1. The rate of the reaction

\[ A \rightarrow X \]

is defined as
A. \(-\Delta[A]/\Delta t\).
B. the time it takes to convert all of A to X.
C. \([A]_{\text{initial}}/\Delta t\).
D. \((X) - [A]/\Delta t\).

2. For the reaction \(H_2(g) + Br_2(g) \rightarrow 2 \text{HBr}\), use the table below to determine the average \(\Delta[Br_2]/\Delta t\) from 20.0 to 30.0 seconds.

<table>
<thead>
<tr>
<th>Time (s)</th>
<th>[H(_2)], M</th>
<th>[Br(_2)], M</th>
<th>[HBr], M</th>
</tr>
</thead>
<tbody>
<tr>
<td>20.0</td>
<td>0.820</td>
<td>0.400</td>
<td>0.120</td>
</tr>
<tr>
<td>30.0</td>
<td>0.560</td>
<td>0.140</td>
<td>0.640</td>
</tr>
</tbody>
</table>

A. \(7.80 \times 10^{-2}\) M/s  
B. \(3.60 \times 10^{-3}\) M/s  
C. \(-2.60 \times 10^{-2}\) M/s  
D. \(-7.80 \times 10^{-2}\) M/s

3. On the graph at right, which plot describes the reaction rate as a function of reactant concentration for a first-order reaction?

A. A  
B. B  
C. C  
D. D
4. Given the following reaction and data set, determine the order of each reactant and the corresponding rate law:

\[ \text{CO(g)} + \text{NO}_2(g) \rightarrow \text{CO}_2(g) + \text{NO(g)} \]

<table>
<thead>
<tr>
<th>Trial</th>
<th>Initial [CO], M</th>
<th>Initial [NO\textsubscript{2}], M</th>
<th>Initial Rate (M/s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.10</td>
<td>0.10</td>
<td>0.00511</td>
</tr>
<tr>
<td>2</td>
<td>0.10</td>
<td>0.40</td>
<td>0.0817</td>
</tr>
<tr>
<td>3</td>
<td>0.20</td>
<td>0.10</td>
<td>0.00502</td>
</tr>
</tbody>
</table>

A. \( \text{rate} = k[\text{CO}] \)  
C. \( \text{rate} = k[\text{NO}_2][\text{CO}] \)
B. \( \text{rate} = k[\text{NO}_2][\text{CO}]^2 \)  
D. \( \text{rate} = k[\text{NO}_2]^2 \)

5. The half-life of the following first-order reaction at 720 K is 5.73 h.

\[ \text{H}_2\text{C} \backslash \text{CH}_2 \rightarrow \text{CH}_3 \backslash \text{CH} \equiv \text{CH}_2 \]

How long does it take to consume 85.0% of the reactant, cyclopropane, at 720 K?

A. 941 min  
C. 89.3 min
B. 765 min  
D. 13.0 min
6. Determine the **reaction order** and **rate constant** for the decomposition reaction 

\[ A \rightarrow B + C. \]

Plotting \(1/[A] \text{ vs. } t\) results in the graph at right with the equation

\[ y = 0.412x + 250. \]

A. zero order, 0.206 M\( \cdot \)s\(^{-1}\)  
B. zero order, 0.412 M\( \cdot \)s\(^{-1}\)  
C. first order, 0.206 s\(^{-1}\)  
D. second order, 0.412 M\(^{-1}\)\( \cdot \)s\(^{-1}\)

7. The rate constant for the first-order reaction,

\[ \text{KClO}_3(aq) \rightarrow \text{KCl}(aq) + \text{O}_2(g), \]

is \(1.37 \times 10^{-3}\) s\(^{-1}\). If the initial concentration of KClO\(_3\) is 5.47 M, what is the concentration after 60.0 s?

A. 5.39 M  
B. 5.04 M  
C. 3.77 M  
D. \(7.49 \times 10^{-3}\) M

8. Which statement regarding the Arrhenius equation is **false**?

A. The frequency factor, \(A\), is the number of times that reactants approach the activation barrier per unit time.

B. The exponential factor, \(e^{-E_a/RT}\), is the fraction of molecular approaches that are successful in overcoming the activation barrier, \(E_a\), to form products.

C. The exponential factor, \(e^{-E_a/RT}\), decreases with increasing temperature.

D. The exponential factor, \(e^{-E_a/RT}\), decreases with increasing activation energy, \(E_a\).
9. Consider the reaction

\[ \text{NO}_2(g) + \text{CO}(g) \rightarrow \text{NO}(g) + \text{CO}_2(g) \]

The activation energy is 145 kJ/mol. Calculate the rate constant at 520 °C, given that the rate constant at 430 °C is 2.57 M\(^{-1}\)·s\(^{-1}\).

A. 1.08 M\(^{-1}\)·s\(^{-1}\)   C. 8.69 M\(^{-1}\)·s\(^{-1}\)
B. 2.01 M\(^{-1}\)·s\(^{-1}\)   D. 42.9 M\(^{-1}\)·s\(^{-1}\)

10. Consider the reaction

\[ \text{N}_2\text{O}(g) \rightarrow \text{N}_2(g) + \text{O}(g) \]

The plot of \(\ln k\) vs. \(1/T\) at right gives a slope of \(-3.019 \times 10^4\) K and a \(y\)-intercept of 27.4. Determine the activation energy \((E_a)\) and frequency factor \((A)\) for the reaction.

A. \(E_a = 5.02 \times 10^5\) J/mol; \(A = 9.54 \times 10^{10}\) s\(^{-1}\)
B. \(E_a = 3.02 \times 10^5\) J/mol; \(A = 9.54 \times 10^{11}\) s\(^{-1}\)
C. \(E_a = 3.02 \times 10^5\) J/mol; \(A = 7.93 \times 10^{11}\) s\(^{-1}\)
D. \(E_a = 2.51 \times 10^5\) J/mol; \(A = 7.93 \times 10^{11}\) s\(^{-1}\)
11. Consider the reaction mechanism:

\[
\begin{align*}
(1) & \quad \text{NO}_2(g) + \text{Cl}_2(g) \rightarrow \text{ClNO}_2(g) + \text{Cl}(g) \\
(2) & \quad \text{NO}_2(g) + \text{Cl}(g) \rightarrow \text{ClNO}_2(g)
\end{align*}
\]

Identify the reaction intermediate.

A. NO\(_2\)(g)  
B. Cl\(_2\)(g)  
C. ClNO\(_2\)(g)  
D. Cl(g)

12. Consider the following reaction mechanism:

\[
\begin{align*}
(1) & \quad 2 \text{ NO}(g) \rightleftharpoons \text{N}_2\text{O}_2(g) \quad \text{fast} \\
(2) & \quad \text{N}_2\text{O}_2(g) + \text{H}_2(g) \rightarrow \text{N}_2\text{O}(g) + \text{H}_2\text{O}(g) \quad \text{slow} \\
(3) & \quad \text{N}_2\text{O}(g) + \text{H}_2(g) \rightarrow \text{N}_2(g) + \text{H}_2\text{O}(g) \quad \text{fast}
\end{align*}
\]

What is the rate law for this overall reaction?

\[
2 \text{ H}_2(g) + 2 \text{ NO}(g) \rightarrow \text{N}_2(g) + 2 \text{ H}_2\text{O}(g)
\]

A. rate = \(k[\text{NO}]^2[\text{H}_2]\)  
B. rate = \(k[\text{NO}]^2\)  
C. rate = \(k[\text{N}_2\text{O}_2][\text{H}_2]\)  
D. rate = \(k[\text{N}_2\text{O}][\text{H}_2]^2\)

13. Select the false statement.

A. A catalyst works by raising the activation energy.  
B. An enzyme is a biological catalyst.  
C. A catalyst increases the rate of a chemical reaction but is not consumed by the reaction.  
D. A homogeneous catalyst exists in the same phase as the reactants.

14. Which statement is always true about the concentrations of the reactants and products at equilibrium?

A. The concentrations of the reactants and products are equal.  
B. The reactant concentration decreases and the product concentration increases.  
C. The concentrations of the reactants and products are constant.  
D. None of the above is true.
15. If $K >> 1$, then

A. products are favored.
B. reactants are favored.
C. reactant and product concentration are about equal.
D. the ratio of products to reactants cannot be estimated.

16. Given the reactions

\[
\begin{align*}
\text{NO(g)} & \rightleftharpoons \frac{1}{2}\text{N}_2(g) + \frac{1}{2}\text{O}_2(g) & K_{c1} &= 1.0 \times 10^{15} \\
2 \text{NOBr(g)} & \rightleftharpoons \text{N}_2(g) + \text{O}_2(g) + \text{Br}_2(g) & K_{c2} &= 5.0 \times 10^{26}
\end{align*}
\]

Calculate $K_{c3}$ for the reaction

\[
2 \text{NO(g)} + \text{Br}_2(g) \rightleftharpoons 2 \text{NOBr(g)}
\]

$K_{c3} =$ ?

A. $4.0 \times 10^{-12}$
B. $2.0 \times 10^3$
C. $1.0 \times 10^9$
D. $5.0 \times 10^{38}$

17. For the reaction

\[
\text{N}_2(g) + 3 \text{H}_2(g) \rightleftharpoons 2 \text{NH}_3(g),
\]

$K_c = 5.9 \times 10^8$ at 25 °C. Calculate $K_p$ at 25 °C.

A. $9.8 \times 10^4$
B. $1.6 \times 10^5$
C. $9.9 \times 10^5$
D. $1.6 \times 10^7$
18. Which of the following is the equilibrium constant expression for this reaction?

\[ \text{NH}_4\text{SH}(s) \rightleftharpoons \text{H}_2\text{S}(g) + \text{NH}_3(g) \]

A. \( K_c = \frac{[\text{H}_2\text{S}][\text{NH}_3]}{[\text{NH}_4\text{SH}]} \)  
B. \( K_c = \frac{[\text{NH}_4\text{SH}]}{[\text{H}_2\text{S}][\text{NH}_3]} \)  
C. \( K_c = [\text{H}_2\text{S}][\text{NH}_3] \)  
D. \( K_c = \frac{1}{[\text{NH}_4\text{SH}]} \)

19. For the reaction

\[ 2 \text{A}(aq) + \text{B}(aq) \rightleftharpoons 3 \text{C}(aq) + 2 \text{D}(aq), \]

\([A] = 1.7 \text{ M}, [B] = 6.4 \text{ M}, [C] = 5.8 \text{ M}, \text{ and } [D] = 2.1 \text{ M} \) at equilibrium.

What is the value of \( K_c \)?

A. 47  
B. 1.1  
C. 0.89  
D. \( 2.1 \times 10^{-2} \)

20. Iron(II) reacts with cyanide to form a complex ion according to the equation:

\[ \text{Fe}^{2+}(aq) + 6 \text{CN}^-(aq) \rightleftharpoons [\text{Fe(CN)}_6]^{4-}(aq) \]

\( K_c = 1.5 \times 10^{35} \) at 298 K

Solutions of Fe\(^{2+}\) and CN\(^-\) are mixed such that \( Q_c = 1.5 \times 10^{20} \). Which statement is true?

A. The reaction is at equilibrium.
B. The reaction is not at equilibrium and will shift toward products to reach equilibrium.
C. The reaction is not at equilibrium and will shift toward reactants to reach equilibrium.
D. The reaction is not at equilibrium and cannot reach equilibrium under these conditions.
21. Consider this reaction and its equilibrium constant at 20 °C.

\[ \text{N}_2\text{O}_4(g) \rightleftharpoons 2 \text{NO}_2(g) \quad K_c = 1.35 \times 10^{-3} \]

A reaction mixture contains \([\text{NO}_2] = 0.0014 \text{ M}\) and \([\text{N}_2\text{O}_4] = 0.035 \text{ M}\). Calculate \(Q_c\) and determine the direction of the reaction at 20 °C.

A. \(Q_c = 5.6 \times 10^{-5}\); the reaction will proceed toward reactants.
B. \(Q_c = 5.6 \times 10^{-5}\); the reaction will proceed toward products.
C. \(Q_c = 3.9 \times 10^{-2}\); the reaction will proceed toward reactants.
D. \(Q_c = 4.0 \times 10^{-2}\); the reaction will proceed toward products.

22. A sealed flask is charged with pure \(\text{BrF}_5(g)\) and the system is allowed to reach equilibrium at 500 K. At equilibrium, the concentration of \(\text{BrF}_5\) is 0.096 M. What is the equilibrium concentration of \(\text{F}_2\)?

\[ \text{BrF}_5(g) \rightleftharpoons \text{BrF}_3(g) + \text{F}_2(g) \quad K_c = 1.2 \times 10^{-36} \text{ at } 500 \text{ K} \]

A. \(1.5 \times 10^{-37} \text{ M}\)  
B. \(3.4 \times 10^{-19} \text{ M}\)  
C. \(5.8 \times 10^{-10} \text{ M}\)  
D. \(0.45 \text{ M}\)

23. Consider the reaction

\[ 2 \text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2 \text{SO}_3(g) \quad K_p = 1.1 \times 10^{-6} \text{ at } 298 \text{ K} \]

If the initial pressure of \(\text{SO}_2\) is 0.84 atm and the initial pressure of \(\text{O}_2\) is 0.54 atm, what is the equilibrium pressure of \(\text{SO}_3\)?

A. \(1.5 \times 10^3 \text{ atm}\)  
B. \(3.6 \times 10^{-4} \text{ atm}\)  
C. \(6.5 \times 10^{-4} \text{ atm}\)  
D. \(4.5 \times 10^{-7} \text{ atm}\)
24. Given the reaction

\[ \text{CO}(g) + 2 \text{H}_2(g) \rightleftharpoons \text{CH}_3\text{OH}(g) \]

which change will cause the greatest shift in the equilibrium toward the products?

A. doubling the volume of the container  
B. doubling \( P_{\text{H}_2} \)  
C. doubling \( P_{\text{CO}} \)  
D. halving \( P_{\text{CO}} \)

25. Consider the following endothermic reaction at 298 K:

\[ 6 \text{H}_2\text{O}(g) + 6 \text{CO}_2(g) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(aq) + 6 \text{O}_2(g) \]

When the temperature is increased at constant pressure, what will happen to the system?

A. The equilibrium constant and the concentrations of reactants and products will all remain unchanged.  
B. The concentrations of products will increase and the equilibrium constant will decrease.  
C. The concentrations of reactants will increase and the equilibrium constant will increase.  
D. The concentrations of products will increase and the equilibrium constant will increase.

26. Given the reactions

(1) \( \text{NaOH}(aq) + \text{HCl}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l) \)

(2) \( \text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s) \),

which statement is false?

A. \( \text{NaOH} \) is both an Arrhenius base and a Brønsted-Lowry base.  
B. \( \text{HCl} \) is an Arrhenius acid in reaction (1) and a Brønsted-Lowry acid in reaction (2).  
C. \( \text{NH}_3 \) is both an Arrhenius base and a Brønsted-Lowry base.  
D. \( \text{NH}_3 \) is a Brønsted-Lowry, but not an Arrhenius, base.

27. Which of the following is not an acid?

A. \( \text{HCl} \)  
B. \( \text{H}_2\text{SO}_4 \)  
C. \( \text{NH}_4^+ \)  
D. \( \text{NaCl} \)
28. Which list correctly classifies the acids?

A. Strong: HClO₄, HNO₃  
   Weak: HF, H₂C₂H₂O₂

B. Strong: HCl, HNO₂  
   Weak: HCN, H₂SO₄

C. Strong: HI, HClO  
   Weak: H₃PO₄, HClO₄

D. Strong: HNO₃, H₂SO₃  
   Weak: H₂CO₃, HBr

29. Select the **strongest** of these acids.

A. HF, \( K_a = 3.5 \times 10^{-4} \)  
   \[ \text{C. } \text{HC}_6\text{H}_5\text{O}, \ K_a = 1.3 \times 10^{-10} \]

B. HNO₂, \( K_a = 4.6 \times 10^{-4} \)  
   \[ \text{D. } \text{HCN}, \ K_a = 4.9 \times 10^{-10} \]

30. Consider the weak acid dissociation

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

The rate laws for acid dissociation (forward reaction) and formation (reverse reaction) are

\[
\text{rate}_{\text{diss}} = k_{\text{diss}}[\text{HCN}]
\]

\[
\text{rate}_{\text{form}} = k_{\text{form}}[\text{H}_3\text{O}^+][\text{CN}^-]
\]

If one exists, determine the relationship between the rate constants and \( K_a \).

A. \( K_a = k_{\text{diss}} \cdot k_{\text{form}} \)

B. \( K_a = \frac{k_{\text{diss}}}{k_{\text{form}}} \)

C. \( K_a = \frac{k_{\text{form}}}{k_{\text{diss}}} \)

D. There is no relationship between the rate constants and \( K_a \).
1. A
2. C
3. C
4. D
5. A
6. D
7. B
8. C
9. D
10. D
11. D
12. A
13. A
14. C
15. A
16. B
17. C
18. C
19. A
20. B
21. B
22. B
23. C
24. B
25. D
26. C
27. D
28. A
29. B
30. B