Q1.A liquid is in equilibrium with its vapours in a sealed container at a fixed temperature. The volume of the 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.

Q2. What is K_c for the following reaction in state of equilibrium?

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

Given: $[SO_2] = 0.6\ M;\ [O_2] = 0.82\ M;$ *and* $[SO_3] = 1.90\ M$

Answer:

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(1.9\ M) \times (1.9\ M)}{(0.6\ M) \times (0.6\ M) \times (0.82\ M)}$$

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

Q3. **Find the value of K_c for each of the following equilibria from the value of K**

(a) $2NOCl\ (g) \rightleftharpoons 2NO\ (g) + Cl_2\ (g)\ ;\ K_p = 1.8 \times 10^{-2}$ *atm at 500 K*

(b) $CaCO_3\ (s) \rightleftharpoons CaO\ (s) + CO_2\ (g)\ ;\ K_p = 167$ *atm at 1073 K.*

Answer:

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

The value of K_c can be calculated as follows:

(a) $2NOCl\ (g) \rightleftharpoons 2NO\ (g) + Cl_2$

$$K_p = 1.8 \times 10^{-2}\ \text{atm},$$

$$\Delta^{ng} = 3 - 2 = 1\ ;\ R = 0.0821\ \text{litre atm}\ K^{-1}\ \text{mol}^{-1}\ ;\ T = 500\ K$$

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

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

(b) $CaCO_3\ (s) \rightleftharpoons CaO\ (s) + CO_2\ (g)$

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

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

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

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

Q4.**Reaction between nitrogen and oxygen takes place as follows:**

$$2N_2\ (g) + O_2\ (g) \rightleftharpoons 2N_2O(g)$$

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 x 10^{-37}, determine the composition of the equilibrium mixture.

Answer: Let x moles of $N_2(g)$ take part in the reaction. According to the equation, x/2 moles of O_2 (g) will react to form x moles of $N_2O(g)$. The molar concentration per litre of different species before the reaction and at the equilibrium point is:

	$2N_2(g)$	+	$O_2(g)$	$\rightleftharpoons$	$2N_2O(g)$
Initial mole/litre:	$\frac{0.482}{10}$		$\frac{0.933}{10}$		Zero

Mole/litre at eqm. point: $\frac{0.482-x}{10}$ $\frac{0.933-\frac{x}{2}}{10}$ $\frac{x}{10}$

The value of equilibrium constant (2.0 x 10^{-37}) is extremely small. This means that only small amounts of reactants have reacted. Therefore, is 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 = $0.1 \times x = 0.1 \times 6.6 \times 10^{-20}$ mol L^{-1}
= **6.6×10^{-21} mol L^{-1}**

Q5.

At 450 K, $K_p = 2.0 \times 10^{10}$ bar^{-1} for the equilibrium reaction:
$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
What is K_c at this temperature?

Answer:

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

K_p = 2.0×10^{10} bar^{-1} ; R = 0.083 L bar K^{-1} mol^{-1} ; T = 450 K; $\Delta^{ng} = 2 - 3 = -1$
K_c = $(2.0 \times 10^{10}$ bar$^{-1}) \times [(0.083$ L bar K^{-1} mol$^{-1}) \times (450\ K)]^{-(-1)}$
= 7.47×10^{11} mol^{-1} L **= 7.47×10^{11} M^{-1}**

Q6. **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**

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ is 1.7×10^{-2}.

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

Answer:

The reaction is: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

$$\text{Concentration quotient } (Q_c) = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(8.13/20 \text{ mol L}^{-1})^2}{(1.57/20 \text{ mol L}^{-1}) \times (1.92/20 \text{ mol L}^{-1})^3}$$

= 2.38×10^3

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.

Q7. If l mole of H_20 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.

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

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 x 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.

follows:

$$\begin{array}{lcccc}
 & H_2O\,(g) \;+ & CO(g) \;\rightleftharpoons & H_2\,(g) \;+ & CO_2(g) \\
\text{Initial moles/litre} & \frac{1}{10} & \frac{1}{10} & 0 & 0 \\
\text{Mole/litre at the equilibrium point} & & & & \\
 & \frac{1-0.4}{10}=\frac{0.6}{10} & \frac{1-0.4}{10}=\frac{0.6}{10} & \frac{0.4}{10} & \frac{0.4}{10}
\end{array}$$

Applying law of chemical equilibrium,

$$\text{Equilibrium constant } (K_c) = \frac{[H_2(g)][CO_2(g)]}{[H_2O(g)][CO(g)]} = \frac{\left(\frac{0.4}{10}\text{mol L}^{-1}\right)\times\left(\frac{0.4}{10}\text{mol L}^{-1}\right)}{\left(\frac{0.6}{10}\text{mol L}^{-1}\right)\times\left(\frac{0.6}{10}\text{mol L}^{-1}\right)}$$

$$= \frac{0.16}{0.36} = \mathbf{0.44}$$

Q8. **What is the equilibrium concentration of each of the substances in the equilibrium when the initial concentration of ICl was 0.78 M?**

$$2ICl(g) \rightleftharpoons I_2\,(g) + Cl_2(g)\,;\, K_c = 0.14$$

Answer:

Suppose at equilibrium, the molar concentration of both I_2 (g) and Cl_2 (g) is x mol L^{-1}.

$$\begin{array}{lccc}
 & 2ICl(g) \;\rightleftharpoons & I_2(g) \;+ & Cl_2(g) \\
\text{Initial molar conc.} & 0.78\text{ M} & 0 & 0 \\
\text{Eqn. molar conc.} & (0.78-2x)\text{M} & x & x
\end{array}$$

$$K_c = \frac{[I_2(g)][Cl_2(g)]}{[ICl\,(g)^2]} = \frac{(x)\times(x)}{(0.78-2x)^2}$$

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

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

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

$$[I_2] = 0.167\text{ M}\,;\,[Cl_2] = \mathbf{0.167\ M}$$

Q9. **K =0.04 atm at 898 K for the equilibrium shown below. What is the equilibrium concentration ok C_2H_6 when it is placed in a flask at 4 atm pressure,and allowed to come to equilibrium.**

$$C_2H_6(g) \rightleftharpoons C_2H_4(g) + H_2(g)$$

Answer:

The equilibrium in the reaction is:

$$\begin{array}{lccc}
 & C_2H_6(g) \;\rightleftharpoons & C_2H_4(g) \;+ & H_2(g) \\
\text{Initial pressure:} & 4\text{ atm} & 0 & 0 \\
\text{Eqn. pressure:} & (4-p)\text{ atm} & p\text{ atm} & p\text{ atm}
\end{array}$$

$$K_p = \frac{p_{C_2H_4}\times p_{H_2}}{p_{C_2H_6}} \quad \text{or} \quad 0.04 = \frac{p^2}{(4-p)}$$

$$p^2 = 0.04\,(4-p) \quad \text{or} \quad p^2 + 0.04\,p - 0.16 = 0$$

$$p = \frac{(-0.04)\pm\sqrt{0.0016-4\,(-0.16)}}{2}$$

$$= \frac{(-0.04)\pm 0.8}{2} = \frac{0.76}{2} = 0.38$$

Equilibrium pressure or concentration of C_2H_6 = (4 – 0.38) = **3.62 atm.**

Q10. 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 0.5×10^{-1} 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 mol

Molar concentration of PCl_5 at equilibrium = 0.05 mol

∴ Moles of PCl_5 decomposed = (x – 0.05) mol

Moles of PCl_3 formed = (x – 0.05) mol

Moles of Cl_2 formed = (x – 0.05) 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,

$$K_c = \frac{[PCl_3][Cl_2]}{[PCl_5]}; \quad 0.0083 = \frac{(x - 0.05) \times (x - 0.05)}{0.05}$$

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

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

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

The molar concentration per litre of PCl_3 at eqm. = 0.07 – 0.05 = **0.02 mol**

The molar concentration per litre of Cl_2 at eqm. = 0.07 – 0.05 = **0.02 mol.**

Q11. Calculate (a) ΔG^{-} and (b) the equilibrium constant for the formation of $N0_2$ from NO and 0_2 at 298 K

$NO(g) + 1/2\ O_2(g) \rightleftharpoons NO_2(g)$ where

$\Delta_f G^{\ominus}(NO_2) = 52.0$ kJ/mol, $\Delta_f G^{\ominus}(NO) = 87.0$ kJ/mol, $\Delta_f G^{\ominus}(O_2) = 0$ kJ/mol,

Answer:

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

$$\Delta G^{\ominus} = \Delta_f G^{\ominus}(NO_2) - [\Delta_f G^{\ominus}(1/2O_2)]$$
$$= 52.0 - (87 + 0) = -35 \text{ kJ mol}^{-1}$$

Step II. Calculation of K_c

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

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

$$K_c = \text{Antilog } 6.314 = \mathbf{1.36 \times 10^6}.$$

Q12.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.

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

If a reaction vessel at 400°C is charged with an equimolar mixture of CO and steam so that $P_{CO} = PH_2O = 4.0$ bar, what will be the partial pressure of H_2 at equilibrium? $K_p = 0.1$ at 400°C.

Answer:

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

	$CO(g)$	+	$H_2O(g)$	$\rightleftharpoons$	$CO_2(g)$	+	$H_2(g)$
Initial pressure:	4.0 bar		4.0 bar		0		0
Eqm. pressure:	$(4-p)$bar		$(4-p)$ bar		p bar		p bar

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

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

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

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

Q13 **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 demate 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.

Q14. **. The pH of a sample of vinegar is 3.76. Calculate the concentration of hydrogen ion in it.**

Answer: pH = – log $[H^+]$ or log $[H^+]$ = – pH = – 3.76 = 4.24

.-. $[H^+]$ = Antilog 4.24 = 1.738 x 10^{-4} = 1.74 x 10^{-4} M

Q15. **The ionization constant of phenol is 1.0 x 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:

$$C_6H_5OH \rightleftharpoons C_6H_5O^- + H^+$$

Initial $\quad 0.05\ M$

After disso. $\quad 0.05 - x \qquad x \quad x$

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

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

In presence of 0.01 C_6H_5ONa, suppose y is the amount of phenol dissociated, then at equilibrium

$$[C_6H_5OH] = 0.05 - y \simeq 0.05,$$
$$[C_6H_5O^-] = 0.01 + y \simeq 0.01\ M, \ [H^+] = y\ M$$

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

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

Q16. **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:

$$HA \rightleftharpoons H^+$$

$$pH = -\log[H^+] \quad \text{or} \quad \log[H^+] = -4.15 = \bar{5}.85$$

$$\therefore \quad [H^+] = 7.08 \times 10^{-5}\ M = 7.08 \times 10^{-5}\ M$$

$$[A^-] = [H^+] = 7.08 \times 10^{-5}\ M$$

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(7.08 \times 10^{-5})(7.08 \times 10^{-5})}{10^{-2}} = \mathbf{5.0 \times 10^{-7}}$$

$$pK_a = -\log K_a = -\log(5.0 \times 10^{-7}) = 7 - 0.699 = \mathbf{6.301}$$

Q17. 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:

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

(*b*) $NaOH + aq \rightarrow Na^+ + OH^-$

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

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

(*d*) $KOH + aq \rightarrow K^+ + OH^-$,

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

Q18 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 $Ca(OH)_2$ 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) l mL of 13.6 M HCl is diluted with water to give 1 litre of the solution.

Answer:

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

$$[OH^-] = [TlOH] = 4.52 \times 10^{-3}\text{ M}$$
$$[H^+] = 10^{-14}/(4.52 \times 10^{-3}) = 2.21 \times 10^{-12}\text{ M}$$
$$\therefore \quad pH = -\log(2.21 \times 10^{-12}) = 12 - (0.3424) = 11.66$$

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

$$Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-$$
$$[OH^-] = 2[Ca(OH)_2] = 2 \times (8.11 \times 10^{-3})\text{ M} = 16.22 \times 10^{-3}\text{ M}$$
$$pOH = -\log(16.22 \times 10^{-3}) = 3 - 1.2101 = 1.79$$
$$pH = 14 - 1.79 = \mathbf{12.21}$$

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

$$[OH^-] = 3.75 \times 10^{-2}\text{ M}$$
$$pOH = -\log(3.75 \times 10^{-2}) = 2 - 0.0574 = 1.43$$
$$pH = 14 - 1.43 = \mathbf{12.57}$$

(*d*) $M_1V_1 = M_2V_2 \therefore 13.6\text{ M} \times 1\text{m L} = M_2 \times 1000\text{ mL} \quad \therefore M_2 = 1.36 \times 10^{-2}$ M

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

Q19. 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:

	$CH_2(Br)COOH$	$\rightleftharpoons$	$CH_2(Br)COO^-$	+	H^+
Initial conc.	C		0		0
Conc. at eqm.	C – Cα		C α		C α

$$K_a = \frac{C\alpha . C\alpha}{C(1-\alpha)} = \frac{C\alpha^2}{1-\alpha} \simeq C\alpha^2 = 0.1 \times (0.132)^2 = 1.74 \times 10^{-3}$$
$$pK_a = -\log(1.74 \times 10^{-3}) = 3 - 0.2405 = \mathbf{2.76}$$
$$[H^+] = C\alpha = 0.1 \times 0.132 = 1.32 \times 10^{-2}\text{ M}$$
$$pH = -\log(1.32 \times 10^{-2}) = 2 - 0.1206 = \mathbf{1.88}$$

Q20. The pH of0.005 M codeine ($C_{18}H_{21}N0_3$) solution is 9.95. Calculate the ionization constant and PK$_b$.

Answer:

$$Cod + H_2O \rightleftharpoons Cod\,H^+ + OH^-$$
$$pH = 9.95 \quad \therefore pOH = 14 - 9.95 = 4.05, \quad i.e., -\log[OH^-] = 4.05$$
or $\log[OH^-] = -4.05 = \bar{5}.95$ or $[OH^-] = 8.913 \times 10^{-5}$ M

$$K_b = \frac{[CodH^+][OH^-]}{[Cod]} = \frac{[OH^-]^2}{[Cod]} = \frac{(8.91 \times 10^{-5})^2}{5 \times 10^{-3}} = \mathbf{1.588 \times 10^{-6}}$$
$$pK_b = -\log(1.588 \times 10^{-6}) = 6 - 0.1987 = \mathbf{5.8}$$