Topics to be covered

1. MEDICS TEST, Revision of Last Class

2. Numericals *& more numericals*

3. Degree of Association

4. Gibbs free energy, Simultaneous equilibrium

5. Magarmach Practice Questions (MPQ) & Home work from modules

Rules to Attend Class

1. Always sit in a peaceful environment with headphone and be ready with your copy and pen.

2. Never ever attend a class from in between or don't join a live class in the middle of the chapter.

3. Make sure to revise the last class before attending the next class & always complete your Magarmach Practice Questions.

4. Never ever engage in chat whether live or recorded on the topic which is not being discussed in current class as by doing so u can be blocked by the admin team or your subscription can be cancelled.

Rules to Attend Class

5. Try to make maximum notes during the class if something is left then u can use the notes pdf after the class to complete the remaining class.

6. Always ask your doubts in doubt section to get answer from faculty. Before asking any doubt please check whether same doubt has been asked by someone or not.

The internal energy of an ideal gas increases during an isothermal process when the gas is

$U = f(T, \cancel{V})$

A Expanded by adding more molecules to it.

B Expanded by adding more heat to it.

C Expanded against zero pressure.

D Compressed by doing work on it.

Maximum work can a gas do, if it is allowed to expand isothermally against

$$w = -P_{ext.} \Delta v$$

A Vacuum.

B High pressure of surrounding.

C Low pressure of surrounding.

D Atmospheric pressure.

If a closed system has adiabatic boundaries, then at least one boundary must be

$$\downarrow\uparrow\downarrow \quad q = 0$$

(A) Permeable

(B) Imaginary

(C) Movable ✓

(D) Fixed

QUESTION

$\uparrow^{Ar}$

The work done in the isothermal reversible expansion of argon gas at 27°C from 4L to 16 L was equal to 4200 cal. What is the amount of argon subjected to such an expansion? (Ar = 40, In4 = 1.4)

$W_{Ar} = ?$

$T = 300 K.$ $W = 4200 \, cal.$

A 5.0 g

$V_1 = 4L$

B 20.0 g

$V_g = 16 L.$

C 200.0 g

D 48.1 g

$$W = -nRT \ln \frac{V_2}{V_1}$$

$$4200 = n \times 2 \times 300 \left(\ln \frac{16}{4} \right)^4$$

$$n = \frac{7}{1.4} = \frac{70}{14} = 5 = \frac{W}{40}$$

$$W = 200 \, g.$$

The molar heat capacity of water in equilibrium with ice at constant pressure is

(A) Zero

(B) Infinity

(C) 40.45 kJ/K-mol

(D) 75.48 J/K-mol

$$C_{p\,m\,(H_2O)} = ?$$

$$H_2O(l) \rightleftharpoons H_2O(s)$$

$$\Delta H = \boxed{n} C_{p,m} \Delta T$$

$$q = m\,\textcircled{s}\,\Delta T$$

$$s = \infty$$
$$= s \times u$$
$$C_{p,m}$$

$$n = 1$$
$$\Delta T = 0$$
$$\Delta H = 0 \checkmark$$
$$C_{p,m} = \frac{\Delta H}{n\,\Delta T}\checkmark$$
$$C_{p,m} = \frac{\Delta H}{n\,0}$$
$$\infty = C_{p,m}$$

The latent heat of vaporization of a liquid at 500 K and 1 atm pressure is 10 kcal/mol. What will be the change in internal energy if 3 moles of the liquid changes to vapour state at the same temperature and pressure?

A   27 kcal

B   13 kcal

C   –27 kcal

D   –13 kcal

$T = 500 K$

$P = 1 atm.$

$l_{vap.} = \Delta H_{vap.} = 10 K \, Cal / mol.$

$\Delta U_{vap} = \Delta H - \Delta n_g R T$

$= 10000^{\times 3} - 3 \times 2 \times 500$

over 1000

$= \dfrac{30000 - 3000}{1000} =$

$3 \, mole \, (l) \Rightarrow 3 \, mole \, (g)$

$\Delta n_g = 3 - 0 = 3$

$\dfrac{27000}{1000} = 27 K Cal$
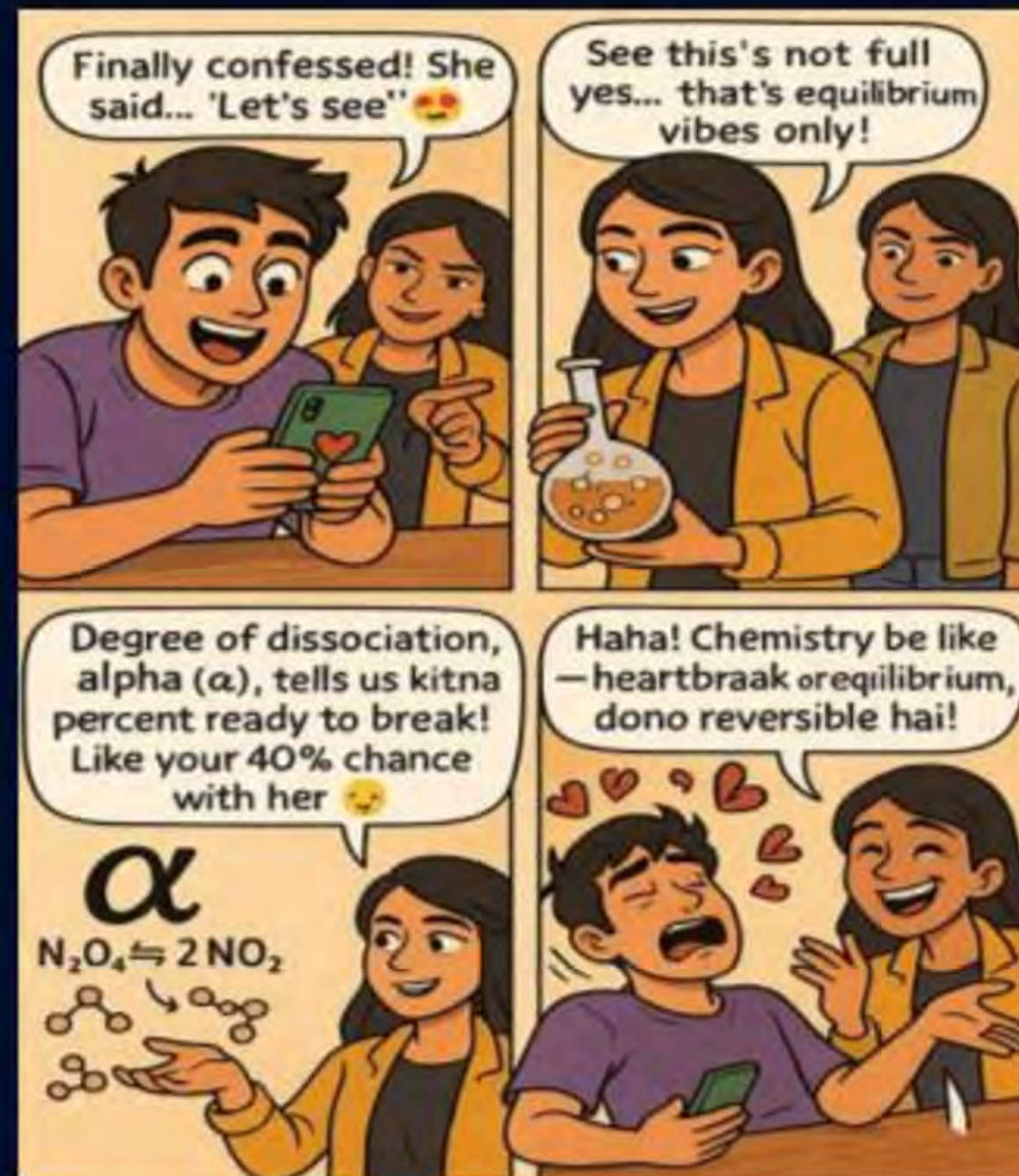
Lec-6 to Lec-10.
↓
Thermodynamics → Tomorrow

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

$$\alpha = \frac{D-d}{d(n-1)} = \frac{M_t - M_0}{M_0(n-1)}$$

-/.age dissociation $= \alpha \times 100$

The reaction $2A(g) + B(g) \rightleftharpoons 3C(g) + D(g)$ is began with the concentrations of A and B both at an initial value of 1.00 M. When equilibrium is reached, the concentration of D is measured and found to be 0.25 M. The value for the equilibrium constant for this reaction is given by the expression

A $\quad [(0.75)^3(0.25)] \div [(0.75)^2(0.25)]$

B $\quad [(0.75)^3(0.25)] \div [(1.00)^2(1.00)]$

C $\quad [(0.75)^3(0.25)] \div [(0.50)^2(0.75)]$

D $\quad [(0.75)^3(0.25)] \div [(0.50)^2(0.25)]$

$$2A(g) + 1B(g) \rightleftharpoons 3C(g) + 1D(g)$$

$$[D]_{eq} = 0.25M$$

$t = 0$ 1 1 0 0

$[D]_{eq} = \dfrac{x}{2} = 0.25$

$t = t_{eq.}$ $1-x$ $1-\dfrac{x}{2}$ $\dfrac{3x}{2}$ $\dfrac{x}{2}$

$1-0.5=0.5$ $1-0.25=0.75$ $\dfrac{3}{2}\times 0.25$ 0.25

                          3×0.25

                          0.75

$\Rightarrow x = 0.5$

$A \quad B$

$2 \to 1 \qquad 2 \to 3$

$1 \to \dfrac{1}{2} \qquad 1 \to \dfrac{3}{2}$

$x \to \dfrac{x}{2} \qquad x \to \dfrac{3x}{2}$

$$K_C = \frac{[C]^3 [D]^1}{[A]^2 [B]^1} = \frac{(0.75)^3 (0.25)^1}{(0.5)^2 (0.75)^1}$$

PCl_5, PCl_3 and Cl_2 are at equilibrium at 500K in a closed container and their concentrations are 0.8×10^{-3} mol L^{-1}, 1.2×10^{-3} mol L^{-1} and 1.2×10^{-3} mol L^{-1} respectively. The value of K_c for the reaction PCl_5 (g) $\rightleftharpoons$ PCl_3 (g) + Cl_2 (g) will be

(A) 1.8×10^{-3} mol L^{-1}

(B) 1.8×10^{-3}

(C) 1.8×10^{-3} L mol^{-1}

(D) 0.55×10^4

$T = 500K.$

$$K_C = \frac{[PCl_3][Cl_2]}{[PCl_5]}$$

$K_c \text{ unit } (mol L^{-1})^{\Delta ng}$

$= mol L^{-1}$

$= \frac{1.2 \times 10^{-3} \times 1.2 \times 10^{-3}}{0.8 \times 10^{-3}}$

$= 18 \times 10^{-4}$

$3O_2(g) \rightleftharpoons 2O_3(g)$ for the given reaction at 298 K, K_c is found to be 3.0×10^{-59}. If the concentration of O_2 at equilibrium is 0.040 M, then concentration of O_3 in M is:

A) 1.20×10^{21}

B) 4.38×10^{-32}

C) 1.90×10^{-63}

D) 2.40×10^{31}

$$K_c = \frac{[O_3]^2}{[O_2]^3} = 3 \times 10^{-59}$$

$$[O_3]^2 = 3 \times 10^{-59} \times (0.04)^3 \qquad 4^2 = 16$$

$$= 192 \times 10^{-65} \qquad 5^2 = 25$$

$$[O_3] = \sqrt{19.2 \times 10^{-64}}$$

$$= 4.38 \times 10^{-32}$$

$PC_5(g) \rightleftharpoons PCl_3(g) + (Cl_2)$

Ⓟ Ⓦ

Consider the following chemical equilibrium of the gas phase reaction at a constant temperature:

$A(g) \rightleftharpoons B(g) + C(g)$; If p being the total pressure, K_p is the pressure equilibrium constant and a is the degree of dissociation, then which of the following is true at equilibrium? $\alpha = a$

$$\alpha \propto P^{\frac{-\Delta n_g}{2 (g)}}$$

$$\alpha \propto P^{\frac{-1}{2}}$$

$$\sqrt{\alpha} \propto \frac{1}{P^{\frac{1}{2}} \cdot T}$$

A If p value is extremely high compared to K_p, $\alpha \approx 1$

B ✓ When p increases α decreases

C If K_p value is extremely high compared to p, α becomes much less than unity

D When p increases α increases

At temperature T, compound $AB_{2(g)}$ dissociates as $AB_{2(g)} \rightleftharpoons AB_{(g)} + \dfrac{1}{2}B_{2(g)}$ having degree of dissociation x (small compared to unity). The correct expression for x in terms K_p and p is

$x <<< 1$

$\alpha <<< 1$

(A) $\sqrt[3]{\dfrac{2K_p}{p}}$

(B) $\sqrt[4]{\dfrac{2K_p}{p}}$

(C) $\sqrt[3]{\dfrac{2K_p^2}{p}}$

(D) $\sqrt{K_p}$

$$1\,AB_2(g) \rightleftharpoons 1\,AB(g) + \frac{1}{2}B_2(g)$$

$$x = c\alpha.$$
$$x <<< 1$$
$$1-x \approx 1 \quad , 1+x \approx 1$$

$t=0$ $1,$ 0 0

$t=t$ $1-x$ $+$ $\underline{x}$ $\rightarrow$ $\frac{x}{2}$

$$P_{AB_2} = x_{AB_2} \times P$$

$$= \frac{1-x}{1+0.5x} \times P$$

$$P_{AB} = \frac{x}{1+0.5x} \times P$$

$$P_{B_2} = \frac{x}{2(1+0.5x)} \times P$$

$$K_P = \frac{\left(\dfrac{x}{1+0.5x}\right)^1 P^1 \times \left(\dfrac{x}{2(1+0.5x)}\right)^{1/2} \times P^{1/2}}{\left(\dfrac{1-x}{1+0.5x}\right)^1 P^1}$$

$$= \frac{x^{3/2} \times P^{1/2}}{(2(1+0.5x))^{1/2}(1-x)^1}$$

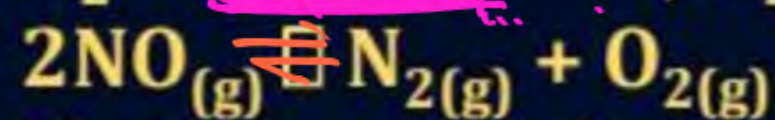$$K_P = \frac{x^{3/2} \times P^{1/2}}{(2)^{1/2}}$$

$$x^{3/2} = \frac{K_p \times 2^{1/2}}{p^{1/2}}$$

$$x = \frac{(K_p)^{2/3} \times \left[(2)^{1/2}\right]^{2/3}}{(p^{1/2})^{2/3}}$$

Consider the following reaction in a sealed vessel at equilibrium with concentrations of
$N_2 = 3.0 \times 10^{-3}$ M, $O_2 = 4.2 \times 10^{-3}$ M and NO $= 2.8 \times 10^{-3}$ M.
$$2NO_{(g)} \rightleftharpoons N_{2(g)} + O_{2(g)}$$
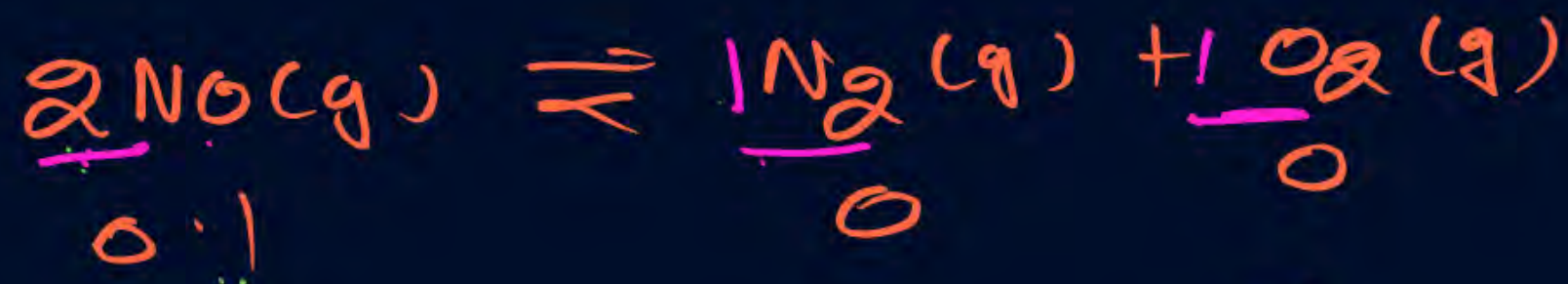If 0.1 mol L^{-1} of $NO_{(g)}$ is taken in a closed vessel, what will be degree of dissociation (α) of $NO_{(g)}$ at equilibrium?

A   0.00889

B   0.0889

C   0.8889

D   0.717

$$2NO(g) \rightleftharpoons 1N_2(g) + 1O_2(g)$$

$$0.1 \qquad\qquad 0 \qquad\qquad 0$$

$$0.1-x$$

$$\frac{x}{2} \qquad\qquad \frac{x}{2}$$

$$\boxed{0.1-0.1\alpha} \qquad \frac{0.1\alpha}{2} \qquad \frac{0.1\alpha}{2}$$

$$\boxed{2.8 \times 10^{-3}} = 0.1 - 0.1\alpha$$

$$C = 0.1$$

$$\alpha = \frac{x}{0}$$

$$x = 0.1\alpha.$$

$$K_C = \frac{[N_2]^1 [O_2]^1}{[NO]^2}$$

$$2.8 \times 10^{-3} = 0.1(1-\alpha)$$

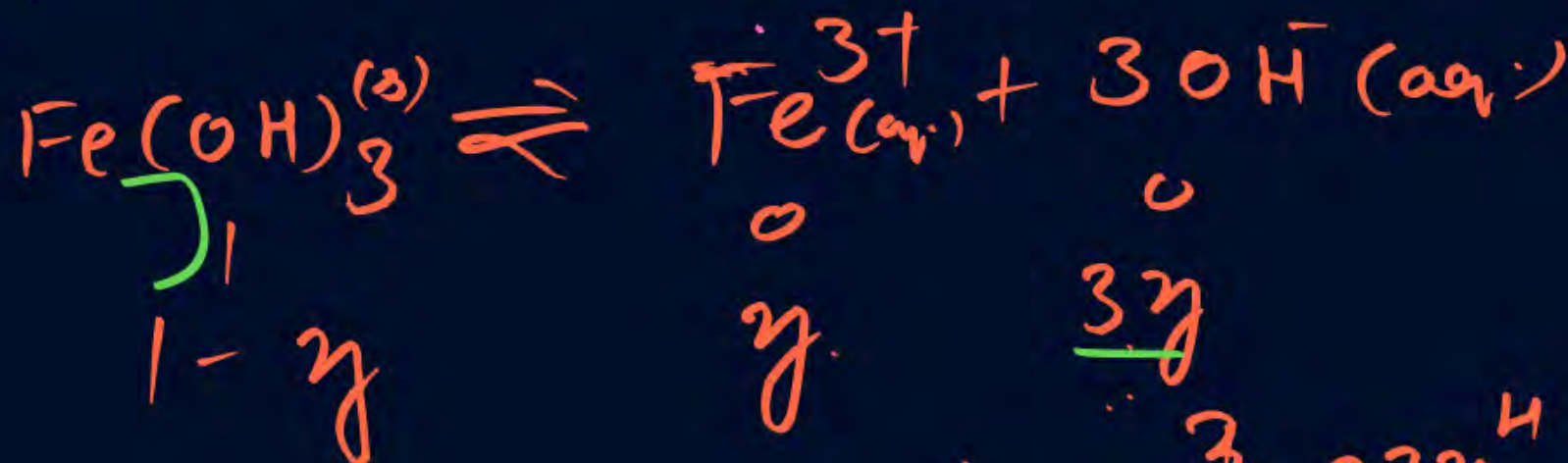$$\alpha = -28 \times 10^{-3} + 1$$

$$= -0.028 + 1$$

$$\alpha = 0.972$$

K_p or K_c will not change unless T change

If the concentration of OH^- ions in the reaction $Fe(OH)_3(s) \rightleftharpoons Fe^{3+}(aq.) + 3OH^-$ is decreased by $\frac{1}{4}$ times, then equilibrium concentration of Fe^{3+} will increase by

$[OH^-]_{old} = x$ $[OH^-]_{new} = x - \frac{x}{4} = \frac{3}{4}x$

(A) 8 times

(B) 16 times

(C) 64 times

(D) 4 times

$$Fe(OH)_3^{(s)} \rightleftharpoons Fe^{3+}_{(aq.)} + 3OH^-_{(aq.)}$$

$$1 \qquad\qquad 0 \qquad\qquad 0$$

$$1-y \qquad\qquad y \qquad\qquad \frac{3y}{3}$$

$$K_c = (y)^1 (3y)^3 = 27y^4$$

$$3y \times \frac{1}{4}$$

$$\frac{3y}{4}$$

$$Fe(OH)_3(s) \rightleftharpoons Fe^{3+}(aq) + 3OH^-(aq)$$

$$[Fe^{3+}]_{new} \quad \frac{3y}{4}$$

$$27y^4 = [Fe^{3+}]_{new} \left(\frac{3y}{4}\right)^3$$

$$27y^4 = [Fe^{3+}]_{new} \frac{27y^3}{64}$$

$$64y = [Fe^{3+}]_{new}$$

Ⓟ Ⓦ

$A + 2B \rightleftharpoons 2C; \quad K = ?$

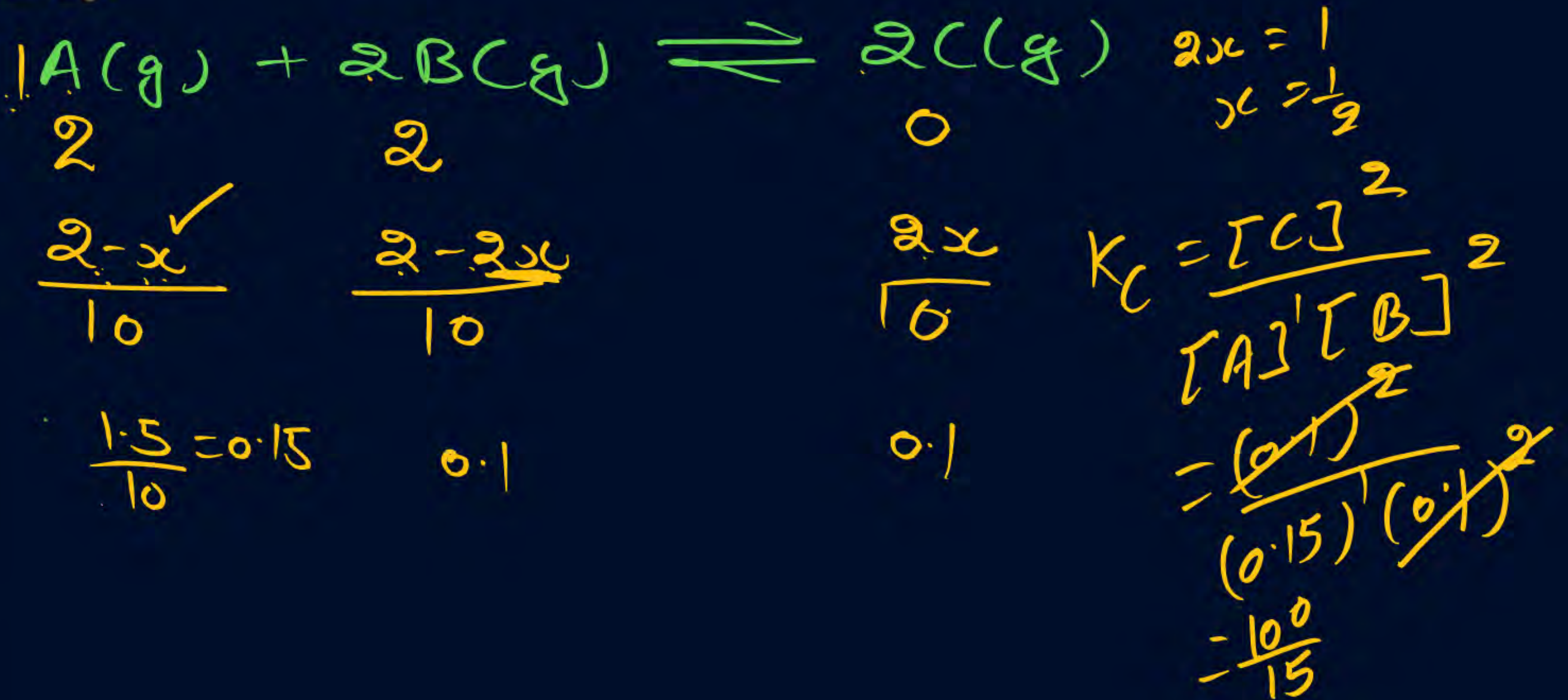2 moles of each A and B are present in 10 lit solution. The product C formed 1 mole. Calculate K_c.

(A) 1.5

(B) 6.67

(C) 0.15

(D) 2.3

$$1 A(g) + 2B(g) \rightleftharpoons 2C(g) \qquad 2x = 1$$

$$\qquad\qquad x = \tfrac{1}{2}$$

$$2 \qquad\qquad 2 \qquad\qquad 0$$

$$\frac{2-x}{10} \qquad \frac{2-2x}{10} \qquad \frac{2x}{10}$$

$$\frac{1.5}{10} = 0.15 \qquad 0.1 \qquad\qquad 0.1$$

$$K_c = \frac{[C]^2}{[A]^1 [B]^2} = \frac{(0.1)^2}{(0.15)^1 (0.1)^2} = \frac{100}{15}$$
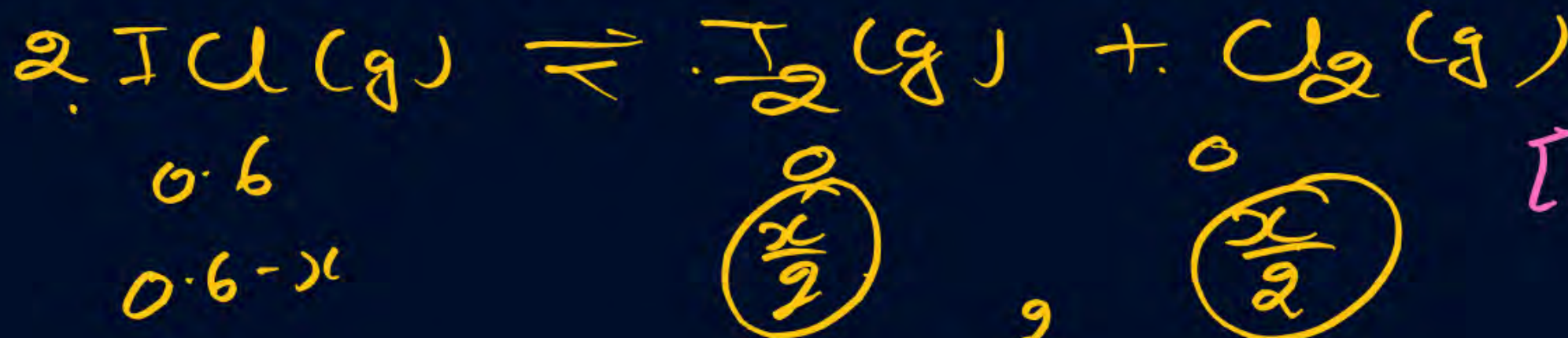
$$2ICl \rightarrow I_2 + Cl_2; K_c = 0.14$$
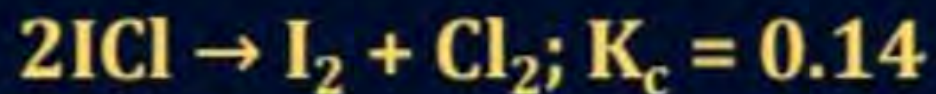
If initial concentration of ICl is 0.6 M, then equilibrium concentration of I_2 is:

(A) 0.37 M

(B) 0.126 M

(C) 0.224 M

(D) 0.748 M

Handwritten working:

$3^2 = 9$ (14)

$9 = 16$ $4^2 = 16$

$$2 ICl (g) \rightleftharpoons I_2 (g) + Cl_2 (g)$$

$$0.6$$

$$0.6 - x \qquad \frac{x}{2} \qquad \frac{x}{2}$$

$$[I_2] = \frac{x}{2}$$

$$K_c = 0.14 = \frac{\left(\frac{x}{2}\right)^2}{(0.6-x)^2}$$

$$\sqrt{0.14} = \frac{x}{2(0.6-x)}$$

$$\left(\frac{14}{100}\right)^{1/2}$$

$$\frac{3.8}{10} = \frac{x}{1.2 - 2x}$$

Consider the reaction $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ for which $K_c = 278$ M^{-1}. 0.001 mole of each reagents $SO_2(g)$, $O_2(g)$ and $SO_3(g)$ are mixed in a 1.0 L flask. Determine the reaction quotient of the system and the spontaneous direction of the system ~~and the spontaneous direction of the system.~~

A $Q_c = 1000$, equilibrium shifts to the right

B $Q_c = 1000$, equilibrium shifts to the left

C $Q_c = 0.001$, equilibrium shifts to the left

D $Q_c = 0.001$, equilibrium shifts to the right

$$Q_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

$$Q_c = \frac{1000}{0.601} = 1000.0$$

$$Q_c > K_c$$

The equilibrium constant K_c for the reaction:
$P_4(g) \rightleftharpoons 2P_2(g)$ is 1.4 at 400°C. Suppose that 3 moles of $P_4(g)$ and 2 moles of $P_2(g)$ are mixed in 2 litre container at 400°C. What is the value of reaction quotient (Q_c)?

$K_c = 1.4$

$$Q_c = \frac{[P_2]^2}{[P_4]} = \left(\frac{2}{2}\right)^2 \left(\frac{2}{3}\right)^1 = \frac{2}{3} < 1$$

$Q_c < K_c$

A) 3/2

B) 2/3

C) 1

D) None of these

The figure shows the change in concentration of species A and B as a function of time. The equilibrium constant K_c for the reaction $A(g) \rightleftharpoons 2B(g)$ is:

(A) $K_c > 1$

(B) $K < 1$

(C) $K = 1$

(D) Data insufficient

$$K_c = \frac{[B]^2}{[A]^1} = \frac{(0.4)^2}{0.1}$$

$$= \frac{0.16}{0.10} = 1.6$$
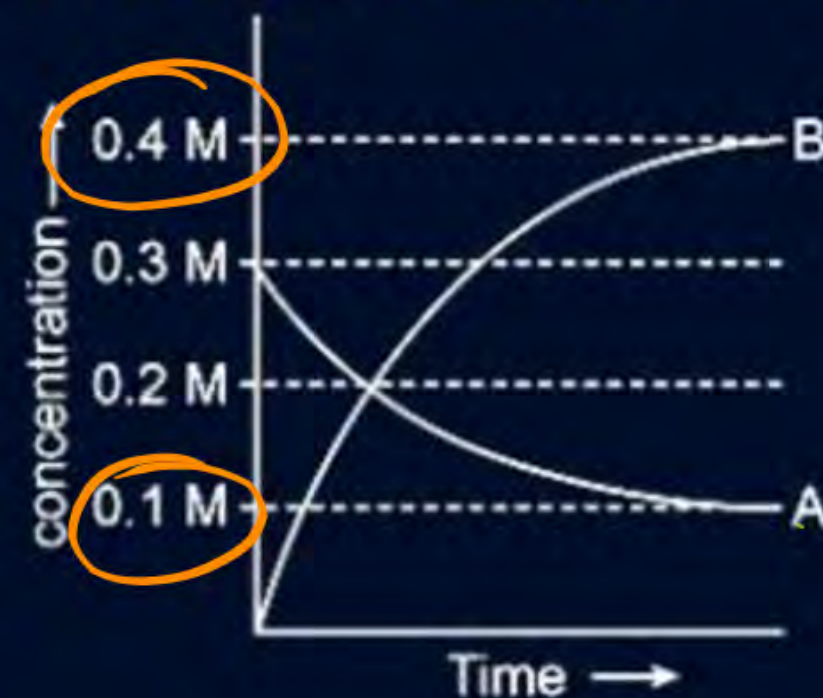
9.2 grams of $N_2O_4(g)$ is taken in a closed one litre vessel and heated till the following equilibrium is reached $N_2O_{4(g)} \rightleftharpoons 2NO_{2(g)}$.

At equilibrium, 50% $N_2O_{4(g)}$ is dissociated. What is the equilibrium constant (in mol litre^{-1})(Molecular mass of $N_2O_4 = 92$)

(A) 0.1

(B) 0.4

(C) 0.2

(D) 2

$V = 1L, \quad W_{N_2O_4} = 9.2g$

$M_{N_2O_4} = 92g$

$n_{N_2O_4} = \frac{9.2}{92} = 0.1 = C$

$\alpha = 50\% \text{ of } C = \frac{50}{100} \times 0.1 = 0.05$

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

$0.1 \quad\quad\quad 0$

$0.1 - x \quad\quad 2x$

$0.1 - 0.05 \quad \frac{2 \times 0.05}{1}$

$K_C = \frac{(0.1)^2}{(0.05)} = \frac{0.1 \times 0.1}{0.05}$

$= \frac{1}{5} = 0.2$
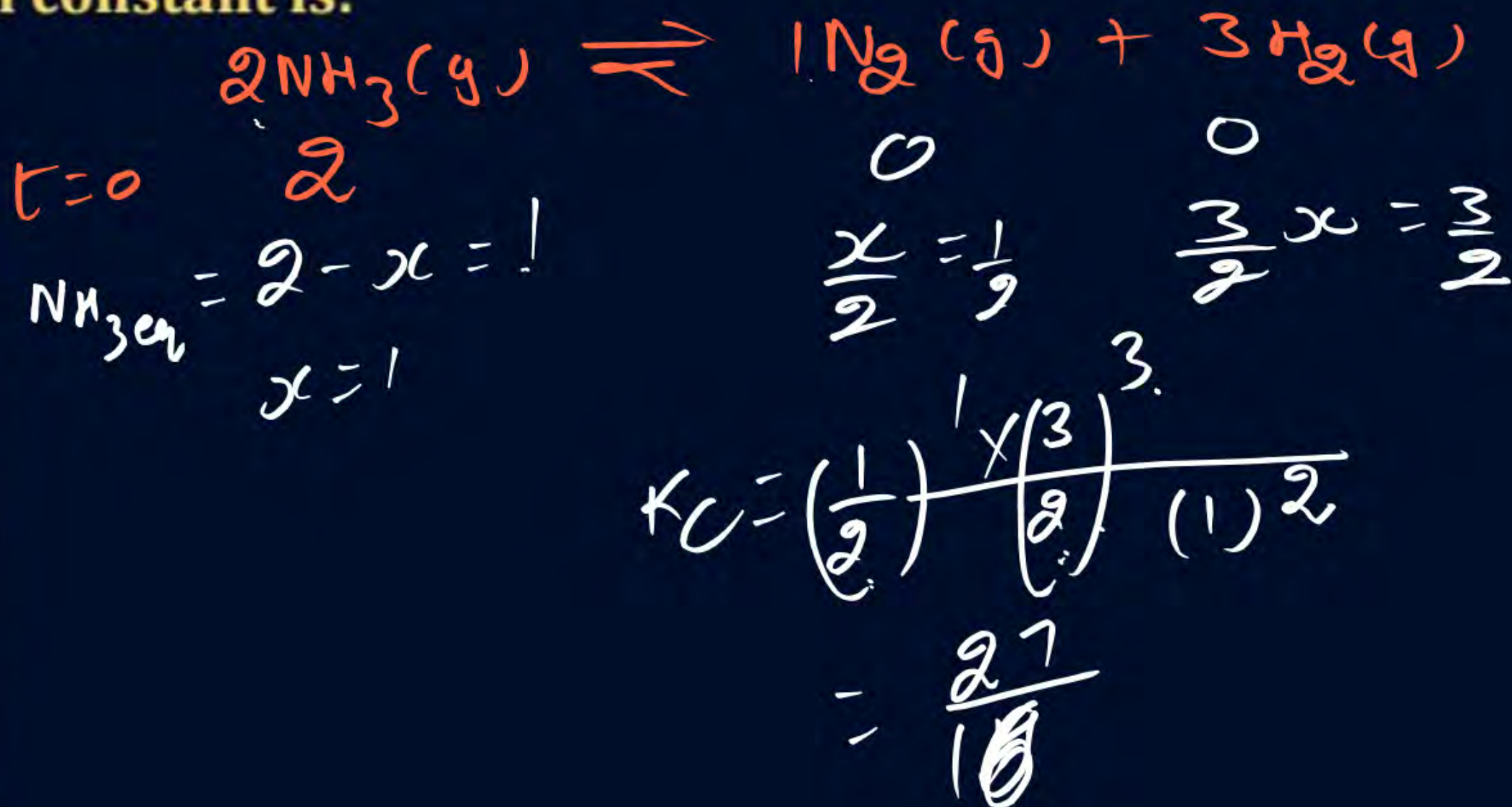
$NH_{3eq} = 1$

$V = 1L$

Two moles of NH_3 when put into a previously evacuated vessel (one litre), partially dissociated into N_2 and H_2. If at equilibrium one mole of NH_3 is present, the equilibrium constant is:

$$2NH_3(g) \rightleftharpoons 1N_2(g) + 3H_2(g)$$

A) $3/4 \ mol^2 \ litre^{-2}$

B) $27/64 \ mol^2 \ litre^{-2}$

C) $27/32 \ mol^2 \ litre^{-2}$

D) $27/16 \ mol^2 \ litre^{-2}$ ✓

$t=0$ 2   0   0

$NH_{3eq} = 2 - x = 1$

$x = 1$

$\dfrac{x}{2} = \dfrac{1}{2}$ $\dfrac{3}{2}x = \dfrac{3}{2}$

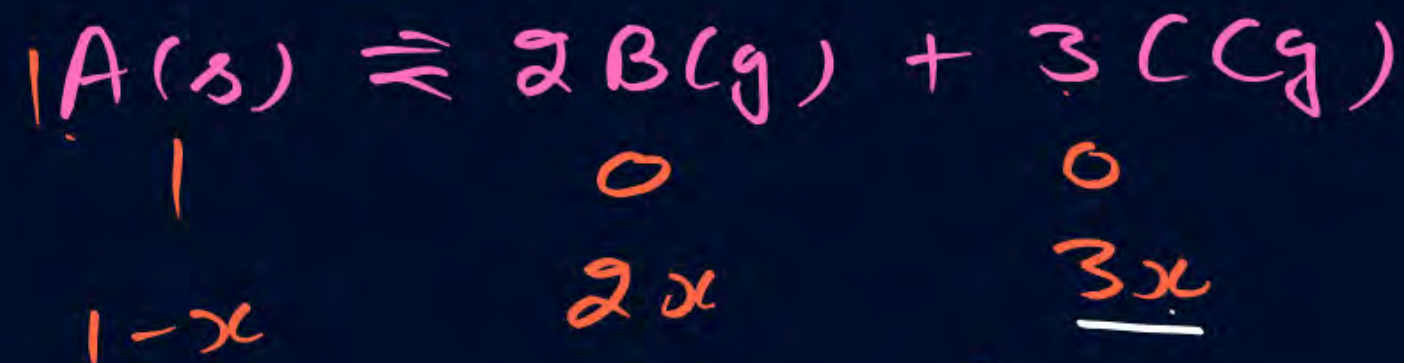$$K_C = \dfrac{\left(\dfrac{1}{2}\right)^1 \times \left(\dfrac{3}{2}\right)^3}{(1)^2}$$

$$= \dfrac{27}{16}$$

In a system $A(s) \rightleftharpoons 2B(g) + 3C(g)$, if the concentration of C at equilibrium is increased by a factor of 2, it will cause the equilibrium concentration of B to change to:

(A) Two times the original value

(B) One half of its original value

(C) $2\sqrt{2}$ times to the original value

(D) $1/2\sqrt{2}$ times the original value

$$A(s) \rightleftharpoons 2B(g) + 3C(g)$$

$$\begin{array}{cccc} & 1 & 0 & 0 \\ & 1-x & 2x & 3x \end{array}$$

$$K_C = [B]^2 [C]^3$$
$$= (2x)^2 (3x)^3 = 108x^5$$

$$[C]_{new} = 6x$$
$$[B]_{new} = ?$$

$$108x^5 = [B]^2_{new} (6x)^3$$

$$\frac{108x^5}{216x^3} = [B]^2_{new}$$
$$\;\;\;\;\;2$$

$$[B]^2_{new} = \frac{x^2}{2}$$

$$[B]_{new} = \frac{x}{\sqrt{2}}$$

$$[B]_{old} = 2x$$

$$[B]_{new} = \frac{x}{\sqrt{2}}$$

$$\frac{2x}{2\sqrt{2}}$$

For A + B $\rightleftharpoons$ C + D, if initially the concentration of A and B are both equal but at equilibrium concentration of D will be twice of that of A then what will be the equilibrium constant of reaction?

$$[A] = [B]$$

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

(A) 4/9

$$[D]_{eq} = 2[A]_{eq}$$

(B) 9/4

$$x = 2(1-x)$$

$$x = 2 - 2x$$

(C) 1/9

$$3x = 2$$

$$x = \frac{2}{3}$$

(D) 4

1	1	0	0
$1-x$	$1-x$	x	x
$1-\frac{2}{3}=\frac{1}{3}$	$1-\frac{2}{3}=\frac{1}{3}$	$\frac{2}{3}$	$\frac{2}{3}$

$$K_c = \frac{\left(\frac{2}{3}\right)^2\left(\frac{3}{1}\right)^2}{} = 4$$

The equilibrium constant K_c for the reaction
$SO_2(g) + NO_2(g) \rightleftharpoons SO_3(g) + NO(g)$ is 16. If 1 mole of each of all the four gases is taken in 1 dm^3 vessel, the equilibrium concentration of NO would be:
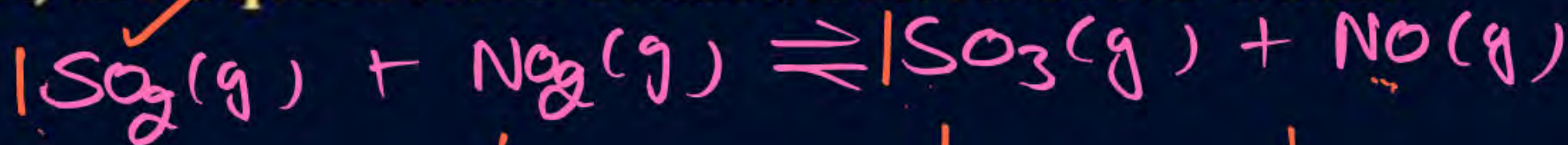
$K_c = 16$

$1 SO_2(g) + NO_2(g) \rightleftharpoons 1 SO_3(g) + NO(g)$

(A) 0.4 M

$V = 1L \qquad t = 0$

$\dfrac{1-x}{1} \qquad \dfrac{1-x}{1} \qquad \dfrac{1+x}{1} \qquad \dfrac{1+x}{1}$

(B) 0.6 M

$1 + x = 4 - 4x$

(C) 1.4 M

$4x + x = 4 - 1 = 3$

$5x = 3$

$x = \dfrac{3}{5}$

(D) 1.6 M

$[NO] = 1 + x$
$= 1 + \dfrac{3}{5} = 1.6 M$

$K_c = \dfrac{(1+x)^2}{(1-x)^2} = 16$

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

At a certain temperature, only 50% HI is dissociated at equilibrium in the following reaction: $2\underline{HI}(g) \rightleftharpoons H_2(g) + I_2(g)$ $x = \dfrac{50}{100} \times \dfrac{2}{2} = 1$

The equilibrium constant for this reaction is:

(A) 0.25

(B) 1.0

(C) 3.0

(D) 0.5

$$2 HI(g) \rightleftharpoons 1 H_2(g) + 1 I_2(g)$$

$$
\begin{array}{cccc}
& 2 HI(g) & \rightleftharpoons & 1 H_2(g) + 1 I_2(g) \\
t=0 & 2 & & 0 \qquad\qquad 0 \\
t=t & 2-x & & \dfrac{x}{2} \qquad \dfrac{x}{2} \\
& 2-1=1 & & 0.5 \qquad 0.5
\end{array}
$$
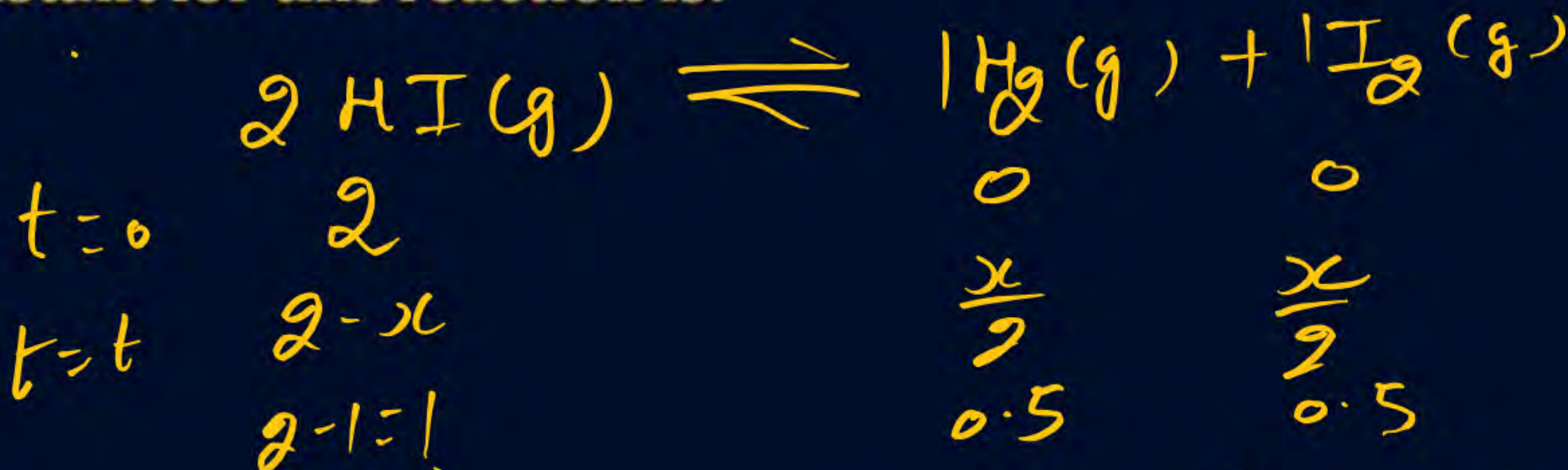
$$K_C = \dfrac{(0.5)^2}{(1)^2} = 0.25$$

The equilibrium constant K_p for the reaction

$$H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g)$$

is 4.0 at 1660°C. Initially 0.80 mole H_2 and 0.80 mole CO_2 are injected into a 5.0 litre flask. What is the equilibrium concentration of $CO_2(g)$?

A) 0.533 M

B) 0.0534 M

C) 0.535 M

D) None of these

$K_p = K_C(RT)^{\Delta n g}$
$u = K_C(\frac{1}{12} \times 1933)^0$

$$H_2(g) + CO_2(g) \rightleftharpoons H_2O(g) + CO(g)$$

$\frac{0.8-x}{5}$ $\frac{0.8-x}{5}$ $\frac{x}{5}$ $\frac{x}{5}$
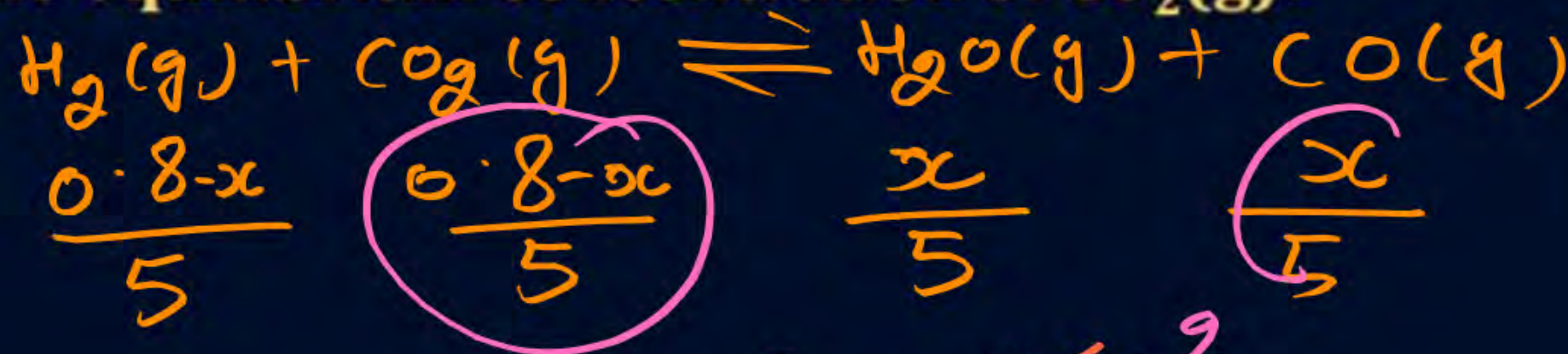
$$4 = \frac{(\frac{x}{5})^2 (\frac{5}{5})^2}{(0.8-x)} = \frac{x^2}{(0.8-x)^2} = 4$$

$$\frac{x}{0.8 - x} = 2$$

$$x = 1.6 - 2x$$

$$3x = 1.6$$

$$x = \frac{16}{30} = 0.53$$

$$[O_2] = \frac{0.8 - 0.53}{5} = 0.053$$

At 1000 K, a sample of pure NO_2 gas decomposes as

$$2NO_2(g) \rightleftharpoons 2NO(g) + O_2(g)$$

The equilibrium constant K_p is 156.25 atm. Analysis shows that the partial pressure of O_2 a 0.25 atm at equilibrium. The partial pressure of NO_2 at equilibrium is:

$P_{O_2} = 0.25\,atm$

$P_{NO} = 0.5\,atm.$

$156.25 = \dfrac{(0.5)^2(0.25)^1}{(P_{NO_2})^2}$

(A) 0.01

(B) 0.02

$(P_{NO_2})^2 = \dfrac{0.0625 \times 10^{-4}}{156.25}$

$(P_{NO_2})^2 = \dfrac{(0.5)^2 \times 0.25}{156.25}$

(C) 0.04

$P_{NO_2} = 2 \times 10^{-2} = 0.02\,atm.$

(D) None of these

For the equilibrium $A \rightleftharpoons B$, the variation of the rate of the forward (a) and reverse (b) reaction with time is given by:
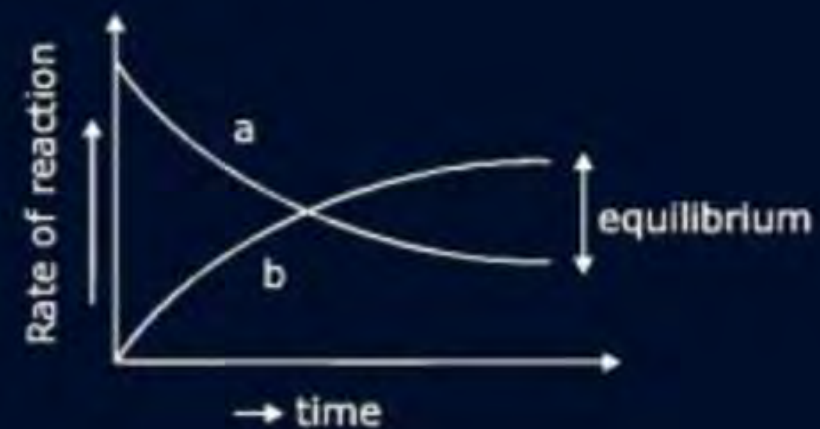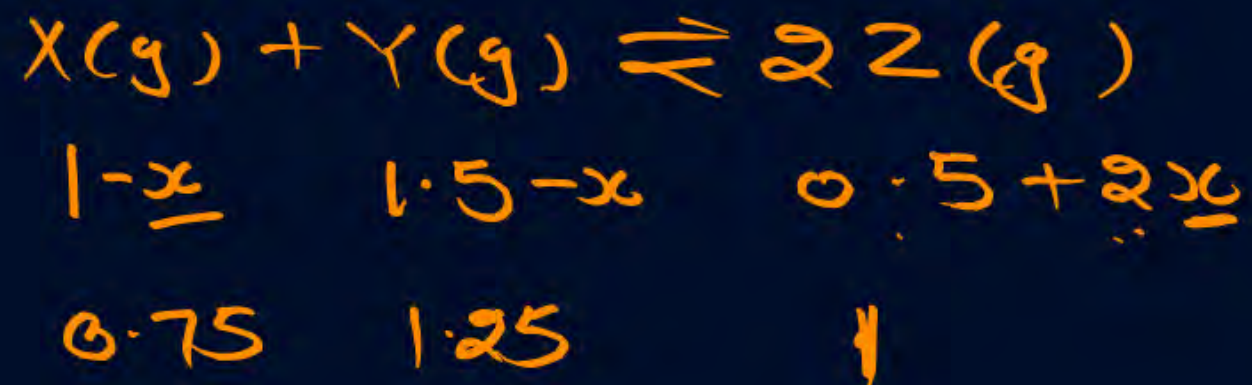
(A)

(B)

(C)

(D)

For a reaction, $X + Y \rightleftharpoons 2Z$; 1.0 mol of X, 1.5 mol of Y, and 0.5 mol of Z, were taken in a 1 L vessel and allowed to react. At equilibrium, the concentration of Z was 1.0 mol L^{-1}. The equilibrium constant of the ⓍⓆ reaction is $\frac{x}{15}$. The value of x is..........

$$X(g) + Y(g) \rightleftharpoons 2Z(g)$$

$$1-x \qquad 1.5-x \qquad 0.5+2x$$

$$0.75 \qquad 1.25 \qquad 1$$

$$K_c = \frac{(1)^2}{(0.75)^1 (1.25)^1} = X$$

$$[Z] = 0.5 + 2x = 1$$

$$2x = 0.5$$

$$x = 0.25$$

$$\frac{\cancel{2000}^{\cancel{400}^{16}}}{\cancel{10000}} = \frac{X}{15} = \frac{16}{15}$$

$$\frac{\cancel{75} \times \cancel{125}}{15 \quad 25}$$

$$X = 16$$

5.1 g NH_4SH is introduced in 3.0 L evacuated flask at 327°C. 30% of the solid NH_4SH decomposed to NH_3 and H_2S as gases. The K_p of the reaction at 327°C is: $(R = 0.082 \text{ atm mol}^{-1}K^{-1}$, molar mass of S = 32 g mol^{-1}, molar mass of N = 14 g mol$^{-1})$
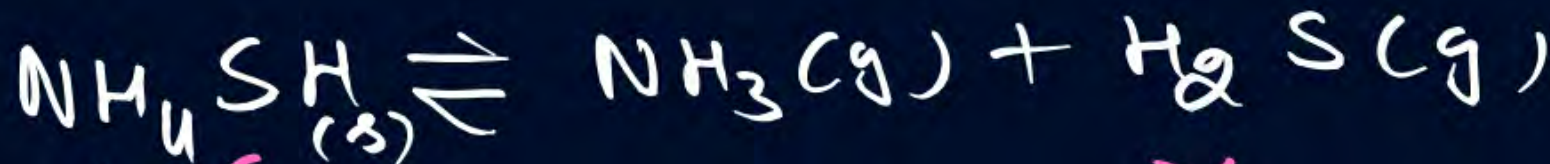
A) 0.242×10^{-4} atm^2

B) 0.242 atm^2

C) 4.9×10^{-3} atm^2

D) 1×10^{-4} atm^2

$V = 3L \qquad T = 600 K \qquad \alpha = \dfrac{30}{100} \times \dfrac{51}{51} \; 0.1 = 0.03$

$$NH_4SH_{(s)} \rightleftharpoons NH_3(g) + H_2S(g)$$

$$0 \quad x \qquad\qquad x \qquad\qquad x$$

$$\qquad\qquad\qquad\qquad \dfrac{0.03}{3} \qquad \dfrac{0.03}{3}$$

$$\qquad\qquad\qquad\qquad 0.01 \qquad 0.01$$

$$K_c = (0.01)^2 = 10^{-4}$$

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

$$= 10^{-4} \left(\frac{1}{12} \times \overset{50}{600} \right)^2$$

$$= 50 \times 50 \times 10^{-4}$$

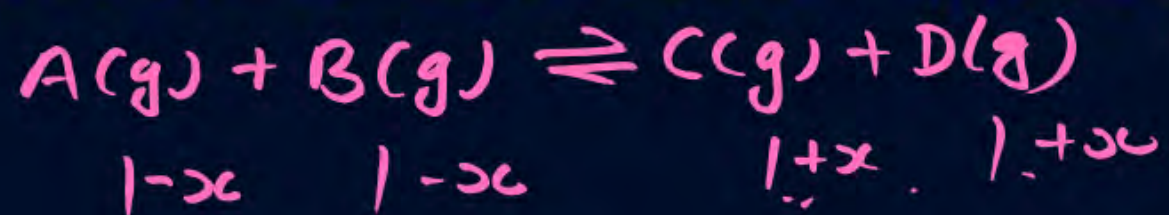$$= 25 \times 10^{-2}$$

$$= 0.25 \, atm^2$$

$K_C = 100$

The equilibrium constant at 298 K for a reaction, $A + B \rightleftharpoons C + D$ is 100. If the initial concentrations of all the four species were 1 M each, then equilibrium concentration of D (in mol L^{-1}) will be

A) 0.818

B) 1.818

C) 1.182

D) 0.182

$$A(g) + B(g) \rightleftharpoons C(g) + D(g)$$
$$1-x \qquad 1-x \qquad\qquad 1+x \qquad 1+x$$

$$K_C = \frac{(1+x)^2}{(1-x)^2} = 100$$

$$\frac{1+x}{1-x} = 10$$

$$1+x = 10 - 10x$$
$$11x = 9$$
$$x = \frac{9}{11}$$

$$[D] = 1 + \frac{9}{11} = \frac{20}{11}$$

THANK YOU