Chemical Equilibrium

Question1

At $T(\mathrm{~K}), K_c$ value of $A\mathrm{O}_2(g)+B\mathrm{O}_2(g) \rightleftharpoons A\mathrm{O}_3(g)+B\mathrm{O}(g)$ is 16 . In a closed 1 L flask, one mole each of AO_2, $BO_2 A\mathrm{O}_3$ and $B\mathrm{O}$ are taken and heated to $T(\mathrm{~K})$. Identify the correct statements about this equilibrium.

I. Total number of moles at equilibrium is 4 .

II. At equilibrium, the ratio of moles of $A\mathrm{O}_2$ and $A\mathrm{O}_3$ is $1: 4$.

III. Total number of moles of $A\mathrm{O}_2$ and $B\mathrm{O}_2$ at equilibrium is 0.8 .

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Options:

A.

I, II only

B.

I, III only

C.

II, III only

D.

I, II, III

Answer: D

Solution:

The given reaction is:

$AO_2(g) + BO_2(g) \rightleftharpoons AO_3(g) + BO(g)$

The equilibrium constant $K_c = 16$.

The volume of the flask is 1 L.

Initial moles of each species are 1 mole.

1. **Calculate the initial concentrations:**

   Since the volume is 1 L, the initial concentrations are numerically equal to the initial moles:

   $[AO_2]_0 = 1\text{ M}$

   $[BO_2]_0 = 1\text{ M}$

   $[AO_3]_0 = 1\text{ M}$

   $[BO]_0 = 1\text{ M}$

2. **Calculate the reaction quotient Q_c to determine the direction of the reaction:**

   $Q_c = \frac{[AO_3]_0[BO]_0}{[AO_2]_0[BO_2]_0} = \frac{(1)(1)}{(1)(1)} = 1$

3. **Compare Q_c with K_c:**

   Since $Q_c = 1$ and $K_c = 16$, we have $Q_c < K_c$. This means the reaction will proceed in the forward direction (to the right) to reach equilibrium.

4. **Set up an ICE table (Initial, Change, Equilibrium) for moles:**

   Let x be the change in moles. Since the reaction proceeds to the right, reactants decrease by x and products increase by x.

   Reaction: $AO_2(g) + BO_2(g) \rightleftharpoons AO_3(g) + BO(g)$

   Initial (mol): 1 1 1 1

   Change (mol): $-x$ $-x$ $+x$ $+x$

   Equilibrium (mol): $(1-x)$ $(1-x)$ $(1+x)$ $(1+x)$

   Since the volume of the flask is 1 L, the equilibrium concentrations are numerically equal to the equilibrium moles.

5. **Use the K_c expression to solve for x:**

   $K_c = \frac{[AO_3]_{eq}[BO]_{eq}}{[AO_2]_{eq}[BO_2]_{eq}}$

   $16 = \frac{(1+x)(1+x)}{(1-x)(1-x)}$

   $16 = \left(\frac{1+x}{1-x}\right)^2$

   Take the square root of both sides:

   $\sqrt{16} = \pm\frac{1+x}{1-x}$

$4 = \pm\frac{1+x}{1-x}$

Since the reaction proceeds in the forward direction, $x > 0$. Also, for the amounts to be positive, $1 - x > 0$, so $x < 1$. Thus, both $(1 + x)$ and $(1 - x)$ must be positive. We take the positive root.

$4 = \frac{1+x}{1-x}$

$4(1 - x) = 1 + x$

$4 - 4x = 1 + x$

$3 = 5x$

$x = \frac{3}{5} = 0.6$

6. **Calculate the equilibrium moles of each species:**

   $n_{AO_2} = 1 - x = 1 - 0.6 = 0.4 \text{ mol}$

   $n_{BO_2} = 1 - x = 1 - 0.6 = 0.4 \text{ mol}$

   $n_{AO_3} = 1 + x = 1 + 0.6 = 1.6 \text{ mol}$

   $n_{BO} = 1 + x = 1 + 0.6 = 1.6 \text{ mol}$

7. **Evaluate each statement:**

   I. Total number of moles at equilibrium is 4.

   Total moles at equilibrium = $n_{AO_2} + n_{BO_2} + n_{AO_3} + n_{BO}$

   Total moles = $0.4 + 0.4 + 1.6 + 1.6 = 4.0$ mol.

   This statement is **correct**. (Note: For this reaction, the number of moles of gaseous reactants equals the number of moles of gaseous products ($\Delta n_g = 0$), so the total number of moles remains constant throughout the reaction).

   II. At equilibrium, the ratio of moles of AO_2 and AO_3 is $1 : 4$.

   Ratio = $n_{AO_2} : n_{AO_3} = 0.4 : 1.6$

   To simplify, divide both sides by 0.4:

   Ratio = $1 : 4$.

   This statement is **correct**.

   III. Total number of moles of AO_2 and BO_2 at equilibrium is 0.8.

   Total moles of AO_2 and $BO_2 = n_{AO_2} + n_{BO_2}$

   Total moles = $0.4 + 0.4 = 0.8$ mol.

   This statement is **correct**.

All three statements (I, II, and III) are correct.

The final answer is $\boxed{D}$

Question2

For the following given equilibrium reaction $\frac{K_c}{K_p}$ is equal to 1076 at $T(\mathrm{~K})$. What is the value of T (in K)?

$$\left(R=0.082 \mathrm{~L}-\mathrm{atm} \mathrm{~K}^{-1} \mathrm{~mol}^{-1}\right)$$

$$\mathrm{N}_2(\mathrm{~g})+3 \mathrm{H}_2(\mathrm{~g}) \rightleftharpoons 2 \mathrm{NH}_3(\mathrm{~g})$$

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Options:

A.

500

B.

600

C.

400

D.

450

Answer: C

Solution:

The equilibrium reaction is

$\mathrm{N}_2(g)+3 \mathrm{H}_2(g) \rightleftharpoons 2 \mathrm{NH}_3(g)$

$\Delta n=$ moles of gaseous product - moles of gaseous reactant

$\Delta n=2-(1+3)=-2$

$$\frac{K_p}{K_c}=(R T)^{\Delta n}$$

or $\frac{K_c}{K_p}=(R T)^{-\Delta t}$

Substituting the values,

$$1076 = (0.082 \times T)^{+2}$$
$$T = 400 \mathrm{~K}$$

Question3

At $T(\mathrm{~K})$, consider the following gaseous reaction, which is in equilibrium.

$$\mathrm{N_2O_5} \rightleftharpoons 2\mathrm{NO_2} + \frac{1}{2}\mathrm{O_2}$$

What is the fraction of $\mathrm{N_2O_5}$ decomposed at constant volume and temperature, if the initial pressure is 300 mm Hg and pressure at equilibrium is 480 mm Hg ? (Assume all gases as ideal)

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Options:

A.

0.2

B.

0.6

C.

0.4

D.

0.8

Answer: C

Solution:

The given reaction is,

$N_2O_5 \longrightarrow 2NO_2 + \frac{1}{2}O_2$

Let $p_0 =$ initial pressure of $N_2O_5 x =$ pressure of N_2O_5 decomposed. Initial state $=$ N_2O_5 has pressure p_0

Equilibrium state $=$ N_2O_5 has pressure

$= p_0 - x$

$NO_2 = 2x$ pressure, $O_2 = \frac{1}{2}x$ pressure

$$\begin{aligned} p_{\text{equilibrium}} &= (p_0 - x) + 2x + \frac{1}{2}x \\ 480 &= p_0 + \frac{3}{2}x \\ \Rightarrow \quad p_0 &= 300\text{mmHg} \\ 180 &= \frac{3}{2}x, x = 120\text{mmHg} \end{aligned}$$

Fraction decomposed

$$= \frac{\text{pressure of } N_2O_5 \text{ decomposed}}{\text{initial pressure of } N_2O_5}$$

$$= \frac{x}{p_0} = \frac{120}{300} = \frac{2}{5} = 0.4$$

Question4

At 298 K , the value of K_p for $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ is 0.113 atm . The partial pressure of N_2O_4 at equilibrium is 0.2 atm . What is the partial pressure (in atm) of NO_2 equilibrium?

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Options:

A.

0.05

B.

0.075

C.

0.30

D.

0.15

Answer: D

Solution:

For reaction,
$N_2O_4(g) \rightleftharpoons 2NO_2(g)$
$K_p = \frac{(p_{NO_2})^2}{p_{N_2O_4}} \Rightarrow 0.113 = \frac{(p_{NO_2})^2}{0.2}$
$p_{NO_2} = 0.15$ atm

Question5

Consider the following gaseous equilibrium reactions (I), (II) and (III) with equilibrium constants K_1, K_2 and K_3 respectively

(I) $\frac{1}{2}N_2 + \frac{3}{2}H_2 \rightleftharpoons NH_3$

(II) $2NO \rightleftharpoons N_2 + O_2$

(III) $H_2 + \frac{1}{2}O_2 \rightleftharpoons H_2O$

The correct expression for the equilibrium constant for the gaseous equilibrium reaction

$2NH_3 + \frac{5}{2}O_2 \rightleftharpoons 2NO + 3H_2O$ is

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Options:

A.

$\frac{K_3^2}{K_1 \times K_2}$

B.

$\frac{K_3^3}{K_1^2 \times K_2}$

C.

$\frac{K_3^2}{K_1^2 \times K_2}$

D.

$\frac{K_3}{K_1^{\frac{1}{2}} \times K_2^2}$

Answer: B

Solution:

I. $\frac{1}{2}\mathrm{~N}_2+\frac{3}{2}\mathrm{H}_2 \rightleftharpoons \mathrm{NH}_3$

Reverse and double the equation I

$$\begin{aligned} & 2\mathrm{NH}_3 \rightleftharpoons \mathrm{~N}_2+3\mathrm{H}_2 \\ \Rightarrow \quad & K_1^{\prime}=\frac{1}{K_1^2} \qquad \ldots(i) \end{aligned}$$

II. $\quad 2\mathrm{NO} \rightleftharpoons \mathrm{N}_2+\mathrm{O}_2$

Reverse the equation II

$$\begin{aligned} & \mathrm{N}_2+\mathrm{O}_2 \rightleftharpoons 2\mathrm{NO} \\ \Rightarrow \quad & K_2^{\prime}=\frac{1}{K_2} \qquad \ldots(ii) \end{aligned}$$

Multiply by equation III

$$\begin{aligned} & 3\mathrm{H}_2+\frac{3}{2}\mathrm{O}_2 \rightleftharpoons 3\mathrm{H}_2\mathrm{O} \\ \Rightarrow \quad & K_3^{\prime}=K_3^3 \qquad \ldots(iii) \end{aligned}$$

Required equation is

$$2\mathrm{NH}_3+\frac{5}{2}\mathrm{O}_2 \rightarrow 2\mathrm{NO}+3\mathrm{H}_2\mathrm{O}$$

$$K=\frac{1}{K_1^2} \cdot \frac{1}{K_2} \times K_3^2=\frac{K_3^3}{K_1^2 K_2}$$

Question6

At $T(\mathrm{~K})$, the following gaseous equilibrium is established.

$$W+X \rightleftharpoons Y+Z$$

The initial concentration of W is two times to the initial concentration of X. The system is heated to $T(\mathrm{~K})$ to establish the equilibrium. At equilibrium the concentration of Y is four times to the concentration of X. What is the value of K_C ?

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Options:

A.

0.375

B.

1.333

C.

2.666

D.

5.333

Answer: C

Solution:

Given reaction:

$W+X \rightleftharpoons Y+Z$

Initial concentrations:

W starts at $2x$, X starts at x, and both Y and Z start at 0.

At equilibrium:

Let the amount of Y and Z formed be y. So, at equilibrium:

- W = $2x-y$
- X = $x-y$
- Y = y
- Z = y

Relationship between Y and X at equilibrium:

At equilibrium, the amount of Y is four times the amount of X. This gives us:

$[Y]_{eq} = 4[X]_{eq}$

So,$y = 4(x - y)$

Solve for y:

$y = 4x - 4y$

$y + 4y = 4x$

$5y = 4x$

$y = \frac{4}{5}x$

Write K_c expression and substitute values:

The equilibrium constant is:

$K_c = \frac{[Y][Z]}{[W][X]}$

Plug in the equilibrium values:

$K_c = \frac{\left(\frac{4}{5}x\right)\left(\frac{4}{5}x\right)}{\left(2x - \frac{4}{5}x\right)\left(x - \frac{4}{5}x\right)}$

Simplify:

Numerator: $\left(\frac{4}{5}x\right)^2 = \frac{16}{25}x^2$

Denominator:

- $2x - \frac{4}{5}x = \frac{10x - 4x}{5} = \frac{6x}{5}$
- $x - \frac{4}{5}x = \frac{5x - 4x}{5} = \frac{x}{5}$

So denominator: $\frac{6x}{5} \cdot \frac{x}{5} = \frac{6x^2}{25}$

Therefore, $K_c = \frac{16x^2/25}{6x^2/25} = \frac{16}{6} = 2.666$

Question7

At $T(\mathrm{~K}), K_C$ value for

$\mathrm{AO}_2(\mathrm{~g}) + \mathrm{BO}_2(\mathrm{~g}) \rightleftharpoons \mathrm{AO}_3(\mathrm{~g}) + \mathrm{BO}(\mathrm{g})$ is 16 . In a closed 1 L flask, one mole each of AO_2, BO_2, AO_3 and BO are taken and heated to $T(\mathrm{~K})$.

What is the concentration (in molL^{-1}) of AO_3 at equilibrium?

Options:

A.

0.4

B.

0.6

C.

1.6

D.

1.4

Answer: C

Solution:

We are given these starting amounts:

$[AO_2] = 1\ \text{mol/L}\ [BO_2] = 1\ \text{mol/L}\ [AO_3] = 1\ \text{mol/L}\ [BO] = 1\ \text{mol/L} K_c = 16$

We need to find the equilibrium concentration of AO_3.

Step 1: Write the reaction equation.

$AO_2\ (g) + BO_2\ (g) \rightleftharpoons AO_3\ (g) + BO\ (g)$

Step 2: Set up an ICE (Initial, Change, Equilibrium) table.

At the start (initial):
$[AO_2] = 1,\ [BO_2] = 1,\ [AO_3] = 1,\ [BO] = 1$

Let x be the amount that reacts.

Change in concentrations: $\begin{array}{cccc} AO_2 & BO_2 & AO_3 & BO \\ -x & -x & +x & +x \end{array}$

At equilibrium:
$[AO_2] = 1 - x$
$[BO_2] = 1 - x$
$[AO_3] = 1 + x$
$[BO] = 1 + x$

Step 3: Write the equilibrium expression and solve for x.

$K_c = \frac{[AO_3][BO]}{[AO_2][BO_2]} = \frac{(1+x)(1+x)}{(1-x)(1-x)} = 16$

Take the square root of both sides: $\frac{1+x}{1-x} = 4$

Solve for x:

$1 + x = 4(1 - x)$

$1 + x = 4 - 4x$

$1 + x + 4x = 4$

$1 + 5x = 4$

$5x = 3$

$x = 0.6$

Step 4: Find the equilibrium concentration of AO_3.

$[AO_3] = 1 + x = 1 + 0.6 = 1.6\ \text{mol/L}$

Question8

At 298 K , the value of K_c for the following reaction is $x\ \text{mol L}^{-1}$.

What is the approximate K_{P} value for this reaction?

$$\left(R = 0.082\ \text{L atm mol}^{-1}\ \text{K}^{-1}\right)$$
$$\text{A}_2\text{O}_4(\text{ g}) \rightleftharpoons 2\text{AO}_2(\text{ g})$$

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Options:

A.

$24.4x$

B.

$122x$

C.

$\frac{x}{24.4}$

D.

$\frac{24.4}{x}$

Answer: A

Solution:

Given reaction:

$A_2O_4(g) \rightleftharpoons 2AO_2(g)$

At 298 K, the equilibrium constant in concentration units is $K_c = x\,\mathrm{mol\,L^{-1}}$.

We need to find the approximate value of K_P.

Relation between K_P and K_C

For a gaseous reaction:

$K_P = K_C(RT)^{\Delta n}$

where

$\Delta n = \text{(moles of gaseous products)} - \text{(moles of gaseous reactants)}$

Step 1: Determine Δn

$$\text{Moles of gaseous products} = 2$$
$$\text{Moles of gaseous reactants} = 1$$
$$\Rightarrow \Delta n = 2 - 1 = +1$$

Step 2: Substitute into the relation

$K_P = K_C(RT)^1 = K_C(RT)$

Step 3: Plug in the given values

At $T = 298\,\mathrm{K}$,

$R = 0.082\,\mathrm{L\,atm\,mol^{-1}\,K^{-1}}$

$RT = 0.082 \times 298 = 24.4\,\mathrm{L\,atm\,mol^{-1}}$

Step 4: Therefore,

$K_P = K_C \times 24.4$

$\boxed{K_P = 24.4x}$

✅ **Final Answer: Option A** — $24.4x$

Question9

At $293 \mathrm{~K}, \Delta_r G^{\circ}$ for the following reaction is $165.469 \mathrm{~kJ} \mathrm{~mol}^{-1}$.

$\frac{3}{2} \mathrm{O}_2(\mathrm{~g}) \longrightarrow \mathrm{O}_3(\mathrm{~g})$

What is the equilibrium constant for this reaction?

$\left(R=83 \mathrm{~J} \mathrm{~mol}^{-1} \mathrm{~K}^{-1}\right)$

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Options:

A.

10^{29}

B.

10^{-29}

C.

5×10^{-27}

D.

$5 \times 10^{+27}$

Answer: B

Solution:

Gibbs free energy is given by

$$\Delta G^{\circ}=-R T \ln K$$

$$\Rightarrow \quad \ln K=-\frac{\Delta G^{\circ}}{R T}$$

$$\ln K=\frac{-165.469}{8.3 \times 298} \approx-66.8$$

$$K=e^{-66.8} \approx 9.75 \times 10^{-30} \approx 10^{-29}$$

So, $K \cong 10^{-29}$

Question10

The following equilibrium is established at STP.

$$B_2(g) \rightleftharpoons 2B(g)$$

Atoms of B occupy 20% of total volume at STP. The total pressure of the system is 1 bar. What is its K_p ? (STP volume $= 22.7$ L)

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Options:

A.

0.05

B.

0.1

C.

0.5

D.

0.025

Answer: A

Solution:

Volume occupied by B atom $= 20\%$ of total volume

$$\begin{aligned} &= 20\% \text{ of } 22.7 \text{ L} = 4.54 \text{ L} \\ \text{Remaining volume } &= (22.7 - 4.54)\text{L} \\ &= 18.16 \text{ L} \end{aligned}$$

Mole fraction of $B = \chi_B = \frac{4.54}{22.7} = 0.2$

Mole fraction of $B_2 = \chi_{B_2} = \frac{18.16}{22.7} = 0.8$

Partial pressure of $B = p_B = 0.2 \times 1 = 0.2$ bar

Partial pressure of

$B_2 = p_{B_2} = 0.8 \times 1 = 0.8$ bar

$$K_p = \frac{p_B^2}{p_{B_2}} = \frac{0.2^2}{0.8} = \frac{0.04}{0.8} = 0.05$$

Question11

At 300 K, K_C for the reaction. $A_2B_2(g) \rightleftharpoons A_2(g) + B_2(g)$ is 100 mol L^{-1}, What is its K_p (in atm) at the same temperature ? $\left(R = 0.082 \text{ L atm mol}^{-1} \text{ K}^{-1}\right)$

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Options:

A. 100

B. 2460

C. 4.06

D. 246

Answer: B

Solution:

To determine K_p for the reaction $A_2B_2(g) \rightleftharpoons A_2(g) + B_2(g)$ at 300 K, we are provided with the following information:

Equilibrium constant $K_C = 100 \text{ mol L}^{-1}$

Temperature $T = 300 \text{ K}$

Gas constant $R = 0.082 \text{ L atm mol}^{-1}\text{K}^{-1}$

The relationship between K_C and K_p is given by the equation:

$K_p = K_C(RT)^{\Delta n}$

Where Δn is the change in the number of moles of gas, calculated as the difference between the total moles of gaseous products and reactants:

$\Delta n = \text{moles of products} - \text{moles of reactants} = 2 - 1 = 1$

Now, substituting these values into the formula:

$K_p = 100 \times (0.082 \times 300)^1$

$K_p = 100 \times 24.6 = 2460\,\text{atm}$

Thus, the value of K_p for the reaction at 300 K is $2460\,\text{atm}$.

Question12

At equilibrium of the reaction,

$$A_2(g) + B_2(g) \rightleftharpoons 2AB(g)$$

The concentrations of A_2, B_2 and AB respectively are $15 \times 10^{-3}\text{M}, 2.1 \times 10^{-3}\text{M}$, and $1.4 \times 10^{-3}\text{M}$ in a sealed vessel at 800 K . What will be K_p for the decomposition of AB at same temperature ?

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Options:

A. 0.62

B. 1.6

C. 0.44

D. 2.27

Answer: B

Solution:

To find the equilibrium constant K_p for the decomposition of AB at 800 K, follow these steps:

Given Equilibrium Reaction:

$A_2(g) + B_2(g) \rightleftharpoons 2AB(g)$

Concentrations at Equilibrium:

$[A_2] = 1.5 \times 10^{-3}\,\text{M}$

$[B_2] = 2.1 \times 10^{-3}\,\mathrm{M}$

$[AB] = 1.4 \times 10^{-3}\,\mathrm{M}$

Decomposition Reaction:

$2AB(g) \rightleftharpoons A_2(g) + B_2(g)$

Expression for K_p:

$K_p = \frac{[A_2][B_2]}{[AB]^2}$

Calculation:

Substitute the concentration values into the expression for K_p:

$K_p = \frac{(1.5\times10^{-3})(2.1\times10^{-3})}{(1.4\times10^{-3})^2}$

Calculate:

$K_p = \frac{(1.5\times2.1)\times10^{-6}}{1.96\times10^{-6}}$

$K_p = \frac{3.15\times10^{-6}}{1.96\times10^{-6}}$

$K_p = 1.6$

Thus, the value of K_p for the decomposition of AB at 800 K is 1.6.

Question13

15 moles of H_2 and 5.2 moles of I_2 are mixed and allowed to attain equilibrium at 773 K . At equilibrium, the number of moles of HI is found to be 10 . The equilibrium constant for the dissociation of HI is

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Options:

A. 2×10^{-2}

B. 50

C. 2×10^{-1}

D. 5.0

Answer: B

Solution:

Given the reaction:

$H_2 + I_2 \rightleftharpoons 2HI$

Initial Conditions:

At $t = 0$, the initial concentrations of the reactants are:

$H_2 = 15$ moles

$I_2 = 5.2$ moles

At Equilibrium:

It is given that the number of moles of HI at equilibrium is 10.

Since the balanced equation shows that 2 moles of HI are formed per mole of reactants, we can set:

$2x = 10$, thus $x = 5$ moles

Therefore, at equilibrium:

$H_2 = 15 - 5 = 10$ moles

$I_2 = 5.2 - 5 = 0.2$ moles

Calculating the Equilibrium Constant:

The equilibrium constant (K_C) for the dissociation of HI is calculated by:

$K_C = \frac{[HI]^2}{[H_2][I_2]}$

Plugging in the equilibrium concentrations:

$K_C = \frac{(10)^2}{10 \times 0.2} = 50$

Therefore, the equilibrium constant for the dissociation of HI is 50.

Question14

K_C for the reaction,

$A_2(g) \overset{T(\mathrm{~K})}{\rightleftharpoons} B_2(\mathrm{~g})$

is 39.0. In a closed one litre flask, one mole of $A_2(g)$ was heated to $T(\mathrm{~K})$. What are the concentrations of $A_2(g)$ and $B_2(g)$ (in mol L $^{-1}$) respectively at equilibrium?

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Options:

A. $0.025, 0.975$

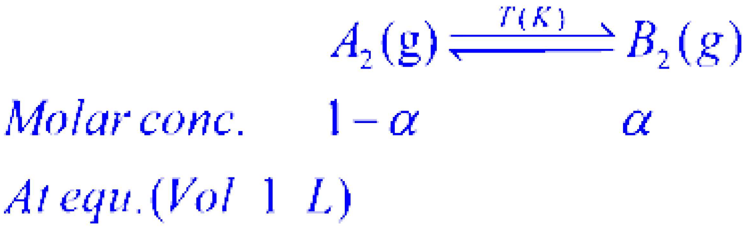
B. $0.975, 0.025$

C. $0.05, 0.95$

D. $0.02, 0.98$

Answer: A

Solution:



$$A_2(g) \xrightleftharpoons{T(K)} B_2(g)$$

Molar conc. At equ.(Vol 1 L) $1-\alpha$ α

$$\begin{aligned} K_C &= \frac{\alpha}{1-\alpha} = 39 \\ 39 - 39\alpha &= \alpha \\ [B_2] = \alpha &= \frac{39}{40} = 0.975 \\ [A_2] &= 1 - \alpha = 0.025 \end{aligned}$$

Question15

At $T(\mathrm{~K})$, the equilibrium constant for the reaction
$H_2(g) + Br_2(\mathrm{~g}) \rightleftharpoons 2HBr(g)$

is 1.6×10^5. If 10 bar of HBr is introduced into a sealed vessel at $T(\mathrm{~K})$, the equilibrium pressure of HBr (in bar) is approximately

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Options:

A. 10.20

B. 10.95

C. 9.95

D. 11.95

Answer: C

Solution:

The reaction given is:

$H_2(g) + Br_2(g) \rightleftharpoons 2HBr(g)$

The equilibrium constant K_C is 1.6×10^5.

When 10 bar of HBr is introduced into a sealed vessel, the equilibrium of the decomposition reaction can be expressed as:

$2HBr(g) \rightleftharpoons H_2(g) + Br_2(g)$

For this reverse reaction, the equilibrium constant K_C becomes:

$K_C = \frac{1}{1.6 \times 10^5}$

Since $\Delta n_g = 0$, we have:

$K_p = K_C = \frac{1}{1.6 \times 10^5}$

For the decomposition:

Initial pressure of HBr = 10 bar

At equilibrium, pressure of HBr = $10 - 2p$

The expression for the equilibrium constant in terms of pressure is:

$K_p = \frac{p^2}{(10-2p)}$

Equating and solving:

$\frac{1}{1.6 \times 10^5} = \frac{p^2}{(10-2p)}$

Solving this gives $p = 0.025$ bar.

Thus, the equilibrium pressure of HBr is:

$p_{HBr} = 10 - 2(0.025) = 9.95$ bar

Question16

K_C for the reaction, $A_2(g) \overset{T(\mathrm{~K})}{\rightleftharpoons} B_2(g)$ is 99.0 . In a 1 L closed flask two moles of $B_2(g)$ is heated to $T(\mathrm{~K})$. What is the concentration of $B_2(g)$ (in molL^{-1}) at equilibrium?

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Options:

A. 0.02

B. 1.98

C. 0.198

D. 1.5

Answer: B

Solution:

The given reaction is:

$A_2(g) \overset{T(K)}{\rightleftharpoons} B_2(g)$

with an equilibrium constant $K_C = 99$.

We are interested in finding the equilibrium concentration of $B_2(g)$ when starting with 2 moles of $B_2(g)$ in a 1 L flask. This is effectively the reverse of the initial reaction:

$B_2(g) \overset{T(K)}{\rightleftharpoons} A_2(g)$

For this reverse reaction, the equilibrium constant is:

$K'_C = \frac{1}{K_C} = \frac{1}{99}$

Initially, at $t = 0$, the concentration of B_2 is:

$[B_2] = 2\,\mathrm{mol/L}$

At equilibrium, let α be the degree of dissociation of B_2 into A_2. Thus, the concentrations become:

$[B_2] = 2 - 2\alpha$

$[A_2] = 2\alpha$

Using the expression for the reverse reaction:

$K'_C = \frac{[A_2]}{[B_2]} = \frac{2\alpha}{2-2\alpha} = \frac{1}{99}$

Solving for α:

$\frac{2\alpha}{2-2\alpha}=\frac{1}{99}$

$\Rightarrow 2\alpha \times 99=2-2\alpha$

$\Rightarrow 198\alpha=2-2\alpha$

$\Rightarrow 200\alpha=2$

$\Rightarrow \alpha=\frac{2}{200}=0.01$

Therefore, the equilibrium concentration of B_2 is:

$[B_2]_{\text{eq}}=2-2\times 0.01=1.98\,\text{mol/L}$

Question17

K_c for the following reaction is 99.0

$$A_2(g) \overset{T(K)}{\rightleftharpoons} B_2(g)$$

In a one litre flask, 2 moles of A_2 was heated to $T(\text{K})$ and the above equilibrium is reached. The concentration at equilibrium of A_2 and B_2 are $C_1\left(A_2\right)$ and $C_2\left(B_2\right)$ respectively. Now, one mole of A_2 was added to flask and heated to $T(\text{K})$ to established the equilibrium again. The concentration of A_2 and B_2 are $C_3\left(A_2\right)$ and $C_4\left(B_2\right)$ respectively. what is the value of $C_3\left(A_2\right)$ in molL^{-1} ?

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Options:

A. 1.98

B. 0.01

C. 0.03

D. 2.97

Answer: C

Solution:

1. Initial Equilibrium

We start with 2 moles of A_2 in a 1 L flask, so the initial concentration of A_2 is 2 M.

Let 'x' be the change in concentration of A_2 as it converts to B_2 at equilibrium.

	$A_2(g) \rightleftharpoons$	$B_2(g)$
Initial	2	0
Change	$-x$	$+x$
Equilibrium	$2-x$	x

The equilibrium constant K_c is given by:

$$K_c = \frac{[B_2]}{[A_2]} = \frac{x}{2-x} = 99.0$$

Solving for x:

$$x = 99(2-x)$$

$$x = 198 - 99x$$

$$100x = 198$$

$$x = 1.98$$

So, at the first equilibrium:

$$C_1(A_2) = 2 - x = 2 - 1.98 = 0.02\,\mathrm{M}$$

$$C_2(B_2) = x = 1.98\,\mathrm{M}$$

2. Adding More A_2 and Re-establishing Equilibrium

Now, we add 1 mole of A_2 to the flask. The concentration of A_2 becomes $0.02 + 1 = 1.02\,\mathrm{M}$. The concentration of B_2 remains $1.98\,\mathrm{M}$.

Let 'y' be the change in concentration of A_2 as it converts to B_2 to reach the new equilibrium.

	$A_2(g) \rightleftharpoons$	$B_2(g)$
Initial	1.02	1.98
Change	$-y$	$+y$
Equilibrium	$1.02-y$	$1.98+y$

The equilibrium constant K_c remains the same:

$$K_c = \frac{[B_2]}{[A_2]} = \frac{1.98+y}{1.02-y} = 99.0$$

Solving for y:

$$1.98 + y = 99(1.02 - y)$$

$$1.98 + y = 100.98 - 99y$$

$$100y = 99$$

$$y = 0.99$$

So, at the second equilibrium:

$C_3(A_2) = 1.02 - y = 1.02 - 0.99 = 0.03\,\mathrm{M}$

$C_4(B_2) = 1.98 + y = 1.98 + 0.99 = 2.97\,\mathrm{M}$

Therefore, the concentration of A_2 at the new equilibrium, $C_3(A_2)$, is 0.03 M.

Answer:

Option C is the correct answer: 0.03

Question18

At $T(\mathrm{~K}), K_c$ for the reaction $A_2(g) \rightleftharpoons B_2(g)$ is 99.0 . Two moles of $A_2(s)$ was heated to $T(\mathrm{~K})$ in a 1 L . closed flask to reach the above equilibrium. What are the concentrations (in mol L^{-1}) of $A_2(g)$ and $B_2(g)$ respectively at equilibrium?

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Options:

A. $1,86,0.0187$

B. $1.98,0.02$

C. $0.0187,1,86$

D. $0.02,1.98$

Answer: D

Solution:

To analyze the equilibrium concentrations for the reaction $A_2(g) \rightleftharpoons B_2(g)$, where the equilibrium constant K_c at temperature $T(\mathrm{K})$ is 99.0, follow these steps:

Given:

$K_c = 99.0$

Initial moles of A_2: 2 moles

Volume of the flask: 1 L

Calculation:

Initial Conditions and Definition:

Since the reaction starts with 2 moles of $A_2(s)$ which sublimes to $A_2(g)$, the initial concentration of A_2 is 2 mol/L (as the flask is 1 L in volume).

Setting up the Equilibrium Expression:

For the equilibrium constant expression:

$K_c = \frac{[B_2]}{[A_2]}$

Let the concentration of B_2 at equilibrium be x mol/L. Therefore, the concentration of A_2 becomes $(2-x)$ mol/L at equilibrium.

Solving the Equilibrium Equation:

Substitute the equilibrium values into the K_c expression:

$99 = \frac{x}{2-x}$

Solving for x:

$99(2-x) = x$

$198 - 99x = x$

$198 = 100x$

$x = 1.98$

Thus, the concentration of B_2 is $[B_2] = 1.98$ mol/L.

Finding the Concentration of A_2:

Calculate $[A_2]$ using the expression $2-x$:

$[A_2] = 2 - 1.98 = 0.02$ mol/L

Conclusion:

The concentrations at equilibrium are:

$[A_2] = 0.02\,\text{mol/L}$

$[B_2] = 1.98\,\text{mol/L}$

Question19

At $27^{\circ}\mathrm{C}$, the degree of dissociation of weak acid ($\mathrm{H}A$) in its 0.5 M aqueous solution is 1%. Its K_e, value is approximately

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Options:

A. 5×10^{-6}

B. 5×10^{-5}

C. 5×10^{-6}

D. 5×10^{-8}

Answer: B

Solution:

Let's work through the problem step by step.

The dissociation of the weak acid is given by:

$\mathrm{HA} \rightleftharpoons \mathrm{H}^{+} + \mathrm{A}^{-}$.

The initial concentration of HA is 0.5 M. With a degree of dissociation, $\alpha = 0.01$ (or 1%), the equilibrium concentrations are:

[HA] = 0.5(1 - 0.01) = 0.5 × 0.99 = 0.495 M

$[H^+]$ = 0.5 × 0.01 = 0.005 M

$[A^-]$ = 0.005 M

The acid dissociation constant, K_a, is defined as:

$K_a = \frac{[\mathrm{H}^{+}][\mathrm{A}^{-}]}{[\mathrm{HA}]}$.

Substitute the equilibrium concentrations:

$K_a = \frac{(0.005)(0.005)}{0.495} = \frac{0.000025}{0.495}$.

Simplifying the expression:

$K_a \approx 5 \times 10^{-5}$.

Thus, the approximate value of K_a (or K_e, as referenced) is 5×10^{-5}.

The correct option is Option B.

Question20

At 500 K , for the reaction $\mathrm{N}_2(\mathrm{~g}) + 3\mathrm{H}_2(\mathrm{~g}) \rightleftharpoons 2\mathrm{NH}_3(\mathrm{~g})$, the K_p is $0.036\ \mathrm{atm}^{-2}$. What is its K_C in $\mathrm{L}^2\ \mathrm{mol}^{-1}$? $\left(R = 0.082\ \mathrm{L}^2\right.$ atom $\left.\mathrm{mol}^{-1}\ \mathrm{K}^{-1}\right)$.

Options:

A. 2.1×10^{-4}

B. 2.1×10^{-5}

C. 60.5

D. 605

Answer: C

Solution:

At 500 K, for the reaction

$N_2(\mathrm{~g}) + 3H_2(\mathrm{~g}) \rightleftharpoons 2NH_3(\mathrm{~g})$

the equilibrium constant K_p is 0.036 atm^{-2}. To find the equilibrium constant K_C in $\mathrm{L}^2\ \mathrm{mol}^{-1}$, given that the gas constant R is 0.082 $\mathrm{L\ atm\ mol}^{-1}\ \mathrm{K}^{-1}$, we utilize the relationship between K_p and K_C. The formula is:

$K_p = K_C(RT)^{\Delta n}$

For this reaction:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

The change in moles of gas, Δn, is calculated as follows:

$\Delta n = (\text{moles of products}) - (\text{moles of reactants}) = 2 - (3 + 1) = 2 - 4 = -2$

Given that:

$K_p = 0.036$ (provided)

We can now use the relationship to find K_C:

$0.036 = K_C(0.082 \times 500)^{-2}$

Solving for K_C gives us:

$K_C = \frac{0.036}{(0.082 \times 500)^{-2}} = 60.516\ \mathrm{L}^2\ \mathrm{mol}^{-2}$

Question21

The formation of ammonia from its constituent elements is an exothermic reaction. The effect of increased temperature on the reaction equilibrium is

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Options:

A. the rate of the forward reaction becomes zero.

B. no effect of temperature.

C. forward reaction is favoured

D. backward reaction is favoured

Answer: D

Solution:

$3\mathrm{H}_2(g)+\mathrm{N}_2(g) \rightleftharpoons 2\mathrm{NH}_3(g) ; \Delta H=+\mathrm{ve}$,

As the formation of ammonia from its constituent elements is an exothermic reaction. Thus, energy is released during the reaction. So, on increasing the temperature, the reaction is favoured in backward direction. As per Le-Chatelier's principle, change made in one direction/site is opposed by movement of reaction in opposite direction.

Question22

At $60^{\circ}\mathrm{C}$, dinitrogen tetroxide is dissociated. Find it's standard free energy change at this temperature and one atmosphere. [Given $\log 1.33=0.1239$]

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Options:

A. $-650 \mathrm{~J} \mathrm{~mol}^{-1}$

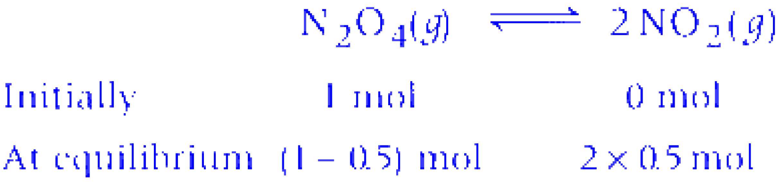
B. -830 J mol^{-1}

C. -790 J mol^{-1}

D. -875 J mol^{-1}

Answer: C

Solution:



	$N_2O_4(g)$	$\rightleftharpoons$	$2NO_2(g)$
Initially	1 mol		0 mol
At equilibrium	$(1-0.5)$ mol		2×0.5 mol

Total number of moles $= 0.5 + 1 = 1.5$ mol

$$p_{N_2O_4} = \frac{0.5}{1.5} \times 1 \text{ atm} = \frac{1}{3} \text{ atm}$$
$$p_{NO_2} = \frac{1}{1.5} \times 1 \text{ atm} = \frac{2}{3} \text{ atm}$$

According to law of chemical equilibrium

$$K_p = \frac{p^2NO_2}{p_{N_2O_4}} = \frac{\left(\frac{2}{3}\right)^2}{\left(\frac{1}{3}\right)} = 1.33 \text{ atm}$$

We know,

$\Delta G\Upsilon = -2.303 RT \log K_p$ ...... (i)

Substituting value in Eq. (i), we get

$$\begin{aligned} \Delta G\Upsilon &= -2.303 \times 0.082 \text{JK}^{-1} \text{ mol}^{-1} \times 333 \text{ K} \\ &= -7.79 \text{ atm litre} \end{aligned}$$

$$\begin{aligned} 1 \text{ atm litre } &= 101.32500 \text{ joules} \\ -7.79 \text{ atm litre } &= -7.79 \times 101.32 \text{ joules} \\ &= -789.2 \text{ J mol}^{-1} \approx -790 \text{ J mol}^{-1} \end{aligned}$$

Question23

Le-Chatelier's principle is not applicable to

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Options:

A. $H_2(g)+I_2(g) \rightleftharpoons 2HI(g)$

B. $Fe(s)+S(s) \rightleftharpoons FeS(s)$

C. $N_2(g)+3H_2(g) \rightleftharpoons 2NH_3(g)$

D. $N_2(g)+O_2(g) \rightleftharpoons 2NO(g)$

Answer: B

Solution:

Le-Chatelier principle is not applicable to pure solids and liquids because of the reason that they experience negligible change in concentration during chemical equilibrium which means that adding or removing a solid form system at equilibrium has no effect on equilibrium.

$Fe(s)+S(s) \rightleftharpoons FeS(s)$ is in solid equilibrium phase, hence Le-Chatelier principle is not applicable.

Question24

Using the data provided, find the value of equilibrium constant for the following reaction at $298\ K$ and $1\ atm$ pressure.

$$NO(g)+\frac{1}{2}O_2(g) \rightleftharpoons NO_2(g)$$

$$\Delta_f HY[NO(g)] = 90.4\ kJ\ mol^{-1}$$

$$\Delta_f HY\ [NO_2(g)] = 32.48\ kJ\ mol^{-1}$$

$$\Delta SY at\ 298\ K = -70.8\ JK^{-1}\ mol^{-1}$$

$$[\text{antilog}(0.50) = 3162\ ]$$

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Options:

A. 3.162×10^4

B. 3.162×10^{-4}

C. 3.162×10^6

D. 3.162×10^7

Answer: C

Solution:

$\mathrm{NO}(g) + \frac{1}{2}\mathrm{O}_2(g) \rightleftharpoons \mathrm{NO}_2(g)$

Given, $\Delta_f H^\circ[\mathrm{NO}(g)] = 90.4 \mathrm{~kJ~mol}^{-1}$

$$\begin{aligned}
&\Delta_f H^\circ\,[\mathrm{NO}_2(g)] = 32.48 \mathrm{~kJ~mol}^{-1} \\
&\Delta H^\circ = \Sigma \Delta H^\circ \text{ product } - \Sigma \Delta_f H^\circ(\mathrm{NO}) \\
&\ = 3248 - 90.4 \\
&\ = -57.92 \mathrm{~kJ~mol}^{-1}
\end{aligned}$$

We know that,

Standard Gibb's free energy, $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$$\begin{aligned}
&= -57.92 - 298 \times (-70.8) \\
= -57.92 + 21.098 &= -36.82 \mathrm{~kJ/mol} \\
&= -38.82 \times 10^3 \mathrm{~J/mol}
\end{aligned}$$

We also know, $\Delta G^\circ = -2.303 RT \log K_{\mathrm{eq}}$

Here, $K_{\mathrm{eq}} =$ Equilibrium constant

$$\begin{aligned}
-36.82 \times 10^3 &= -2303 \times 8.314 \times 298 \times \log K_{\mathrm{eq}} \\
\log K_{\mathrm{eq}} &= \frac{36.82}{5.70584} \\
\log K_{\mathrm{eq}} &= 0.501 \\
K_{\mathrm{eq}} &= \text{ antilog } 0.501 \\
&= 3.162 \times 10^6
\end{aligned}$$

Question25

Standard entropies of X_2, Y_2 and XY_3 are 60, 40 and $50 \mathrm{JK}^{-1} \mathrm{~mol}^{-1}$ respectively. At what temperature, the following reaction will be at equilibrium? [given: $\Delta H \Upsilon=-30 \mathrm{~kJ}$]

$$\frac{1}{2} X_2+\frac{3}{2} Y_2 \rightleftharpoons X Y_3$$

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Options:

A. 500 K

B. 750 K

C. 1000 K

D. 1250 K

Answer: B

Solution:

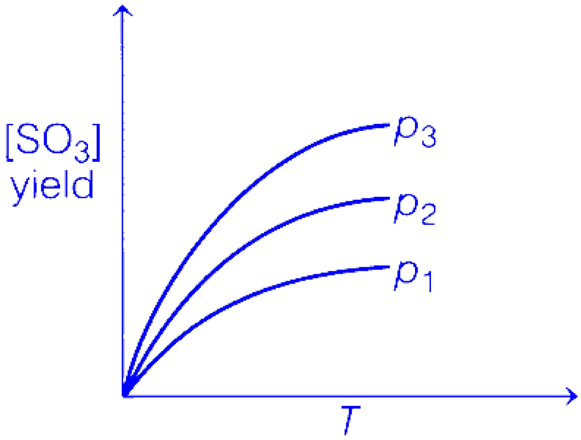
$$\begin{aligned}
& \frac{1}{2} X_2+\frac{3}{2} Y_2 \rightleftharpoons X Y_3 ; \Delta H=-30 \mathrm{~kJ} \\
& \Delta S_{\text {reaction }} \quad=\sum_{i=1} \Delta S_{\text {product }} \quad-\sum \Delta S_{\text {reactant }} \\
& X_2+3 Y_2 \rightleftharpoons 2 X Y_3 \\
& \Delta S_{\text {reaction }} \quad=2 \times 50-3 \times 40-1 \times 60 \\
& =-80 \mathrm{JK}^{-1} \mathrm{~mol}^{-1}
\end{aligned}$$

Using Gibb's free energy, $\Delta G=\Delta H-T \Delta S$

$$\begin{aligned}
\Delta G & =0 \\
0 & =\Delta H-T \Delta S \Rightarrow \Delta H=T \Delta S \\
1000 \times(-60) & =T \times(-80) \Rightarrow T=750 \mathrm{~K} .
\end{aligned}$$

Question26

For the reaction $\mathrm{SO_2}(g) + \frac{1}{2}\mathrm{O_2}(g) \rightleftharpoons \mathrm{SO_3}(g)$, the percentage yield of product at different pressure is shown in the figure. Then, which among the following is true?



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Options:

A. Pressure has no effect

B. $p_1 < p_2 < p_3$

C. $p_1 > p_2 > p_3$

D. $p_1 = p_2 = p_3 \neq 0$

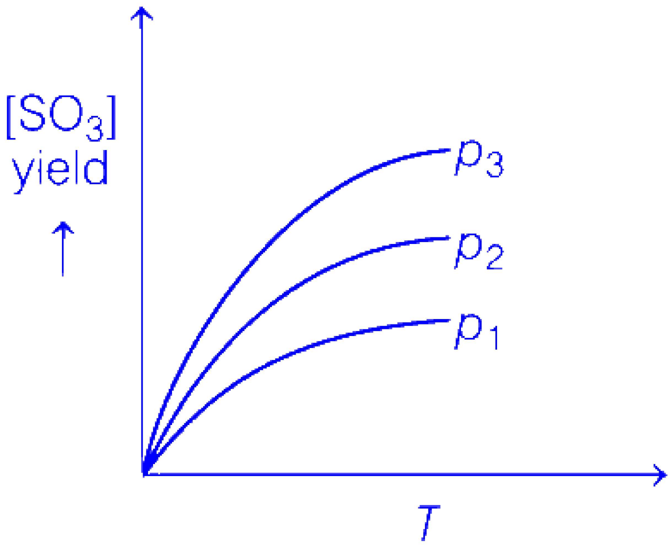
Answer: C

Solution:

$\mathrm{SO_2} + \frac{1}{2}\mathrm{O_2}(g) \rightleftharpoons \mathrm{SO_3}(g)$

The above equilibrium reaction is exothermic reaction. When pressure is lowered, the forward reaction is favoured.

On going from p_1 to p_3 yield of $\mathrm{SO_3}$ is increasing because lower pressure favours forward reaction.



Question27

Which among the following denotes the correct relationship between K_p and K_c for the reaction, $2A(g) \rightleftharpoons B(g) + C(g)$

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Options:

A. $K_p > K_c$

B. $K_c > K_p$

C. $K_c = (K_p)^2$

D. $K_p = K_c$

Answer: D

Solution:

The relationship between K_p and K_c is

$K_p = K_c(RT)^{\Delta n}$ ..... (i)

Here, Δn = (number of moles of product of gas) − (number of moles of reactant of gas)

K_p = equilibrium constant at constant pressure

K_c = equilibrium constant at constant concentration

R = rate constant

$2A(g) \rightleftharpoons B(g) + C(g)$

The value of Δn is $= 2 - 2 = 0$

Put value of Δn in Eq. (i) we get,

$K_p = K_C(RT)^0 \Rightarrow [K_p = K_C]$
