell reaction.

$F=96500\text{J/V mol}$



19.6.5. Using the given information, determine the ΔG^0 in J for the following cell reaction.

$F=96500\text{J/V mol}$

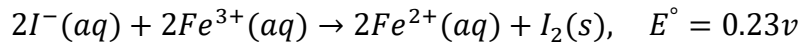


19: Electron Transfer Reactions

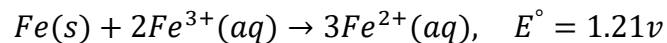
19.7: Electrochemical Cells under Nonstandard Conditions

The Nernst Equation

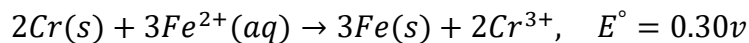
19.7.1. Determine the cell voltage E at 25°C , when $[\text{Fe}^{2+}] = [\text{I}^-] = 0.01\text{M}$, $[\text{Fe}^{3+}] = 0.02\text{M}$.



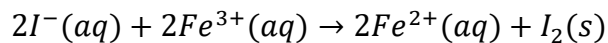
19.7.2. Determine the cell voltage E at 25°C , when $[\text{Fe}^{2+}] = 0.01\text{M}$, $[\text{Fe}^{3+}] = 0.02\text{M}$



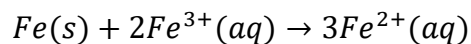
19.7.3. Determine the cell voltage E at 25°C , when $[\text{Fe}^{2+}] = 0.01\text{M}$, $[\text{Cr}^{3+}] = 0.005\text{M}$



19.7.4. Calculate the equilibrium constant for the reaction at 25°C . $R = 8.314\text{J/Kmol}$, $F = 96500\text{J/Vmol}$



19.7.5. For the reaction, if the E_{cell} is 1.50v and $[\text{Fe}^{3+}] = 0.02\text{M}$, determine the concentration of Fe^{2+} .

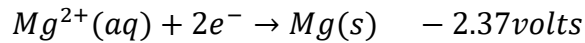
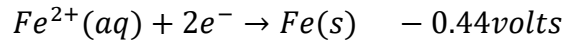


19: Electron Transfer Reactions

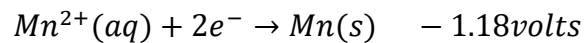
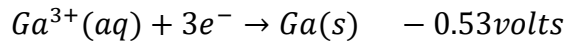
Galvanic Cell Potentials

Figure 19.3: Use the following reactions to answer questions 19.7.6-19.7.11.

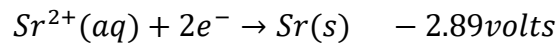
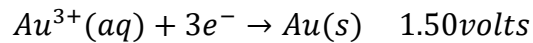
Cell 1:



Cell 2:



Cell 3:



19.7.6. Find the following values for cell 1.

a. What is the E°_{cell} for cell 1?

b. What is the $\Delta G^{\circ}_{\text{rxn}}$ for cell 1?

19.7.7. What is the E_{cell} for cell 1 (0.4M Mg^{2+} and 1.4M Fe^{2+})?

19.7.8. Find the following values for cell 2.

a. What is the E°_{cell} for cell 2?

b. What is the $\Delta G^{\circ}_{\text{rxn}}$ for cell 2?

19.7.9. What is the E_{cell} for cell 2 (0.5M Mn^{2+} and 1.5M Ga^{3+})?

19: Electron Transfer Reactions

19.7.10. Find the following values for cell 3.

- What is the E°_{cell} for cell 3?
- What is the $\Delta G^\circ_{\text{rxn}}$ for cell 3?

19.7.11. What is the E_{cell} for cell 3 (1.2M Sr^{2+} and 0.6M Au^{3+})?

19.9: Electrolysis

19.9.1. How many minutes will it take to plate out 2.50g of chromium metal from Cr^{3+} solution using a current of 32.0 amps?

19.9.2. What current (in A) is needed to plated out 1.50g of chromium metal from Cr^{3+} solution in 30sec?

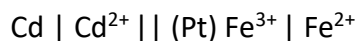
19.9.3. How many grams of copper can be obtained if a current of 10amps passes through a solution of copper (II) sulfate for 20min?

19.9.4. In an electrolytic cell, 0.05g of copper was produced from the copper (II) sulfate solution. How many grams of potassium would be plated out if the same current was applied through molten KCl?

19.9.5. Calculate the number of kilowatt-hrs of electricity required to produce 1.00kg of Al by electrolysis of Al^{3+} if 5.00V of emf is applied.

General Questions

19.1. In the voltaic cell that is represented as



the electron flow will be from

- Pt to Cd^{2+}
- Pt to Cd
- Fe^{2+} to Cd^{2+}
- Cd^{2+} to Fe^{2+}
- Cd to Fe^{3+}

19: Electron Transfer Reactions

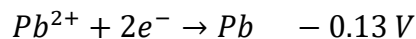
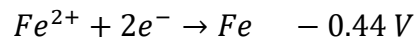
19.2. In the voltaic cell that is represented as



Which of the following statements is false?

- The mass of the zinc electrode decreases during discharge.
- The copper electrode is the anode.
- Electrons flow through the external circuit from the zinc electrode to the copper electrode.
- Reduction occurs at the copper electrode during discharge.
- The concentration of Cu^{2+} decreases during discharge.

19.3. For a galvanic cell using $\text{Fe} \mid \text{Fe}^{2+}(1.0 \text{ M})$ and $\text{Pb} \mid \text{Pb}^{2+}(1.0 \text{ M})$ half-cells, which of the following statements is correct?



- The mass of the iron electrode decreases during discharge.
- Electrons leave the lead electrode to pass through the external circuit during discharge.
- The concentration of Pb^{2+} increases during discharge.
- The iron electrode is the cathode.
- When the cell has completely discharged (to zero voltage), the concentration of Pb^{2+} is zero.

19.4. What is the cell reaction of the voltaic cell $\text{Cr(s)} \mid \text{Cr}^{3+}(\text{aq}) \parallel \text{Cl}^{-}(\text{aq}) \mid \text{Cl}_2(\text{g}) \mid \text{Pt}$?

19.5. When Au is obtained by electrolysis from NaAuCl_4 , what is the minimum number of coulombs required to produce 1.00 mol of gold?

19.6. How many faradays are required to convert a mole of $\text{Cr}_2\text{O}_7^{2-}$ to Cr^{3+} ?

19: Electron Transfer Reactions

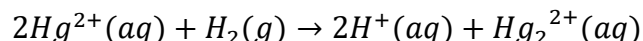
19.7. Which of the following cell reactions would require the use of an inert electrode?

1. $Zn(s) + 2MnO_2(s) + 2NH_4^+(aq) \rightarrow Zn^{2+}(aq) + Mn_2O_3(s) + 2NH_3(aq) + H_2O(l)$
2. $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$
3. $3Cu(s) + 2Au^{3+}(aq) \rightarrow 3Cu^{2+}(aq) + 2Au(s)$
4. $Cl_2(g) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(s)$

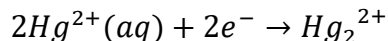
- a. 1 only
b. 1 and 3 only
c. 2 and 3 only
d. 1 and 4 only
e. 3 and 4 only

19.8. How many moles of electrons are produced from a current of 15.0 A in 1.00 hr?

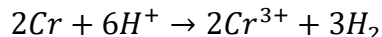
19.9. The following has a potential of 0.92 V:



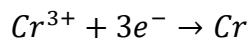
If the concentration of the ions were 1.0 molar and the pressure of H_2 were 1.0 atmosphere, then what would be the E° for the half-reaction?



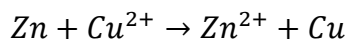
19.10. A cell with the potential of 0.74 V has the cell reaction



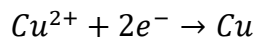
If the concentrations of the ions were 1.0 molar and the pressure of H_2 were 1.0 atmosphere, then what would be the E° of the half-reaction?



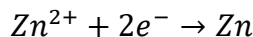
19.11. A standard cell that consisted of a strip of zinc dipped into 1.0 M zinc ion and a strip of copper dipped into 1.0 M copper ion and which was connected by a salt bridge had a potential of 1.10 V



If a potential of 0.60 V were assigned to the half-reaction

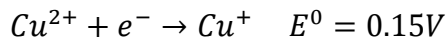
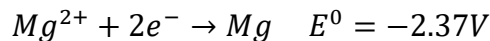


instead of 0.34 V, which is the potential given in a standard reduction potential table, the potential for the reaction would be



19: Electron Transfer Reactions

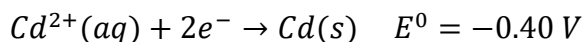
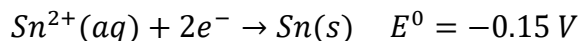
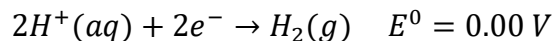
19.12. Consider the following electrode potentials:



Which one of the reactions below will proceed spontaneously from left to right?

- | | |
|--|--|
| a. $\text{Mg}^{2+} + \text{V} \rightarrow \text{V}^{2+} + \text{Mg}$ | c. $\text{V}^{2+} + 2\text{Cu}^{+} \rightarrow \text{V} + \text{Cu}^{2+}$ |
| b. $\text{Mg}^{2+} + 2\text{Cu}^{+} \rightarrow 2\text{Cu}^{2+} + \text{Mg}$ | d. $\text{V} + 2\text{Cu}^{2+} \rightarrow \text{V}^{2+} + 2\text{Cu}^{+}$ |
| | e. none of these |

19.13. Consider the following standard reduction potentials:

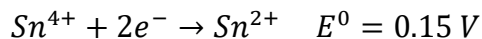
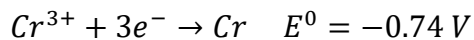


which pair of substances react spontaneously

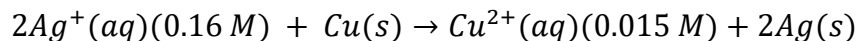
- | | |
|---|-----------------------------|
| a. Sn^{2+} with Cd^{2+} | d. Cd with Sn |
| b. Sn with Cd^{2+} | e. Cd with Sn^{2+} |
| c. Sn^{2+} with H^{+} | |

19.14. Calculate the maximum electrical work obtained when 7.10 grams of Cl_2 gas are consumed in the reaction $\text{Cd}(\text{s}) + \text{Cl}_2(\text{g}) \rightleftharpoons \text{Cd}^{2+}(\text{aq}) + 2\text{Cl}^{-}(\text{aq})$. ($E_{\text{cell}} = 1.76\text{V}$)

19.15. Calculate E° for the cell reaction $2\text{Cr} + 3\text{Sn}^{4+} \rightarrow 3\text{Sn}^{2+} + 2\text{Cr}^{3+}$.



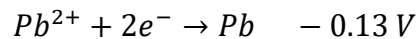
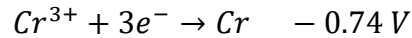
19.16. At 25°C , calculate the voltage of the cell



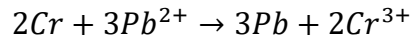
if $E^{\circ}_{\text{cell}} = 0.460\text{V}$.

19: Electron Transfer Reactions

Figure 19.4: Use the following information for Questions 19.17-19.18

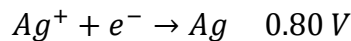
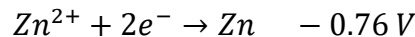


19.17. What is the standard cell potential for the reaction?



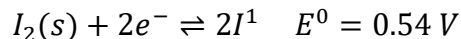
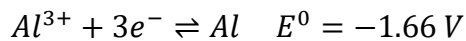
19.18. Calculate the Gibbs free energy change for the reaction above the initial concentration of Cr^{3+} and Pb^{2+} are 1.00M.

19.19. What is the value of the reaction quotient, Q, for the cell that is constructed from the two half-reactions?



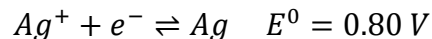
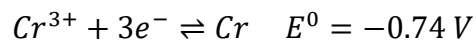
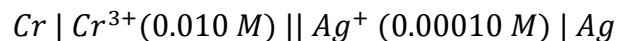
when the Zn^{2+} concentration is 0.0100 M and the Ag^{+} concentration is 1.25M?

19.20. Calculate ΔG° (in joules) for the reaction $2\text{AlI}_3(\text{aq}) \rightleftharpoons 2\text{Al}(\text{s}) + 3\text{I}_2(\text{s})$.



19.21. What is the reduction potential for the half-reaction $\text{Al}^{3+}(\text{aq}) + 3e^{-} \rightarrow \text{Al}(\text{s})$ at 25°C is $[\text{Al}^{3+}] = 0.10 \text{ M}$ and $E^{\circ} = -1.66 \text{ V}$?

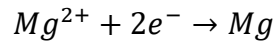
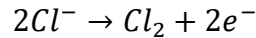
19.22. What is the emf at 25°C for the following cell?



19.23. What is the Cu^{2+} concentration at 25°C in the cell $\text{Zn}(\text{s}) \mid \text{Zn}^{2+}(1.0 \text{ M}) \parallel \text{Cu}^{2+}(\text{aq}) \mid \text{Cu}(\text{s})$? The cell emf is 1.03 V. The standard cell emf is 1.10 V.

19: Electron Transfer Reactions

19.24. Molten magnesium chloride is electrolyzed using inert electrodes and reactions represented by the following equations



concerning this electrolysis which of the following statements is true

- oxidation occurs at the cathode
 - Mg^{2+} ions are reduced at the anode
 - electrons pass through the metallic part of the circuit from Mg^{2+} ions to the Cl^{-} ion
 - Cl^{-} ions are oxidizing agents
 - the cations and the electrolyte undergo reduction
- 19.25. A piece of iron half immerse in a sodium chloride solution will corrode more rapidly than a piece of iron half immersed in pure water because
- the sodium iron oxidized the iron atoms
 - the chloride ions oxidized the iron atoms
 - the chloride ions form a precipitate with iron
 - the chloride ions increase the pH of the solution
 - the sodium ions and chloride ions carry a current through the solution

19.26. A reaction is spontaneous when

- ΔG° is negative or E° is positive
- ΔG° is negative or E° is negative
- ΔG° is positive or E° is negative
- ΔG° is positive or E° is positive
- ΔG° is negative, ΔH° is negative, and E° is negative

19.27. What reaction occurs at the cathode during electrolysis of aqueous $CuSO_4$?

- $2H_2O + 2e^{-} \rightarrow H_2 + 2OH^{-}$
- $Cu \rightarrow Cu^{2+} + 2e^{-}$
- $2H_2O \rightarrow O_2 + 4H^{+} + 4e^{-}$
- $2H^{+} + 2e^{-} \rightarrow H_2$
- $Cu^{2+} + 2e^{-} \rightarrow Cu$

19: Electron Transfer Reactions

19.28. For a certain oxidation-reduction reaction, ΔG° is positive. This means that

- a. ΔG° is negative and K is less than 1
- b. ΔG° is negative and K is greater than 1
- c. ΔG° is zero and K is greater than 1
- d. ΔG° is positive and K is greater than 1
- e. ΔG° is positive and K is less than 1

19.29. Cathodic protection results when

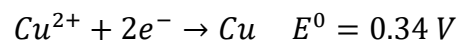
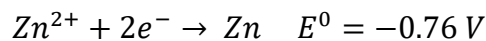
- a. iron is attached to a more active metal
- b. iron is amalgamated with Mercury
- c. iron is tin plated for use as a tin can
- d. iron is painted to protect it from corrosion
- e. iron is made amphoteric

19.30. What mass of chromium could be deposited by electrolysis of an aqueous solution of $\text{Cr}_2(\text{SO}_4)_3$ for 60.0 minutes using constant current of 10.0 amperes? (One faraday = 96,500 coulombs.)

19.31. In a galvanic cell or an electrolysis cell, the cathode is always

- a. the positive electrode
- b. the negative electrode
- c. the positive electrode and the negative electrode, respectively
- d. the electrode at which some species gain electrons
- e. the electrode at which some species lose electrons

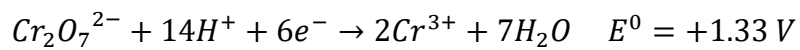
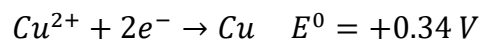
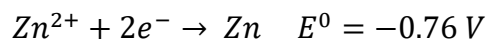
19.32. The voltage of the cell at 25°C would be



- a. between 0.76 and 1.10V
- b. between 0.34 and 0.76 V
- c. between 0.00 and 0.76 V
- d. less than 0.42 V
- e. greater than 1.10 V

19: Electron Transfer Reactions

19.33. On the basis of the following standard electrode potentials, determine which of the following is the strongest oxidizing agent.



- a. Zn^{2+}
- b. Zn
- c. Cu
- d. $\text{Cr}_2\text{O}_7^{2-}$
- e. Cr^{3+}

19.34. Which of the following species is the best reducing agent?

- a. Cd
- b. Mg^{2+}
- c. I^{-}
- d. Sn
- e. F^{-}