Use the following information to answer questions 9 and 10.

The photos of a reaction along with the molecular-level representations are shown below.

9. What is the correct balanced net ionic equation for the reaction illustrated?
(A) KCl + AgNO₃ → KNO₃ + AgCl
(B) K⁺ + NO₃⁻ → KNO₃
(C) Ag⁺ + NO₃⁻ → AgNO₃
(D) Ag⁺ + Cl⁻ → AgCl

10. Which species are the spectator ions in this reaction?
(A) K⁺ and NO₃⁻
(B) K⁺ and Cl⁻
(C) Ag⁺ and NO₃⁻
(D) Ag⁺ and Cl⁻

4. Consider the reaction shown below:

Given the following information, calculate Δ_H°_rxn for the reaction represented above, where each molecule represents 1 mole of that substance. Assume that all states are those that are listed below.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Δ_H° (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>CO₂(g)</td>
<td>-393.5</td>
</tr>
<tr>
<td>H₂O(g)</td>
<td>-285.8</td>
</tr>
<tr>
<td>CH₃OH(g)</td>
<td>-201.0</td>
</tr>
</tbody>
</table>

(A) -2332.2 kJ
(B) -1528.2 kJ
(C) -478.3 kJ
(D) 45.8 kJ

5. Which of the following graphs describes a pathway of reaction that is exothermic with a high activation energy?
(A) 
(B) 
(C) 
(D)
11. A student performs a neutralization reaction involving an acid and a base in an open polystyrene coffee-cup calorimeter. How would the calculated value of $\Delta H$ differ from the actual value if there was significant heat loss to the surroundings?

(A) $\Delta H_{\text{calc}}$ would be negative, but more negative than the actual value.
(B) $\Delta H_{\text{calc}}$ would be negative, but less negative than the actual value.
(C) $\Delta H_{\text{calc}}$ would be positive, but more positive than the actual value.
(D) $\Delta H_{\text{calc}}$ would be positive, but less positive than the actual value.

12. Which of the following is endothermic?

(A) Water freezes to form ice.
(B) Steam condenses on a bathroom mirror.
(C) Ice cream melts.
(D) Coffee cools as it sits.

15. A popular chemistry demonstration is to drop a piece of sodium metal into water. The products are sodium hydroxide and hydrogen gas. Determine $\Delta H_{\text{rxn}}$ for this reaction for 1.00 mol of hydrogen gas being produced, given

$\Delta H_{\text{f[H}_2\text{O(1)]}} = -286 \text{ kJ/mol}$

$\Delta H_{\text{f[NaOH(aq)]}} = -470 \text{ kJ/mol}$

(A) $-368 \text{ kJ}$
(B) $-184 \text{ kJ}$
(C) $184 \text{ kJ}$
(D) $368 \text{ kJ}$

1. Magnesium reacts with hydrochloric acid to produce hydrogen gas. An experiment was set up to determine the rate of production of the hydrogen gas by measuring its change in volume over the first 60 seconds of the reaction. Magnesium ribbon was cut into squares of 0.5 cm$^2$ and 1.0 cm$^2$. The volume of hydrochloric acid solution and the total mass of magnesium were held constant in each trial. Which set of conditions would produce the highest rate of production of hydrogen gas?

(A) 1 M HCl, 0.5-cm$^2$ pieces of magnesium
(B) 1 M HCl, 1.0-cm$^2$ pieces of magnesium
(C) 3 M HCl, 1.0-cm$^2$ pieces of magnesium
(D) 3 M HCl, 0.5-cm$^2$ pieces of magnesium

Questions 2–4 refer to the following:

Given the reaction: $A + 2B \rightarrow C$

<table>
<thead>
<tr>
<th>Trial</th>
<th>$[A]_0$ (mol L$^{-1}$)</th>
<th>$[B]_0$ (mol L$^{-1}$)</th>
<th>Rate of formation of C (mol L$^{-1}$ s$^{-1}$)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>$1.2 \times 10^{-2}$</td>
<td>$1.2 \times 10^{-2}$</td>
<td>$1.8 \times 10^{-4}$</td>
</tr>
<tr>
<td>2</td>
<td>$1.2 \times 10^{-2}$</td>
<td>$2.4 \times 10^{-2}$</td>
<td>$7.2 \times 10^{-4}$</td>
</tr>
<tr>
<td>3</td>
<td>$2.4 \times 10^{-2}$</td>
<td>$1.2 \times 10^{-2}$</td>
<td>$3.6 \times 10^{-4}$</td>
</tr>
</tbody>
</table>

2. Based on the experimental data given in the table, what is the order of the reaction with respect to A?

(A) zero
(B) first
(C) second
(D) third

3. Based on the experimental data given in the table, what is the order of the reaction with respect to B?

(A) zero
(B) first
(C) second
(D) third

4. What is the rate of disappearance of B in trial 2?

(A) $2.4 \times 10^{-2}$ mol L$^{-1}$ s$^{-1}$
(B) $1.4 \times 10^{-3}$ mol L$^{-1}$ s$^{-1}$
(C) $3.6 \times 10^{-4}$ mol L$^{-1}$ s$^{-1}$
(D) $7.2 \times 10^{-4}$ mol L$^{-1}$ s$^{-1}$
6. In a chemical reaction that has first-order kinetics, which is true at constant temperature?
   (A) Half-life and \( k \) are both constant.
   (B) Neither half-life nor \( k \) is constant.
   (C) Half-life is constant, but \( k \) changes.
   (D) Half-life changes, but \( k \) is constant.

7. The reaction \( Q + R_2 \rightarrow R_2Q \) is found to be first order in \( R_2 \) and zero order in \( Q \). Which of these mechanisms is most likely for this reaction?
   (A) \( Q + R_2 \rightarrow R_2Q \) (slow)
   (B) \( 2Q \rightarrow Q_2 \) (slow)
   \( Q_2 + R_2 \rightarrow R_2Q + Q \) (fast)
   (C) \( Q + Z \rightarrow QZ \) (slow)
   \( QZ + R \rightarrow QR + Z \) (fast)
   \( QR + R \rightarrow R_2Q \) (fast)
   (D) \( R_2 \rightarrow 2R \) (slow)
   \( R + Q \rightarrow RQ \) (fast)
   \( RQ + R \rightarrow R_2Q \) (fast)

8. Which is true of the reaction diagram for an elementary reaction, such as the one shown?

(A) It consists of zero transition states and zero intermediates.
(B) It has one transition state and zero intermediates.
(C) It has one transition state and one intermediate.
(D) It has two transition states and two intermediates.

9. Which is a correct explanation for why the reaction rate increases with increasing temperature?
   (A) The reaction becomes more exothermic.
   (B) The enthalpy change of the reaction increases.
   (C) The activation energy of the reaction decreases.
   (D) More of the colliding particles have the activation energy.

10. How would the net-ionic equation written for a catalyzed reaction differ from the net-ionic equation written for the same reaction without a catalyst?
   (A) The net-ionic equation would be different because an additional reactant is needed.
   (B) The net-ionic equation would be different because the catalyst should be shown on both the reactant and product side.
   (C) The net-ionic equation would be the same because the catalyst is a spectator ion and would be left out of a net ionic equation.
   (D) The net-ionic equation would be the same because the catalyst is neither a reactant nor a product.
15. Which energy profile represents the reaction that would occur at the lowest rate at a given temperature?

(A)  

(B)  

(C)  

(D)  

2. For the reaction \( A \rightleftharpoons B \), for which value of \( K \) is the concentration of the reactants greater than the concentration of the products?
(A) \( K \) equals 0
(B) \( K \) equals 1
(C) \( K \) is less than 1
(D) \( K \) is greater than 1

4. Hydrogen gas at a pressure of 9.0 atm is added to a container containing 5.0 atm of iodine vapor. The resulting reaction is allowed to come to equilibrium, and it is found that the final pressure of iodine vapor is 2.0 atm. What is the \( K_p \) value for the following reaction?

\[ H_2(g) + I_2(g) \rightleftharpoons 2HI(g) \]

(A) 0.25
(B) 0.33
(C) 3.0
(D) 4.0

Questions 8–12 refer to the following:

Hydrogen gas, iodine vapor, and hydrogen iodide gas are added to an evacuated flask until the concentrations of \( H_2 \) and \( I_2 \) are both 1.0 \( M \) and that of \( HI \) equals 8.0 \( M \).

\[ H_2(g) + I_2(g) \rightleftharpoons 2 HI(g) \]

(At 425°C, \( K = 9.00 \))

8. Determine the value of \( Q \).
(A) 2.0
(B) 4.0
(C) 8.0
(D) 16

9. The reaction vessel is heated to 675°C, where the equilibrium concentration of the HI is determined to be 6.0 M. What is the \( K \) value at 675°C?
(A) 6.0
(B) 9.0
(C) 24
(D) 36

10. Once the system reaches equilibrium at 675°C, additional \( H_2 \) is added and the temperature is held constant. Which statement best describes how this affects the value of \( K \)?
(A) The value of \( K \) increases because the formation of products is favored.
(B) The value of \( K \) decreases because the formation of reactants is favored.
(C) The value of \( K \) remains unchanged because the temperature remained constant.
(D) The value of \( K \) cannot be determined until the equilibrium concentrations are determined.
Questions 14 and 15 refer to the following reaction at equilibrium:

\[ 2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g) \quad \Delta H = -198 \text{ kJ} \]

14. Which of the following occurs after the addition of \( \text{O}_2 \)?
(A) The rate of the forward reaction increases, and the value of \( K \) increases.
(B) The rate of the forward reaction decreases, and the value of \( K \) increases.
(C) The rate of the forward reaction decreases, and the value of \( K \) remains the same.
(D) The rate of the forward reaction increases, and the value of \( K \) remains the same.

15. Which of the following changes to the system at equilibrium would result in an increase in the concentration of \( \text{SO}_3 \)?
(A) an increase in the concentration of \( \text{SO}_2 \) and an increase in the temperature
(B) an increase in the concentration of \( \text{SO}_2 \) and a decrease in the temperature
(C) an increase in the concentration of \( \text{SO}_3 \) and a decrease in the temperature
(D) an increase in the concentration of \( \text{SO}_3 \) and an increase in the temperature

16. It is found that the value of the reaction quotient is 12 for a given reaction, whereas the equilibrium constant is 0.8. Which statement best describes how the system will respond?
(A) Since \( Q > K \), the forward reaction is favored and the equilibrium shifts to the right.
(B) Since \( Q < K \), the forward reaction is favored and the equilibrium shifts to the left.
(C) Since \( Q < K \), the reverse reaction is favored and the equilibrium shifts to the right.
(D) Since \( Q > K \), the reverse reaction is favored and the equilibrium shifts to the left.

2. If the concentration of the HF in the reaction above is 0.1 \( M \), what is an approximate pH of the solution?
(A) 1
(B) 2
(C) 7
(D) 9

4. Which of the following correctly ranks the species from weakest to strongest base?
- \( \text{HF} \) \( (K_a = 7.1 \times 10^{-4}) \)
- \( \text{HCN} \) \( (K_a = 6.2 \times 10^{-10}) \)
- \( \text{HClO}_2 \) \( (K_a = 1.2 \times 10^{-7}) \)
- \( \text{NH}_4^+ \) \( (K_a = 5.6 \times 10^{-10}) \)
(A) \( \text{CN}^- < \text{NH}_3 < \text{F}^- < \text{ClO}_2^- < \text{H}_2\text{O} < \text{NO}_2^- \)
(B) \( \text{CN}^- < \text{NH}_3 < \text{F}^- < \text{ClO}_2^- < \text{NO}_2^- < \text{H}_2\text{O} \)
(C) \( \text{H}_2\text{O} < \text{NO}_2^- < \text{ClO}_2^- < \text{F}^- < \text{NH}_3 < \text{CN}^- \)
(D) \( \text{NO}_2^- < \text{H}_2\text{O} < \text{ClO}_2^- < \text{F}^- < \text{NH}_3 < \text{CN}^- \)

5. At the normal body temperature of a human, 37°C, the equilibrium constant for the dissociation of water is higher than it is at 25°C, as shown in the equilibrium expression.

\[ K_w = [\text{H}^+]\cdot[\text{OH}^-] = 2.42 \times 10^{-14} \]

What is true of pure water at 37°C?
(A) Water is neutral because the hydrogen and hydroxide ion concentrations are equal.
(B) Water is basic because the hydroxide ion concentration is greater than \( 1 \times 10^{-7} \) \( M \).
(C) Water is acidic because the hydrogen ion concentration is greater than \( 1 \times 10^{-7} \) \( M \).
(D) Water is neutral because the hydrogen and hydroxide ion concentrations are each \( 1 \times 10^{-7} \) \( M \).

Questions 7–9 refer to the following solutions:

![Beaker #1: HCl(aq)
Beaker #2: HNO_2(aq)
Beaker #3: HI(aq)](image)

Each beaker contains 100 mL of the indicated acidic solution, all with a pH of 3.00.

\[ K_a \] for \( \text{HNO}_2 = 4.0 \times 10^{-4} \]

7. Which beaker would have the highest hydrogen ion concentration?
(A) Beaker 1 has the highest hydrogen ion concentration.
(B) Beaker 2 has the highest hydrogen ion concentration.
(C) Beaker 1 and 3 are equally high.
(D) All have the same hydrogen ion concentration.

8. Which solution has the highest percent ionization?
(A) Beaker 1 has the highest percent ionization.
(B) Beaker 2 has the highest percent ionization.
(C) Beakers 1 and 3 are equally high.
(D) All have the same percent ionization.

9. If 50 mL of 0.1 \( M \) \( \text{NaOH} \) was added to each beaker, which resulting solution would have the lowest pH?
(A) Beaker 1 would have the lowest pH.
(B) Beaker 2 would have the lowest pH.
(C) Beakers 1 and 3 are equally low.
(D) All would have the same pH.
10. A solution is prepared by mixing 500 mL of 1 M HCl with 500 mL of 0.01 M HNO₃. What is the pH of the resulting solution?
   (A) 0.0
   (B) between 0.0 and 1.0
   (C) 1.0
   (D) greater than 1.0

11. In a research project, a scientist adds 0.1 mole of HCN, 0.1 mole of H₂O⁺, and 0.1 mole of CN⁻ to water to make a total volume of 1 L. Will this reaction proceed to a greater extent in the forward direction or in the reverse direction?
   \[ \text{HCN} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CN}^- \quad K = 6.2 \times 10^{-10} \]
   (A) Forward; acids always dissociate in water.
   (B) Forward; the \( K \) value is less than 1.
   (C) Reverse; the \( K \) value is less than 1.
   (D) Reverse; water cannot be a reactant.

12. What is the conjugate acid of HPO₄²⁻?
   (A) H₂O⁺
   (B) PO₄³⁻
   (C) H₃PO₄
   (D) H₂PO₄⁻

49. Calculate the pH of a solution that is 0.40 M H₂NNH₃⁺ and 0.80 M H₂NNH₂NO₃. In order for this buffer to have pH = \( pK_a \), would you add HCl or NaOH? What quantity (moles) of which reagent would you add to 1.0 L of the original buffer so that the resulting solution has pH = \( pK_a \)?

51. Which of the following mixtures would result in buffered solutions when 1.0 L of each of the two solutions are mixed?
   a. 0.1 M KOH and 0.1 M CH₃NH₂Cl
   b. 0.1 M KOH and 0.2 M CH₃NH₂
   c. 0.2 M KOH and 0.1 M CH₃NH₂Cl
   d. 0.1 M KOH and 0.2 M CH₃NH₂Cl

115. The titration of Na₂CO₃ with HCl has the following qualitative profile:

   Identify the major species in solution at points A–F.
   Calculate the pH at the halfway points to equivalence, B and D. (Hint: Refer to Exercise 113.)

Questions 4 and 5 refer to the following reaction:

\[ \text{CH}_3\text{NH}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{CH}_3\text{NH}_3^+(aq) + \text{OH}^-(aq) \quad K_b = 4.00 \times 10^{-4} \]

4. Determine the [OH⁻] for 0.25 M methylamine, CH₃NH₂.
   (A) 2.0 \times 10^{-9} M
   (B) 6.0 \times 10^{-9} M
   (C) 1.0 \times 10^{-4} M
   (D) 1.0 \times 10^{-2} M

5. Determine the pH of the methylamine solution.
   (A) 4.00
   (B) 6.00
   (C) 10.00
   (D) 12.00

6. Solution A consists of 1.0 M HC₃H₅O₂ (\( K_a = 1.8 \times 10^{-5} \)), and Solution B consists of 1.0 M HC₃H₅O₂ and 1.0 M NaC₃H₅O₄. Which of the following statements is true?
   (A) HC₃H₅O₂ undergoes a greater percent dissociation in Solution A than Solution B, and the pH of Solution A is lower than that of Solution B.
   (B) HC₃H₅O₂ undergoes a greater percent dissociation in Solution A than Solution B, and the pH of Solution A is higher than that of Solution B.
   (C) HC₃H₅O₂ undergoes a lower percent dissociation in Solution A than Solution B, and the pH of Solution A is lower than that of Solution B.
   (D) HC₃H₅O₂ undergoes a lower percent dissociation in Solution A than Solution B, and the pH of Solution A is higher than that of Solution B.

7. Consider 100.0 mL of 0.100 M aqueous solutions of each of the following acids: HCN (\( K_a = 6.2 \times 10^{-10} \)), HF (\( K_a = 7.2 \times 10^{-4} \)), HCl, HC₃H₅O₂ (\( K_a = 1.8 \times 10^{-5} \)). Each acid is titrated with 0.100 M NaOH(aq). Once the equivalence point is reached for each solution, rank the pH values from highest to lowest.
   (A) HCl, HCN, HF, HC₃H₅O₂
   (B) HCN, HC₃H₅O₂, HF, HCl
   (C) HCl, HF, HC₃H₅O₂, HCN
   (D) HC₃H₅O₂, HCN, HCl, HF
15. A 0.10 M HCl solution may be titrated with two different bases: 0.10 M NaOH or 0.10 M Ca(OH)$_2$. Which statement best describes the volumes of the bases necessary to reach the equivalence point?
   (A) Half as much NaOH is needed.
   (B) Equal volumes of NaOH and Ca(OH)$_2$ are needed.
   (C) Twice as much Ca(OH)$_2$ is needed.
   (D) Twice as much NaOH is needed.

4. Three ionic salts, AX, BX$_2$, and CX$_3$, each have a solubility of $2 \times 10^{-3}$ M. Which of the following correctly ranks the $K_{sp}$ of the three salts from least to greatest?
   (A) CX$_3$ < BX$_2$ < AX
   (B) AX < BX$_2$ < CX$_3$
   (C) BX$_2$ < CX$_3$ < AX
   (D) AX < CX$_3$ < BX$_2$

6. Barium iodate has a $K_{sp}$ of $4.0 \times 10^{-9}$. What is the iodate ion concentration in a saturated solution of barium iodate?
   (A) $1.0 \times 10^{-3}$ M
   (B) $2.0 \times 10^{-3}$ M
   (C) $6.3 \times 10^{-5}$ M
   (D) $2.0 \times 10^{-9}$ M

7. Which of these salts would have increased solubility in a 0.10 M solution of HNO$_3$, compared with their solubilities in pure water at the same temperature?
   (A) AgBr
   (B) HBrCl$_3$
   (C) Mg(OH)$_2$
   (D) PbI$_2$

9. Ten grams of Ca(OH)$_2(s)$ is added, with stirring, to 100 mL of water, producing a saturated solution with solid Ca(OH)$_2$ on the bottom of the beaker. Which is true?
   (A) The addition of 30 drops of 3 M HCl will cause the $K_{sp}$ of the reaction to decrease.
   (B) The addition of 0.1 mole of solid Ca(NO$_3$)$_2$ will cause some additional Ca(OH)$_2$ to form.
   (C) The addition of 0.1 mole of solid NaOH will cause some additional Ca(OH)$_2$ to dissolve.
   (D) The addition of 0.1 mole of solid Ca(OH)$_2$ will cause the $K_{sp}$ of the reaction to increase.

10. Solid sodium chloride was added to pure water slowly, with stirring, at 25°C until no more sodium chloride would dissolve. Solid remained on the bottom of the container, even after the solution was stirred repeatedly over 8 hours. After the solid was allowed to settle, some of the clear, colorless solution was carefully poured off into a flask. This flask was placed in a refrigerator at a temperature of 4°C for 2 hours. When the flask was removed from the refrigerator, a small amount of white solid was found to have appeared in the flask. What is the best explanation for this?
   (A) Solids settle out of solutions over time.
   (B) NaCl has an exothermic heat of solution.
   (C) The $K_{sp}$ of NaCl decreases with decreasing temperature.
   (D) The $K_{sp}$ of NaCl increases with decreasing temperature.

11. Which of the following statements is true concerning the relative solubility of silver chloride? All samples are at the same temperature.
   (A) AgCl is less soluble in 1.0 M AgNO$_3$ than it is in pure water.
   (B) AgCl is less soluble in 1.0 M HNO$_3$ than it is in pure water.
   (C) AgCl is more soluble in 1.0 M AgNO$_3$ than it is in pure water.
   (D) AgCl is more soluble in 1.0 M HNO$_3$ than it is in pure water.
1. Consider the freezing of liquid water at -10°C and 1 atm. For this process, what are the signs for ΔH, ΔS, and ΔG?

<table>
<thead>
<tr>
<th></th>
<th>ΔH</th>
<th>ΔS</th>
<th>ΔG</th>
</tr>
</thead>
<tbody>
<tr>
<td>(A)</td>
<td>+</td>
<td>-</td>
<td>0</td>
</tr>
<tr>
<td>(B)</td>
<td>-</td>
<td>+</td>
<td>0</td>
</tr>
<tr>
<td>(C)</td>
<td>+</td>
<td>+</td>
<td>-</td>
</tr>
<tr>
<td>(D)</td>
<td>-</td>
<td>-</td>
<td>-</td>
</tr>
</tbody>
</table>

2. A 100-mL sample of water is placed in a coffee-cup calorimeter. Solid NaCl is then dissolved in the water. The temperature of the water decreases from 20.5°C to 19.7°C and is then allowed to return to room temperature (20.5°C). Determine the signs for ΔH and ΔS for the process of dissolving NaCl and ΔG for the entire process at constant temperature.

<table>
<thead>
<tr>
<th></th>
<th>ΔH</th>
<th>ΔS</th>
<th>ΔG</th>
</tr>
</thead>
<tbody>
<tr>
<td>(A)</td>
<td>+</td>
<td>-</td>
<td>0</td>
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<tr>
<td>(B)</td>
<td>-</td>
<td>+</td>
<td>0</td>
</tr>
<tr>
<td>(C)</td>
<td>+</td>
<td>+</td>
<td>-</td>
</tr>
<tr>
<td>(D)</td>
<td>-</td>
<td>-</td>
<td>-</td>
</tr>
</tbody>
</table>

3. Which reaction would have the most positive ΔS°?
(A) CO(g) + 2H₂(g) → CH₃OH(l)
(B) 2CH₂OH(g) + 3O₂(g) → 2CO₂(g) + 4H₂O(g)
(C) HCl(g) + NH₃(g) → NH₄Cl(s)
(D) Ba(OH)₂ · 8H₂O(s) → Ba(NO₃)₂(s) + 2NH₃(g) + 10H₂O(l)

4. Solutions A and B are both clear and colorless. When Solution A is mixed with Solution B, the temperature of the mixture increases and a yellow precipitate is observed. What can be concluded from these observations?
(A) The reaction is thermodynamically favored (spontaneous) at all temperatures.
(B) The reaction is thermodynamically favored (spontaneous) only at high temperatures.
(C) The reaction is thermodynamically favored (spontaneous) only at low temperatures.
(D) The reaction is not thermodynamically favored (spontaneous) at any temperature.

Questions 5–8 refer to the following information:

The removal of copper from metallic ores has been a challenge since ancient times. Strong heating of copper(II) carbonate results in the formation of copper(II) oxide.

Although the decomposition of copper(II) oxide into copper metal and oxygen gas is not thermodynamically favorable under standard conditions, it may be made favorable by adding carbon to the mixture.

The resulting reaction is: 2CuO(s) + C(s) → 2Cu(s) + CO₂(g).

CuO(s) → Cu(s) + ½O₂(g) \[ ΔG° = 128 \text{ kJ/mol} \]

C(s) + O₂(g) → CO₂(g) \[ ΔG° = -394 \text{ kJ/mol} \]

5. Calculate ΔG° for the overall reaction.
(A) -138 kJ/mol
(B) -266 kJ/mol
(C) 522 kJ/mol
(D) 650 kJ/mol

6. Predict and justify the sign of ΔS° for the reaction CuO(s) → Cu(s) + ½O₂(g).
(A) ΔS° is positive because ΔG° is positive.
(B) ΔS° is negative because ΔG° is positive.
(C) ΔS° is positive because the products have more moles of gas than the reactant.
(D) ΔS° is negative because the products have more moles of gas than the reactant.

7. Why is the overall process thermodynamically favorable?
(A) Carbon serves as a catalyst to lower the activation energy of the reaction.
(B) The process involves coupling a very thermodynamically favorable reaction with one that is not thermodynamically favored.
(C) When two reactions are coupled, free energy is always released.
(D) The addition of carbon speeds up the reaction because powdered carbon has a large surface area.

8. Which statement is correct about the value of K for the reaction 2CuO(s) + C(s) → 2Cu(s) + CO₂(g) at standard conditions?
(A) K will be less than 1.
(B) K will be 0.
(C) K will be between 0 and 1.
(D) K will be greater than 1.