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
   a. the direction in which a reaction is spontaneous.
   b. the extent to which a reaction occurs.
   c. the rate of reaction.
   d. the temperature at which a reaction is spontaneous.
   e. the enthalpy change of a reaction.

2. Which of the following involves a decrease in entropy?
   a. the sublimation of carbon dioxide
   b. the dissolution of NaCl in water
   c. the decomposition of N₂O₄(g) to NO₂(g)
   d. the evaporation of ethanol
   e. the freezing of liquid water into ice

3. Of the following product-favored processes, which are endothermic?
   1. the combustion of methane to produce water and carbon dioxide
   2. the expansion of an ideal gas
   3. the melting of ice at temperatures greater than 0ºC.
   a. 1 only
   b. 2 only
   c. 3 only
   d. 1 and 2
   e. 2 and 3

4. Calculate the standard molar entropy change for the combustion of methane.

   \[ \text{CH}_4(g) + 2 \text{O}_2(g) \rightarrow \text{CO}_2(g) + 2 \text{H}_2\text{O}(g) \]

<table>
<thead>
<tr>
<th>Species</th>
<th>( S^\circ ) (J/K·mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH(_4)(g)</td>
<td>186.3</td>
</tr>
<tr>
<td>O(_2)(g)</td>
<td>205.1</td>
</tr>
<tr>
<td>CO(_2)(g)</td>
<td>213.7</td>
</tr>
<tr>
<td>H(_2)O(g)</td>
<td>188.8</td>
</tr>
</tbody>
</table>

   a. -5.2 J/K
   b. -1.0 J/K
   c. +1.0 J/K
   d. +5.2 J/K
   e. +11.1 J/K

5. Predict the signs of \( \Delta H, \Delta S, \) and \( \Delta G \) for the melting of ice at 50ºC.
   a. \( \Delta H < 0, \Delta S < 0, \Delta G < 0 \)
   b. \( \Delta H < 0, \Delta S > 0, \Delta G < 0 \)
   c. \( \Delta H < 0, \Delta S > 0, \Delta G < 0 \)
   d. \( \Delta H > 0, \Delta S < 0, \Delta G < 0 \)
   e. \( \Delta H > 0, \Delta S > 0, \Delta G < 0 \)
6. Above what temperature would you expect a reaction to become spontaneous if \( \Delta H = +322 \text{ kJ} \) and \( \Delta S = +531 \text{ J/K} \)?
   a. 171 K
   b. 209 K
   c. 606 K
   d. The reaction will be spontaneous at any temperature.
   e. The reaction will NOT be spontaneous at any temperature.

7. Calculate \( \Delta G^\circ_{\text{nn}} \) for the reaction below at 25.0ºC

\[
2 \text{H}_2\text{S}(g) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}(g) + \text{S}(s)
\]

given \( \Delta H^\circ_{\text{nn}} = -442.4 \text{ kJ} \), and \( \Delta S^\circ_{\text{nn}} = -175.4 \text{ J/K} \).
   a. -438.0 kJ
   b. -390.1 kJ
   c. -321.9 kJ
   d. +3943 kJ
   e. +5182 kJ

8. Calculate \( \Delta G^\circ \) for the reaction below at 25.0ºC.

\[
4 \text{Fe}(s) + 3 \text{O}_2(g) \rightarrow 2 \text{Fe}_2\text{O}_3(s)
\]

<table>
<thead>
<tr>
<th>Species</th>
<th>( \Delta H^\circ_f \text{(kJ/mol)} )</th>
<th>( S^\circ_f \text{(J/K·mol)} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe(s)</td>
<td>0</td>
<td>27.78</td>
</tr>
<tr>
<td>O(_2)(g)</td>
<td>0</td>
<td>205.14</td>
</tr>
<tr>
<td>Fe(_2)O(_3)(s)</td>
<td>-824.2</td>
<td>87.40</td>
</tr>
</tbody>
</table>

   a. -1629 kJ
   b. -1484 kJ
   c. -780.8 kJ
   d. -659.7 kJ
   e. +1629 kJ

9. The free energy change for the formation of the complex ion AlF\(_6\)\(^{3-}\) is -140. kJ at 25ºC. What is the equilibrium constant for the reaction?
   a. \( 2.9 \times 10^{25} \)
   b. \( 5.65 \times 10^4 \)
   c. \( 3.5 \times 10^{24} \)
   d. \( 5.2 \times 10^{29} \)
   e. \( 2.3 \times 10^{56} \)

10. In the following reaction,

\[
\text{Fe}^{2+}(aq) + \text{Ag}^+(aq) \rightarrow \text{Fe}^{3+}(aq) + \text{Ag}(s)
\]
   a. \( \text{Fe}^{2+} \) is oxidized and \( \text{Fe}^{3+} \) is reduced.
   b. \( \text{Fe}^{2+} \) is oxidized and \( \text{Ag}^+ \) is reduced.
   c. \( \text{Ag}^+ \) is oxidized and \( \text{Ag}(s) \) is reduced.
   d. \( \text{Ag}^+ \) is oxidized and \( \text{Fe}^{2+} \) is reduced.
   e. \( \text{Ag}^+ \) is oxidized and \( \text{Fe}^{3+} \) is reduced.

11. Write a balanced chemical equation for the following reaction in a basic solution.
H₂O₂(aq) + Cr(OH)₃(s) → H₂O(ℓ) + CrO₄²⁻(aq)

a. 2 H₂O₂(aq) + 3 Cr(OH)₃(s) → H₂O(ℓ) + 3 CrO₄²⁻(aq) + 11/2 H⁺(aq)
b. 2 H₂O₂(aq) + Cr(OH)₃(s) → H₂O(ℓ) + CrO₄²⁻(aq) + 2 OH⁻(aq)
c. H₂O₂(aq) + 2 Cr(OH)₃(s) → H₂O(ℓ) + 2 CrO₄²⁻(aq) + 4 H₂O(ℓ)
d. 3 H₂O₂(aq) + 2 Cr(OH)₃(s) + 4 OH⁻(aq) → 2 CrO₄²⁻(aq) + 8 H₂O(ℓ)
e. 4 H₂O₂(aq) + 2 Cr(OH)₃(s) → 2 H₂O(ℓ) + 2 CrO₄²⁻(aq) + 4 OH⁻(aq)

12. All of the following statements concerning voltaic cells are true EXCEPT
   a. the two half-cells are connected by a salt bridge.
   b. electrons flow from the anode to the cathode.
   c. oxidation occurs at the cathode.
   d. voltaic cells can be used as a source of energy.
   e. a voltaic cell consists of two-half cells.

13. What is the correct cell notation for the reaction below?

   Cu²⁺(aq) + Pb(s) → Cu(s) + Pb²⁺(aq)

   a. Pb | Pb²⁺(aq) || Cu²⁺(aq) | Cu
   b. Pb | Cu²⁺(aq) || Pb²⁺(aq) | Cu
   c. Pb | Cu(s) || Pb²⁺(aq) | Cu²⁺(aq)
   d. Cu | Pb²⁺(aq) || Cu²⁺(aq) | Pb
   e. Cu | Cu²⁺(aq) || Pb²⁺(aq) | Pb

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

   Pt | Sn²⁺(aq), Sn⁴⁺(aq) || Cd²⁺(aq) | Cd(s)

   a. Cd²⁺(aq) + Sn⁴⁺(aq) → Cd(s) + Sn²⁺(aq)
   b. Cd²⁺(aq) + Sn²⁺(aq) → Cd(s) + Sn⁴⁺(aq)
   c. Cd(s) + Sn⁴⁺(aq) → Cd²⁺(aq) + Sn²⁺(aq)
   d. Cd(s) + Cd²⁺(aq) → Sn²⁺(aq) + Sn⁴⁺(aq)
   e. Cd(s) + Sn²⁺(aq) → Cd²⁺(aq) + Sn⁴⁺(aq)

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

   Cl₂(g) + 2 e⁻ → 2 Cl⁻(aq) \quad E° = +1.36 \, V
   Ag⁺(aq) + e⁻ → Ag(s) \quad E° = +0.80 \, V
   Fe²⁺(aq) + 2 e⁻ → Fe(s) \quad E° = -0.44 \, V

   a. Cl₂
   b. Cl⁻
   c. Ag
   d. Fe²⁺
   e. Fe

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

   Hg²⁺(aq) + 2 e⁻ → Hg(ℓ) \quad E° = +0.855 \, V
   Cu²⁺(aq) + 2 e⁻ → Cu(s) \quad E° = +0.337 \, V
   Cd²⁺(aq) + 2 e⁻ → Cd(s) \quad E° = -0.40 \, V

   a. Hg²⁺
b. Hg(ℓ)
c. Cu^{2+}
d. Cd^{2+}
e. Cd  

17. Consider the following half-reactions:

- \( \text{Fe}^{3+}(aq) + e^- \rightarrow \text{Fe}^{2+}(aq) \) \( E^\circ = +0.77 \text{ V} \)
- \( \text{Sn}^{2+}(aq) + 2 e^- \rightarrow \text{Sn}(s) \) \( E^\circ = -0.14 \text{ V} \)
- \( \text{Fe}^{2+}(aq) + 2 e^- \rightarrow \text{Fe}(s) \) \( E^\circ = -0.44 \text{ V} \)
- \( \text{Al}^{3+}(aq) + 3 e^- \rightarrow \text{Al}(s) \) \( E^\circ = -1.66 \text{ V} \)
- \( \text{Mg}^{2+}(aq) + 2 e^- \rightarrow \text{Mg}(s) \) \( E^\circ = -2.37 \text{ V} \)

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

a. Fe^{3+} and Sn^{2+}
b. Fe^{3+}, Sn^{2+}, and Fe^{2+}
c. Fe^{2+}, Sn, and Fe
d. Mg and Mg^{2+}
e. Mg^{2+} only

18. Given the following two half-reactions, determine which overall reaction is spontaneous and calculate the cell potential.

- \( \text{Mg}^{2+}(aq) + 2 e^- \rightarrow \text{Mg}(s) \) \( E^\circ = -2.37 \text{ V} \)
- \( \text{Ni}^{2+}(aq) + 2 e^- \rightarrow \text{Ni}(s) \) \( E^\circ = -0.25 \text{ V} \)

a. \( \text{Mg}^{2+}(aq) + \text{Ni}(s) \rightarrow \text{Mg}(s) + \text{Ni}^{2+}(aq) \) \( E_{cell}^\circ = +2.12 \text{ V} \)
b. \( \text{Mg}^{2+}(aq) + \text{Ni}(s) \rightarrow \text{Mg}(s) + \text{Ni}^{2+}(aq) \) \( E_{cell}^\circ = -2.62 \text{ V} \)
c. \( \text{Mg}^{2+}(aq) + \text{Ni}^{2+}(aq) \rightarrow \text{Mg}(s) + \text{Ni}(s) \) \( E_{cell}^\circ = +2.62 \text{ V} \)
d. \( \text{Mg}(s) + \text{Ni}^{2+}(aq) \rightarrow \text{Ni}(s) + \text{Mg}^{2+}(aq) \) \( E_{cell}^\circ = +2.12 \text{ V} \)
e. \( \text{Mg}^{2+}(aq) + \text{Ni}^{2+}(aq) \rightarrow \text{Ni}(s) + \text{Mg}(s) \) \( E_{cell}^\circ = -2.12 \text{ V} \)

19. Calculate \( E_{cell}^\circ \) for the following reaction:

\( \text{F}_2(g) + 2 \text{Cl}^-(aq) \rightarrow 2 \text{F}^-(aq) + \text{Cl}_2(g) \)

given the following standard reduction potentials.

- \( \text{F}_2(g) + 2 e^- \rightarrow 2 \text{F}^-(aq) \) \( E^\circ = +2.87 \text{ V} \)
- \( \text{Cl}_2(g) + 2 e^- \rightarrow 2 \text{Cl}^-(aq) \) \( E^\circ = +1.36 \text{ V} \)

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

\( \text{Ag}^+ | \text{Ag(aq, 0.150 M)} || \text{Sn}^{2+}(aq, 0.500 M), \text{Sn}^{4+}(aq, 0.500 M) | \text{Pt} \)
given the following standard reduction potentials.

\[
\begin{align*}
\text{Ag}^+(aq) + e^- &\rightarrow \text{Ag}(s) \quad E^\circ = +0.80 \text{ V} \\
\text{Sn}^{4+}(aq) + 2 e^- &\rightarrow \text{Sn}^{2+}(aq) \quad E^\circ = +0.14 \text{ V}
\end{align*}
\]

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

21. \(E^\circ_{\text{cell}}\) for the following redox reaction is +0.236 V.

\[2 \text{ Fe}^{3+}(aq) + 2 \text{ I}^- (aq) \rightarrow 2 \text{ Fe}^{2+}(aq) + \text{ I}_2(s)\]

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. \(\text{Al}^{3+}\) is reduced to \(\text{Al}(s)\) 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 \(K_{\text{sp}}\) for PbCl\(_2\)?

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

24. The \(K_{\text{sp}}\) for BaF\(_2\) is \(1.7 \times 10^{-6}\). What is the concentration of Ba\(^{2+}\) in a saturated solution of BaF\(_2\)?

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

25. What is the molar solubility of Ag\(_2\)CrO\(_4\) in 0.20 M K\(_2\)CrO\(_4\)? The value of \(K_{\text{sp}}\) for Ag\(_2\)CrO\(_4\) is \(9.0 \times 10^{-12}\).

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

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

a. \(6.3 \times 10^{-21}\) mol/L
27. The concentration of Ca\(^{2+}\) in a solution is \(4.7 \times 10^{-4}\) M. What concentration of CO\(_3^{2-}\) is required to just begin precipitating CaCO\(_3\)? The \(K_{sp}\) for CaCO\(_3\) is \(3.8 \times 10^{-9}\).

a. \(8.1 \times 10^{-6}\) M
b. \(5.1 \times 10^{-5}\) M
c. \(3.8 \times 10^{-4}\) M
d. \(4.7 \times 10^{-4}\) M
e. \(1.4 \times 10^{-2}\) 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?

<table>
<thead>
<tr>
<th>Compound</th>
<th>(K_{sp})</th>
</tr>
</thead>
<tbody>
<tr>
<td>AgCl</td>
<td>(1.8 \times 10^{10})</td>
</tr>
<tr>
<td>Ag(_2)CrO(_4)</td>
<td>(9.0 \times 10^{12})</td>
</tr>
<tr>
<td>AgBr</td>
<td>(3.3 \times 10^{13})</td>
</tr>
<tr>
<td>AgI</td>
<td>(1.5 \times 10^{16})</td>
</tr>
</tbody>
</table>

a. AgCl→Ag\(_2\)CrO\(_4\)→AgBr→AgI
b. AgI→AgBr→Ag\(_2\)CrO\(_4\)→AgCl
c. Ag\(_2\)CrO\(_4\)→AgCl AgBr→AgI
d. Ag\(_2\)CrO\(_4\)→AgI AgBr→AgCl
e. AgI→AgBr→AgCl→Ag\(_2\)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.
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.