1. Which of the following linear alkane molecules is likely to have the largest standard entropy?

A. CH₄
B. CH₃CH₃
C. CH₃CH₂CH₃
D. CH₃CH₂CH₂CH₃
E. CH₃CH₂CH₂CH₂CH₃

Ans.

2. All of the following statements concerning entropy are true EXCEPT

A. entropy values for substances are greater than or equal to zero.
B. entropy is a thermodynamic state function.
C. entropy is zero for elements under standard conditions.
D. a positive change in entropy denotes a change toward greater disorder.

Ans.

3. Which reaction is likely to have a negative change in entropy?

A. 2 NH₃(g) \rightarrow N₂(g) + 3 H₂(g)
B. CaCO₃(s) \rightarrow CaO(s) + CO₂(g)
C. N₂O₄(g) \rightarrow 2 NO₂(g)
D. 2 CO(g) \rightarrow 2 C(s) + O₂(g)

Ans.

4. The dissolution of ammonium nitrate (NH₄NO₃) occurs spontaneously in water at 25 °C. As it dissolves, the temperature of the water decreases. What are the signs of \(\Delta H\), \(\Delta S\), and \(\Delta G\) for this process?

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. \(\Delta H < 0\), \(\Delta S < 0\), \(\Delta G < 0\)

Ans. The process is endothermic, \(\Delta H > 0\) (T drops as NH₄NO₃ dissolves), and the process is spontaneous (\(\Delta G < 0\)). Consequently, \(\Delta S\) must be > 0 for the process to be spontaneous (since \(\Delta H > 0\)).
5. What is the correct equilibrium constant expression for the balanced reaction shown below?

\[ \text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightleftharpoons 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l}) \]

\[ K = \frac{[\text{CO}_2]^3[\text{H}_2\text{O}]^4}{[\text{C}_3\text{H}_8][\text{O}_2]^5} \]

A. \[ K = \frac{[\text{CO}_2][\text{H}_2\text{O}]}{[\text{C}_3\text{H}_8][\text{O}_2]} \]

B. \[ K = \frac{[\text{CO}_2]}{[\text{C}_3\text{H}_8][\text{O}_2]} \]

C. \[ K = \frac{[\text{CO}_2]^3}{[\text{C}_3\text{H}_8][\text{O}_2]^5} \]

D. \[ K = \frac{[\text{CO}_2]^3}{[\text{C}_3\text{H}_8][\text{O}_2]^5} \]

Ans. (As a pure liquid, H\text{}_2\text{O} is not included in \( K \))

6. At what temperatures will a reaction be spontaneous if \( \Delta H = +158 \text{ kJ} \) and \( \Delta S = +411 \text{ J/K} \)?

A. All temperatures below 384 K

B. All temperatures above 384 K

C. Temperatures between 158 K and 411 K

D. The reaction will be spontaneous at any temperature.

E. The reaction will never be spontaneous.

Ans. An endothermic process at low T is nonspontaneous. As T increases, the entropy loss of the surroundings (manifested as \( \Delta H \)) is outweighed by the entropy increase of the system.

7. Calculate \( \Delta G^\circ \) for the following reaction at 425 °C, at which temperature \( K_{\text{eq}} = 56 \).

\[ \text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2 \text{HI}(\text{g}) \]

\( R = 8.314 \text{ J/K·mol} \)

A. -23.4 kJ

B. -14.2 kJ

C. -10.1 kJ

D. -6.18 kJ

E. +14.2 kJ

Ans.

\[ \Delta G^\circ = -RT\ln K_{\text{eq}} = -(0.008314 \text{ kJ/mol K})(425 + 273)\ln(56) = -23.4 \text{ kJ} \]
8. In the reduction of HClO₄ to Cl⁻ (HClO₄ → Cl⁻), the oxidation state of Cl changes by

A. -7  
B. +7  
C. -8  
D. +8  

Ans. (Oxidation of final state – Oxidation of initial state: -1 – 7 = -8)

9. Write a balanced half-reaction for the reduction of ClO₃⁻(aq) to Cl₂(g) in an acidic solution.

A. 2 ClO₃⁻(aq) + 6 H⁺(aq) + 10 e⁻ → Cl₂(g) + 6 OH⁻(aq)  
B. 2 ClO₃⁻(aq) + 12 H⁺(aq) + 5 e⁻ → Cl₂(g) + 6 H₂O(l)  
C. 2 ClO₃⁻(aq) + 10 e⁻ → Cl₂(g) + 6 H₂O(l) + 3 O₂(g)  
D. 2 ClO₃⁻(aq) + 12 H⁺(aq) + 10 e⁻ → Cl₂(g) + 6 H₂O(l)  
E. none of these are balanced  

Ans. (only one with mass and charge balanced)

10. Use the standard reduction potentials below to determine which element or ion is the best reducing agent.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Potential</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pd²⁺(aq) + 2 e⁻ → Pd(s)</td>
<td>E° = +0.90 V</td>
</tr>
<tr>
<td>Sn²⁺(aq) + 2 e⁻ → Sn(s)</td>
<td>E° = -0.14 V</td>
</tr>
<tr>
<td>Cr²⁺(aq) + 2 e⁻ → Cr(s)</td>
<td>E° = -0.91 V</td>
</tr>
</tbody>
</table>

A. Pd²⁺  
B. Cr(s)  
C. Sn²⁺  
D. Cr²⁺  
E. Pd(s)  

Ans. (this species is most readily oxidized)

11. Which of the following statements about the reaction is correct?
(Cr³⁺ + 3e⁻ → Cr, E° = -0.74 V) (Cd²⁺ + 2e⁻ → Cd, E° = -0.403 V)

3Cd²⁺(aq) + 2Cr(s) → 3Cd(s) + 2Cr³⁺(aq)

A. E°₁ < 0    E°₂ = -0.403 + 0.74 = 0.337 V; therefore E°₁ > 0  
B. ΔG° < 0   ΔG° = -nF E°₁ < 0  
C. 6 moles of electrons are transferred in the reaction  True  
D. A and C  
E. B and C  

Ans.
12. In an electrochemical cell . . .

A. the oxidation process occurs at the anode.
B. the reduction process occurs at the cathode.
C. the salt bridge enables electrical connection through the cell by making a physical connection between the anode and cathode compartments.
D. All of the above.  \textbf{Ans. (follows from electrochemical cell definitions)}
E. Just A and B.

13. Given the following two half-reactions, determine which of the overall reactions is spontaneous and calculate its standard cell potential.

\[
\begin{align*}
\text{Al}^{3+}(aq) + 3 \text{e}^- &\rightarrow \text{Al(s)} \quad E^\circ = -1.66 \text{ V} \\
\text{Cu}^{2+}(aq) + 2\text{e}^- &\rightarrow \text{Cu(s)} \quad E^\circ = +0.34 \text{ V}
\end{align*}
\]

A. \(2 \text{Al}^{3+}(aq) + 3 \text{Cu(s)} \rightarrow 2 \text{Al(s)} + 3 \text{Cu}^{2+}(aq)\) \(E^\circ_{\text{cell}} = -2.00 \text{ V}\)
B. \(2 \text{Al}^{3+}(aq) + 3 \text{Cu(s)} \rightarrow 2 \text{Al(s)} + 3 \text{Cu}^{2+}(aq)\) \(E^\circ_{\text{cell}} = +2.00 \text{ V}\)
C. \(2 \text{Al(s)} + 3 \text{Cu}^{2+}(aq) \rightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Cu(s)}\) \(E^\circ_{\text{cell}} = +2.00 \text{ V}\)  \textbf{Ans.}
D. \(2 \text{Al(s)} + 3 \text{Cu}^{2+}(aq) \rightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Cu(s)}\) \(E^\circ_{\text{cell}} = -1.32 \text{ V}\)

14. Calculate \(E\) for the following electrochemical cell reaction at 298 K

\[
\text{Sn}^{2+}(aq) + 2\text{AgI(s)} \rightarrow \text{Sn}^{4+}(aq) + 2\Gamma(aq) + \text{Ag(s)}
\]

where

\[
\begin{align*}
[	ext{Sn}^{2+}] &= 0.50 \text{ M} \quad [	ext{Sn}^{4+}] = 0.50 \text{ M} \quad [\Gamma] = 0.15 \text{ M} \\
\text{AgI(s)} + \text{e}^- &\rightarrow \text{Ag(s)} + \Gamma(aq) \quad E^\circ = -0.15 \text{ V} \\
\text{Sn}^{4+}(aq) + 2 \text{e}^- &\rightarrow \text{Sn}^{2+}(aq) \quad E^\circ = +0.15 \text{ V}
\end{align*}
\]

A. +0.05 V  B. -0.25 V  C. -0.30 V  D. -0.32 V  E. -0.35 V  \textbf{Ans.}

\[
E = E^\circ - \frac{RT}{nF} \ln Q = -0.30 - (0.0257/2) \ln ([\Gamma]^2)[\text{Sn}^{4+}]/[\text{Sn}^{2+}]) = -0.30 - 0.01286 \ln [0.15]^2 \\
E = -0.30 + 0.01286 \ln [0.15]^2 = -0.30 + 0.0487 = -0.25 \text{ V}
\]
EXTRA CREDIT (5 points each)

15. Citric Acid (Cit) binds reversibly to Citric Acid binding protein (CBP) with an equilibrium constant of $2.96 \times 10^6$. The enthalpy of binding is $-102.4 \text{ kJ mol}^{-1}$.

\[
\text{CBP} + \text{Cit} \rightleftharpoons \text{CBP:Cit}
\]

What is the entropy of binding at 25 °C?

A. -36.9 J/K  
B. -220 J/K  
C. -346 J/K  
D. -3970 J/K

\[
\Delta G^\circ = -(0.008314 \times 298) \ln(2.96 \times 10^6) = -36.9 \text{ kJ}
\]

\[
\Delta S^\circ = \frac{\Delta G^\circ - \Delta H^\circ}{T} = \frac{(-36.9 \text{ kJ} + 102.4 \text{ kJ})}{298} = -0.220 \text{ J/K}
\]

16. To generate 1 kilogram (1000 grams) of Al from Al$_2$O$_3$ in an electrolytic cell in 1 hour, what is the electrical current (in amperes) that must be supplied to the cell?

A. $1.07 \times 10^7$  
B. $3.58 \times 10^6$  
C. 2980  
D. 990

\[
\frac{(1000 \text{ grams Al})}{(26.98 \text{ grams/mole Al})} = 37.06 \text{ moles Al}
\]

\[
(37.06 \text{ moles Al}) \times (3 \text{ moles e}^- \text{ per mol Al}) = 111.2 \text{ moles e}^- \text{ transferred per kg Al}
\]

\[
(111.2 \text{ moles e}^-) \times (96485 \text{ Coulombs per mole e}^-) = 1.073 \times 10^7 \text{ Coulombs}
\]

\[
\text{Amps} = \frac{(1.073 \times 10^7 \text{ Coulombs})}{(3600 \text{ seconds})} = 2980 \text{ Amps}
\]