Chapter – 7 (Equilibrium)

Exercise Questions:

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

a. **What is the initial effect of the change on vapour pressure.**
b. **How do rates of evaporation and condensation change initially?**
c. **What happens when equilibrium is restored finally and what will be the final vapour pressure?**

Answer:

a.) On increasing the volume of the container, the vapour pressure will initially decrease because the same amount of vapours are distributes over a large space.

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

c.) Finally, equilibrium will be reversed when the rates of forward and backward processes becomes 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 Kc for the following equilibrium when the equilibrium concentration of each substance is : (SO2) = 0.60M, (O2) = 0.82 M and (SO3) = 1.90M?

2SO2(g) + O2(g) –> 2SO3(g)

Answer:

As per the question,

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

$K_c = [SO3]^2 / [SO2]^2[O2]$

$= (1.9)^2M^2 / (0.6)^2(0.82)M^3$

$= 12.229\ M^{-1}$(approx.)

Hence, K for the equilibrium is 12.229 / M.

Question :3 At a certain temperature and total pressure of 10*5 PA , iodine vapour contains 40% by volume of I atoms

I2(g) -> 2I(g)

Calculate Kp for the equilibrium.

Answer:

Kp value at equilibrium for reaction,

I2 (g) → 2I (g) is given as,

Given, 40% volume is occupied by I atoms.

∴ 60 % volume is occupied by I2.

Let V be total volume, then volume occupied by I and I2 are 0.4V and 0.6V respectively.

Partial pressure of I2 = $(60/100) \times 10^5$ = 60 kPa

Partial pressure of I = $(40/100) \times 10^5$ = 40 kPa

Substituting in equation for Kp,

Kp = = 2.67×10^4 KPa

Question :4 Write the expression for the equilibrium constant, Kc for each of the following reactions :

I. **2NOCl (g) -> 2NO (g) + Cl2 (g)**

II. **2Cu(NO3)2 (s) -> 2CuO(s) + 4NO2(g) + O2 (g)**

III. **CH3COOC2H5 (aq) + H2O (l) -> CH3COOH (aq) + C2H5OH (aq)**

IV. **Fe*3+ (aq) + 3OH*- (aq) ->Fe(OH) (s)**

V. **I2 (s) + 5 F2 -> 2 IF5**

Answer:

i.) $K_c = [NO(g)]^2[Cl2(g)] / [NOCl(g)]$

ii.) $K_c = \frac{[Cu(s)]^2[NO2(g)]^4[O2(g)]}{[Cu(NO3)_{2(g)}]^2}$

$= [NO2(g)]^4[O2(g)]$

iii.) $K_c = \frac{[CH3COOH(aq)][C2H5OH(aq)]}{[CH3COOC2H5(aq)][H2O(l)]}$

$= \frac{[CH3COOH(aq)][C2H5OH(aq)]}{[CH3COOC2H5(aq)]}$

iv.) $K_c = \frac{[Fe(OH)_{3(aq)}]}{[Fe^{3+}(aq)][OH^-(aq)]}$

$= 1 / [Fe^{3+}(aq)][OH^-(aq)]$

v.) $K_C = \frac{[IF5]^2}{[I_{2(s)}][F2]^5}$

$= [IF5]^2 / [F2]^5$

Question :5 Find out the value of Kc for each of the following equilibria from the value of Kp :

- **I. 2NOCl (g) -> 2 NO (g) + Cl2 (g); Kp = 1.8 x 10*-2 at 500K**
- **II. CaCO3 (s) -> CaO (s) + CO2 (g); Kp = 167 at 1073 K**

Answer:

(i) 2NOCl(g) <--> 2NO(g) + Cl2(g) ; Kp = 1.8 × 10^-2 at 500K

we know,

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

Δng = np - nr

= (2 + 1) - (2) = 1

Now, 1.8 ×10^-2 = $Kc(RT)^1$

Kc = 1.8 × 10^-2/(0.0821 × 500)

Kc = 4.38 × 10^-4

(ii) CaCO3(s) → CaO(s) + CO2(g); Kp=167 at 1073K

Δng = np - nr

= 1 - 0 = 1

Now, Kp = $Kc(RT)^1$

167 = Kc (0.0821 × 1073)

Kc = 167/0.0821 × 1073

Kc =1.89

Question :6 For the following equilibrium, Kc = 6.3 x 10*14 at 1000 K

NO (g) + O2 (g) -> NO2 (g) + O2(g)

Both the forward and reverse reaction in the equilibrium are elementary bimolecular reactions. What is Kc for the reverse reaction?

Answer:

For the reverse reaction, $K_c = 1 / K_c$

$= 1 / 6.3 \times 10^{14}$

$= 1.59 \times 10^{-15}$

Question :7 Explain why pure liquids and solids can be ignored while writing the equilibrium constant expression?

Answer:

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

Mole concentration = Number of moles / Volume

= Mass / molecular mass / Volume
= Mass / volume x Molecular mass
= Density / Molecular mass
Though the density of the solid and pure liquid is fixed and molar is also fixed.
: Molar concentration are constant.

Question :8 Reaction between N2 and O2 takes place as follows :

2N2 (g) + O2 (g) -> 2N2O (g)

If a mixture of 0.482 MIL N2 and 0.933 O2 is placed in a 10 L reaction vessel and allowed to form N2O at a temperature for which Kc = 2.0 x 10*-37, determine the composition of equilibrium mixture.

Answer:

Let the concentration of N2O at equilibrium be x. The given reaction is

	2N2 (g)	+	O2 (g)	↔	2N2O (g)
Intial Concentration:	0.482 mol		0.933 mol		0 mol
At equilibrium	(0.482-x) mol		(0.933 -x) mol		x mol

Therefore, at equilibrium, in the 10 L vessel:
N2 = 0.482-x / 10
O2 = 0.933-x/2 / 10
N2O = x / 10

The value of equilibrium constant i.e. Kc = 2.0 × 10-37 is very small. Therefore, the amount of N2 and O2 reacted is also very small. Thus, x can be neglected from the expressions of molar concentrations of N2 and O2.

Then,
N2 = 0.482/10 = 0.0482 molL-1 and O2 = 0.933/10 = 0.0933 molL-1
Now,
$K_c = [N2O(g)]^2 / [N2(g)]^2[O2(g)]$
$2.0 \text{ x } 10^{-37} = 0.01x^2 / (0.482 / 10)^2 \text{ x } (0.933 / 10)$
$2.0 \text{ x } 10^{-37} = 0.01x^2 / 2.1676 \text{ x } 10^{-4}$
$x^2 = 43.352 \text{ x } 10^{-40}$
or $x = 6.6 \text{ x } 10^{-20}$
[N2O] = x / 10
$= 6.6 \text{ x } 10^{-20} / 10$
$= 6.6 \text{ x } 10^{-21}$

Question :9 Nitric oxide reacts with Br2 and gives nitrosyl bromide as per reaction given below:

$2NO (g) + Br_2 \rightarrow 2NOBr (g)$

When 0.087 mol of NO and 0.0437 mol of Br2 are mixed in a closed container at constant temperature, 0.0518 mol of NOBr is obtained at equilibrium. Calculate equilibrium amount of NO and Br2.

Answer:

The given equation is:

$$2NO(g) + Br_2(g) \rightarrow 2NOBr(g)$$

At t = 0	0.087	0.0437	0
At eql	(0.087 – 2x)	(0.0437 – x)	2x

Given, moles of NOBr at equilibrium = 0.518

e.g., 2x = 0.0518

x = 0.0259

Moles of NO at equilibrium = 0.087 - 2x

= 0.087 - 2 × 0.0259

= 0.087 - 0.0518

= 0.0352 mol

Moles of Br2 at equilibrium = 0.0437 - x

= 0.0437 - 0.0259

= 0.0178 mol

Question :10 At 450K Kp = 2.0 x 10*10/ bar for the given reaction at equilibrium

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

What is Kc at this temperature?

Answer:

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

For the given reaction:

$\Delta n = 2 – 3 = -1$

T = 450K

R = 0.0831 bar L bar / K/ mol

$K_P = 2.0 \times 10^{10}$ / bar

We know,

$K_p = K_c (RT)^\wedge \Delta n_g$

Kc = Kp/(RT)^(-1) [because , $\Delta ng = -1$]

= Kp(RT)

= 2 × 10^10 × 0.0831 × 450

= 7.479×10^{11} L/mol

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

2HI (g) -> H2 (g) + I2 (g)

Answer:

The initial concentration of HI is 0.2 atm. At equilibrium, it has a partial pressure of 0.04 atm. Therefore, a decrease in the presence of HI is 0.2 – 0.04 = 0.16. The given reaction is:

	$2HI(g) \rightarrow$	$H2(g)$ +	$I2(g)$
Initial conc.	0.2 atm	0	0
At equilibrium	0.4 atm	0.16 / 2	0.16 / 2
		0.08 atm	0.08 atm

Therefore,

$K_p = p_{H2} \text{ x } p_{I2} / p^2_{HI}$

$= 0.08 \text{ x } 0.08 / (0.04)^2$

$= 0.0064 / 0.0016$

$= 4.0$

Hence, the value of K_p for the given equilibrium is 4.0 .

Question :12 A mixture of 1.57 Mol of N2, 1.92 mol of H2 and 8.13 Molly of NH3 is introduced into a 20L reaction vessel at 500 K. At this temperature, the equilibrium constant, Kc for the reaction N2 (g) + 3H2 (g) -> 2 NH3 (g) is 1.7 x 10*2. Is the reaction mixture at equilibrium? If not, what is the direction of the net reaction?

Answer:

The given reaction is:

$N2(g) + 3H2(g) \rightarrow 2NH3(g)$

The given concentration of various species is

$[N2] = 1.57 / 20 \text{ mol / L}$

$[H2] = 1.92 / 20 \text{ mol /L}$

$[NH3] = 8.31 / 20 \text{ mol /L}$

Now, reaction quotient Q_c is:

$Q = [NH3]^2 / [N2][H2]^3$

$= (8.13 / 20)^2 / (1.57 / 20)(1.92 / 20)^3$

$= 2.4 \text{ x } 10^3$

Since, $Q_c \neq K_c$, the reaction mixture is not at equilibrium.

Again, $Q_c > K_c$. Hence, the reaction will proceed in the reverse direction.

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

Kc = $[NH3]^4[O2]^5$ / $[NO]^4[H2O]^6$

Write the balanced chemical equation corresponding to this expression.

Answer:

The balanced chemical equation corresponding to the given expression can be written as:

4NO(g) + 6H2O(g) → 4NH3(g) + 5O2(g)

Question :14 One mole of H2O and one mole of CO are taken in 10 L vessel and heated to 725 K. At equilibrium 40% of water reacts with CO according to the equation,

H2O (g) + CO (g) -> H2 (g) + CO2 (g)

Calculate the equilibrium constant for the reaction.

Answer:

The given reaction is :

H2O(g) + CO(g) → H2(g) + CO2(g)

Compound	H2O	CO	H2	CO2
Initial Conc.	0.1 M	0.1 M	0	0
Equilibrium Conc.	0.06 M	0.06 M	0.04 M	0.04 M

Therefore, the equilibrium constant for the reaction,

K_c = ([H2][CO2]) / ([H2O][CO])

= (0.4 x 0.4) / (0.6 x 0.6)

= 0.444

Question :15 At 700 K, equilibrium constant for the reaction :

H2 (g) + I2 (g) -> 2 HI (g)

Is 54.8. If 0.5 mol / L of HI (g) is present at equilibrium at 700 K, what are the concentration of H2 (g) and I2 (g) assuming that we initially started with HI (g) and allowed it to reach equilibrium at 700 K?

Answer:

H2(g) + I2(g) → 2HI(g) ; Kc = 54.8

when reaction will be reverse

2HI(g) → H2(g) + I2(g) , then, equilibrium constant of reverse reaction is K'c = 1/Kc = 1/54.8

A/C to question,

[HI] at equilibrium = 0.5 mol/L

Let at equilibrium conc. of [H2] = [I2] = x mol/L

K_c = [H2][I2] / $[HI]^2$

$1/54.8 = x \, . \, x/(0.5)^2$

$1/54.8 = x^2/(0.5\,)^2$

$x^2 = 0.25/54.8 = 0.00456$

x = 0.0675 M

[H2] = [I2] = 0.0675 M

Question :16 What is the equilibrium concentration of each of the substances in the equilibrium when the initial concentration of ICl was 0.78 M?

2 ICl (g) -> I2 (g) + Cl2 (g); Kc = 0.14

Answer:

The given reaction is:

	2ICl (g) →	I2 (g) +	Cl2 (g)
Initial conc.	0.78 M	0	0
At equilibrium	(0.78 - 2x) M	x M	x M

Now we can write, Kc = I2 x Cl2 / (ICl)2

$= x \text{ X } x / (0.78 – 2x)^2 = 0.14$

$x^2/ (0.78 – 2x)^2 = 0.14$

x / 0.78 – 2x = 0.374

x = 0.292 – 0.748x

1.748x = 0.292

x = 0.167

Hence, at equilibrium,

[H2] = [I2] = 0.167 M

[HI] = (0.78 – 2 x 0.167)M

= 0.446 M.

Question :17 Kp = 0.04 ATM at 899 K for the equilibrium shown below. What is the equilibrium concentration of C26 when it is placed in a flask at 4.0 ATM pressure and allowed to come to equilibrium?

C2H6 (g) -> C2H4 (g) + H2 (g)

Answer:

	C2H6(g) →	C2H4(g) +	H2(g)
At t = 0	4 atm	0	0
At equil.	(4 – p) atm	p	p

We know,

$$K_p = \frac{P_{C2H4}.P_{H2}}{P_{C2H6}}$$

$K_p = P. P/(4 - P)$
$0.04 = P^2/(4 - P)$
$1/25 = P^2/(4 - P)$
$4 - P = 25P^2$
$25P^2 + P - 4 = 0$
$P = \{-1 \pm \sqrt{(1 + 400)}\}/50$
$P = \{ -0.02 \pm 0.4\}$
$P = 0.38$
Hence, pressure of $C_2H_6 = 4 - p = 4 - 0.38 = 3.62$ atm

Question :18 Ethyl acetate is formed by the reaction between ethanol and acetic acid and the equilibrium is represented as :

$CH_2COOH (l) + C_2H_5OH (l) \rightarrow CH_3COOC_2H_5 (l) + H_2O (l)$

- **I. Write the concentration ratio Qc, for this reaction.**
- **II. At 293 K, if one starts with 1.00 mol of acetic acid and 0.18 mol of ethanol, there is 0.171 mol of Ethyl acetate in the final equilibrium mixture Calculate the equilibrium constant.**
- **III. Starting with 0.5 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 sometime. Has equilibrium been reached?**

Answer:

$CH_3COOH (l) + C_2H_5OH (l) \rightarrow CH_3COOC_2H_5 (l) + H_2O (l)$

(i) Reaction quotient (Qc) = $[CH_3COOC_2H_5][H_2O] / [C_2H_5OH][CH_3COOH]$

Here, H_2O is not excess that's why its concentration isn't constant.

(ii)

	$CH_3COOH (l)$	+	$C_2H_5OH (l)$	→	$CH_3COOC_2H_5 (l)$	+	$H_2O (l)$
At t = 0	1 mol		0.180 mol		0		0
At equil.	(1 – x) mol		(0.180 – x) mol		x		x

Given,
$[CH_3COOC_2H_5]$ at equilibrium = 0.171 mol = x
So, x = 0.171 mol
Hence, $[CH_3COOH] = 1 - x = 1 - 0.171 = 0.829$
$[C_2H_5OH] = 0.180 - x = 0.180 - 0.171 = 0.009$
$[CH_3COOC_2H_5] = [H_2O] = x = 0.171$
Now, $K_c = [CH_3COOC_2H_5][H_2O]/[CH_3COOH][C_2H_5OH]$
$= 0.171 \times 0.171/0.829 \times 0.009$

$= 3.919 \approx 3.92$

Hence, Kc = 3.92

(iii) Similarly,

At time t ,

$[CH_3COOH] = 1 - x = 0.5 - 0.214 = 0.786$ mol

$[C_2H_5OH] = 0.5 - x = 0.5 - 0.214 = 0.286$ mol

$[CH_3COOC_2H_5] = x = 0.214$ mol

$[H_2O] = x = 0.214$ mol

Reaction quotient (Qc) = $[CH_3COOC_2H_5][H_2O]/[CH_3COOH][C_2H_5OH]$

$= 0.214 \times 0.214/0.786 \times 0.286$

$= 0.2037 \approx 0.204$

Here, Qc ≠ Kc hence, equilibrium has not been reached.

Question :19 A sample of pure PCl_5 was introduced into an evacuated vessel at 473 K. After equilibrium was attained, concentration of PCl_5 was found to be 0.5 x 10*-1 mol/L . If value of Kc is 8.3 x 10*-3 ,what are the concentrations of PCL_3 and Cl_2at equilibrium?

PCl_5 (g) -> PCl_3 (g) + Cl_2 (g)

Answer:

	$PCl_5(g)$ →	PCl_3 (g) +	Cl_2 (g)
At equil.	0.05	x	x

Given,

$[PCl_5]$ at equilibrium = 0.05 M

$Kc = 8.3 \times 10^{-3}$

But we know,

$Kc = [PCl_3][Cl_2]/[PCl_5]$

$8.3 \times 10^{-3} = x . x /0.05$

$8.3 \times 10^{-3} \times 0.05 = x^2$

$x^2 = 0.415 \times 10^{-3} = 4.15 \times 10^{-4}$

Now, take square root both sides,

$x = 2.04 \times 10^{-2}$ M

Hence, $[PCl_3] = [Cl_2] = 2.04 \times 10^{-2}$ M

Question :20 One of the reaction that takes place in producing steel from iron or is the reduction of iron (II) oxide by carbon monoxide to give iron metal and CO_2.

FeO(s) + CO(g) -> Fe (s) + CO_2 (g); Kp = 0.265 ATM at 1050k

What are the equilibrium partial pressures of CO and CO2 at 1050 K if the initial pressure are:

Answer:

FeO(s) + CO (g) $\rightarrow$ Fe (s) + CO2 (g); $K_p = 0.265$ at 1050K

At t = 0 1.4 atm 0.8 atm

Qp = Pco2/Pco

= 0.8/1.4 = 0.571

We can see that, Kp < Qp , the reaction will go in reverse direction. Due to this pressure of CO2 will decreases and that of CO will increase to attain equilibrium .

Hence, Pco2 = (0.8 - P) and Pco = (1.4 + p)

Kp = Pco2/Pco

0.265 = (0.8 - p)/(1.4 + p)

0.265(1.4 + p) = 0.8 - p

0.265p + 0.265 × 1.4 = 0.8 - p

1.265p = 0.8 - 0.371 = 0.429

p = 0.429/1.265 = 0.339 atm

Hence, at equilibrium,

Pressure of CO2 (Pco2) = 0.8 - 0.339 = 0.461 atm.

Pressure of CO (Pco) = 1.4 + 0.339 = 1.739 atm

Question :21 Equilibrium constant, Kc for the reaction

N2 (g) + 3H2 (g) -> 2 NH3 of) at 500 K is 0.061

At a particular time, the analysis shows that composition of the reaction mixture is 3.0 mol/L H2 and 0.5 mol/L NH3. Is the reaction at equilibrium? If not in which direction does the reaction tend to proceed to reach equilibrium?

Answer:

N2 (g) + 3H2(g) $\rightarrow$ 2NH3

At a particular time: 3.0 mol /L 2.0 mol/L 0.5 mol/L

So,

$Q_c = [NH3]^2 / [N2][H2]^3$

$= (0.5)^2 / (3.0)(2.0)^3$

= 0.0104

It is given that $K_c = 0.061$

:$Q_c \neq K_c$, the reaction is not at equilibrium.

:$Q_c < K_c$,

The reaction proceed in the forward direction to reach at equilibrium.

Question :22 Bromine monochloride, BrCl decomposes into bromine and chlorine and reaches the equilibrium :

2BrCl (g) -> Br2 (g) + Cl2 (g)

For which Kc = 32 at 500 K. If initially pure BrCl is present at a concentration of 3.3 x 10*-3 mol/L ,what is it's molar concentration in the mixture at equilibrium?

Answer:

Let the amount of bromine and chlorine formed at equilibrium be x. The given reaction is:

	$2BrCl (g) \rightarrow$	$Br_2 (g)$ +	$Cl_2 (g)$
Initial Conc.	3.3x10-3	0	0
At equilibrium	3.3x10-3 -2x	x	x

Now, we can write,

$Kc = [Br2][Cl2] / [BrCl]^2$

$(x) \times (x) / (3.3\times10^{-3} -2x)^2 = 32$

$x / (3.3\times10^{-3} -2x) = 5.66$

$x = 18.678 \times10^{-3} - 11.32x$

$x + 11.32x = 18.678 \times10^{-3}$

$12.32x = 18.678 \times10^{-3}$

$x = 1.5 \times 10^{-3}$

Therefore, at equilibrium,

$[BrCl] = 3.3\times10^{-3} - (2 \times 1.5 \times 10^{-3})$

$= 3.3\times10\text{-}3 - 3.0 \times10^{-3}$

$= 0.3 \times 10^{-3}$

$= 3.0 \times 10^{-4}$ mol L-1

Question :23 At 1127 K and 1 ATM pressure, a gaseous mixture of CO and CO2 in equilibrium with solid carbon has 90.55% CO by mass

C (s) + CO2 (g) -> 2CO(g)

Calculate Kc for this reaction at the above temperature.

Answer:

Let weight of mixture is 100g , then 90.55% CO by mass means 90.55g CO and 9.45g CO2 present in the mixture.

No of moles of CO = Given weight of CO/molar mass of CO

= 90.55/28 = 3.234 mol [molar mass of CO = 28 g/mol]

No of moles of CO2 = given weight of CO2/molar mass of CO2

= 9.45/44 = 0.215 mol [molar mass of CO2 = 44g/mol]

Mole fraction of $CO(x_{CO}) = n_{CO} / n_{CO} + n_{CO2}$

= 3.234 / 3.234 + 0.215

= 0.938

Mole fraction of CO2 = $(1 - x_{CO}) = 1 - 0.938 = 0.062$
Now, partial pressure of CO = mole fraction of CO × total pressure
$= 0.938 \times 1$ atm
$= 0.938$ atm
Similarly, partial pressure of CO2 = mole fraction of CO2 × total pressure
$= 0.062 \times 1$ atm
$= 0.062$ atm
Now,
$C(s) + CO_2(g) = 2CO(g)$
$K_p = P_{[CO]}^2 / P_{[CO2]}$
$Kp = (0.938)^2/(0.062) = 14.19$
Δng = number of mole of gaseous products- number of mole of gaseous reactants
$= 2 - 1 = 1$
Now, $K_P = K_C (RT)^{\wedge} \Delta n_g$
$Kc = Kp/RT$
$= 14.19/0.0821 \times 1127$ [T = 1127 K and R = 0.0821 L.atm/K/mol]
$= 0.15336 \approx 0.153$

Question :24 Calculate a. And b. The equilibrium constant for the formation of NO2 from NO and O2 at 298 K
NO (g) + ½ O2 (g) -> NO2 (g)
Where,
$\Delta_f G^0$ (NO2) = 52.0 kJ / mol
$\Delta_f G^0$ (NO) = 87.0 kJ / mol
$\Delta_f G^0$ (O2) = 0 kJ / mol

Answer:
a.) For the given reaction, we have
$\Delta G^0 = \Delta G^0$(Products) - ΔG^0(Reactants)
$\Delta G^0 = 52.0 - (87.0 + 0)$
$= -35.0$ KJ / mol
b.) We know that,
$\Delta G^0 = RT \log K_c$
$\Delta G^0 = 2.303 RT \log K_c$
$K_c = -35.0 \times 10^{-3} / -2.303 \times 8.314 \times 298$
$= 6.134$
$:K_c = \text{antilog}(6.134)$
$= 1.36 \times 10^6$

Therefore, the equilibrium constant for the given reaction Kc is 1.36×10^6.

Question :25 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?

a. $PCl_5\ (g) \rightarrow PCl_3\ (g) + Cl_2\ (g)$

b. $CaO\ (s) + CO_2\ (g) \rightarrow CaCO_2\ (s)$

c. $3Fe\ (s) + 4H_2O\ (g) \rightarrow Fe_3O_4\ (s) + 4\ H_2\ (g)$

Answer:

a.) The number of moles of reaction products will increase. According to Le Chatelier's principle, if pressure is decreased, then the equilibrium shifts in the direction in which the number of moles of gases is more. In the given reaction, the number of moles of gaseous products is more than that of gaseous reactants. Thus, the reaction will proceed in the forward direction. As a result, the number of moles of reaction products will increase.

b.) The number of moles of reaction products will decrease.

c.) The number of moles of reaction products remain the same.

Question :26 Which of the following reactions will get affected by increasing the pressure. Also, mention whether change will cause the reaction to go into forward or backward direction.

a. $COCl_2\ (g) \rightarrow CO\ (g) + Cl_2\ (g)$

b. $CH_4\ (g) + 2S_2\ (g) \rightarrow CS_2\ (g) + 2\ H_2S\ (g)$

c. $CO_2\ (g) + C\ (S) \rightarrow 2CO(g)$

d. $2H_2\ (g) + CO(g) \text{->} CH_3OH\ (g)$

e. $CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$

f. $4\ NH_3\ (g) + 5O_2\ (g) \rightarrow 4NO\ (g) + 6\ H_2O\ (g)$

Answer:

When pressure is increased:

The reaction given in (i), (iii), (iv), (v) and (vi) will get affected.

Since the number of moles of gaseous reactants is more than that of gaseous products; the reaction given in (iv) will proceed in the forward direction.

Since the number of moles of gaseous reactants is less than that of gaseous products, the reactions are given in (i), (iii), (v) and (vi) will shift in the backward direction.

Question :27 The equilibrium constant for the following reactions is 1.6 x 10*5 at 1024K

H_2 (g) + Br_2 (g) → 2HBr(g)

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

Answer:

$H_2 + Br_2 \rightarrow 2HBr$; $K_p = 1.6 \times 10^5$ at 1024K

When reaction will be reverse ,

$2HBr \rightarrow H_2 + Br_2$; $K'_p = 1/K_p = 1/1.6 \times 10^5$ at 1024K

At initial time , pressure of HBr is 10 bar

Pressure of H_2 and Br_2 are 0 bar .

At equilibrium , pressure of HBr is (10-x) bar

Pressure of H_2 and Br_2 are x/2 bar.

$K'_p = P(H_2) \times P(Br_2) / P^2(HBr)$

$= (x/2).(x/2)/(10-x)^2$

$1/1.6 \times 10^5 = x^2/4(10-x)^2$

$0.625 \times 10^{-5} = x^2/4(10-x)^2$

Because values of K'p is so small so, $(10-x) \approx 10$

$0.625 \times 10^{-5} = x^2/4 \times 100$

$625 \times 4 \times 10^{-5} = x^2$

$x = 0.05$

$x/2 = 0.025 = 2.5 \times 10^{-2}$ bar

Hence, $P(H_2) = P(Br_2) = 2.5 \times 10^{-2}$ bar

So, $P(HBr) = 10-x = 10-0.05 \approx 10$bar

Question :28 Dihydrogen gas is obtained from natural gas by partial oxidation with steam as per following endothermic reaction:

CH_4 (g) + H_2O (g) → CO (g) + 3 H_2 (g)

a. **Write as expression for Kp for the above reaction.**
b. **How will the values of Kp and composition of equilibrium mixture be affected by**
   I. **Increasing the pressure**
   II. **Increasing the temperature**
   III. **Using a catalyst?**

Answer:

a.) For the given reaction,

$K_p = p_{CO} \times p^3_{H2} / p_{CH4} \times p_{H2O}$

b.) i.) According to Le Chatelier's principle, the equilibrium will shift in the backward direction.
ii.) According to Le Chatelier's principle, as the reaction is endothermic, the equilibrium will shift in the forward direction.
iii.) The equilibrium of the reaction is not affected by the presence of a catalyst. A catalyst only increase the rate of a reaction. Thus, equilibrium will be attained quickly.

Question :29 Describe the effect of:

a. **Addition of H2**
b. **Addition of CH3OH**
c. **Removal of CO**
d. **Removal of CH3OH**

On the equilibrium of the reaction :

$2H_2 (g) + CO (g) \rightarrow CH_3OH (g)$

Answer:

a.) According to Le Chatelier's principle, on the addition of H2, the equilibrium of the given reaction will shift in the forward direction.
b.) On addition of CH3OH, the equilibrium will shift in the backward direction.
c.) On removing CO, the equilibrium will shift in the backward direction.
d.) On removing CH3OH, the equilibrium will shift in the forward direction.

Question :30 At 473 K, equilibrium constant Kc for decomposition of phosphorous pentachloride, PCl5 is 8.3 x 10*-3. If decomposition is depicted as:

$PCl_5 (g) \rightarrow PCl_3 (g) + Cl_2 (g)$

a. **Write an expression for Kc for the reaction**
b. **What is the value of Kc for the reaction reverse reaction at the same temperature?**
c. **What would be the effect on Kc if**
    I. **More PCl5 is added**
    II. **Pressure is increased**
    III. **The temperature is increased**

Answer:

a.) Kc = [PCl3(g)]{Cl2(g)] / [PCl3(g)]
b.) Value of Kc for the reverse reaction at the same temperature is:

$K'_c = 1 / K_c$

$= 1 / 8.3 \times 10^{-3} = 1.2048 \times 10^2$

= 120.48

c.) (i) Kc would remain the same because in this case, the temperature remains the same.

(ii) Kc is constant at constant temperature. Thus, in this case, Kc would not change.

(iii) In an endothermic reaction, the value of Kc increases with an increase in temperature. Since the given reaction in an endothermic reaction, the value of Kc will increase if the temperature is increased.

Question :31 Dihydrogen gas used in Haber's process is produced by reacting methane from natural gas with high temperature steam. The first stage of two reaction involves the formation of CO and H2. In second stage, CO formed in first stage is reacted with more steam in water gas shift reaction.

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

If a reaction vessel at 400 degree C is charged with an equivalent mixture of CO and steam such that pCO = pH2O = 4.0 bar, what will be the partial pressure of H2 at equilibrium? Kp = 10.1 at 400degree C

Answer:

$CO + H_2O \rightarrow CO_2 + H_2$

At t = 0, pressure of CO = 4 bar

Pressure of H2O = 4 bar

At eqlib., pressure of CO = (4 -x)

Pressure of H2O =(4- x) bar

Pressure of CO2 = x bar

Pressure of H2 = x bar

Now, equilibrium constant (Kc) = [CO2][H2]/[CO][H2O]

$Kp = x^2/(4 - x)^2$

$10.1 = x^2/(4 - x)^2$

$\sqrt{10.1} = x/(4 - x)$

3.17(4 - x) = x

12.68 - 3.17x = x

12.68 = 4.17x

x = 12.68/4.17 = 3.04 bar

Hence, pressure of H2 = x bar = 3.04 bar

Question :32 Predict which of the following reactions will have appreciable concentration of reactants and products :

a. **Cl_2 (g) → 2Cl (g) Kc = 5 x 10*-39**
b. **Cl_2 (g) + 2 NO (g) → 2NOCl (g) Kc =3.7 x 10*8**
c. **Cl_2 (g) + $2NO_2$ (g) → $2NO_2Cl$ (g) Kc = 1.8**

Answer:

If the value of Kc lies between 10^{-3} and 10^{3}, a reaction has an appreciable concentration of reactants and products. Thus, the reaction given in (c) will have an appreciable concentration of reactants and products.

Question :33 The value of Kc for the reaction $3O_2$ (g) → $2O_3$ (g) is 2.0 x 10*-50 at 250 degree C. If the equilibrium concentration of O_2 in air at 25 degreiC is 1.6 x 10*-2 ,what is the concentration of O_3?

Answer:

The given reaction is:

$3O_2$ (g) → $2O_3$ (g)

Then, Kc = $[O_3 (g)]^2 / [O_2 (g)]^3$

It is given that Kc = 2.0×10^{-50} and O_2 (g) = 1.6×10^{-2}

Then we have,

$2.0 \times 10^{-50} = [O_3 (g)]2 / [1.6 \times 10^{-2}]^3$

$[O_3 (g)]^2 = [2.0 \times 10^{-50}] \text{ x } [1.6 \times 10^{-2}]^3$

$[O_3 (g)]^2 = 8.192 \text{ x } 10^{-56}$

O_3 (g) = $2.86 \text{ x} 10^{-28}$ M

Hence, the concentration of O_2 (g) = $2.86 \text{ x} 10^{-28}$ M

Question :34 The reaction, CO (g) + $3H_2$ (g) → CH_4 (g) + H_2O (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 Kc for the reaction at given temperature is 3.90.

Answer:

Let the concentration of methane at equilibrium be x.

	CO(g) +	$3H_2$(g) ↔	CH_4(g) +	H_2O(g)
At equilibrium	0.3/1 M	0.1/1 M	x	0.02/1 M

It is given that Kc = 3.90.

Therefore,
Kc = [CH4(g)] [H2O(g)] / [CO(g)] [H2(g)]3
[x] [0.02] / (0.3) $(0.1)^3$ = 3.90
x = (3.90) (0.3) $(0.1)^3$ / [0.02]
x = 0.00117 / 0.02
= 0.0585 M
= 5.85 x 10^{-2}
Hence, the concentration of CH4 at equilibrium is 5.85 × 10^{-2} M.

Question :35 What is meant by the conjugate acid – base pair? Find the conjugate acid/base for the following species :
HNO2, CN-, HClO4, F-, OH-, CO3*2-, S2-

Answer:
A conjugate acid – base a pair is a pair that has difference of only one proton.
The conjugate acid – base pair of the following are as follows:
HNO2 – $NO2^-$ (Base)
CN^- - HCN (Acid)
HClO4 – $ClO4^-$(Base)
F^- - HF (Acid)
OH^- - H2O (Acid) O^{2-}(Base)
$CO3^{2-}$ - $HCO3^-$(Acid)
S^{2-} - HS^-(Acid)

Question :36 Which of the following are Lewis acids? H2O, BF3, H+, NH4+

Answer:
Lewis acids are the acids which can accept a pair of electrons.
H2O – Lewis base
BF3 – Lewis acid
H+ - Lewis acid
$NH4^+$ - Lewis acid.

Question :37 What will be the conjugate base of the Bronsted acids: HF, H2SO4, HCO3*-?

Answer:
The following shows the conjugate bases for the Bronsted acids:
HF – F^-
H2SO4 – $HSO4^-$

$HCO3^-$ - $CO3^{2-}$

Question :38 Write the conjugate acids for the following Bronsted bases: NH2*-, NH3 and HCOO-.

Answer:

The following shows the conjugate acids for the Bronsted base:

$NH2^-$ - NH3

NH3 – $NH4^+$

$HCOO^-$ - HCOOH

Question :39 The species : H2O, HCO3*-, HSO4*- and NH3 can act as both Bronsted acids and bases. For each case give the corresponding conjugate acid and base.

Answer:

The conjugate acids and conjugate bases for the given species are listed in the table below:

Species	Conjugate acid	Conjugate base
H2O	$H3O^+$	OH^-
$HCO3^{2-}$	H2CO3	$CO3^{2-}$
$HSO4^-$	H2SO4	$SO4^{2-}$
NH3	$NH4^+$	$NH2^-$

Question :40 Classify the following species into Lewis acids and Lewis bases and show how these act as Lewis Acid/base :

- **a. OH-**
- **b. F-**
- **c. H+**
- **d. BCl3**

Answer:

a.) OH^-

It is Lewis base as it has a tendency to lose a pair of electrons.

b.) F^-

It is a Lewis base as it has a tendency to lose its lone pair of electrons.

c.) H^+

It is a Lewis acid as it has a tendency to accept a pair of electrons.

d.) $BCl3^-$

It is a Lewis acid as it has tendency to accept a pair of electrons.

Question :41 The concentration of hydrogen ions in a sample of soft drink is 3.8 x 10*-3 M. What is it's pH?

Answer:

$pH = -\log[H^+]$

$= -\log(3.8 \times 10^{-3})$

$= -\log 3.8 - \log(10^{-3})$

$= -\log 3.8 + 3$

$= -0.5798 + 3$

$= 2.423$

Question :42 The pH of a sample of vinegar is 3.76. Calculate the concentration of hydrogen ion in it.

Answer:

$pH = -\log(H^+]$

$\log[H^+] = -$ Ph

$[H^+] = \text{antilog}(-pH)$

$= \text{Antilog}(-3.76)$

$= 0.000178$

$= 1.78 \times 10^{-4}$

: 1.78×10^{-4} is the concentration of white vinegar sample.

Question :43 The ionization constant of HF, HCOOH and HCN at 298 K are 6.8 x 10*-4, 1.8 x 10*-4 and 4.8 x 10*-9 respectively. Calculate the ionization constants of the corresponding conjugate base.

Answer:

For F^-, $Kb = K_w / K_a$

$= 10^{-14} / (6.8 \times 10^{-4})$

$= 1.47 \times 10^{-11}$

For $HCOO^-$, Kb

$= 10^{-14} / 1.8 \times 10^{-4}$

$= 5.6 \times 10^{-11}$

For CN^-,

$Kb = 10^{-14} / (4.8 \times 10^{-19})$

$= 2.08 \times 10^{-6}$

Question :44 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 it's degree of ionization if the solution is also 0.05 M in sodium phenolate?

Answer:

$C_6H_5OH + H_2O \rightarrow C_6H_5O^- + H_3O^+$

$Ka = [C_6H_5O^-][H_3O^+] / [C_6H_5OH]$

According to rule of ostwald law,

$[C_6H_5O^-] = [H^3O^+] = \sqrt{k_a c}$

Where, C is the concentration of phenol solution

e.g., C = 0.01 M

Now, $[C_6H_5O^-] = [H^3O^+] = \sqrt{\{10^{-10} \times 0.05\}} = 2.24 \times 10^{-6}$ M

Now, when mixture contains 0.05M phenol and 0.01M phenolate ion then,

$Ka = [C_6H_5O^-][H_3O^+]/[C_6H_5OH]$

$10^{-10} = 0.01M \times [H_3O^+]/0.05M$

$[H_3O^+] = 0.05 \times 10^{-10}/0.01 = 5 \times 10^{-10}$ M

Now, we know, $[H^3O^+] = C\alpha$

$5 \times 10^{-10} = 0.05\ \alpha$

$\alpha = 5 \times 10^{-10}/0.05 = 10^{-8}$

Hence, degree of ionization = 10^{-8}

Question :45 The first ionization constant of H2S is 9.1 x 10*-8 , Calculate the concentration of HS- ion in its 0.1 M solution. How will this concentration be affected if the solution is 0.1 M in HCl also? If the second dissociation constant of H2S is 1.2 x 10*-13 , calculate the concentration of S*2- under both conditions.

Answer:

(i) To calculate the concentration of HS- ion:

Case I (in the absence of HCl):

Let the concentration of HS- be x M.

	H2S →	H+ +	HS-
Ci	0.1	0	0
Cf	0.1-x	x	x

Then Ka1 = [H+] [HS-] / H2S

$9.1 \times 10^{-8} = xx / 0.1-x$

$(9.1 \times 10^{-8})\ (0.1-x) = x2$

Taking 0.1 - x M ; 0.1M, we have

$(9.1 \times 10^{-8})\ (0.1) = x2$

$9.1 \times 10^{-9} = x2$

$x = \sqrt{9.1\ x\ 10^{\wedge} - 9}$

$= 9.54 \times 10^{-5}$

$= [HS] = 9.54 \text{ X } 10^{-5}M$

Case II (in the presence of HCl):

In the presence of 0.1 M of HCl, let [HS-] be y M.

Then, $H2S \rightarrow H+ \ + \ HS-$

Ci 0.1 0 0

Cf 0.1-y y y

Also, $HCI \rightarrow H+ \ + \ CI-$

0.1 0.1

Now, Ka1 = [H+] [HS-] / H2S

Ka1 = [y] [0.1+y] / [0.1-y]

$9.1 \times 10^{-8} = y \times 0.1 / 0.1$ (∵ 0.1-y; 0.1M) (and 0.1+y; 0.1M)

$9.1 \times 10^{-8} = y$

[HS-] = 9.1 × 10–8

(ii) To calculate the concentration of [S2-]:

Case I (in the absence of 0.1 M HCl):

$HS- \rightarrow H+ \ + S2-$

HS- = 9.54 x 10-5 M (From first ionization, case I)

Let S2- be X.

Also, [H+] = 9.54 x 10-5 M (From first ionization, case I)

Ka2 = (9.54 x 10-5) (X) / (9.54 x 10-5)

1.2x10-13 = X = S2-

Case II (in the presence of 0.1 M HCl):

Again, let the concentration of HS- be X' M.

$[HS^-] = 9.1 \text{ X } 10^{-8}M$

[H+] = 0.1 M (From HCl, case II)

$[S^{2-}] = X'$

Then, $Ka2 = [H^+][S^{2-}] / [HS^-]$

$1.2 \text{ X } 10^{-13} = (0.1)(X') / 9.1 \text{ X } 10^{-8}$

$10.92 \text{ X } 10^{-21} = 0.1\ X'$

$X' = 1.092 \text{ X } 10^{-20} / 0.1$

$= 1.092 \text{ X } 10^{-19}$

$Ka1 = 1.74 \times 10^{-5}$

Question :46 The ionization constant of acetic acid is 1.74 x 10*-5 . Calculate the degree of dissociation of acetic acid in its 0.05 M solution. Calculate the concentration of acetate ion in the solution and it's pH.

Answer:

CH3COOH = CH3COO$^-$ + H$^+$, Ka = [CH3COO$^-$][H$^+$] / [CH3COOH]

= [H$^+$]^2 / [CH3COOH]

[H$^+$] = $\sqrt{Ka[CH3COOH]}$ = $\sqrt{(1.74\,X\,10^{-5})(5\,X\,10^{-2})}$ = 9.33 X 10^{-4}M

[CH3COO$^-$] = [H$^+$] = 9.33 X 10^{-4}M

pH = -log(9.33 x 10^{-4}) 4 – 0.9699 = 4 – 0.97 = 3.03

Question :47 It has been found that 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 it's pKa.

Answer:

HA → H$^+$ + A$^-$

pH = -loh[H+]

log[H+] = -4.15

[H+] = 7.08 X 10^{-5}M

[A-] = [H+] = 7.08 X 10^{-5}M

Ka = [H+][A-] / [HA] = (7.08 X 10^{-5})(7.08 X 10^{-5}) / 10^{-2} = 5.0 X 10^{-7}

p$_{Ka}$ = -logKa = -log(5.0 x 10^{-7}) = 7 – 0.699 = 6.301

Question :48 Assuming complete dissociation, calculate the pH of the following solutions :

a. **0.003 M HCl**
b. **0.005M NaOH**
c. **0.002 M HBr**
d. **0.002 M KOH**

Answer:

(i) 0.003MHCl:

H2O + HCl ↔ H3O+ + Cl-

Since HCl is completely ionized,

[H3O+] = [HCl]

[H3O+] = = 0.003

Now

pH = -log [H3O+] = -log (0.003)
= 2.52
Hence, the pH of the solution is 2.52.
(b) 0.005 M NaOH
NaOH(aq) → Na+(aq) + HO-(aq)
[NaOH] = [HO-]
[HO-] = 0.05
pOH = -log[HO-] = -log (0.05)
= 2.30
: pH = 14 - 2.30 = 11.70
Hence, the pH of the solution is 11.70.
(c) 0.002 M HBr
HBr + H2O → H3O+ + Br-
[HBr] = [H3O+]
[H3O+] = 0.002
:pH = -log [H3O+] = -log (0.002)
= 2.69
Hence, the pH of the solution is 2.69.
(d) 0.002 M KOH
KOH(aq) → K+(aq) + OH-(aq)
[OH-] = [KOH]
[OH-] = 0.002
Now pOH = -log[OH-] = -log (0.002)
= 2.69
:pH = 14-2.69 = 11.31
Hence, the pH of the solution is 11.31.

Question :49 Calculate the pH of the following solutions :

a. **2 g of TlOH dissolved in water to give 2 litre of solution**
b. **0.3 g of Ca(OH)2 dissolved in water to give 500 mL of solution.**
c. **0.3 g of NaOH dissolved in water to give 200 mL of solution.**
d. **1 mL of 13.6 M of HCl is diluted with water to give 1 litre of solution.**

Answer:
For 2g of TlOH dissolved in water to give 2 L of solution:
[TlOH(aq)] = 2/2 g/L
= 2/2 x 1/221 M
= 1/221 M

TIOH(aq) → TI+(aq) + OH-(aq)

OH-(aq) = TIOH(aq) = 1/221M

Kw = [H+] [OH-]

10^{-14} = [H+] [1/221]

[H+] = 221x10^{-14}

pH = -log [H+] = -log (221x10^{-14})

= 11.65

(b) For 0.3 g of Ca(OH)2 dissolved in water to give 500 mL of solution:

Ca(OH)2 → Ca2+ + $2OH^-$

[Ca(OH)2] = 0.3 x 1000/500 = 0.6M

OH^- (aq) = 2 x [Ca(OH)2(aq)] = 2 x 0.6 = 1.2M

[H+] = Kw / OH-(aq)

= 10^{-14}/1.2 M

= 0.833 x 10^{-14}

pH = -log(0.833 x 10^{-14})

= -log(8.33 x 10^{-13})

= (-0.902 + 13)

= 12.098

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

NaOH → Na +(aq) + OH-(aq)

[NaOH] = 0.3 x 1000/200 = 1.5M

[OH-(aq)] = 1.5M

Then [H+] = 10^{-14} / 1.5

= 6.66 x 10^{-13}

pH = -log (6.66 x 10^{-13})

= 12.18

(d) For 1mL of 13.6 M HCl diluted with water to give 1 L of solution:

13.6 x 1 mL = M2 x 1000 mL

(Before dilution) (after dilution)

13.6 x 10^{-3} = M2 x 1L

M2 = 1.36 x 10^{-2}

[H+] = 1.36 × 10^{-2}

pH = - log (1.36 × 10^{-2})

= (- 0.1335 + 2)

= 1.866 = 1.87

Question :50 The degree of ionization of a 0.1 M bromoacetic acid solution is 0.132. Calculate the pH of the solution and the pKa of bromoacetic acid.

Answer:

Degree of ionization, α = 0.132
Concentration, c = 0.1 M

Thus, the concentration of H3O+ = c.α
= 0.1 × 0.132
= 0.0132
pH = -log [H+]
= -log (0.0132)
= 1.879 : 1.88
Now,

Ka = Cα2
= 0.1 x (0.132)2
Ka = 0.0017

pKα = 2.75

Question : 51 The pH of 0.005 M codeine (C18H21NO3) solution is 9.95. Calculate it's ionization constant and pKb.

Answer:
c = 0.005
pH = 9.95
pOH = 4.05
pH = -log (4.05)
4.05 = - log [OH-]
[OH-] = 8.91x 10^{-5}

cα = 8.91x 10^{-5}

α = 8.91x 10^{-5} / 5 x 10^{-3} = 1.782 x 10^{-2}

Thus, Kb = cα2
= 0.005 x $(1.782)^2$ x 10^{-4}
= 0.005 x 3.1755 x 10^{-4}
= 0.0158 x 10^{-4}
Kb = 1.58 x 10^{-6}
PKb = -logKb
= -log (1.58 x 10^{-6})
= 5.80

Question :52 What is pH of 0.001 M aniline solution? The ionization constant of aniline can be taken from Table 7.7 . Calculate the degree of ionization of aniline in the solution. Also calculate the ionization constant of the conjugate acid of aniline.

Answer:

$Kb = 4.27 \times 10^{-10}$

$c = 0.001M$

$pH = ?$

$\alpha = ?$

$Kb = c\alpha2$

$4.27 \times 10^{-10} = 0.001 \times \alpha2$

$4270 \times 10^{-10} = \alpha2$

$\alpha = 65.34 \times 10^{-4}$

Then (anion) = $c\alpha = 0.001 \times 65.34 \times 10^{-4}$

$= 0.65 \times 10^{-5}$

$pOH = -\log (0.65 \times 10^{-5})$

$= 6.187$

$pH = 7.813$

Now

$Ka \times Kb = Kw$

:$4.27 \times 10^{-10} \times Ka = Kw$

$Ka = 10^{-14} / 4.27 \times 10^{-10}$

$= 2.34 \times 10^{-5}$

Thus, the ionization constant of the conjugate acid of aniline is 2.34×10^{-5}.

Question :53 Calculate the degree of ionization of 0.05 M acetic acid if it's pKa value is 4.74 .How is the degree of dissociation affected when it's solution also contains

- **a. 0.01 M**
- **b. 0.1 M of HCl?**

Answer:

$c = 0.05$ M

$pKa = 4.74$

$pKa = -\log (Ka)$

$Ka = 1.82 \times 10^{-5}$

$Ka = c\,\alpha^2$

$\alpha = \sqrt{Ka/c}$

$\alpha = \sqrt{1.82\,x10^{-5}/5x10\hat{}-2}$

= 1.908×10^{-2}

When HCl is added to the solution, the concentration of H+ ions will increase. Therefore, the equilibrium will shift in the backward direction i.e., dissociation of acetic acid will decrease.

Question :54 The ionization constant of dimethylamine is 5.4 x 10*-4 . Calculate it's degree of ionization in its 0.02 M solution. What percentage of dimethyl anime is ionized if the solution is also 0.1 M in NaOH?

Answer:

Kb = 5.4×10^{-4}

c = 0.02M

Then,

$\alpha = \sqrt{Kb/c}$

$= \sqrt{5.4\ x\ 10^{-4}/0.02}$

= 0.1643

Now, if 0.1 M of NaOH is added to the solution, then NaOH (being a strong base) undergoes complete ionization.

NaOH(aq) → Na+(aq) + OH- (aq)

0.1M 0.1M

and

$(CH_3)_2NH + H_2O \leftrightarrow (CH_3)_2NH_2^+ + OH$

0.02-x x x

Then $(CH_3)_2NH_2^+$ = x

[OH-] = x + 0.1 ; 0.1

Kb = $[(CH_3)_2NH_2^+][OH^-] / [(CH_3)_2NH]$

5.4×10^{-4} = x x 0.1 / 0.02

x = 0.0054

It means that in the presence of 0.1 M NaOH, 0.54% of dimethylamine will get dissociated.

Question :55 Calculate the hydrogen ion concentration in the following biological fluids whose pH are given below :

a. **Human muscle – fluid, 6.83**
b. **Human stomach fluid, 1.2**
c. **Human blood, 7.38**
d. **Human saliva, 6.4.**

Answer:

i.) Human saliva, 6.4

pH = 6.4

= -log[H+][H+] = 3.98 X 10^{-7}

ii.) Human stomach fluid, 1.2

pH = 1.2

= -log[H+]

:[H+] = 0.063

iii.) Human muscle fluid, 6.83

pH = 6.83

pH= -log[H+][H+]

6.4 = 1.48 X 10^{-7}M

iv.) Human blood, 7.38

pH = 7.38 = -log[H+]

; [H+] = 4.17 X 10^{-8}M

Question :56 The pH of milk, black coffee, tomato juice, lemon juice and egg white are 6.8, 5.0, 4.2, 2.2 , 7.8 respectively. Calculate corresponding hydrogen ion concentration in each.

Answer:

The hydrogen ion concentration in the given substances can be calculated by using the given relation:

pH = -log [H+]

(i) pH of milk = 6.8

Since, pH = -log [H+]

6.8 = -log [H+]

log [H+] = -6.8

[H+] = anitlog(-6.8)

= $1.5\text{x}10^{-7}$ M

(ii) pH of black coffee = 5.0

Since, pH = -log [H+]

5.0 = -log [H+]

log [H+] = -5.0

[H+] = anitlog(-5.0)

= 10^{-5} M

(iii) pH of tomato juice = 4.2

Since, pH =-log [H+]

4.2 =-log [H+]
log [H+] = -4.2
[H+] = anitlog(-4.2)
= 6.31×10^{-5} M

(iv) pH of lemon juice = 2.2
Since, pH = -log [H+]
2.2 = -log [H+]
log [H+] = -2.2
[H+] = anitlog(-2.2)
=6.31×10^{-3} M

(v) pH of egg white = 7.8
Since, pH = -log [H+]
7.8 = -log[H+]
log[H+] = -7.8
[H+] = anitlog(-7.8)
= 1.58×10^{-8} M

Question :57 If 0.561 f of KOH is dissolved in water to give 200 mL of solution at 298 K. Calculate the concentration of potassium, hydrogen and hydroxyl ions. What is it's pH

Answer:
[KOH(aq)] = 0.561 / (1/5) g/L
= 2.805 X (1 / 56.11)
= 0.05M
$KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$
$[OH^-] = 0.05M = [K^+][H^+][OH^-] = K_W$
$[H^+] = K_W / [OH^-][H^+] = 10^{-14} / 0.05$
$[H^+] = 2X\ 10^{-13}M$
pH = -log[H+]
$pH = -\log[2 \times 10^{-13}]$
pH = 12.70

Question :58 The solubility of Sr(OH)2 at 298 K is 19.23 g/L of solution. Calculate the concentration of strontium and hydroxyl ions and the pH of the solution.

Answer:
Given, solubility of Sr(OH)2 is 19.23g/L

Molecular weight of Sr(OH)2 is 121.6 g/mol
We know,
Molarity = solubility/molecular weight
= 19.23 g/L/121.6g/mol
= 0.1581 M

$Sr(OH)2 \rightarrow Sr^{2+} + 2OH^{-}$
0.1581M   0.1581M   2×0.1581M = 0.3162M

$[Sr^{2+}] = 0.1581M$ and $[OH^{-}] = 0.3162M$

We know, ionic product(Kw) = $[H^{+}][OH^{-}] = 10^{-14}$

$[H^{+}] = 10^{-14} / [OH^{-}] = 10^{-14} / 0.3162 = 3.16 \times 10^{-14}$

Now Arrhenius formula,

pH = $-\log[H^{+}] = -\log(3.16 \times 10^{-14}) = 14 - 0.4997 = 13.5003$

Question :59 The ionization constant of probation acid is 1.32 x 10*-5 . Calculate the degree of ionization of the acid in its 0.05 M solution and also it's pH. What will be it's degree of ionization of the solution is 0.01 M in HCl also?

Answer:

$CH3CH2COOH + H20 \rightarrow CH3CH2COO^{-} + H3O^{+}; Ka = 1.32 \times 10^{-5}$

At t = 0   0.05   0   0

At eqlib. $(0.05 - C\alpha)$ $C\alpha$ $C\alpha$

We know, from Ostwald's dilution law,

$Ka = C\alpha^{2} \Rightarrow \alpha = \sqrt{\{Ka/C\}}$

$= \sqrt{\{1.32 \times 10^{-5}/0.05\}} = 0.016248$

$[H30+] = C\alpha = 0.05 \times 0.016248 = 8.124 \times 10^{-4}$ M

Now use Arrhenius formula for pH

pH = $-\log[H30+] = -\log(8.124 \times 10^{-4}) = 4 - 0.9098 = 3.09$

Now, when the solution contains 0.01 M HCl

$Ka = [CH3CH2COO^{-}][H3O^{+}]/[CH3CH2COOH]$

$1.32 \times 10^{-5} = C\alpha \times 0.01/(0.05 - C\alpha)$

$(0.05 - C\alpha) \approx 0.05$

$1.32 \times 10^{-5} = C\alpha \times 0.01/0.05$

$C\alpha = 5 \times 1.32 \times 10^{-5} = 6.6 \times 10^{-5}$ M

Hence, degree of ionization (α) = $6.6 \times 10^{-5}/0.05 = 1.32 \times 10^{-5}$

Question :60 The pH of 0.1 M solution of cyanide acid (HCNO) is 2.34 . Calculate the ionization constant of the acid and it's degree of ionization in the solution.

Answer:

$c = 0.1M$

$pH = 2.34$

$-\log [H+] = pH$

$-\log [H+] = 2.34$

$[H+] = 4.5 \text{ X } 10^{-3}$

Also,

$[H+] = c\ \alpha$

$4.5 \times 10^{-3} = 0.1 \times \alpha$

$\alpha = 0.1/ (4.5 \times 10^{-3})$

$\alpha = 0.045$

$Ka = c\ \alpha^2$

$Ka = 0.1 \times (0.045)^2$

$Ka = 0.0002025$

$Ka = 2.025 \times 10^{-4}$

Question : 61 The ionization constant of nitrous acid is 4.5 x 10*-4 . Calculate the pH of 0.04 M sodium nitrite solution and also it's degree of hydrolysis.

Answer:

NaNO2 is the salt of a strong base (NaOH) and a weak acid (HNO2).

$NO2^- + H2O \rightarrow HNO2 + OH-$

$Kh = [HNO2] [OH-] / [NO-2]$

$Kw / ka = 10^{-14} / 4.5 \times 10^{-4} = 0.22 \times 10^{-10}$

Now, If x moles of the salt undergo hydrolysis, then the concentration of various species present in the solution will be:

$[NO2^-] = 0.04 - x ; 0.04$

$[HNO2] = x$

$[OH-] = x$

$Kh = x2 / 0.04 = 0.22 \times 10^{-10}$

$x2 = 0.0088 \times 10^{-10}$

$x = 0.093 \times 10^{-5}$

$: [OH-] = 0.093 \times 10^{-5}$ M

$[H3O+] = 10^{-14} / 0.093 \times 10^{-5} = 10.75 \times 10^{-9}$ M

$pH = -\log(10.75 \times 10^{-9})$

$= 7.96$

Therefore, degree of hydrolysis

= x / 0.04 = (0.093 x 10^{-5}) / 0.04 = 2.325 x 10^{-5}

Question :62 A 0.02 M solution of pyridinium hydrochloride has pH = 3.44. Calculate the ionization constant of pyridine.

Answer:

pH = 3.44 As

We know,

pH = log[H+]

: [H+] = 3.63 x 10^{-4}

Now, Kh = 3.63 x 10^{-4} / 0.02

(Given that concentration = 0.02M)

Kh = 6.6 x 10^{-6}

As we know that,

Kh = Kw / Ka

Ka = Kw / Kh = 10^{-14} / 6.6 x 10^{-6}

= 1.51 x 10^{-9}

Question :63 Predict if the solutions of the following slats are neutral, acidic or basic NaCl, KBr, NaCN, NH4NO3, NaNO2, KF.

Answer:

The solution of a salt of strong acid and strong base is neutral while that of weak acid and strong base is basic. and A solution of strong acid and a weak base is acidic.

(i) NaCl

NaCl is formed by HCl { strong acid } and NaOH { strong base } . and we know, A solution of strong acid and strong base is neutral.

Hence, NaCl is neutral.

(ii) KBr is formed by KOH { strong base } and HBr { strong acid } hence, KBr is neutral.

(iii) NaCN is formed by NaOH { strong base } and HCN { weak acid } hence, NaCN is basic.

(iv) NH4NO3 is formed by HNO3 {strong acid } and NH4OH { weak base } hence , NH4NO3 is acidic.

(v)NaNO2 is formed by NaOH { strong base } and HNO2 { weak acid comparatively} hence, solution is basic.

(vi) KF is formed by KOH { strong base } and HF { weak acid comparatively }
because Hydrolysis of it
F- + H2O → HF + OH- , hence, solution is basic.

Question :64 The ionization constant of chloroacetic acid is 1.35 x 10*-3 . What will be the pH of 0.1 M acid and it's 0.1 M sodium salt solution?

Answer:
For finding PH of 0.1 M sodium salt solution use the formula ,
PH = -1/2 [log Kw + log Ka - log Kb]
Where, Kw is the ionic product
Ka is Ionisation constant of acid
Kb is the Ionisation constant of basic.
Given, Here,
Kw = 10^-14
Ka = 1.35 × 10^-3
Kb = 0.1
Now, PH = -1/2 [log(10^-14) + log(1.35 × 10^-3) - log(0.1)]
= -1/2 [-14 +(-3 + 0.1303) -(-1)]
= -1/2 [-14 -3 + 0.1303 +1]
= -1/2 [-15.8697] = 7.93485 ≈ 7.94
Hence, PH of 0.1 M acid and sodium salt solution is 7.94

Question :65 Ionic product of water at 310 K is 2.7 x 10*-14. What is the pH of neutral water at this temperature?

Answer:
Ionic product, Kw = [H+][OH-]
Assuming, $[H^+]$ = y
As, $[H^+] = [OH^-]$
Kw = y^2, Kw at 310K is
2.7×10^{-14}
: $2.7 \times 10^{-14} = y^2$
Y = 1.64×10^{-7}
[H+] = 1.64×10^{-7}
pH = -log[H+]
= - log[1.64×10^{-7}] = 6.78. Thus, the pH of neutral water at 310 K temperature is 6.78.

Question :66 Calculate the pH of the resultant mixture :

a. **10 mL of 0.2 M $Ca(OH)_2$ + 25 mL of 0.1 M HCl**
b. **10 mL of 0.01 M H_2SO_4 + 10 mL of 0.01 M $Ca(OH)_2$**
c. **10 mL of 0.1 M H_2SO_4 + 10 mL of 0.1 M KOH**

Answer:
(a) Here, for base
M1 = [OH-] = 2 × 0.2 = 0.4 M
V1 = 10 ml so, M1V1 = 0.4 × 10 = 4 Mml
For acid,
M2 = [H+] = 0.1 M
V2 = 25 ml so, M2V2 = 0.1 × 25ml = 2.5 Mml
Here, strength of base (M1V1) > strength of acid (M2V2) so, solution is basic .
So, [OH-] = {M1V1 - M2V2}/(V1 + V2)
= (4 - 2.5)/(10 + 25) = 0.043 M
Now, use Arrhenius formula,
POH = -log[OH-] = -log(0.043) = 2 - 0.6335 = 1.3665
So, PH = 14 - POH = 14 - 1.3665 = 12.6335
Similarly we can solve next all questions
(b) For acid (H_2SO_4),
M1 = [H+] = 2 × 0.01M = 0.02 M , V1 = 10 ml
So, M1V1 = 0.02 × 10 = 0.2 Mml
For base $Ca(OH)_2$,
M2 =[OH-] = 2 × 0.01 M = 0.02M , V1 = 10 ml
So, M2V2 = 0.02M × 10ml = 0.2 Mml
Because strength of H_2SO_4 = strength of base = 0.2 Mml
So, solution is neutral .
Hence, pH of solution is 7 .
(c) For acid (H_2SO_4),
M1 = [H+] = 2 × 0.1 M = 0.2 M ,V1 = 10ml
M1V1 = 0.2 M × 10 ml = 2Mml
For base (KOH),
M2 = 0.1M , V2 = 10ml
So, M2V2 = 0.1M × 10 Mml= 1 Mml
Because strength of acid (H_2SO_4) > strength of base (KOH)
So, solution is acidic.
Now, [H+] = (M1V1 - M2V2)/(V1+V2)
= (2 - 1)/(10+10) = 0.05 M
pH = -log[H+] = -log(5 × 10^-2) = 2 - 0.6990
= 1.301 ≈ 1.3

Question :67 Determine the solubilities of silver chromate, barium chromate, ferric hydroxide, lead chloride and Mercury's iodide at 298 K from their solubility product constant given in table 7.9 . Determine also the molarities of individual ions.

Answer:

(a) Silver chromate

Ag2CrO4 <=> 2Ag+ + CrO42- ; Ksp = 1.1 × 10^-12

But solubility product (Ksp) = [2Ag+]^2 [CrO42-]

Let solubility of silver chromate is s

Ksp = [2s]^2 [s] = 4s^3

s^3 = Ksp/4 = 1.1 × 10^-12/4 = 2.75 × 10^-13

Take log both sides,

3logs = log(2.75 × 10^-13) = -12.5607

s = 6.503 × 10^-5 M { using antilog concept }

So, [Ag+] = 2s = 1.3006 × 10^-5 M

[CrO4^{2}-] = s = 6.503 × 10^-5 M

(b) BaCrO4 <=> Ba^{2}+ + CrO4^{2}- ; Ksp = 1.2 × 10^-10

Let solubility of BaCrO4 is s M

Ksp = [Ba2+][CrO4^{2}-]

Ksp = s.s = s^2

s = √Ksp = √{1.2 × 10^-10} = 1.1 × 10^-5 M

So, [Ba^{2}+] = [CrO4^{2}-] = 1.1 × 10^-5 M

(c) Fe(OH)3 <=> Fe^{3}+ + 3OH- ;Ksp =10^-38

Let solubility of Fe(OH)3 is s M

Ksp = [Fe^{3}+][3OH-]^3

Ksp = s.27s^3

27s^4 = Ksp

s^4 = Ksp/27 = 10^-38/27 = 0.037 × 10^-38

Take log both sides,

4logs = log(0.037 × 10^-38) = -39.4318

logs = -9.8579 = 10(bar).1421

s = 1.387 × 10^-10 { by using antilog concept }

So, [Fe^{3}+] = 1.387 × 10^-10 M

[OH-] = 3s = 3 × 1.387 × 10^-10 M

= 4.161 × 10^-10 M

(d) PbCl2 <=> Pb^{2}+ + 2Cl- ;Ksp = 1.6 × 10^-5

Let solubility of PbCl2 is s M

Ksp = [Pb^{2}+][2Cl-]^2

Ksp = s.4s²
4s³ = Ksp
s³ = Ksp/4 = 1.6 × 10^-5/4 = 0.4 × 10^-5 M
3logs = log(4 × 10^-6) = -6 + 0.6021 = -5.3979
logs = -1.7993 = 2(bar).2007
Take antilog both sides,
s = 1.585 × 10^-2 M
Hence, [Pb²+] = 1.585 × 10^-2 M
[Cl-] = 2s = 2 × 1.585 × 10^-2 M = 3.17 × 10^-2 M

(e) Hg2I2 <=> Hg2²+ + 2I- ; Ksp = 4.5 × 10^-29
Let solubility of Hg2I2 is s M
Ksp = [Hg2²+][2I-]²
Ksp = s.4s² = 4s³
s³ = Ksp/4 = 4.5 × 10^-29/4 = 1.125 × 10^-29
3logs = log(1.125 × 10^-29) = -29+0.0512 =-28.9488
logs = -9.6496 = 10(bar).3504
s = 2.241 × 10^-10 M{by using antilog }
[Hg2²+] = s = 2.241 × 10^-10 M
[I-] = 2s = 2 × 2.241 × 10^-10 = 4.482 × 10^-10 M

Question :68 The solubility product constant of Ag2CrO4 and AgBr are 1.1 x 10*-12 and 5.0 x 10*-13 respectively. Calculate the ratio of the molarities of their saturated solutions.

Answer:
Silver chromate :
$Ag2CrO4 \rightarrow 2Ag^{+} + CrO4^{2-}$
$[Ag+] = 2s1, CrO4^{2-} = s1$
$Ksp = (2s1)^{2}(s1) = 4s^{3} = 1.1 \text{ x } 10^{-12}$
$S1 = 6.5 \text{ x } 10^{-5}$ …(1)
Silver bromide :$AgBr \rightarrow Ag^{+} + Br^{-}$
$[Ag^{+}] = [Br^{-}] = s2$
$Ksp = (s2) \text{ x } (s2) = s2^{2} = 5.0 \text{ x } 10^{-13}$
$S2 = 7.07 \text{ x } 10^{-7}$ …(2)
Divide equation (1) by equation (2) to obtain the ratio of the molarities of saturated solutions:
$S1 / s2 = 6.50 \text{ x } 10^{-5} / 7.07 \text{ x } 10^{-7} = 91.9$

Question :69 Equal volumes of 0.002 M solutions of sodium iodate and cupric chlorate are mixed together Will it lead to precipitation of copper iodate?

Answer:

$2NaIO3 + Cu(ClO3)2 \rightarrow 2NaClO3 + Cu(IO3)2$

Equal volume of 0.002M solutions of NaIO3 and Cu(IO3)2 are mixed together. So, molar concentration of each species will be halved.

so, conc. Of each = 0.002M/2 = 0.001 M

e.g., [Cu2+]mix = [IO3-]mix = 0.001 M

Now, Cu(IO3)2 <=> Cu2+ +2IO3-

Solubility product {Ksp} = [Cu2+][IO3]^2

For precipitation, ionic product > Ksp

e.g., [Cu2+][IO3]^2 > Ksp

But here,

0.001 × 0.001 = 10^-9 < 7.4 × 10^-8

It is less than Ksp. hence, there won't be precipitation.

Question :70 The ionization constant of benzoic acid is 6.46 x 10*-5 and Ksp for silver benzoate is 2.5 x 10*-13. How many times is silver benzoate more soluble in a buffer of pH 3.19 compared to its solubility in pure water

Answer:

Given:

The ionization constant of benzoic acid (Ka) is 6.46×10^{-5}

Ksp for silver benzoate is 2.5×10^{-13}

pH = 3.19

Ionization of silver benzoate:

$C6H5COOAg \rightarrow C6H5COO- + Ag+$

Solubility in water: Let solubility in water is x mol/l Then

[C6H5COO-] = [Ag+] = x mol/l

As we know that,

Ksp = [A+] [B-]

Where A and B are the ions dissolved

In the above reaction,

[A+] = C6H5COO-

[B-] = Ag+

: Ksp = [C6H5COO-][Ag+]

As Ksp = 2.5×10^{-13} (given)

2.5×10^{-13} = x2

x = 5×10^{-7}

Solubility in buffer of pH =3.19

As we know that,

pH = –log [H+]

: –log [H+] = 3.19

By taking antilog of both the sides, we get

[H+] = antilog -3.19

[H+] = 6.457 × 10-4 M

C6H5COO- ions combine with the H= ions o from benzoic acid but [H+] remains almost constant because we have buffer solution. Now,

C6H5COOH $\rightarrow$ C6H5COO- + H+

As we know that,

Ka = (the concentration of products) / (the concentration of reactants)

Ka = [C6H5COO$^-$][H$^+$] / [C6H5COOH]

= [C6H5COOH]^2 / [C6H5COO$^-$] = H/Ka

[H+] = 6.457 × 10^{-4} M (calculated above)

Ka = 6.46 × 10^{-5} (given)

:[C6H5COOH] / [C6H5COO$^-$] = H$^+$ / Ka = 6.457 x 10^{-4} / 6.46 x 10^{-5} = 10

[C6H5COOH] = 10 [C6H5COO-]

Suppose solubility in the buffer solution is y mol/l.

Then as most of the benzoate ion is converted into benzoic acid molecules(which remains almost ionized) we have

Y = [Ag+] = [C6H5COO-] + [C6H5COOH]

As [C6H5COOH] = 10 [C6H5COO-] (calculated above)

: y = [C6H5COO-] + 10 [C6H5COO-]

y = 11 [C6H5COO-]

[C6H5COO-] = y / 11

As we know that,

Ksp = [C6H5COO-][Ag+]

As Ksp = 2.5× 10^{-13} (given)

[Ag+] = y, [C6H5COO-] = y / 11

: 2.5× 10-13 = y / 11 × y

Y2 = 2.75 × 10^{-12}

Y = $\sqrt{2.75 \times 10^{-12}}$

Y = 1.66 × 10^{-6}

Now,

(Solubility in buffer solution / solubility in pure water)

= (y / x) = (1.66 x 10 – 6 / 5 x 10^{-7}) = 3.32

Thus, the silver benzoate is 3.32 times more soluble in a buffer of pH compared to its solubility in pure water

Question :71 What is the maximum concentration of equimolar solutions of ferrous sulphate and sodium sulphide so that when mixed in equal volumes., there is no precipitation of iron sulphide?

Answer:

FeSO4 + Na2S → FeS + Na2SO4

Let F ml of molar solution are mixed so that no precipitation of FeS is noticed. Let molarity of FeSO4 and Na2S be α M

Thus,

$[Fe^{2+}][S^{2-}] = K_{SP}$

$= 6.3 \text{ x } 10^{-18}$

= [a x V / 2V][a x V / 2v]

$= 6.3 \text{ x } 10^{-18}$

$a = 5.02 \text{ x } 10^{-9}$M

Question :72 What is the minimum volume of water required to dissolve 1 g of calcium sulphate at 298 K?

Answer:

CaSO4 → Ca^{2}+ + $SO4^{2}$-

Let s is the solubility of CaSO4

Then, we know, Ksp = [Ca^{2}+][$SO4^{2}$-]

Ksp = s.s = s^2

A/C to question,

Ksp = 9.1 × 10^-6

So, 9.1 × 10^-6 = s^2

s = √{9.1 × 10^-6} = 3.017 × 10^-3 M

So, solubility of CaSO4 = 3.017 × 10^-3 M

= 3.017 × 10^-3 mol/L

= 3.017 × 10^-3 × 136g/L [weight = molecule weight × mole . and molar weight = 136g/mol]

= 410.3 × 10^-3 g/L

= 0.4103 g/L

It means 0.4103 g CaSO4 is dissolved in 1 L

Therefore, 1g CaSO4 is dissolved in 1/0.4103

= 2.437 L

Question :73 The concentration of sulphide ion in 0.1 M HCl solution saturated with hydrogen sulphide is 1.0 x 10*-19 M. If 10 mL of this is added to 5 mL of 0.04 M solution of the following : FeSO4, MnCl2 , ZnCl 2, and CdCl2, in which of these solutions precipitation will takes place?

Answer:

Here given,

$[S^{2-}]$ (conc. of sulphide ion) = 10^-19 M

A/C to question,

10mL of sulphide ion is mixed with 5mL of 0.04M solution of different solute so, that final volume of solution is 15mL.

So, $[S^{2-}]$mix = 10 × 10^-19/15 = 6.67 × 10^-19M

$[M^{2+}]$ = 5 × 0.04/15 = 1.33 × 10^-2 M

Here, M^{2+} shows Fe^{2+}, Mn^{2+}, Zn^{2+} or Cd^{2+}

Now , ionic product of $[M^{2+}][S^{2-}]$ = 1.33 × 10^-2 × 6.67 × 10^-19 = 8.87 × 10^-22 M

But Ksp of Cds and Zns is less than ionic product of $[M^{2+}][S^{2-}]$

Hence , CdCl2 and ZnCl2 are precipitated as CdS and ZnS