

Exam3Fall2009thermoelectro

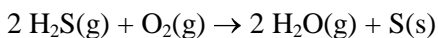
Multiple Choice

Identify the letter of the choice that best completes the statement or answers the question.

- _____ 1. Thermodynamics can be used to determine all of the following EXCEPT
- the direction in which a reaction is spontaneous.
 - the extent to which a reaction occurs.
 - the rate of reaction.
 - the temperature at which a reaction is spontaneous.
 - the enthalpy change of a reaction.
- _____ 2. Which of the following involves a decrease in entropy?
- the sublimation of carbon dioxide
 - the dissolution of NaCl in water
 - the decomposition of $\text{N}_2\text{O}_4(\text{g})$ to $\text{NO}_2(\text{g})$
 - the evaporation of ethanol
 - the freezing of liquid water into ice
- _____ 3. Of the following product-favored processes, which are endothermic?
- the combustion of methane to produce water and carbon dioxide
 - the expansion of an ideal gas
 - the melting of ice at temperatures greater than 0°C .
- 1 only
 - 2 only
 - 3 only
 - 1 and 2
 - 2 and 3
- _____ 4. Calculate the standard molar entropy change for the combustion of methane.
- $$\text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$$
- | Species | S° (J/K·mol) |
|--------------------------------|---------------------|
| $\text{CH}_4(\text{g})$ | 186.3 |
| $\text{O}_2(\text{g})$ | 205.1 |
| $\text{CO}_2(\text{g})$ | 213.7 |
| $\text{H}_2\text{O}(\text{g})$ | 188.8 |
- 5.2 J/K
 - 1.0 J/K
 - +1.0 J/K
 - +5.2 J/K
 - +11.1 J/K
- _____ 5. Predict the signs of ΔH , ΔS , and ΔG for the melting of ice at 50°C .
- $\Delta H < 0$, $\Delta S < 0$, $\Delta G < 0$
 - $\Delta H < 0$, $\Delta S > 0$, $\Delta G < 0$
 - $\Delta H < 0$, $\Delta S > 0$, $\Delta G > 0$
 - $\Delta H > 0$, $\Delta S < 0$, $\Delta G < 0$
 - $\Delta H > 0$, $\Delta S > 0$, $\Delta G < 0$

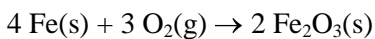
- _____ 6. Above what temperature would you expect a reaction to become spontaneous if $\Delta H = +322 \text{ kJ}$ and $\Delta S = +531 \text{ J/K}$?
- 171 K
 - 209 K
 - 606 K
 - The reaction will be spontaneous at any temperature.
 - The reaction will NOT be spontaneous at any temperature.

- _____ 7. Calculate $\Delta G_{\text{rxn}}^\circ$ for the reaction below at 25.0°C



given $\Delta H_{\text{rxn}}^\circ = -442.4 \text{ kJ}$, and $\Delta S_{\text{rxn}}^\circ = -175.4 \text{ J/K}$.

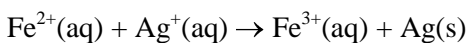
- 438.0 kJ
 - 390.1 kJ
 - 321.9 kJ
 - +3943 kJ
 - +5182 kJ
- _____ 8. Calculate ΔG° for the reaction below at 25.0°C .



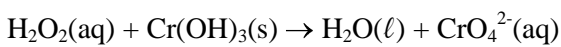
| Species | $\Delta H_f^\circ(\text{kJ/mol})$ | $S_f^\circ(\text{J/K}\cdot\text{mol})$ |
|------------------------------------|-----------------------------------|--|
| Fe(s) | 0 | 27.78 |
| O ₂ (g) | 0 | 205.14 |
| Fe ₂ O ₃ (s) | -824.2 | 87.40 |

- 1629 kJ
 - 1484 kJ
 - 780.8 kJ
 - 659.7 kJ
 - +1629 kJ
- _____ 9. The free energy change for the formation of the complex ion AlF_6^{3-} is -140. kJ at 25°C . What is the equilibrium constant for the reaction?
- 2.9×10^{-25}
 - 5.65×10^1
 - 3.5×10^{24}
 - 5.2×10^{29}
 - 2.3×10^{56}

- _____ 10. In the following reaction,

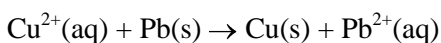


- Fe²⁺ is oxidized and Fe³⁺ is reduced.
 - Fe²⁺ is oxidized and Ag⁺ is reduced.
 - Ag⁺ is oxidized and Ag(s) is reduced.
 - Ag⁺ is oxidized and Fe²⁺ is reduced.
 - Ag⁺ is oxidized and Fe³⁺ is reduced.
- _____ 11. Write a balanced chemical equation for the following reaction in a basic solution.



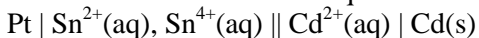
- $2 \text{H}_2\text{O}_2(\text{aq}) + 3 \text{Cr}(\text{OH})_3(\text{s}) \rightarrow \text{H}_2\text{O}(\ell) + 3 \text{CrO}_4^{2-}(\text{aq}) + 11/2 \text{H}^+(\text{aq})$
- $2 \text{H}_2\text{O}_2(\text{aq}) + \text{Cr}(\text{OH})_3(\text{s}) \rightarrow \text{H}_2\text{O}(\ell) + \text{CrO}_4^{2-}(\text{aq}) + 2 \text{OH}^-(\text{aq})$
- $\text{H}_2\text{O}_2(\text{aq}) + 2 \text{Cr}(\text{OH})_3(\text{s}) \rightarrow \text{H}_2\text{O}(\ell) + 2 \text{CrO}_4^{2-}(\text{aq}) + 4 \text{H}_2\text{O}(\ell)$
- $3 \text{H}_2\text{O}_2(\text{aq}) + 2 \text{Cr}(\text{OH})_3(\text{s}) + 4 \text{OH}^-(\text{aq}) \rightarrow 2 \text{CrO}_4^{2-}(\text{aq}) + 8 \text{H}_2\text{O}(\ell)$
- $4 \text{H}_2\text{O}_2(\text{aq}) + 2 \text{Cr}(\text{OH})_3(\text{s}) \rightarrow 2 \text{H}_2\text{O}(\ell) + 2 \text{CrO}_4^{2-}(\text{aq}) + 4 \text{OH}^-(\text{aq})$

- ___ 12. All of the following statements concerning voltaic cells are true EXCEPT
- the two half-cells are connected by a salt bridge.
 - electrons flow from the anode to the cathode.
 - oxidation occurs at the cathode.
 - voltaic cells can be used as a source of energy.
 - a voltaic cell consists of two-half cells.
- ___ 13. What is the correct cell notation for the reaction below?



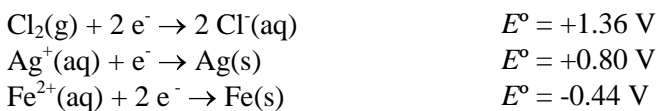
- $\text{Pb} | \text{Pb}^{2+}(\text{aq}) || \text{Cu}^{2+}(\text{aq}) | \text{Cu}$
- $\text{Pb} | \text{Cu}^{2+}(\text{aq}) || \text{Pb}^{2+}(\text{aq}) | \text{Cu}$
- $\text{Pb} | \text{Cu}(\text{s}) || \text{Pb}^{2+}(\text{aq}) | \text{Cu}^{2+}$
- $\text{Cu} | \text{Pb}^{2+}(\text{aq}) || \text{Cu}^{2+}(\text{aq}) | \text{Pb}$
- $\text{Cu} | \text{Cu}^{2+}(\text{aq}) || \text{Pb}^{2+}(\text{aq}) | \text{Pb}$

- ___ 14. Write a balanced chemical equation for the overall reaction represented by the cell notation below.



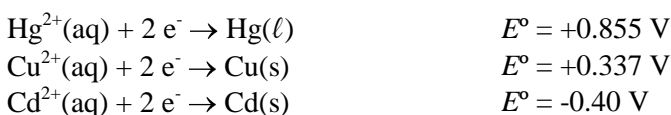
- $\text{Cd}^{2+}(\text{aq}) + \text{Sn}^{4+}(\text{aq}) \rightarrow \text{Cd}(\text{s}) + \text{Sn}^{2+}(\text{aq})$
- $\text{Cd}^{2+}(\text{aq}) + \text{Sn}^{2+}(\text{aq}) \rightarrow \text{Cd}(\text{s}) + \text{Sn}^{4+}(\text{aq})$
- $\text{Cd}(\text{s}) + \text{Sn}^{4+}(\text{aq}) \rightarrow \text{Cd}^{2+}(\text{aq}) + \text{Sn}^{2+}(\text{aq})$
- $\text{Cd}(\text{s}) + \text{Cd}^{2+}(\text{aq}) \rightarrow \text{Sn}^{2+}(\text{aq}) + \text{Sn}^{4+}(\text{aq})$
- $\text{Cd}(\text{s}) + \text{Sn}^{2+}(\text{aq}) \rightarrow \text{Cd}^{2+}(\text{aq}) + \text{Sn}^{4+}(\text{aq})$

- ___ 15. Use the standard reduction potentials below to determine which compound or ion is the best oxidizing agent?



- Cl_2
- Cl^-
- Ag
- Fe^{2+}
- Fe

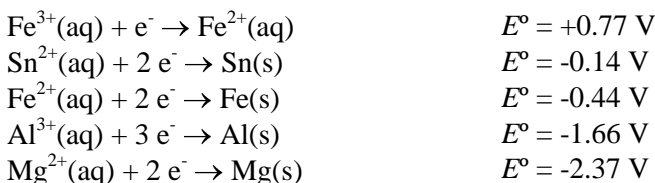
- ___ 16. Use the standard reduction potentials below to determine which compound or ion is the best reducing agent?



- Hg^{2+}

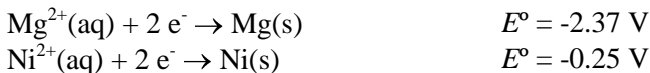
- b. Hg(ℓ)
- c. Cu²⁺
- d. Cd²⁺
- e. Cd

_____ 17. Consider the following half-reactions:

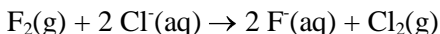


Which of the above metals or metal ions are able to oxidize Al(s)?

- a. Fe³⁺ and Sn²⁺
 - b. Fe³⁺, Sn²⁺, and Fe²⁺
 - c. Fe²⁺, Sn, and Fe
 - d. Mg and Mg²⁺
 - e. Mg²⁺ only
- _____ 18. Given the following two half-reactions, determine which overall reaction is spontaneous and calculate the cell potential.



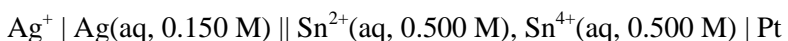
- a. $\text{Mg}^{2+}(\text{aq}) + \text{Ni}(\text{s}) \rightarrow \text{Mg}(\text{s}) + \text{Ni}^{2+}(\text{aq})$ $E_{\text{cell}}^\circ = +2.12 \text{ V}$
 - b. $\text{Mg}^{2+}(\text{aq}) + \text{Ni}(\text{s}) \rightarrow \text{Mg}(\text{s}) + \text{Ni}^{2+}(\text{aq})$ $E_{\text{cell}}^\circ = -2.62 \text{ V}$
 - c. $\text{Mg}^{2+}(\text{aq}) + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Mg}(\text{s}) + \text{Ni}(\text{s})$ $E_{\text{cell}}^\circ = +2.62 \text{ V}$
 - d. $\text{Mg}(\text{s}) + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Ni}(\text{s}) + \text{Mg}^{2+}(\text{aq})$ $E_{\text{cell}}^\circ = +2.12 \text{ V}$
 - e. $\text{Mg}^{2+}(\text{aq}) + \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Ni}(\text{s}) + \text{Mg}(\text{s})$ $E_{\text{cell}}^\circ = -2.12 \text{ V}$
- _____ 19. Calculate E_{cell}° for the following reaction:



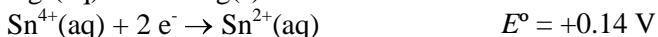
given the following standard reduction potentials.



- a. -4.23 V
 - b. -1.51 V
 - c. 0.76 V
 - d. +1.51 V
 - e. +4.23 V
- _____ 20. Calculate E for the following electrochemical cell at 25°C

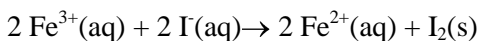


given the following standard reduction potentials.



- a. -0.915 V
- b. +0.61 V
- c. +0.89 V
- d. +0.915 V
- e. +0.99 V

___ 21. E_{cell}° for the following redox reaction is +0.236 V.



What is $\Delta^\circ G$ for this reaction?

- a. -91.0 kJ
- b. -78.1 kJ
- c. -47.2 kJ
- d. -45.5 kJ
- e. -22.8 kJ

___ 22. Al^{3+} is reduced to $\text{Al}(\text{s})$ at an electrode. If a current of 1.00 ampere is passed for 24 hours, what mass of aluminum is deposited at the electrode? Assume 100 % current efficiency.

- a. 1.87 g
- b. 8.05 g
- c. 24.2 g
- d. 54.1 g
- e. 72.5 g

___ 23. The solubility of PbCl_2 is 0.016 mol/L. What is the value of K_{sp} for PbCl_2 ?

- a. 4.1×10^{-6}
- b. 1.6×10^{-5}
- c. 2.6×10^{-4}
- d. 5.1×10^{-4}
- e. 4.8×10^{-2}

___ 24. The K_{sp} for BaF_2 is 1.7×10^{-6} . What is the concentration of Ba^{2+} in a saturated solution of BaF_2 ?

- a. $5.7 \times 10^{-7} \text{ M}$
- b. $1.7 \times 10^{-6} \text{ M}$
- c. $1.0 \times 10^{-2} \text{ M}$
- d. $1.3 \times 10^{-3} \text{ M}$
- e. $7.5 \times 10^{-3} \text{ M}$

___ 25. What is the molar solubility of Ag_2CrO_4 in 0.20 M K_2CrO_4 ? The value of K_{sp} for Ag_2CrO_4 is 9.0×10^{-12} .

- a. $4.5 \times 10^{-11} \text{ mol/L}$
- b. $1.3 \times 10^{-10} \text{ mol/L}$
- c. $3.6 \times 10^{-8} \text{ mol/L}$
- d. $6.7 \times 10^{-6} \text{ mol/L}$
- e. $1.3 \times 10^{-4} \text{ mol/L}$

___ 26. What is the molar solubility of Fe^{3+} in a solution that is buffered at a pH of 4.00? The K_{sp} for $\text{Fe}(\text{OH})_3$ is 6.3×10^{-38} .

- a. $6.3 \times 10^{-23} \text{ mol/L}$

- b. 2.3×10^{-9} mol/L
- c. 6.3×10^{-8} mol/L
- d. 2.3×10^{-6} mol/L
- e. 1.4×10^{-4} mol/L

- _____ 27. The concentration of Ca^{2+} in a solution is 4.7×10^{-4} M. What concentration of CO_3^{2-} is required to just begin precipitating CaCO_3 ? The K_{sp} for CaCO_3 is 3.8×10^{-9} .
- a. 8.1×10^{-6} M
 - b. 5.1×10^{-5} M
 - c. 3.8×10^{-4} M
 - d. 4.7×10^{-4} M
 - e. 1.4×10^{-2} M
- _____ 28. The following anions can be separated by precipitation as silver salts: Cl^- , Br^- , I^- , CrO_4^{2-} . If Ag^+ is added to a solution containing the four anions, each at a concentration of 0.10 M, in what order would they precipitate?

| Compound | K_{sp} |
|---------------------------|-----------------------|
| AgCl | 1.8×10^{-10} |
| Ag_2CrO_4 | 9.0×10^{-12} |
| AgBr | 3.3×10^{-13} |
| AgI | 1.5×10^{-16} |

- a. $\text{AgCl} \rightarrow \text{Ag}_2\text{CrO}_4 \rightarrow \text{AgBr} \rightarrow \text{AgI}$
- b. $\text{AgI} \rightarrow \text{AgBr} \rightarrow \text{Ag}_2\text{CrO}_4 \rightarrow \text{AgCl}$
- c. $\text{Ag}_2\text{CrO}_4 \rightarrow \text{AgCl} \rightarrow \text{AgBr} \rightarrow \text{AgI}$
- d. $\text{Ag}_2\text{CrO}_4 \rightarrow \text{AgI} \rightarrow \text{AgBr} \rightarrow \text{AgCl}$
- e. $\text{AgI} \rightarrow \text{AgBr} \rightarrow \text{AgCl} \rightarrow \text{Ag}_2\text{CrO}_4$

Essay

29. (6 Pts) Historically, to prevent the oxidation of the iron hulls in ocean vessels, large zinc plates were often affixed to the outside of the hull below the waterline. How does the zinc protect the iron hull?
30. (6 Pts) Explain the function of a salt bridge in a voltaic cell.

Exam3Fall2009thermoelectro
Answer Section

MULTIPLE CHOICE

1. ANS: C
2. ANS: E
3. ANS: E
4. ANS: A
5. ANS: E
6. ANS: C
7. ANS: B
8. ANS: B
9. ANS: C
10. ANS: B
11. ANS: D
12. ANS: C
13. ANS: A
14. ANS: B
15. ANS: A
16. ANS: E
17. ANS: B
18. ANS: D
19. ANS: D
20. ANS: B
21. ANS: D
22. ANS: B
23. ANS: B
24. ANS: E
25. ANS: D
26. ANS: C
27. ANS: A
28. ANS: E

ESSAY

29. ANS:
The iron, zinc, and seawater form a voltaic cell. The zinc is more easily oxidized than the iron, so it serves as the anode of the cell. Oxidation occurs at the zinc anode, rather than the iron hull.
30. ANS:
The salt bridge physically separates the redox couples in an electrochemical cell while enabling electrical conduction through the cell. In addition, the salt bridge helps maintain electroneutrality as reduction and oxidation occur at the cathode and anode. As oxidation occurs, anions flow from the salt bridge toward the anode and cations flow from the anode to the salt bridge. Likewise, cations flow from the salt bridge toward the cathode and anions from the cathode to the salt bridge.