1. Given the graphed data at right, if you are asked to calculate the concentration change for HI(g) during the time interval from 30 s to 75 s, you would be calculating the

A. rate law.  
B. rate constant.  
C. instantaneous rate.  
D. average rate.

2. For the reaction A + B → 2 C, use the table below to determine \( \frac{\Delta [B]}{\Delta t} \).

<table>
<thead>
<tr>
<th>Time (s)</th>
<th>[A] (M)</th>
<th>[B] (M)</th>
<th>[C] (M)</th>
</tr>
</thead>
<tbody>
<tr>
<td>10.0</td>
<td>0.910</td>
<td>0.500</td>
<td>0.250</td>
</tr>
<tr>
<td>20.0</td>
<td>0.730</td>
<td>0.320</td>
<td>0.610</td>
</tr>
</tbody>
</table>

A. \( 1.80 \times 10^{-2} \) M/s  
B. \( 9.00 \times 10^{-3} \) M/s  
C. \( -1.80 \times 10^{-3} \) M/s  
D. \( -3.60 \times 10^{-2} \) M/s

3. According to the plot, which of these reactions demonstrates zero-order kinetics?

A. Reaction 1  
B. Reaction 2  
C. Reaction 3  
D. Reaction 4
4. For the reaction $3 \text{AB} \rightarrow 2 \text{C}$, use the table below to determine the **order** of reaction:

<table>
<thead>
<tr>
<th>[AB] (M)</th>
<th>Initial Rate (M/s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.046</td>
<td>0.0042</td>
</tr>
<tr>
<td>0.023</td>
<td>0.0011</td>
</tr>
</tbody>
</table>

A. 0  
B. 1  
C. 2  
D. 3

5. For the reaction

$$\text{H}_2\text{SeO}_3(aq) + 6 \text{I}^- (aq) + 4 \text{H}^+(aq) \rightarrow \text{Se(s)} + 2 \text{I}_3^- (aq) + 3 \text{H}_2\text{O(l)}$$

rate = $k[\text{H}_2\text{SeO}_3]^a[\text{H}^+]^b[\text{I}^-]^c$

it is experimentally determined that the reaction rate doubles when $[\text{H}_2\text{SeO}_3]$ is doubled, the rate quadruples when $[\text{H}^+]$ is doubled, and the rate increases by a factor of eight when $[\text{I}^-]$ is doubled. What is the **overall order** of the reaction?

A. third order  
B. fourth order  
C. sixth order  
D. eighth order

6. The decomposition of $\text{N}_2\text{O}_5(g)$ is a first-order reaction with $k = 7.0 \times 10^{-2} \text{ s}^{-1}$. If the initial concentration of $\text{N}_2\text{O}_5(g)$ is 1.5 M, what is the concentration of $\text{N}_2\text{O}_5$ after 4.2 minutes?

A. 7.8 M  
B. $4.2 \times 10^{-4}$ M  
C. $3.3 \times 10^{-8}$ M  
D. $2.1 \times 10^{-9}$ M
7. The reaction of butadiene, C_4H_6, to form C_8H_{12} is second order in butadiene.

\[2 \text{C}_4\text{H}_6(\text{g}) \rightarrow \text{C}_8\text{H}_{12}(\text{g})\]

\[k = 6.14 \times 10^{-2} \text{ M}^{-1} \text{s}^{-1}\]

**How long** has the reaction run if the initial butadiene concentration was 0.278 M and the current butadiene concentration is 0.074M?

A. 18 minutes  
B. 11 minutes  
C. 2.7 minutes  
D. 1.3 minutes

8. The decomposition of A is a first-order reaction with a half-life of 738 seconds. The initial concentration of A is 0.503 M. Determine the rate constant, \( k \).

\[\text{A} \rightarrow \text{products}\]

A. \(1.04 \times 10^3 \text{ s}^{-1}\)  
B. \(5.36 \times 10^{-2} \text{ s}^{-1}\)  
C. \(1.94 \times 10^{-3} \text{ s}^{-1}\)  
D. \(9.39 \times 10^{-4} \text{ s}^{-1}\)

9. What effect does doubling the Kelvin temperature of a reaction mixture have on the rate constant, \( k \)?

A. The value of the rate constant doubles.  
B. The value of the rate constant increases, but not necessarily doubles.  
C. The value of the rate constant decreases by half.  
D. The value of the rate constant decreases, but not necessarily by half.
10. According to the collision model for a chemical reaction, the frequency factor $A$ in the Arrhenius equation depends on the

A. orientation factor and collision frequency.  
B. temperature and concentration. 
C. activation energy and temperature. 
D. activation energy and orientation factor.

11. The rate of the overall reaction

$$2 \text{H}_2\text{O}_2(aq) \rightarrow 2 \text{H}_2\text{O}(l) + 2 \text{O}_2(g)$$

can be increased by adding bromine. According to these elementary steps, what are the functions of $\text{Br}_2$, $\text{Br}^-$, and $\text{H}^+$ in the overall reaction?

$$\text{Br}_2(aq) + \text{H}_2\text{O}_2(aq) \rightarrow 2 \text{Br}^-(aq) + 2 \text{H}^+(aq) + \text{O}_2(g)$$
$$2 \text{Br}^-(aq) + 2 \text{H}^+(aq) + \text{H}_2\text{O}_2(aq) \rightarrow 2 \text{H}_2\text{O}(l) + \text{O}_2(g) + \text{Br}_2(aq)$$

A. $\text{Br}_2$, $\text{Br}^-$, and $\text{H}^+$ are all catalysts. 
B. $\text{Br}^-$ and $\text{H}^+$ are catalysts; $\text{Br}_2$ is a reaction intermediate. 
C. $\text{Br}_2$ is a catalyst; $\text{Br}^-$ and $\text{H}^+$ are reaction intermediates. 
D. $\text{Br}_2$, $\text{Br}^-$, and $\text{H}^+$ are all reaction intermediates.

12. For the reaction

$$\text{H}_2(g) + 2 \text{IBr}(g) \rightarrow 2 \text{HBr}(g) + \text{I}_2(g),$$

the following mechanism has been proposed.

Step 1: $\text{H}_2(g) + \text{IBr}(g) \rightarrow \text{HBr}(g) + \text{HI}(g)$
Step 2: $\text{HI}(g) + \text{IBr}(g) \rightarrow \text{HBr}(g) + \text{I}_2(g)$

If the rate law is $\text{rate} = k[\text{H}_2][\text{IBr}]$, which is the rate-determining step in the proposed mechanism?

A. Step 1 is the rate-determining reaction. 
B. Step 1 and Step 2 both determine the rate. 
C. Step 2 is the rate-determining step. 
D. The rate-determining step cannot be identified from the information given.
13. Catalysts
   A. increase the rate of reaction by lowering the activation energy, $E_a$.
   B. increase the rate of the reaction by increasing the reaction temperature, $T$.
   C. increase the rate of reaction by adding energy to the reaction.
   D. must be in the same phase as the reactants and products.

14. At dynamic equilibrium,
   A. the concentrations of the reactants and products are equal.
   B. the concentrations of the reactants and products remain constant.
   C. the rates of the forward and reverse reactions are equal.
   D. both B and C are correct.

15. Initially, 0.800 M of $\text{SO}_2$ and 0.200 M of $\text{O}_2$ are present in a reaction vessel. Given the balanced chemical equation and $K_c$, what is true about the equilibrium concentrations of reactants and products?

   \[
   2 \text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2 \text{SO}_3(g) \quad K_c = 1.98 \text{ at } 1303 \text{ K}
   \]

   A. The equilibrium strongly favors neither reactants nor products, so appreciable amounts of $\text{SO}_3$, $\text{SO}_2$ and $\text{O}_2$ will all be present at equilibrium.
   B. The equilibrium lies far to the right and the concentration of $\text{SO}_3$ will be significantly higher than the concentrations of $\text{SO}_2$ and $\text{O}_2$.
   C. The equilibrium lies far to the right and the concentration of $\text{SO}_3$ will be significantly lower than the concentrations of $\text{SO}_2$ and $\text{O}_2$.
   D. The equilibrium lies far to the left and the concentration of $\text{SO}_3$ will be significantly lower than the concentrations of $\text{SO}_2$ and $\text{O}_2$.

16. Given the reactions

   \[
   \begin{align*}
   2 \text{NO}(g) & \rightleftharpoons \text{N}_2(g) + \text{O}_2(g) \quad K_{c1} = 4.4 \times 10^{18} \\
   2 \text{NO}(g) + \text{O}_2(g) & \rightleftharpoons 2 \text{NO}_2(g) \quad K_{c2} = 3.0 \times 10^6 
   \end{align*}
   \]

   calculate $K_{c3}$ for the reaction

   \[
   \text{N}_2(g) + 2 \text{O}_2(g) \rightleftharpoons 2 \text{NO}_2(g) \quad K_{c3} = ?
   \]

   A. $6.8 \times 10^{-13}$
   B. $2.2 \times 10^{-6}$
   C. $3.1 \times 10^{11}$
   D. $8.4 \times 10^{24}$
17. Which of the following is the equilibrium constant expression for this reaction?

\[ \text{FeO}(s) + \text{H}_2(g) \rightleftharpoons \text{Fe}(s) + \text{H}_2\text{O}(g) \]

A. \( K_c = \frac{[\text{H}_2\text{O}]^2}{[\text{FeO}_2][\text{H}_2]} \)  
B. \( K_c = \frac{[\text{Fe}][\text{H}_2\text{O}]}{[\text{FeO}][\text{H}_2]} \)
C. \( K_c = \frac{[\text{FeO}][\text{H}_2]}{[\text{Fe}][\text{H}_2\text{O}]} \)
D. \( K_c = \frac{[\text{H}_2\text{O}]}{[\text{H}_2]} \)

18. \( \text{N}_2\text{O}_4(g) \) and \( \text{NO}_2(g) \) are in equilibrium according to this equation.

\[ \text{N}_2\text{O}_4(g) \rightleftharpoons 2 \text{NO}_2(g) \quad K_c = ? \]

\( \text{N}_2\text{O}_4(g) \) at an initial concentration of 0.100 M produces a 0.132 M equilibrium concentration of \( \text{NO}_2(g) \) at 500 K. Determine \( K_c \) at 500 K.

A. 0.0972  
B. 0.210  
C. 0.314  
D. 0.512

19. The reaction

\[ 2 \text{NOCl}(g) \rightleftharpoons 2\text{NO}(g) + \text{Cl}_2(g) \]

has an equilibrium constant, \( K_P = 1.8 \times 10^{-2} \) at 800 °C. If the reaction quotient, \( Q = 9.3 \times 10^{-4} \) at 800 °C,

A. the reaction will proceed towards the reactants (to the left).
B. the reaction has reached equilibrium and no change in concentration occurs.
C. the reaction will proceed towards the products (to the right).
D. more reactants are needed for the reaction to continue.
20. If the equilibrium concentrations of PCl\(_3\) and Cl\(_2\) are both 0.0852 M, what is the equilibrium concentration of PCl\(_5\)?

PCl\(_5\)(g) ⇌ PCl\(_3\)(g) + Cl\(_2\)(g) \quad K_c = 0.030

A. 0.013 M  
B. 0.24 M  
C. 0.76 M  
D. 3.6 M

21. A reaction vessel is filled with 0.400 M of N\(_2\)(g) and 0.200 M of O\(_2\)(g). What is the equilibrium concentration of N\(_2\)O at 350 K?

2 N\(_2\)(g) + O\(_2\)(g) ⇌ 2 N\(_2\)O(g) \quad K_c = 7.3 \times 10^{-36} \text{ at } 350 \text{ K}

A. 2.3 \times 10^{-30} M  
B. 4.8 \times 10^{-19} M  
C. 3.7 \times 10^{-9} M  
D. 1.6 \times 10^{-5} M

22. Given the reaction,

\[ \text{CO}(g) + \text{H}_2\text{O}(g) \rightleftharpoons \text{CO}_2(g) + \text{H}_2(g) \quad \Delta H = -41 \text{ kJ; } K_c = 9.03 \text{ at } 698 \text{ K,} \]

which change will cause the greatest shift in the equilibrium toward \textit{products}?  

A. decreasing the pressure  
B. lowering the temperature  
C. adding Ar(g)  
D. adding H\(_2\)(g)
23. For the reaction,

\[ \text{P}_4(g) + 6 \text{Cl}_2(g) \rightleftharpoons 4 \text{PCl}_3(g) \]

at equilibrium at a fixed temperature, which of the following is correct?

A. Decreasing the volume causes the reaction to shift to the left (reactants).
B. Decreasing the volume causes the reaction to shift to the right (products).
C. Increasing the volume causes the reaction to shift to the right (products).
D. Increasing the volume has no effect on the composition of the equilibrium system.

24. Which acid-base theory describes the reactions below? What are the functions of the reactants?

Reaction 1: \(\text{NaOH}(aq) + \text{HCl}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)\)
Reaction 2: \(\text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s)\)

A. Reaction 1 is best described by the Arrhenius theory; NaOH is the acid and HCl is the base.
   Reaction 2 is best described by the Brønsted-Lowry theory; NH\(_3\) is the acid and HCl is the base.
B. Reaction 1 is best described by the Arrhenius theory; NaOH is the base and HCl is the acid.
   Reaction 2 is best described by the Brønsted-Lowry theory; NH\(_3\) is the base and HCl is the acid.
C. Reaction 1 is best described by the Brønsted-Lowry theory; NaOH is the acid and HCl is the base.
   Reaction 2 is best described by the Arrhenius theory; NH\(_3\) is the acid and HCl is the base.
D. Reaction 1 is best described by the Brønsted-Lowry theory; NaOH is the base and HCl is the acid.
   Reaction 2 is best described by the Arrhenius theory; NH\(_3\) is the base and HCl is the acid.

25. Considering the acid equilibrium reaction

\[ \text{HA}(aq) \rightleftharpoons \text{H}^+(aq) + \text{A}^-(aq) \]

A. the weaker the acid, the weaker its conjugate base.
B. the stronger the acid, the weaker its conjugate base.
C. the stronger the acid, the stronger its conjugate base.
D. a large acid ionization constant indicates a very weak acid.
26. Calculate the pH of pure water at 50 °C. \( K_w = 5.48 \times 10^{-14} \) at 50 °C.

A. 6.631  
B. 6.916  
C. 7.091  
D. 7.362

27. Which statement is true about the following acid ionization reactions?

\[
\text{HNO}_3(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{NO}_3^-(aq)
\]

\[
\text{HCHO}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{CHO}_2^-(aq)
\]

A. HNO_3 is a strong acid and [H_3O^+] is not equal to [HNO_3].  
B. HNO_3 is a weak acid and [H_3O^+] is equal to [HNO_3].  
C. HCHO_2 is a strong acid and [H_3O^+] is not equal to [HCHO_2].  
D. HCHO_2 is a weak acid and [H_3O^+] is not equal to [HCHO_2].

28. Which 0.100 M acid solution will have the highest percent ionization?

hypochlorous acid, HClO  
acetic acid, HC_2H_3O_2  
lactic acid, HC_3H_5O_3  
iodic acid, HI_3O

\[
K_a = 2.9 \times 10^{-8} 
\]

\[
K_a = 1.8 \times 10^{-5} 
\]

\[
K_a = 1.4 \times 10^{-4} 
\]

\[
K_a = 1.7 \times 10^{-1} 
\]

A. hypochlorous acid  
B. acetic acid  
C. lactic acid  
D. iodic acid
29. Determine the pH of a mixture of weak acids that is 0.200 M in HNO₂ and 0.150 M in HCN.

\[
\begin{align*}
\text{HNO}_2(\text{aq}) + \text{H}_2\text{O}(l) &\rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{NO}_2^-(\text{aq}) & K_a = 4.6 \times 10^{-4} \\
\text{HCN}(\text{aq}) + \text{H}_2\text{O}(l) &\rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{CN}^-(\text{aq}) & K_a = 4.9 \times 10^{-10} \\
\text{H}_2\text{O}(l) + \text{H}_2\text{O}(l) &\rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq}) & K_w = 1.0 \times 10^{-14}
\end{align*}
\]

A. 2.02  
B. 3.24  
C. 4.18  
D. 7.00

30. Determine the pH of a 0.27 M pyridine (C₅H₅N) solution.

\[
\text{C}_5\text{H}_5\text{N}(\text{aq}) + \text{H}_2\text{O}(l) \rightleftharpoons \text{C}_5\text{H}_5\text{NH}^+(\text{aq}) + \text{OH}^-(\text{aq}) \\
K_b = 1.7 \times 10^{-9}
\]

A. 4.66  
B. 8.77  
C. 9.33  
D. 13.43
Answer Key:
1. D
2. C
3. A
4. C
5. C
6. C
7. C
8. D
9. B
10. A
11. C
12. A
13. A
14. D
15. A
16. A
17. D
18. D
19. C
20. B
21. B
22. B
23. B
24. B
25. B
26. A
27. D
28. D
29. A
30. C