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CLEARLY SHOW THE METHOD USED AND THE STEPS INVOLVED IN ARRIVING AT YOUR ANSWERS. It is to your advantage to do this, since you may obtain partial credit if you do and you will receive little or no credit if you do not. Attention should be paid to significant figures.

Be sure to write all your answers to the questions on the lined pages following each question in this booklet. Do NOT write your answers on the lavender insert.

Answer Question 1 below. The Section II score weighting for this question is 20 percent.

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

1. After a 1.0 mole sample of HI(g) is placed into an evacuated 1.0 L container at 700. K, the reaction represented above occurs. The concentration of HI(g) as a function of time is shown below.

(a) Write the expression for the equilibrium constant, \( K_c \), for the reaction.

(b) What is [HI] at equilibrium?
(c) Determine the equilibrium concentrations of $H_2(g)$ and $I_2(g)$.

(d) On the graph above, make a sketch that shows how the concentration of $H_2(g)$ changes as a function of time.

(e) Calculate the value of the following equilibrium constants at 700 K.

   (i) $K_c$

   (ii) $K_p$

(f) At 1,000 K, the value of $K_c$ for the reaction is $2.6 \times 10^{-2}$. In an experiment, 0.75 mole of $HI(g)$, 0.10 mole of $H_2(g)$, and 0.50 mole of $I_2(g)$ are placed in a 1.0 L container and allowed to reach equilibrium at 1,000 K. Determine whether the equilibrium concentration of $HI(g)$ will be greater than, equal to, or less than the initial concentration of $HI(g)$. Justify your answer.

\[ a) \quad K_c = \frac{[I_2][H_2]}{[HI]^2} \]

\[ b) \quad [HI] = 0.80 \, \text{M} \]

\[ c) \quad n(H_2) = n(I_2) = n(HI) \div 2 = CV/2 = \frac{(0.20)(1.0)}{2} = 0.10 \, \text{mol} \]

\[ \Rightarrow [I_2] = [H_2] = \frac{n}{V} = \frac{0.10}{1.0} = 0.10 \, \text{M} \]

\[ d) \quad i) \quad K_c = \frac{[0.10][0.10]}{[0.30]^2} = 0.015625 = 16 \times 10^{-3} \]

\[ ii) \quad K_p = K_c \cdot (RT)^n = 0.015625 \cdot (0.0821 \times 700)^{1-2} = 16 \times 10^{-3} \]

\[ e) \quad Q = \frac{(0.50)(0.10)}{(0.75)^2} = \frac{4}{45} = 8.9 \times 10^{-2} \]

$Q > K_c$, therefore the backward reaction is favoured to reach equilibrium. $[HI]$ will increase, $[HI] >$ initial.
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Answer Question 1 below. The Section II score weighting for this question is 20 percent.

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

1. After a 1.0 mole sample of \text{HI}(g) is placed into an evacuated 1.0 L container at 700. K, the reaction represented above occurs. The concentration of \text{HI}(g) as a function of time is shown below.

(a) Write the expression for the equilibrium constant, \( K_c \), for the reaction.

(b) What is \([\text{HI}]\) at equilibrium?
(c) Determine the equilibrium concentrations of \( H_2(g) \) and \( I_2(g) \).

(d) On the graph above, make a sketch that shows how the concentration of \( H_2(g) \) changes as a function of time.

(e) Calculate the value of the following equilibrium constants at 700. K.

   (i) \( K_c \)

   (ii) \( K_p \)

(f) At 1,000 K, the value of \( K_c \) for the reaction is \( 2.6 \times 10^{-2} \). In an experiment, 0.75 mole of \( HI(g) \), 0.10 mole of \( H_2(g) \), and 0.50 mole of \( I_2(g) \) are placed in a 1.0 L container and allowed to reach equilibrium at 1,000 K. Determine whether the equilibrium concentration of \( HI(g) \) will be greater than, equal to, or less than the initial concentration of \( HI(g) \). Justify your answer.

\[
A) \quad K_c = \frac{[H_2][I_2]}{[HI]^2} \\
\]

B) Using the graph, the concentration of \( HI \) is \( 0.80 \text{ mol L}^{-1} \)

\[
C) \quad [H_2] = [I_2] = x \\
\]

\[
K_c = \frac{(0.3)^2}{(0.8)^2} \\
\]

\[
K_p = K_c \frac{(RT)^n}{(P)^2} \\
K_p = 6 \times 10^{-2} \frac{(8.31 \cdot 700 \text{ K})^{2-2}}{x^{-2}} \\
K_c = \frac{[H_2][I_2]}{[HI]^2} \\
K_c = \frac{0.10 \times 0.50}{(0.45)^2} \\
K_c = 0.089 \\
K_c = 8.9 \times 10^{-2} \\
\]

Thus, the concentration of \( HI \) is greater than the initial concentration.
CHEMISTRY
Section II
(Total time—90 minutes)

Part A
Time—40 minutes
YOU MAY USE YOUR CALCULATOR FOR PART A.

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

1. After a 1.0 mole sample of HI(g) is placed into an evacuated 1.0 L container at 700. K, the reaction represented above occurs. The concentration of HI(g) as a function of time is shown below.

(a) Write the expression for the equilibrium constant, \( K_c \), for the reaction.

(b) What is [HI] at equilibrium?
(c) Determine the equilibrium concentrations of \( \text{H}_2(g) \) and \( \text{I}_2(g) \).

(d) On the graph above, make a sketch that shows how the concentration of \( \text{H}_2(g) \) changes as a function of time.

(e) Calculate the value of the following equilibrium constants at 700 K.
   
   (i) \( K_c \)
   
   (ii) \( K_p \)

(f) At 1,000 K, the value of \( K_c \) for the reaction \( 2 \text{HI}(g) \rightarrow \text{H}_2(g) + \text{I}_2(g) \) is \( 2.6 \times 10^{-2} \). In an experiment, 0.75 mole of \( \text{HI}(g) \), 0.10 mole of \( \text{H}_2(g) \), and 0.50 mole of \( \text{I}_2(g) \) are placed in a 1.0 L container and allowed to reach equilibrium at 1,000 K. Determine whether the equilibrium concentration of \( \text{HI}(g) \) will be greater than, equal to, or less than the initial concentration of \( \text{HI}(g) \). Justify your answer.

\[
\text{a)} \quad K_c = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]} \\
\text{b)} \quad [\text{HI}] \text{ at equilibrium} = 0.80 \text{ M} \\
\text{c)} \quad 2\text{HI} \rightleftharpoons \text{H}_2 + \text{I}_2 \quad x = 0.2 \text{ M} \\
\text{x = 0.2 M} \quad \therefore [\text{H}_2] \text{ and } [\text{I}_2] = 0.2 \text{ M} \\
\text{at equilibrium.} \\
\text{e)} \quad (i) \quad K_c = \frac{[0.2 \text{ M}][0.2 \text{ M}]}{1.0 \text{ M} - 0.2 \text{ M}} = \frac{0.05}{0.80} = K_c \\
\text{ii) } K_p = K_c (RT)^{\Delta n} \quad \Delta n = (2 \text{ moles}) - (2 \text{ moles}) = 0 \\
K_p = (0.05)(8.31 \text{ J/mol K})(700 \text{ K}) \quad K_p = (0.05)(1) = 0.05 = K_p \text{ also} \\
K_p = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} \\
\text{f)} \quad 2\text{HI} \rightleftharpoons \text{H}_2 + \text{I}_2 \quad 0.75 - x \quad 2.6 \times 10^{-2} = x^2 \quad x = 0.140 \text{ M} \\
\text{This would decrease the molality because 0.140 M would be subtracted from 0.75 M.}

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GO ON TO THE NEXT PAGE.