



NCERT Solutions for 11th Class Chemistry: Chapter 7-Equilibrium



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NCERT Solutions for 11th Class Chemistry: Chapter 7-Equilibrium

Class 11: Chemistry Chapter 7 solutions. Complete Class 11 Chemistry Chapter 7 Notes.

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Question 1. A liquid is in equilibrium with its vapours in a sealed container at a fixed temperature. The volume of the

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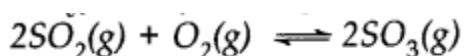
container is suddenly increased, (i) What is the initial effect of the change on the vapour pressure? (ii) How do the rates of evaporation and condensation change initially? (iii) What happens when equilibrium is restored finally and what will be the final vapour pressure?

Answer: (i) On increasing the volume of the container, the vapour pressure will initially decrease because the same amount of vapours are now distributed over a larger space.

(ii) On increasing the volume of the container, the rate of evaporation will increase initially because now more space is available. Since the amount of the vapours per unit volume decrease on increasing the volume, therefore, the rate of condensation will decrease initially.

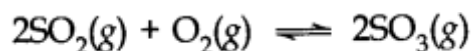
(iii) Finally, equilibrium will be restored when the rates of the forward and backward processes become equal. However, the vapour pressure will remain unchanged because it depends upon the temperature and not upon the volume of the container.

Question 2. What is K_c for the following reaction in state of equilibrium?



Given: $[\text{SO}_2] = 0.6 \text{ M}$; $[\text{O}_2] = 0.82 \text{ M}$; and $[\text{SO}_3] = 1.90 \text{ M}$

Answer:



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$$K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 [\text{O}_2]} = \frac{(1.9 \text{ M}) \times (1.9 \text{ M})}{(0.6 \text{ M}) \times (0.6 \text{ M}) \times (0.82 \text{ M})}$$

$$= 12.229 \text{ M}^{-1} = 12.229 \text{ L mol}^{-1}$$

Question 3.

At a certain temperature and total pressure of 10^5 Pa , iodine vapours contain 40% by volume of iodine atoms in the equilibrium $\text{I}_2(\text{g}) \rightleftharpoons 2\text{I}(\text{g})$. Calculate K_p for the equilibrium.

Answer:

According to available data:

$$\text{Total pressure of equilibrium mixture} = 10^5 \text{ Pa}$$

$$\text{Partial pressure of iodine atoms (I)} = \frac{40}{100} \times (10^5 \text{ Pa}) = 0.4 \times 10^5 \text{ Pa}$$

$$\text{Partial pressure of iodine molecules (I}_2\text{)} = \frac{60}{100} \times (10^5 \text{ Pa}) = 0.6 \times 10^5 \text{ Pa}$$



$$K_p = \frac{p_{\text{I}^2}}{p_{\text{I}}} = \frac{(0.4 \times 10^5 \text{ Pa})^2}{(0.6 \times 10^5 \text{ Pa})} = 2.67 \times 10^4 \text{ Pa}$$

Question 4. Write the expression for the equilibrium constant for each of the following reactions

- (i) $2\text{NOCl}(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{Cl}_2(\text{g})$
- (ii) $2\text{Cu}(\text{NO}_3)_2(\text{s}) \rightleftharpoons 2\text{CuO}(\text{s}) + 4\text{NO}_2(\text{g}) + \text{O}_2(\text{g})$
- (iii) $\text{CH}_3\text{COOC}_2\text{H}_5(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{COOH}(\text{aq}) + \text{C}_2\text{H}_5\text{OH}(\text{aq})$
- (iv) $\text{Fe}^{3+}(\text{aq}) + 3\text{OH}^-(\text{aq}) \rightleftharpoons \text{Fe}(\text{OH})_3(\text{s})$
- (v) $\text{I}_2(\text{s}) + 5\text{F}_2(\text{g}) \rightleftharpoons 2\text{IF}_5(\text{l})$

Answer:

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- (i)
$$K_c = \frac{[\text{NO}(g)]^2 [\text{Cl}_2(g)]}{[\text{NOCl}(g)]^2}$$
- (ii)
$$K_c = \frac{[\text{CuO}(g)]^2 [\text{NO}_2(g)]^4 [\text{O}_2(g)]}{[\text{Cu}(\text{NO}_3)_2(s)]^2} = [\text{NO}_2(g)]^4 [\text{O}_2(g)]$$
- (iii)
$$K_c = \frac{[\text{CH}_3\text{COOH}(aq)][\text{C}_2\text{H}_5\text{OH}(aq)]}{[\text{CH}_3\text{COOC}_2\text{H}_5(aq)][\text{H}_2\text{O}(l)]}$$
$$= \frac{[\text{CH}_3\text{COOH}(aq)][\text{C}_2\text{H}_5\text{OH}(aq)]}{[\text{CH}_3\text{COOC}_2\text{H}_5(aq)]}$$
- (iv)
$$K_c = \frac{[\text{Fe}(\text{OH})_3(s)]}{[\text{Fe}^{3+}(aq)][\text{OH}^-(aq)]^3} = \frac{1}{[\text{Fe}^{3+}(aq)][\text{OH}^-(aq)]^3}$$
- (v)
$$K_c = \frac{[\text{IF}_5(l)]^2}{[\text{I}_2(s)][\text{F}_2(g)]^5} = \frac{[\text{IF}_5(l)]^2}{[\text{F}_2(g)]^5}$$

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Question 5. Find the value of K_c for each of the following equilibria from the value of K

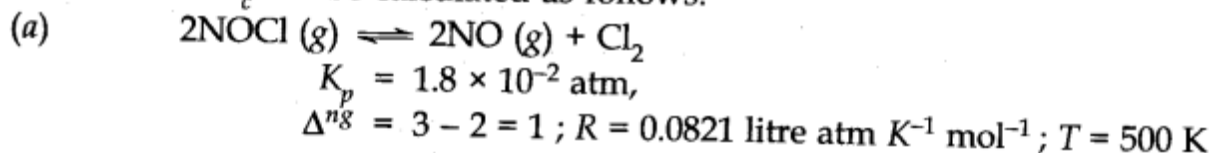
- (a) $2\text{NOCl}(g) \rightleftharpoons 2\text{NO}(g) + \text{Cl}_2(g)$; $K_p = 1.8 \times 10^{-2} \text{ atm at } 500 \text{ K}$
(b) $\text{CaCO}_3(s) \rightleftharpoons \text{CaO}(s) + \text{CO}_2(g)$; $K_p = 167 \text{ atm at } 1073 \text{ K}.$

Answer:

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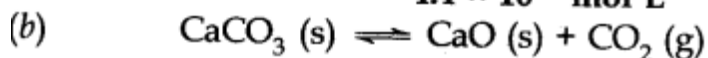
K_p and K_c are related to each other as $K_p = K_c (RT)^{\Delta n_g}$

The value of K_c can be calculated as follows:



$$\therefore K_c = \frac{K_p}{(RT)^{\Delta n_g}} = \frac{(1.8 \times 10^{-2} \text{ atm})}{(0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 500 \text{ K})^1}$$

$$= 4.4 \times 10^{-4} \text{ mol L}^{-1}$$



$$K_p = 167 \text{ atm}, \quad \Delta n_g = 1$$

$$R = 0.0821 \text{ liter atm K}^{-1} \text{ mol}^{-1}; T = 1073 \text{ K}$$

$$K_c = \frac{K_p}{(RT)^{\Delta n_g}} = \frac{(167 \text{ atm})}{(0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 1073 \text{ K})^1}$$

$$= 1.9 \text{ mol L}^{-1}$$

Question 6. For the following equilibrium, $K = 6.3 \times 10^{14}$ at 1000 K. $\text{NO(g)} + \text{O}_3 \rightleftharpoons \text{NO}_2\text{(g)} + \text{O}_2\text{(g)}$ Both the forward and reverse reactions in the equilibrium are elementary bimolecular reactions. What is K_c for the reverse reaction?

Answer:

$$\text{For the reverse reaction } K_c = \frac{1}{K_c} = \frac{1}{6.3 \times 10^{14}} = 1.59 \times 10^{-15}.$$

Question 7. Explain why pure liquids and solids can be ignored while writing the value of equilibrium constants.

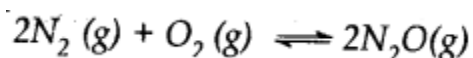
Answer: This is because molar concentration of a pure solid or liquid is independent of the amount present.

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$$\text{Molar concentration} = \frac{\text{No. of moles}}{\text{volume}} \times \frac{\text{Mass}}{\text{volume}} \times \text{Density}$$

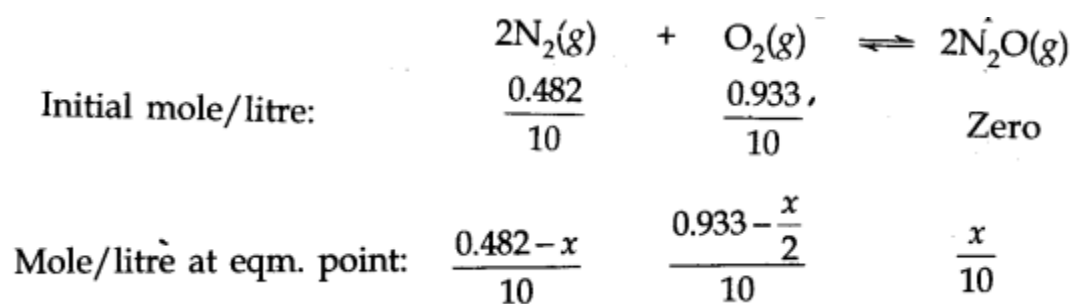
Since density of pure liquid or solid is fixed and molar mass is also fixed. Therefore molar concentration are constant.

Question 8. Reaction between nitrogen and oxygen takes place as follows:



If a mixture of 0.482 mol of N_2 and 0.933 mol of O_2 is placed in a reaction vessel of volume 10 L and allowed to form N_2O at a temperature for which $K_c = 2.0 \times 10^{-37}$, determine the composition of the equilibrium mixture.

Answer: Let x moles of $\text{N}_2(\text{g})$ take part in the reaction. According to the equation, x/2 moles of $\text{O}_2(\text{g})$ will react to form x moles of $\text{N}_2\text{O}(\text{g})$. The molar concentration per litre of different species before the reaction and at the equilibrium point is:



The value of equilibrium constant (2.0×10^{-37}) is extremely small. This means that only small amounts of reactants have reacted. Therefore, is

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extremely small and can be omitted as far as the reactants are concerned.

Applying Law of Chemical Equilibrium $K_c = \frac{[N_2O(g)]^2}{[N_2(g)]^2 [O_2(g)]}$

$$2.0 \times 10^{-37} = \frac{\left(\frac{x}{10}\right)^2}{\left(\frac{0.482}{10}\right)^2 \times \left(\frac{0.933}{10}\right)} = \frac{0.01 x^2}{2.1676 \times 10^{-4}}$$
$$x^2 = 43.352 \times 10^{-40} \quad \text{or} \quad x = 6.6 \times 10^{-20}$$

As x is extremely small, it can be neglected.

Thus, in the equilibrium mixture

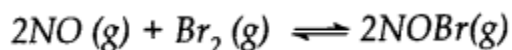
Molar conc. of N_2 = **0.0482 mol L⁻¹**

Molar conc. of O_2 = **0.0933 mol L⁻¹**

Molar conc. of N_2O = $0.1 \times x = 0.1 \times 6.6 \times 10^{-20} \text{ mol L}^{-1}$
= **$6.6 \times 10^{-21} \text{ mol L}^{-1}$**

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Question 9. Nitric oxide reacts with bromine and gives nitrosyl bromide as per reaction given below:



When 0.087 mole of NO and 0.0437 mole of Br₂ are mixed in a closed container at constant temperature, 0.0518 mole of NOBr is obtained at equilibrium. Determine the compositions of the equilibrium mixture.

Answer: The balanced chemical equation for the reaction is:

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According to the equation, 2 moles of NO (g) react with 1 mole of Br₂ (g) to form 2 moles of NOBr (g). The composition of the equilibrium mixture can be calculated as follows:

No. of moles of NOBr (g) formed at equilibrium = 0.0518 mol (given)

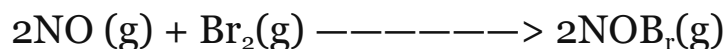
No. of moles of NO (g) taking part in reaction = 0.0518 mol

No. of moles of NO (g) left at equilibrium = 0.087 – 0.0518 = 0.0352 mol

No. of moles of Br₂ (g) taking part in reaction = 1/2 x 0.0518 = 0.0259 mol

No. of moles of Br₂ (g) left at equilibrium = 0.0437 – 0.0259 = 0.0178 mol

The initial molar concentration and equilibrium molar concentration of different species may be represented as:



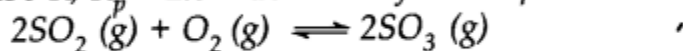
Initial moles 0.087 0.0437 0

Moles at eqm. point: 0.0352 0.0178 0.0518

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Question 10.

At 450 K, $K_p = 2.0 \times 10^{10} \text{ bar}^{-1}$ for the equilibrium reaction:



What is K_c at this temperature?

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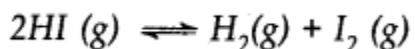
Answer:

$$K_p = K_c (RT)^{\Delta n_g} \text{ or } K_c = \frac{K_p}{(RT)^{\Delta n_g}} = K_p (RT)^{-\Delta n_g}$$

$$\begin{aligned} K_p &= 2.0 \times 10^{10} \text{ bar}^{-1}; R = 0.083 \text{ L bar K}^{-1} \text{ mol}^{-1}; T = 450 \text{ K}; \Delta n_g = 2 - 3 = -1 \\ K_c &= (2.0 \times 10^{10} \text{ bar}^{-1}) \times [(0.083 \text{ L bar K}^{-1} \text{ mol}^{-1}) \times (450 \text{ K})]^{-(-1)} \\ &= 7.47 \times 10^{11} \text{ mol}^{-1} \text{ L} = 7.47 \times 10^{11} \text{ M}^{-1} \end{aligned}$$

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Question 11. A sample of HI (g) is placed in a flask at a pressure of 0.2 atm. At equilibrium partial pressure of HI (g) is 0.04 atm. What is K_p for the given equilibrium?



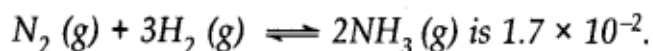
Answer:

$$p_{\text{HI}} = 0.04 \text{ atm}, p_{\text{H}_2} = 0.08 \text{ atm}; p_{\text{I}_2} = 0.08 \text{ atm}$$

$$K_p = \frac{p_{\text{H}_2} \times p_{\text{I}_2}}{p_{\text{HI}}^2} = \frac{(0.08 \text{ atm}) \times (0.08 \text{ atm})}{(0.04 \text{ atm}) \times (0.04 \text{ atm})} = 4.0$$

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Question 12. A mixture of 1.57 mol of N_2 , 1.92 mol of H_2 and 8.13 mol of NH_3 is introduced into a 20 L reaction vessel at 500 K. At this temperature, the equilibrium constant K_c for the reaction



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Is this reaction at equilibrium? If not, what is the direction of net reaction?

Answer:

The reaction is: $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$

$$\text{Concentration quotient } (Q_c) = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(8.13/20 \text{ mol L}^{-1})^2}{(1.57/20 \text{ mol L}^{-1}) \times (1.92/20 \text{ mol L}^{-1})^3} \\ = 2.38 \times 10^3$$

The equilibrium constant (K_c) for the reaction = 1.7×10^{-2}

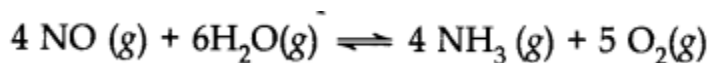
As $Q_c \neq K_c$; this means that the reaction is not in a state of equilibrium.

Question 13. The equilibrium constant expression for a gas reaction is,

$$K_c = \frac{[\text{NH}_3]^4 [\text{O}_2]^5}{[\text{NO}]^4 [\text{H}_2\text{O}]^6}$$

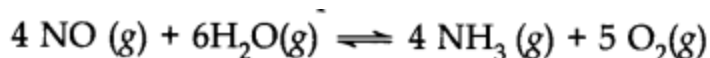
Write the balanced chemical equation corresponding to this expression.

Answer: Balanced chemical equation for the reaction is 4



Question 14. If 1 mole of H_2O and 1 mole of CO are taken in a 10 litre vessel and heated to 725 K, at equilibrium point 40 percent of water (by mass) reacts with carbon monoxide according to equation.

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Calculate the equilibrium constant for the reaction.

Answer: Number of moles of water originally present = 1 mol

Percentage of water reacted = 40%

Number of moles of water reacted = $1 \times 40/100 = 0.4$ mol

Number of moles of water left = $(1 - 0.4) = 0.6$ mole According to the equation, 0.4 mole of water will react with 0.4 mole of carbon monoxide to form 0.4 mole of hydrogen and 0.4 mole of carbon dioxide.

Thus, the molar concentration, per litre of the reactants and products before the reaction and at the equilibrium point are as follows:

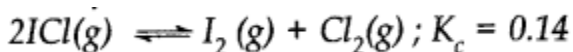
	$\text{H}_2\text{O (g)}$	+	CO(g)	$\rightleftharpoons$	$\text{H}_2 \text{ (g)}$	+	$\text{CO}_2\text{(g)}$
Initial moles/litre	$\frac{1}{10}$		$\frac{1}{10}$		0		0
Mole/litre at the equilibrium point	$\frac{1-0.4}{10} = \frac{0.6}{10}$		$\frac{1-0.4}{10} = \frac{0.6}{10}$		$\frac{0.4}{10}$		$\frac{0.4}{10}$

Applying law of chemical equilibrium,

$$\begin{aligned} \text{Equilibrium constant (K}_c\text{)} &= \frac{[\text{H}_2\text{(g)}][\text{CO}_2\text{(g)}]}{[\text{H}_2\text{O(g)}][\text{CO(g)}]} = \frac{\left(\frac{0.4}{10} \text{ mol L}^{-1}\right) \times \left(\frac{0.4}{10} \text{ mol L}^{-1}\right)}{\left(\frac{0.6}{10} \text{ mol L}^{-1}\right) \times \left(\frac{0.6}{10} \text{ mol L}^{-1}\right)} \\ &= \frac{0.16}{0.36} = 0.44 \end{aligned}$$

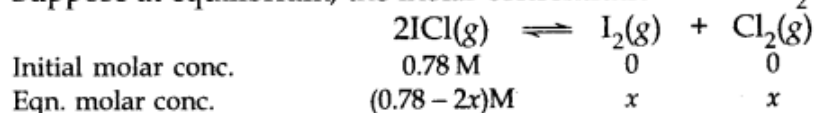
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Question 15. What is the equilibrium concentration of each of the substances in the equilibrium when the initial concentration of ICl was 0.78 M?



Answer:

Suppose at equilibrium, the molar concentration of both $\text{I}_2(g)$ and $\text{Cl}_2(g)$ is $x \text{ mol L}^{-1}$.



$$K_c = \frac{[\text{I}_2(g)][\text{Cl}_2(g)]}{[\text{ICl}(g)]^2} = \frac{(x) \times (x)}{(0.78 - 2x)^2}$$

$$\frac{x}{(0.78 - 2x)} = (0.14)^{1/2} = 0.374 \quad \text{or} \quad x = 0.374 (0.78 - 2x)$$

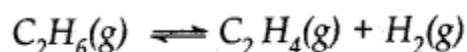
$$x = 0.292 - 0.748x \quad \text{or} \quad 1.748x = 0.292; x = \frac{0.292}{1.748} = 0.167$$

$$[\text{ICl}] = (0.78 - 2 \times 0.167) = (0.78 - 0.334) = \mathbf{0.446 \text{ M}}$$

$$[\text{I}_2] = 0.167 \text{ M}; [\text{Cl}_2] = \mathbf{0.167 \text{ M}}$$

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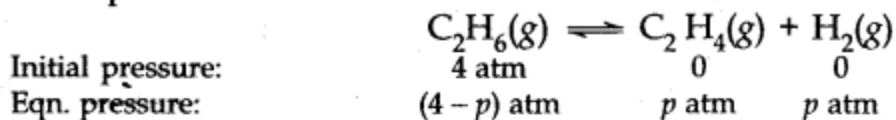
Question 16. $K = 0.04 \text{ atm}$ at 898 K for the equilibrium shown below. What is the equilibrium concentration of C_2H_6 when it is placed in a flask at 4 atm pressure, and allowed to come to equilibrium.



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Answer:

The equilibrium in the reaction is:

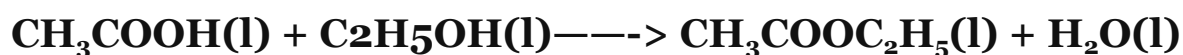


$$\begin{aligned} K_p &= \frac{p_{\text{C}_2\text{H}_4} \times p_{\text{H}_2}}{p_{\text{C}_2\text{H}_6}} \quad \text{or} \quad 0.04 = \frac{p^2}{(4-p)} \\ p^2 &= 0.04(4-p) \quad \text{or} \quad p^2 + 0.04p - 0.16 = 0 \\ p &= \frac{(-0.04) \pm \sqrt{0.0016 - 4(-0.16)}}{2} \\ &= \frac{(-0.04) \pm 0.8}{2} = \frac{0.76}{2} = 0.38 \end{aligned}$$

Equilibrium pressure or concentration of $\text{C}_2\text{H}_6 = (4 - 0.38) = 3.62 \text{ atm}$.

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Question 17. The ester, ethyl acetate is formed by the reaction of ethanol and acetic acid and the equilibrium is represented as:



(i) Write the concentration ratio (concentration quotient) Q for this reaction. Note that water is not in excess and is not a solvent in this reaction.

(ii) At 293 K, if one starts with 1.000 mol of acetic acid and 0.180 mol of ethanol, there is 0.171 mol of ethyl acetate in the final equilibrium mixture. Calculate the equilibrium constant.

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(iii) Starting with 0.50 mol of ethanol and 1.0 mol of acetic acid and maintaining it at 293 K, 0.214 mol of ethyl acetate is found after some time. Has equilibrium been reached?

Answer:

(i) The concentration ratio (Concentration quotient) Q_c for the reaction is:

$$Q_c = \frac{[\text{CH}_3\text{COOC}_2\text{H}_5(l)][\text{H}_2\text{O}(l)]}{[\text{CH}_3\text{COOH}(l)][\text{C}_2\text{H}_5\text{OH}(l)]}$$

(ii) $\text{CH}_3\text{COOH}(l) + \text{C}_2\text{H}_5\text{OH}(l) \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5(l) + \text{H}_2\text{O}(l)$

Initial molar conc.	1.0 mol	0.18 mol	0	0
Molar conc. at equilibrium point	$(1 - 0.171) = 0.829 \text{ mol}$	$(0.18 - 0.171) = 0.009 \text{ mol}$	0.171 mol	0.171 mol

Applying Law of Chemical equilibrium,

$$K_c = \frac{[\text{CH}_3\text{COOC}_2\text{H}_5(l)][\text{H}_2\text{O}(l)]}{[\text{CH}_3\text{COOH}(l)][\text{C}_2\text{H}_5\text{OH}(l)]}$$

$$= \frac{(0.171 \text{ mol}) \times (0.171 \text{ mol})}{(0.829 \text{ mol}) (0.009 \text{ mol})} = 3.92$$

(iii) $\text{CH}_3\text{COOH}(l) + \text{C}_2\text{H}_5\text{OH}(l) \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5(l) + \text{H}_2\text{O}(l)$

Initial molar conc.	1.0 mol	0.5 mol	0.214 mol	0.214 mol
Molar conc. at equilibrium	$1.0 - 0.214 = 0.786 \text{ mol}$	$0.5 - 0.214 = 0.286 \text{ mol}$		

$$Q_c = \frac{[\text{CH}_3\text{COOC}_2\text{H}_5(l)][\text{H}_2\text{O}(l)]}{[\text{CH}_3\text{COOH}(l)][\text{C}_2\text{H}_5\text{OH}(l)]}$$

$$= \frac{(0.214 \text{ mol}) \times (0.214 \text{ mol})}{(0.786 \text{ mol}) (0.286 \text{ mol})} = 0.204$$

Since Q_c is less than K_c this means that the equilibrium has not been reached. The reactants are still taking part in the reaction to form the products.

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Question 18. A sample of pure PCl_5 was introduced into an evacuated vessel at 473 K. After equilibrium was reached, the concentration of PCl_5 was found to be $0.5 \times 10^{-1} \text{ mol L}^{-1}$. If K_c is 8.3×10^{-3} what are the concentrations of PCl_3 and Cl_2 at equilibrium?

Answer: Let the initial molar concentration of PCl_5 per litre = $x \text{ mol}$

Molar concentration of PCl_5 at equilibrium = 0.05 mol

$\therefore$ Moles of PCl_5 decomposed = $(x - 0.05) \text{ mol}$

Moles of PCl_3 formed = $(x - 0.05) \text{ mol}$

Moles of Cl_2 formed = $(x - 0.05) \text{ mol}$

The molar conc./litre of reactants and products before the reaction and at the equilibrium point are:

	PCl_5	$\rightleftharpoons$	PCl_3	+	Cl_2
Initial moles/litre	x		0		0
Moles/litre at eqm. point	0.05		$(x - 0.05)$		$(x - 0.05)$
Equilibrium constant (K_c)	$= 8.3 \times 10^{-3} = 0.0083$				
Applying Law of chemical equilibrium,					

$$K_c = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]}; \quad 0.0083 = \frac{(x - 0.05) \times (x - 0.05)}{0.05}$$

$$(x - 0.05)^2 = 0.0083 \times 0.05 = 4.15 \times 10^{-4}$$

$$(x - 0.05) = (4.15 \times 10^{-4})^{1/2} = 2.037 \times 10^{-2} = 0.02 \text{ moles}$$

$$x = 0.05 + 0.02 = 0.07 \text{ mol}$$

The molar concentration per litre of PCl_3 at eqm. = $0.07 - 0.05 = 0.02 \text{ mol}$

The molar concentration per litre of Cl_2 at eqm. = $0.07 - 0.05 = 0.02 \text{ mol}$.

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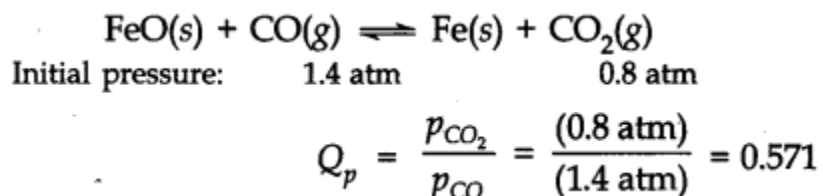
Question 19. One of the reactions that takes place in producing steel from iron ore is the reduction of iron

(II) oxide by carbon monoxide to give iron metal and Co_2

$\text{FeO(s)} + \text{CO(g)} \longrightarrow \text{Fe(s)} + \text{Co}_2\text{(g)}$; $K_p = 0.265$ atm at 1050 K

What are the equilibrium partial pressures of CO and Co_2 at 1050 K if the initial pressures are: $P_{\text{CO}} = 1.4$ atm and $P_{\text{Co}_2} = 0.80$ atm?

Answer:



Since $Q_p > K_p (0.265)$, this means that the reaction will move in the backward direction to attain the equilibrium. Therefore, partial pressure of Co_2 will decrease while that of CO will increase so that the equilibrium may be attained again. Let p atm be the decrease in the partial pressure of Co_2 . Therefore, the partial pressure of CO will increase by the same magnitude i.e., p atm.

$$p_{\text{CO}_2} = (0.8 - p) \text{ atm}; p_{\text{CO}}(\text{g}) = (1.4 + p) \text{ atm}$$

At equilibrium, $K_p = \frac{p_{\text{CO}_2}}{p_{\text{CO}}} = \frac{(0.8 - p) \text{ atm}}{(1.4 + p) \text{ atm}} = \frac{(0.8 - p)}{(1.4 + p)}$

or $0.265 = \frac{(0.8 - p)}{(1.4 + p)}$

$$0.371 + 0.265 p = 0.8 - p \quad \text{or} \quad 1.265 p = 0.8 - 0.371 = 0.429$$

$$p = 0.429/1.265 = 0.339 \text{ atm}$$

$$(p_{\text{CO}})_{\text{eq}} = (1.4 + 0.339) = 1.739 \text{ atm}$$

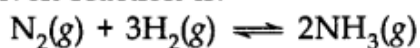
$$(p_{\text{CO}_2})_{\text{eq}} = (0.8 - 0.339) = 1.461 \text{ atm}$$

Question 20.

Equilibrium constant K_c for the reaction, $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ at 500 K is 0.061. At particular time, the analysis shows that the composition of the reaction mixture is: 3.0 mol L^{-1} of N_2 ; 2.0 mol L^{-1} of H_2 ; 0.50 mol L^{-1} of NH_3 . Is the reaction at equilibrium? If not, in which direction does the reaction tend to proceed to reach the equilibrium?

Answer:

The given reaction is:



According to available data.

$$\text{N}_2 = [3.0]; \text{H}_2 = [2.0]; \text{NH}_3 = [0.50]$$

$$Q_c = \frac{[\text{NH}_3(\text{g})]^2}{[\text{N}_2(\text{g})][\text{H}_2(\text{g})]^3} = \frac{[0.50]^2}{[3.0][2.0]^3} = \frac{0.25}{24} = 0.0104.$$

Since the value of Q_c is less than that of K_c (0.061), the reaction is not in a state of equilibrium. It will proceed in the **forward direction** till Q_c becomes the same as K_c .

Question 21. Bromine monochloride (BrCl) decomposes into bromine and chlorine and reaches the equilibrium:

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The value of K_c is 32 at 500 K. If initially pure Br_2Cl is present at a concentration of $3.3 \times 10^{-3} \text{ mol L}^{-1}$ what is its molar concentration in the mixture at equilibrium?

Answer: Let x moles of Br_2Cl decompose in order to attain the equilibrium. The initial molar concentration and the molar concentration at equilibrium point of different species may be represented as follows:

	2BrCl (g)	$\rightleftharpoons$	$\text{Br}_2(\text{g})$	+	$\text{Cl}_2(\text{g})$
Initial moles/litre	0.0033		0		0
Moles/litre at eqm. point	$0.0033 - x$		$x/2$		$x/2$

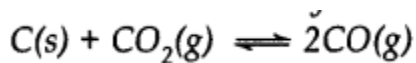
Applying Law of chemical equilibrium, $K_c = \frac{[\text{Br}_2][\text{Cl}_2]}{[\text{BrCl}]^2}$ or $32 = \frac{(x/2) \times (x/2)}{(0.0033 - x)^2}$

On taking the square root, $5.656 = \frac{x/2}{(0.0033 - x)}$

$$\frac{x}{(0.0033 - x)} = 11.31 \quad \text{or} \quad 12.31x = 0.037 \quad ; \quad x = \frac{0.037}{12.31} = 0.003$$

$\therefore$ Molar concentration of BrCl at equilibrium point = $0.0033 - 0.003$
 $= 0.0003 \text{ mol L}^{-1} = 3 \times 10^{-4} \text{ mol L}^{-1}$

Question 22. At 1127 K and 1 atmosphere pressure, a gaseous mixture of CO and CO_2 in equilibrium with solid carbon has 90.55% CO by mass.



Calculate K_c for the reaction at the above temperature.

Answer:

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Step I: Calculation of K_p for the reaction

Let the total mass of the gaseous mixture = 100 g

Mass of CO in the mixture = 90.55 g

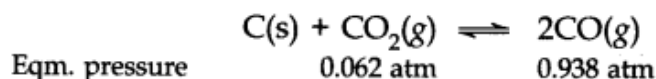
Mass of CO_2 in the mixture = $(100 - 90.55) = 9.45$ g

$$\text{No. of moles of CO} = \frac{90.55\text{g}}{(28\text{g mol}^{-1})} = 3.234 \text{ mol}$$

$$\text{No. of moles of CO}_2 = \frac{9.45}{(44\text{g mol}^{-1})} = 0.215 \text{ mol}$$

$$p_{\text{CO}} \text{ in the mixture} = \frac{(3.234 \text{ mol})}{(3.234 + 0.215)} \times 1 \text{ atm} = \frac{(3.234 \text{ mol})}{(3.449 \text{ mol})} \times 1 \text{ atm} = 0.938 \text{ atm}$$

$$p_{\text{CO}_2} \text{ in the mixture} = \frac{(0.215 \text{ mol})}{(3.449 \text{ mol})} \times 1 \text{ atm} = 0.062 \text{ atm}$$



$$K_p = \frac{p^2_{\text{CO}}}{p_{\text{CO}_2}} = \frac{(0.938 \text{ atm})^2}{(0.062 \text{ atm})} = 14.19 \text{ atm}$$

Step II. Calculation of K_c for the reaction.

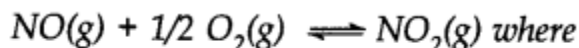
$$K_c = \frac{K_p}{(RT)^{\Delta n_g}}$$

$$K_p = 14.19 \text{ atm}, \quad R = 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}, \quad T = 1127 \text{ K}; \quad \Delta n_g = 2 - 1 = 1$$

$$K_c = \frac{(14.19 \text{ atm})}{(0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}) \times (1127 \text{ K})^1} = 6.46$$

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Question 23. Calculate (a) $\Delta G^\ominus$ and (b) the equilibrium constant for the formation of NO_2 from NO and O_2 at 298 K



$$\Delta_f G^\ominus (\text{NO}_2) = 52.0 \text{ kJ/mol}, \quad \Delta_f G^\ominus (\text{NO}) = 87.0 \text{ kJ/mol}, \quad \Delta_f G^\ominus (\text{O}_2) = 0 \text{ kJ/mol},$$

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Answer:

Step I. Calculation of $\Delta G^\ominus$

$$\begin{aligned}\Delta G^\ominus &= \Delta_f G^\ominus (\text{NO}_2) - [\Delta_f G^\ominus (1/2\text{O}_2)] \\ &= 52.0 - (87 + 0) = -35 \text{ kJ mol}^{-1}\end{aligned}$$

Step II. Calculation of K_c

$$\Delta G^\ominus = -2.303 RT \log K_c$$

$$\log K_c = -\frac{\Delta G^\ominus}{2.303 RT} = -\frac{(-35 \times 10^3 \text{ J mol}^{-1})}{2.303 \times (8.314 \text{ kJ mol}^{-1}) \times (298 \text{ K})} = 6.134$$

$$K_c = \text{Antilog } 6.134 = 1.36 \times 10^6.$$

Question 24. Does the number of moles of reaction products increase, decrease or remain same when each of the following equilibria is subjected to a decrease in pressure by increasing the volume?

- (i) $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$
- (ii) $\text{CaO}(\text{s}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{CaCO}_3(\text{s})$
- (iii) $3\text{Fe}(\text{s}) + 4\text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{Fe}_3\text{O}_4(\text{s}) + 4\text{H}_2(\text{g})$

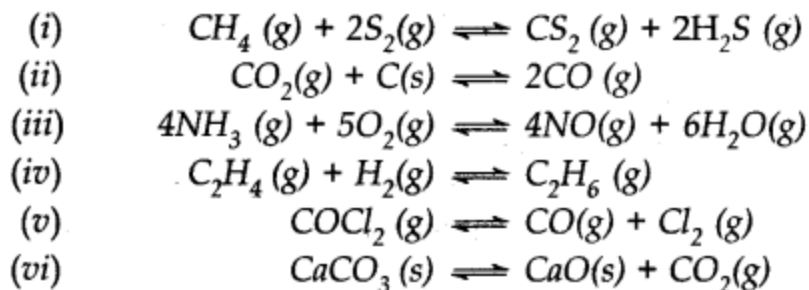
Answer: (i) Pressure will increase in the forward reaction and number of moles of products will increase.

(ii) Pressure will increase in backward reaction and number of moles of products will decrease.

(iii) The change in pressure will have no effect on the equilibrium constant and there will be no change in the number of moles.

Question 25. Which of the following reactions will get affected by increase in pressure? Also mention whether the change will cause the reaction to go to the right or left direction.

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Answer: Only those reactions will be affected by increasing the pressure in which the number of moles of the gaseous reactants and products are different ($n_p \neq n_r$) (gaseous). With the exception of the reaction (1); all the remaining five reactions will get affected by increasing the pressure. In general,

- The reaction will go to the left if $n_p > n_r$.
- The reaction will go to the right if $n_r > n_p$.

Keeping this in mind,

(i) Increase in pressure will not affect equilibrium because $n_p = n_r = 3$.

(ii) Increase in pressure will favour backward reaction because $n_p(2) > n_r(1)$

(iii) Increase in pressure will favour backward reaction because $n_p(10) > n_r(9)$

(iv) Increase in pressure will favour forward reaction because $n_p(1) < n_r(2)$

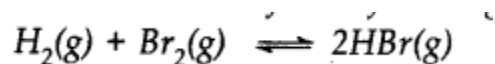
(v) Increase in pressure will favour backward reaction because $n_p(2) > n_r(1)$

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(vi) Increase in pressure will favour backward reaction because $n_p (1) > n_r (0)$.

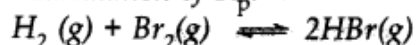
Question 26. The equilibrium constant for the following reaction is 1.6×10^5 at 1024 K.

Find the equilibrium pressure of all gases if 10.0 bar of HBr is introduced into a sealed container at 1024 K.



Ans.

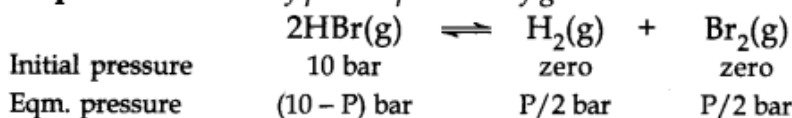
Step I. Calculation of K_p .



$$K_p = K_c(RT)^{\Delta n} = K_c(RT)^0 \quad (\because \Delta n = 2 - 2 = \text{zero})$$

$$K_p = K_c = 1.6 \times 10^5$$

Step II. Calculation of partial pressure of gases



$$K_p = \frac{p_{\text{H}_2} \times p_{\text{Br}_2}}{p_{\text{HBr}}^2} \text{ or } \frac{1}{1.6 \times 10^5} = \frac{P/2 \times P/2}{(10 - P)^2} = \frac{P^2}{4(10 - P)^2}$$

On taking square root; $\frac{P^2}{4(10 - P)^2} = \left(\frac{1}{1.6 \times 10^5} \right)^{1/2} \text{ or } \frac{2(10 - P)}{P} = (1.6 \times 10^5)^{1/2}$

$$= 4 \times 10^2$$

$$20 - 2P = 4 \times 10^2 P \text{ or } P(4 \times 10^2 + 2) = 20$$

or $P = \frac{20}{(400 + 2)} = \frac{20}{402} = 0.050 \text{ bar}$

$$p_{\text{H}_2} = 0.025 \text{ bar; } p_{\text{Br}_2} = 0.025 \text{ bar; } p_{\text{HBr}} = 10 - 0.05 = 9.95 \text{ bar} \approx 10.0 \text{ bar.}$$

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Question 27. Hydrogen gas is obtained from the natural gas by partial oxidation with steam as per the following endothermic reaction:



Write the expression for K_p for the above reaction

How will the value of K_p and composition of equilibrium mixture be affected by:

(i) increasing the pressure, (ii) increasing the temperature, (iii) using a catalyst?

Answer: The expression for K_p for the reaction is:

$$K_p = \frac{(p_{\text{CO}}) \times (p_{\text{H}_2})^3}{(p_{\text{CH}_4}) \times (p_{\text{H}_2\text{O}})}$$

(i) By increasing the pressure, the number of moles per unit volume will increase. In order to decrease the same, the equilibrium gets shifted to the left or in the backward direction. As a result, more of reactants will be formed and the value of K_p will decrease.

(ii) If the temperature is increased, according to Le Chatelier's principle, the forward reaction will be favoured as it is endothermic. Therefore, the equilibrium gets shifted to the right and the value of K_p will increase.

(iii) The addition of catalyst will not change the equilibrium since it influences both the forward and the backward reactions to the same extent. But it will be attained more quickly.

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Question 28. What is the effect of:

- (i) addition of H_2 (ii) addition of CH_3OH
(iii) removal of CO (iv) removal of CH_3OH

On the equilibrium $2H_2(g) + CO(g) \rightleftharpoons CH_3OH(g)$

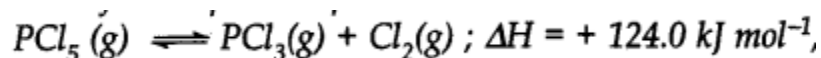
Answer: (i) Equilibrium will be shifted in the forward direction.

(ii) Equilibrium will be shifted in the backward direction.

(iii) Equilibrium will be shifted in the backward direction.

(iv) Equilibrium will be shifted in the forward direction.

Question 29. At 473 K, the equilibrium constant K_c for the decomposition of phosphorus pentachloride (PCl_5) is 8.3×10^{-3} . if decomposition proceeds as:



(a) Write an expression for K_c for the reaction

(b) What is the value of K_c for the reverse reaction at the same temperature.

(c) What would be the effect on K_c if

(i) More of PCl_5 is added (ii) Temperature is increased.

Answer:

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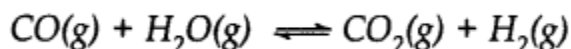
(a) The expression for $K_c = \frac{[PCl_3(g)][Cl_2(g)]}{[PCl_5(g)]}$,

(b) For reverse reaction (K_c') = $\frac{PCl_5(g)}{[PCl_3(g)][Cl_2(g)]} = \frac{1}{8.3 \times 10^{-3}} = 120.48$

(c) (i) By adding more of PCl_5 , value of K_c will remain constant because there is no change in temperature.

(ii) By increasing the temperature, the forward reaction will be favoured since it is endothermic in nature. Therefore, the value of equilibrium constant will increase.

Question 30. Dihydrogen gas used in Haber's process is produced by reacting methane from natural gas with high temperature steam. The first stage of two stage reaction involves the formation of CO and H_2 . In second stage, CO formed in first stage is reacted with more steam in water gas shift reaction.

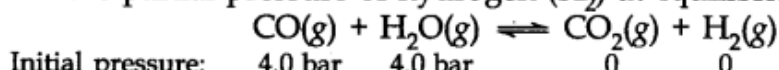


If a reaction vessel at 400°C is charged with an equimolar mixture of CO and steam so that $P_{CO} = P_{H_2O} = 4.0$ bar, what will be the partial pressure of H_2 at equilibrium? $K_p = 0.1$ at 400°C .

Answer:

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Let the partial pressure of hydrogen (H_2) at equilibrium point = p bar



Initial pressure: 4.0 bar 4.0 bar 0 0

Eqm. pressure: $(4-p)$ bar $(4-p)$ bar p bar p bar

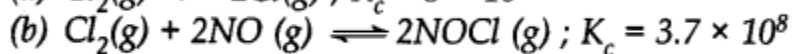
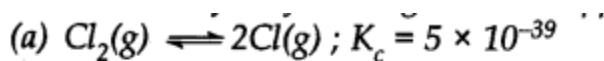
$$K_p = \frac{p_{CO_2} \times p_{H_2}}{p_{CO} \times p_{H_2O}} \quad \text{or} \quad 0.1 = \frac{(p \text{ bar}) \times (p \text{ bar})}{(4-p) \text{ bar} \times (4-p) \text{ bar}}$$

$$\frac{p^2}{(4-p)^2} = 0.1 \quad \text{or} \quad \frac{p}{(4-p)} = (0.1)^{1/2} = 0.316$$

$$p = 0.316 (4-p) \quad \text{or} \quad p = 1.264 - 0.316 p$$

$$1.316 p = 1.264 \quad \text{or} \quad p = \frac{1.264}{1.316} = 0.96 \text{ bar}$$

Question 31. Predict which of the following will have appreciable concentration of reactants and products:



Answer: Following conclusions can be drawn from the values of K_c .

(a) Since the value of K_c is very small, this means that the molar concentration of the products is very small as compared to that of the reactants.

(b) Since the value of K_c is quite large, this means that the molar concentration of the products is very large as compared to that of the reactants.

(c) Since the value of K_c is 1.8, this means that both the products and reactants have appreciable concentration.

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Question 32. The value of K_c for the reaction $3\text{O}_2(\text{g}) \rightleftharpoons 2\text{O}_3(\text{g})$ is 2.0×10^{-50} at 25°C . If equilibrium concentration of O_2 in air at 25°C is 1.6×10^{-2} , what is the concentration of O_3 ?

Answer:



$$K_c = \frac{[\text{O}_3]^2}{[\text{O}_2]^3} \quad \text{or} \quad (2.0 \times 10^{-50}) = \frac{[\text{O}_3]^2}{(1.6 \times 10^{-2})^3}$$

$$\begin{aligned} \text{or} \quad [\text{O}_3]^2 &= (2.0 \times 10^{-50}) \times (1.6 \times 10^{-2})^3 \\ [\text{O}_3]^2 &= 8.192 \times 10^{-56} \quad \text{or} \quad [\text{O}_3] = (8.192 \times 10^{-56})^{1/2} = 2.86 \times 10^{-28} \text{ M.} \end{aligned}$$

Question 33.

The reaction $\text{CO}(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons \text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g})$ is at equilibrium at 1300 K in a 1L flask. It also contain 0.30 mol of CO, 0.10 mol of H_2 and 0.02 mol of H_2O and an unknown amount of CH_4 in the flask. Determine the concentration of CH_4 in the mixture. The equilibrium constant, K_c for the reaction at the given temperature is 3.90.

Answer:



According to available data

$$K_c = \frac{[\text{CH}_4] \times [\text{H}_2\text{O}]}{[\text{CO}] \times [\text{H}_2]^3} \quad \text{or} \quad 3.90 = \frac{[\text{CH}_4] \times [0.02]}{[0.30] \times [0.1]^3}$$

$$[\text{CH}_4] = \frac{(3.9) \times (0.30) \times (0.001)}{(0.02)} = 5.85 \times 10^{-2} \text{ M}$$

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Question 34. What is meant by conjugate acid-base pair?

Find the conjugate acid/base for the following species:

HNO_2 , CH^- , HClO_4 , OH^- , CO_3^{2-} , S^{2-}

Answer: An acid-base pair which differs by a proton only ($\text{HA} \rightleftharpoons \text{A}^- + \text{H}^+$) is known as conjugate acid-base pair.

Conjugate acid: HCN , H_2O , HCO_3^- , HS^- .

Conjugate base: NO_2^- , ClO_4^- , O_2^-

Question 35. Which of the following are Lewis Acids?

H_2O , BF_3 , H^+ and NH_4^+ ,

Answer: BF_3 , H^+ ions are Lewis acids.

Question 36. What will be the conjugate bases for the

Bronsted acids? HF , H_2SO_4 and H_2CO_3 ?

Answer: Conjugate bases: F^- , HSO_4^- , HCO_3^- .

Question 37. Write the conjugate acids for the following

Bronsted bases:

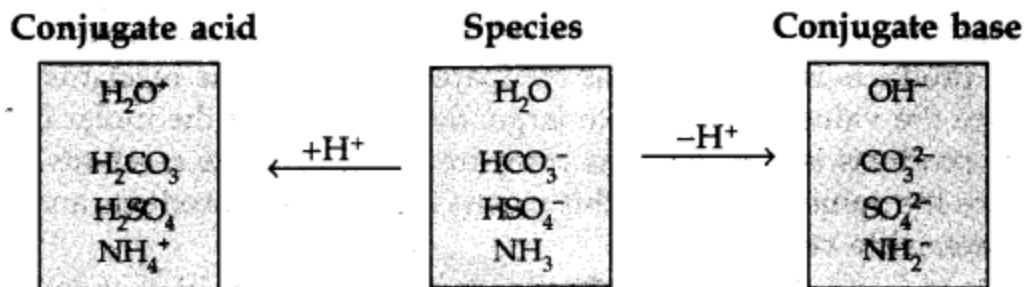
NH_2 , NH_3 and HCOO^-

Answer: NH_3 , NH_4^+ and HCOOH

Question 38. The species H_2O , HCO_3^- , HSO_4^- and NH_3 can act both as Bronsted acid and base. For each case, give the corresponding conjugate acid and base.

Answer:

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Question 39. Classify the following species into Lewis acids and Lewis bases and show how these can act as Lewis acid/Lewis base?

(a) OH^- ions (b) F^- (c) H^+ (d) BCl_3

Answer: (a) OH^- ions can donate an electron pair and act as Lewis base.

(b) F^- ions can donate an electron pair and act as Lewis base.

(c) H^+ ions can accept an electron pair and act as Lewis acid.

(d) BCl_3 can accept an electron pair since Boron atom is electron deficient. It is a Lewis acid.

Question 40. The concentration of hydrogen ions in a sample of soft drink is $3.8 \times 10^{-3} \text{ M}$. What is the pH value?

Answer: $\text{pH} = -\log [\text{H}^+] = -\log (3.8 \times 10^{-3}) = -\log 3.8 + 3 = 3 - 0.5798 = 2.4202 = 2.42$

Question 41. The pH of a sample of vinegar is 3.76. Calculate the concentration of hydrogen ion in it.

Answer: $\text{pH} = -\log [\text{H}^+]$ or $\log [\text{H}^+] = -\text{pH} = -3.76 = 4.24$

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$$\therefore [\text{H}^+] = \text{Antilog } 4.24 = 1.738 \times 10^{-4} = 1.74 \times 10^{-4} \text{ M}$$

Question 42. The ionization constant of HF, HCOOH and HCN at 298 K are 6.8×10^{-4} , 1.8×10^{-4} and 4.8×10^{-9} respectively, Calculate the ionization constant of the corresponding conjugate base.

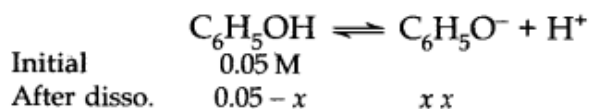
Answer: For F^- , $K_b = K_w / K_a = 10^{-14} / (6.8 \times 10^{-4}) = 1.47 \times 10^{-11} = 1.5 \times 10^{-11}$.

$$\text{For } \text{HCOO}^-, K_b = 10^{-14} / (1.8 \times 10^{-4}) = 5.6 \times 10^{-11}$$

$$\text{For } \text{CN}^-, K_b = 10^{-14} / (4.8 \times 10^{-9}) = 2.08 \times 10^{-6}$$

Question 43. The ionization constant of phenol is 1.0×10^{-10} . What is the concentration of phenolate ion in 0.05 M solution of phenol? What will be its degree of ionization if the solution is also 0.01 M in sodium phenolate?

Answer:



$$\therefore K_a = \frac{x \times x}{0.05 - x} = 1.0 \times 10^{-10} \text{ (Given)} \quad \text{or} \quad \frac{x^2}{0.05} = 1.0 \times 10^{-10}$$

$$\text{or } x^2 = 5 \times 10^{-12} \quad \text{or } x = 2.2 \times 10^{-6} \text{ M}$$

In presence of 0.01 $\text{C}_6\text{H}_5\text{ONa}$, suppose y is the amount of phenol dissociated, then at equilibrium

$$[\text{C}_6\text{H}_5\text{OH}] = 0.05 - y \approx 0.05,$$

$$[\text{C}_6\text{H}_5\text{O}^-] = 0.01 + y \approx 0.01 \text{ M}, [\text{H}^+] = y \text{ M}$$

$$\therefore K_a = \frac{(0.01)(y)}{0.05} = 1.0 \times 10^{-10} \text{ (Given)} \quad \text{or} \quad y = 5 \times 10^{-10}$$

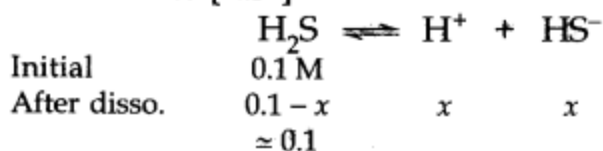
$$\therefore \alpha = \frac{y}{c} = \frac{5 \times 10^{-10}}{5 \times 10^{-2}} = 10^{-8}.$$

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Question 44. The-first ionization constant of H_2S is 9.1×10^{-8} . Calculate the concentration of HS^- ions in its 0.1 M solution and how will this concentration be affected if the solution is 0.1 M in HCl also? If the second dissociation constant of H_2S is 1.2×10^{-13} , calculate the concentration of S^{2-} under both conditions.

Answer:

To calculate $[\text{HS}^-]$

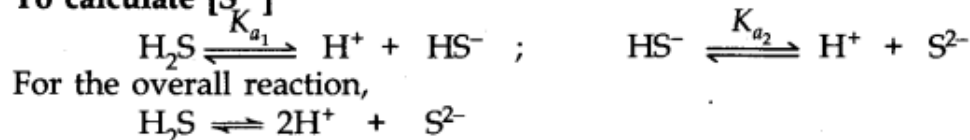


$$K_a = \frac{x \times x}{0.1} = 9.1 \times 10^{-8} \quad \text{or} \quad x^2 = 9.1 \times 10^{-9} \quad \text{or} \quad x = 9.54 \times 10^{-5}.$$

In presence of 0.1 M HCl, suppose H_2S dissociated is y . Then at equilibrium, $[\text{H}_2\text{S}] = 0.1 - y \approx 0.1$, $[\text{H}^+] = 0.1 + y \approx 0.1$, $[\text{HS}^-] = y$ M

$$K_a = \frac{0.1 \times y}{0.1} = 9.1 \times 10^{-8} \text{ (Given)} \quad \text{or} \quad y = 9.1 \times 10^{-8} \text{ M}$$

To calculate $[\text{S}^{2-}]$



$$K_a = K_{a_1} \times K_{a_2} = 9.1 \times 10^{-8} \times 1.2 \times 10^{-13} = 1.092 \times 10^{-20}$$

$$K_a = \frac{[\text{H}^+]^2 [\text{S}^{2-}]}{[\text{H}_2\text{S}]}$$

In the absence of 0.1 M HCl, $[\text{H}^+] = 2 [\text{S}^{2-}]$

Hence, if $[\text{S}^{2-}] = x$, $[\text{H}^+] = 2x$

$$\therefore \frac{(2x)^2 x}{0.1} = 1.092 \times 10^{-20} \quad \text{or} \quad 4x^3 = 1.092 \times 10^{-21} = 273 \times 10^{-24}$$

$$3 \log x = \log 273 - 24 = 2.4362 - 24$$

$$\log x = 0.8127 - 8 = \bar{8}.8127,$$

$$\text{or} \quad x = \text{Antilog } \bar{8}.8127 = 273 \times 10^{-24} = 6.497 \times 10^{-8} = 6.5 \times 10^{-8} \text{ M.}$$

In presence of 0.1 M HCl, suppose $[\text{S}^{2-}] = y$, then

$$[\text{H}_2\text{S}] = 0.1 - y \approx 0.1 \text{ M}, \quad [\text{H}^+] = 0.1 + y \approx 0.1 \text{ M}$$

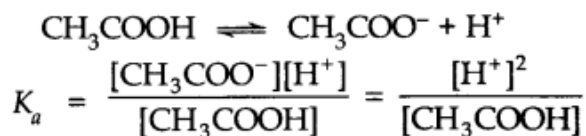
$$K_a = \frac{(0.1)^2 \times y}{0.1} = 1.09 \times 10^{-20} \quad \text{or} \quad y = 1.09 \times 10^{-19} \text{ M.}$$

Question 45. The ionization constant of acetic acid is 1.74×10^{-5} . Calculate the degree of dissociation of acetic acid in its

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0.05 M solution. Calculate the concentration of acetate ions in the solution and its pH.

Answer:



or $[\text{H}^+] = \sqrt{K_a [\text{CH}_3\text{COOH}]} = \sqrt{(1.74 \times 10^{-5})(5 \times 10^{-2})} = 9.33 \times 10^{-4} \text{ M}$

$$[\text{CH}_3\text{COO}^-] = [\text{H}^+] = 9.33 \times 10^{-4} \text{ M}$$
$$\text{pH} = -\log (9.33 \times 10^{-4}) = 4 - 0.9699 = 4 - 0.97 = 3.03$$

Question 46. It has been found that the pH of a 0.01 M solution of an organic acid is 4.15. Calculate the concentration of the anion, the ionization constant of the acid and its PK_a .

Answer:

$$\text{HA} \rightleftharpoons \text{H}^+$$
$$\text{pH} = -\log [\text{H}^+] \quad \text{or} \quad \log [\text{H}^+] = -4.15 = \bar{5}.85$$
$$\therefore [\text{H}^+] = 7.08 \times 10^{-5} \text{ M} = 7.08 \times 10^{-5} \text{ M}$$
$$[\text{A}^-] = [\text{H}^+] = 7.08 \times 10^{-5} \text{ M}$$
$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{(7.08 \times 10^{-5})(7.08 \times 10^{-5})}{10^{-2}} = 5.0 \times 10^{-7}$$
$$\text{pK}_a = -\log K_a = -\log (5.0 \times 10^{-7}) = 7 - 0.699 = 6.301$$

Question 47. Assuming complete dissociation, calculate the pH of the following solutions:

(a) 0.003 M HCl (b) 0.005 M NaOH (c) 0.002 M HBr (d) 0.002 M KOH

Answer:

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(a) $\text{HCl} + \text{aq} \rightarrow \text{H}^+ + \text{Cl}^-$, $\therefore [\text{H}^+] = [\text{HCl}] = 3 \times 10^{-3} \text{ M}$, $\text{pH} = -\log (3 \times 10^{-3}) = 2.52$

(b) $\text{NaOH} + \text{aq} \rightarrow \text{Na}^+ + \text{OH}^-$

$$\therefore [\text{OH}^-] = 5 \times 10^{-3} \text{ M}, [\text{H}^+] = 10^{-14} / (5 \times 10^{-3}) = 2 \times 10^{-12} \text{ M}$$
$$\text{pH} = -\log (2 \times 10^{-12}) = 11.70$$

(c) $\text{HBr} + \text{aq} \rightarrow \text{H}^+ + \text{Br}^-$, $\therefore [\text{H}^+] = 2 \times 10^{-3} \text{ M}$, $\text{pH} = -\log (2 \times 10^{-3}) = 2.70$

(d) $\text{KOH} + \text{aq} \rightarrow \text{K}^+ + \text{OH}^-$,

$$\therefore [\text{OH}^-] = 2 \times 10^{-3} \text{ M}, [\text{H}^+] = 10^{-14} / (2 \times 10^{-3}) = 5 \times 10^{-12}$$
$$\text{pH} = -\log (5 \times 10^{-12}) = 11.30$$

Question 48. Calculate the pH of the following solutions:

(a) 2g of TlOH dissolved in water to give 2 litre of the solution

(b) 0.3 g of $\text{Ca}(\text{OH})_2$ dissolved in water to give 500 mL of the solution

(c) 0.3 g of NaOH dissolved in water to give 200 mL of the solution

(d) 1 mL of 13.6 M HCl is diluted with water to give 1 litre of the solution.

Answer:

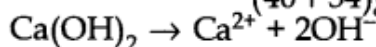
$$(a) \text{ Molar conc. of TIOH} = \frac{2g}{(204 + 16 + 1) g \text{ mol}^{-1}} \times \frac{1}{2 \text{ L}} = 4.52 \times 10^{-3} \text{ M}$$

$$[\text{OH}^-] = [\text{TIOH}] = 4.52 \times 10^{-3} \text{ M}$$

$$[\text{H}^+] = 10^{-14} / (4.52 \times 10^{-3}) = 2.21 \times 10^{-12} \text{ M}$$

$$\therefore \text{pH} = -\log (2.21 \times 10^{-12}) = 12 - (0.3424) = 11.66$$

$$(b) \text{ Molar conc. of Ca(OH)}_2 = \frac{0.3 g}{(40 + 34) g \text{ mol}^{-1}} \times \frac{1}{0.5 \text{ L}} = 8.11 \times 10^{-3} \text{ M}$$



$$[\text{OH}^-] = 2[\text{Ca(OH)}_2] = 2 \times (8.11 \times 10^{-3}) \text{ M} = 16.22 \times 10^{-3} \text{ M}$$

$$\text{pOH} = -\log (16.22 \times 10^{-3}) = 3 - 1.2101 = 1.79$$

$$\text{pH} = 14 - 1.79 = 12.21$$

$$(c) \text{ Molar conc. of NaOH} = \frac{0.3 g}{40 g \text{ mol}^{-1}} \times \frac{1}{0.2 \text{ L}} = 3.75 \times 10^{-2} \text{ M}$$

$$[\text{OH}^-] = 3.75 \times 10^{-2} \text{ M}$$

$$\text{pOH} = -\log (3.75 \times 10^{-2}) = 2 - 0.0574 = 1.43$$

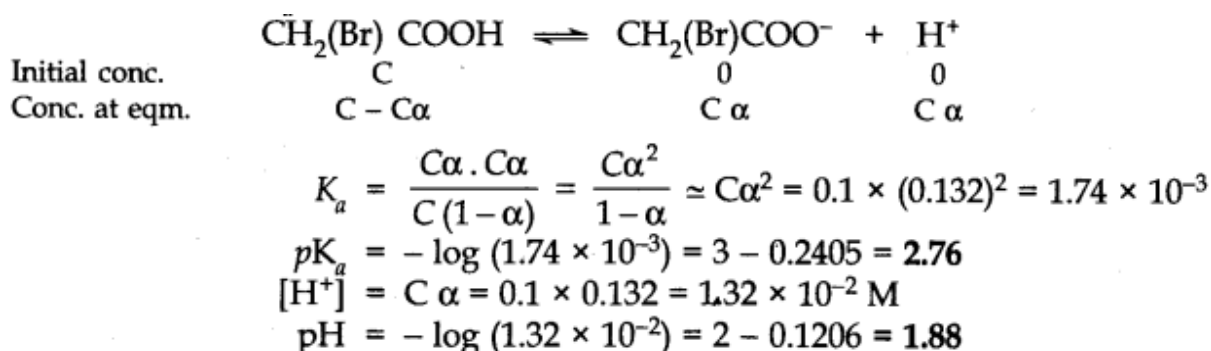
$$\text{pH} = 14 - 1.43 = 12.57$$

$$(d) M_1 V_1 = M_2 V_2 \therefore 13.6 \text{ M} \times 1 \text{ mL} = M_2 \times 1000 \text{ mL} \therefore M_2 = 1.36 \times 10^{-2} \text{ M}$$

$$[\text{H}^+] = [\text{HCl}] = 1.36 \times 10^{-2} \text{ M}, \text{pH} = -\log (1.36 \times 10^{-2}) = 2 - 0.1335 \approx 1.87$$

Question 49. The degree of ionization of a 0.1 M bromoacetic acid solution is 0.132. Calculate the pH of the solution and the PK_a of bromoacetic acid.

Answer:



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Question 50. The pH of 0.005 M codeine ($C_{18}H_{21}NO_3$) solution is 9.95. Calculate the ionization constant and pK_b .

Answer:

$$\begin{aligned} \text{Cod} + \text{H}_2\text{O} &\rightleftharpoons \text{CodH}^+ + \text{OH}^- \\ \text{pH} &= 9.95 \quad \therefore \text{pOH} = 14 - 9.95 = 4.05, \text{ i.e., } -\log [\text{OH}^-] = 4.05 \\ \text{or } \log [\text{OH}^-] &= -4.05 = \bar{5}.95 \quad \text{or } [\text{OH}^-] = 8.913 \times 10^{-5} \text{ M} \\ K_b &= \frac{[\text{CodH}^+][\text{OH}^-]}{[\text{Cod}]} = \frac{[\text{OH}^-]^2}{[\text{Cod}]} = \frac{(8.91 \times 10^{-5})^2}{5 \times 10^{-3}} = 1.588 \times 10^{-6} \\ pK_b &= -\log (1.588 \times 10^{-6}) = 6 - 0.1987 = 5.8 \end{aligned}$$

NCERT 11th Chemistry Chapter 7, class 11 Chemistry Chapter 7 solutions



Chapterwise NCERT Solutions for Class 11 Chemistry:

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- Chapter 2-Structure of Atom
- Chapter 3-Classification of Elements and Periodicity in Properties
- Chapter 4-Chemical Bonding and Molecular Structure
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- Chapter 6-Thermodynamics
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- Chapter 12-Organic Chemistry Some Basic Principles and Techniques
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