PART 1

1&2. Which is the correct $K_c$ expression for the equilibrium: $[\text{Co(CO)}_4]_2(\text{s}) \rightleftharpoons 2\text{Co(s)} + 8\text{CO(g)}$?

(a) $K_c = \frac{[\text{Co}]^2[\text{CO}]^8}{[\text{Co(CO)}_4]^2}$
(b) $K_c = \frac{[\text{Co}]^2[\text{CO}]^8}{[\text{Co(CO)}_4]^2}$
(c) $K_c = \frac{[\text{CO}]^8}{[\text{Co(CO)}_4]^2}$
(d) $K_c = \frac{2[\text{Co}] + 8[\text{CO}]}{[\text{Co(CO)}_4]^2}$
(e) $K_c = [\text{CO}]^8$

3&4. Consider the equilibrium: $\text{H}_2(g) + \text{Br}_2(g) \rightleftharpoons 2\text{HBr(g)}$ $K_c = 7.9 \times 10^{-11}$

What is the equilibrium constant for the equilibrium: $4\text{HBr(g)} \rightleftharpoons 2\text{H}_2(g) + 2\text{Br}_2(g)$?

(a) $1.5 \times 10^{-10}$
(b) $2.2 \times 10^{12}$
(c) $1.6 \times 10^{20}$
(d) $3.1 \times 10^{15}$
(e) $6.8 \times 10^{21}$

5&6. Consider the conversion of a substance between liquid and a gas: liquid $\rightleftharpoons$ gas.

When the two phases are in equilibrium at one atmosphere pressure and at the boiling point of the substance,

(a) $\Delta H = 0$ for the process.
(b) $\Delta G = 0$ for the process.
(c) $\Delta E = 0$ for the process.
(d) $\Delta S = 0$ for the process.
(e) both (a) and (c)
7&8. Consider this reaction: $A_2 + B \rightarrow \text{products}$. The experimentally-derived rate law expression is:

$$\text{Rate} = k[A_2]^2[B].$$

If, during a reaction, the concentration of $A_2$ is halved and the concentration of $B$ is doubled, the rate of the reaction will _____ by a factor of _____.

(a) decrease, 2  
(b) increase, 2  
(c) increase, 4  
(d) decrease, 4  
(e) remain unchanged

9&10. The following mechanism is suggested for the reaction between NO and Br$_2$.

\begin{align*}
\text{Step 1} & \quad 2\text{NO} \rightleftharpoons \text{N}_2\text{O}_2 \quad \text{fast equilibrium (} K = \text{equilibrium constant)} \\
\text{Step 2} & \quad \text{N}_2\text{O}_2 + \text{Br}_2 \rightarrow 2\text{NOBr} \quad \text{slow (} k = \text{rate constant for forward reaction)}
\end{align*}

What is the appropriate rate law expression? Watch out for the correct constant combination!

(a) Rate = $K[N_2O_2][Br_2]$  
(b) Rate = $k[NO]^2$  
(c) Rate = $k[NO][Br_2]$  
(d) Rate = $kK[NO]^2[Br_2]$  
(e) Rate = $kK[NO]$  

11&12. Which of the following statements is/are FALSE concerning the action of catalysts?

(1) The activation energy of the rate-determining step is lowered and the reaction slows down.  
(2) Their presence does change the mechanism of the reaction.  
(3) When added to equilibrium reactions, a catalyst will NOT cause the equilibrium constant to change and the equilibrium to shift.

(a) 2, 3 only  
(b) 1 only  
(c) 2 only  
(d) 1, 2, 3  
(e) 1, 2 only
13&14. Consider the following rate law expression for the reaction: \( X + 2Y \rightarrow 2Z \)

\[
\text{Rate} = k[X]^2
\]

The units of the specific rate constant, \( k \), are:

(a) time\(^{-1}\)  (b) M\(^{-1}\)time\(^{-1}\)  (c) M\(^{-2}\)time\(^{-1}\)  (d) M\(^{3}\)time\(^{-1}\)  (e) Mtime\(^{-1}\)

15&16. Which set of curves represents the changes in concentration of A and B with time for a reaction:
\( A(g) \xrightarrow{\text{K}} 2B(g) \) in which K is much less than 1. Assume that initially only A is in the container at 1 M concentration.


17&18. Consider the following gas phase reaction: \( \text{H}_2(g) + \text{Cl}_2(g) \xleftrightarrow{\text{K}} 2\text{HCl}(g) \). The equilibrium constant at a particular temperature is 16. At that temperature the following is measured:

\[
[\text{H}_2] = [\text{Cl}_2] = 0.10 \text{ M} \\
[\text{HCl}] = 2.00 \text{ M}
\]

Which of the following statements is TRUE?

(a) \( Q > K \); the equilibrium is shifting so that more \( \text{H}_2(g) + \text{Cl}(g) \) will be produced.

(b) \( Q > K \); the equilibrium is shifting so that more \( \text{HCl}(g) \) will be produced.

(c) \( Q < K \); the equilibrium is shifting so that more \( \text{H}_2(g) + \text{Cl}_2(g) \) will be produced.

(d) \( Q < K \); the equilibrium is shifting so that more \( \text{HCl}(g) \) will be produced.

(e) \( Q = K \); the system is at equilibrium.
19&20. Rate data were collected for the following reaction at a particular temperature. What is the correct rate law expression?

\[ \text{2X} + \text{Y} \rightarrow \text{Z} \]

<table>
<thead>
<tr>
<th>Experiment</th>
<th>[X]\text{initial}</th>
<th>[Y]\text{initial}</th>
<th>Initial Rate of Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.60 M</td>
<td>0.20 M</td>
<td>0.050 M/s</td>
</tr>
<tr>
<td>2</td>
<td>0.30 M</td>
<td>0.20 M</td>
<td>0.025 M/s</td>
</tr>
<tr>
<td>3</td>
<td>0.60 M</td>
<td>0.40 M</td>
<td>0.20 M/s</td>
</tr>
</tbody>
</table>

(a) Rate = k[X][Y]  
(b) Rate = k[X]^2[Y]  
(c) Rate = k[X][Y]^2  
(d) Rate = k[X]^2[Y]^2  
(e) Rate = k[X]^2

21&22. Consider the following reaction and choose the correct statement:

\[ 2 \text{H}_2\text{O}(l) + \text{O}_2(g) \rightarrow 2 \text{H}_2\text{O}_2(l) \quad \Delta H^\circ = +196 \text{ kJ} \]
\[ \Delta S^\circ = -126 \text{ J/K} \]

(a) The reaction is spontaneous at all temperatures.  
(b) The reaction is nonspontaneous at all temperatures.  
(c) The reaction becomes spontaneous as the temperature increases.  
(d) The reaction becomes spontaneous as the temperature decreases.  
(e) The temperature does not influence the spontaneity of any reaction.
23&24. Consider the equilibrium at a certain temperature: \( Z \rightleftharpoons X + 3Y \)

A reaction begins with 6.0 moles of \( Z \) in a 3.0 L container. When the system reaches equilibrium, there are 3.0 moles of \( X \) present. What is the value of the equilibrium constant, \( K_c \)?

(a) 0.037   (b) 6   (c) 0.11   (d) 27   (e) 0.081

25&26. Consider the reaction below at 25°C for which \( \Delta G^\circ = +159 \text{ kJ/mol rxn} \) and \( \Delta H^\circ = +164 \text{ kJ/mol rxn} \). Calculate \( \Delta S^\circ \) at 25°C.

\[
\text{CH}_4(\text{g}) + \text{N}_2(\text{g}) \rightarrow \text{HCN}(\text{g}) + \text{NH}_3(\text{g})
\]

(a) +2.0 J/K   (b) +440 J/K   (c) +68 J/K   (d) +110 J/K   (e) +17 J/K

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27&28. Consider the equilibrium at a certain temperature: \( \text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g) \quad K_c = 63.0 \)

If 2.00 mol of HI are placed in a 1.00 L container, what is the concentration of HI at equilibrium?

(a) 1.60 M    (b) 1.86 M    (c) 0.85 M    (d) 0.27 M    (e) 0.51 M
Consider the equilibrium: \[ W(g) + 2X(s) \rightleftharpoons Y(s) + 3Z(g) \]
The reaction as written is strongly endothermic. Predict how the following changes will affect
(i) the moles of \( Z \),
(ii) the value of the equilibrium constant, and
(iii) the activation energy, \( E_a \), for the forward reaction.

The possible answers are: increase (I), decrease (D), or remain unchanged (U)

(a) The concentration of \( X \) is doubled.
(b) The temperature is decreased.
(c) A catalyst is introduced.
(d) The volume of the container is doubled.

Consider the following first order decomposition reaction: \[ A(g) \rightarrow B(g) + C(g) \]
You are given the following data:
rate constant at 0.0°C = \( 8.0 \times 10^{-7} \) s\(^{-1} \)
rate constant at 50.0°C = \( 8.9 \times 10^{-4} \) s\(^{-1} \)

\[ \Delta E_{\text{rxn}} = \Delta H_{\text{rxn}} = +51.5 \text{ kJ/mol} \]

(a) Calculate the activation energy, \( E_a \), (in kJ/mol rxn) for the reaction.

\[
\ln \left( \frac{k_2}{k_1} \right) = \frac{E_a}{R} \left( \frac{T_2 - T_1}{T_1 T_2} \right) \quad \text{or} \quad \ln \left( \frac{k_2}{k_1} \right) = \frac{E_a}{R} \left( \frac{1}{T_1} - \frac{1}{T_2} \right) \quad R = 8.314 \text{ J/mol} \cdot \text{K}
\]
(4 pts) (b) Sketch a potential energy diagram for this reaction. Label the axes. Use 50 kJ/mol for $E_a$ if you did not do part (b).

(1 pt) (c) Does the reaction release or absorb energy? (Look at the original data).

(3 pts) (d) Determine the activation energy for the reverse reaction.

(5 pts) (e) How many grams of a 20.0 g sample of A remain unreacted at 0.0°C after 2.00 days?

(1 pt) 31. Grammar bonus: Each time I go to an aquarium, I always check to see if there is an ________ (octopus or octopi).