Equilibrium Important Questions With Answers

NEET Chemistry 2023

1. What is the minimum concentration of SO_4^{2-} required to precipitate $BaSO_4$ in a solution containing 1 x 10^{-4} mole of Ba^{2+}? (K_{sp} for $BaSO_4$ = 4 x 10^{-10})
   a) 4 x 10^{-10} M b) 2 x 10^{-10} M **c) 4 x 10^{-6} M** d) 2 x 10^{-3} M

   Solution : -
   $BaSO_4 \rightleftharpoons Ba^{2+} + SO_4^{2-}$
   K_{sp} = 4 x 10^{-10}
   4 x 10^{-10} = 1 x 10^{-4} x s
   s = $\frac{4\times10^{-10}}{1\times10^{-4}}$ = 4 x 10^{-6} M

2. Which of the following reactions will not get affected on increasing the pressure?
   a) $2H_{2(g)} + CO_{(g)} \rightleftharpoons CH_3OH_{(g)}$ b) $4NH_{3(g)} + 5O_{2(g)} \rightleftharpoons 4NO_{(g)} + 6H_2O_{(g)}$
   c) $CH_{4(g)} + 2S_{2(g)} \rightleftharpoons CS_{2(g)} + 2H_2S_{(g)}$ d) $PCl_{s(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$

   Solution : -
   $CH_{4(g)}+2S_{2(g)} \rightleftharpoons CS_{2(g)}+2H_2S(g)$
   Since the number of moles of gaseous reactants and products are same (Δn_g= 0), the reaction will not be affected by changing the pressure.

3. For the reversible reaction:
   $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + \text{heat}$
   The equilibrium shifts in forward direction:
   a) By increasing the concentration of $NH_3(g)$ b) By decreasing the pressure
   c) By decreasing concentration of $N_2(g)$ and $H_2(g)$
   d) By increasing pressure and decreasing temperature.

   Solution : -
   For the reversible reaction:
   $N_2(g)+3H_2(g) \rightleftharpoons 2NH_3(g)$+ heat The equilibrium shifts in forward direction by increasing pressure and decreasing temperature.
   According to Le-Chatalier principle when equilibrium is disturbed by a change, the system will try to nullify the effect of change to restore the equilibrium.
   Thus, when pressure is increased, the equilibrium will shift in a direction in which there is a decrease in the number of moles of gaseous substances. This will nullify the effect of increased pressure. This happens in the forward direction.
   Similarly for an exothermic reaction, when the temperature is decreased, the equilibrium shifts in the forward direction.

4. If K_1 and k_2 are the respective equilibrium constants for the two reactions
   $XeF_6(g) + H_2O(g) \rightleftharpoons XeOF_4(g) + 2HF(g)$
   $XeO_4(g) + XeF_6(g) \rightleftharpoons XeOF_4(g) + XeO_3F_2(g)$

The equilibrium constant of the reaction,

$$XeO_4(g) + 2HF(g) \rightleftharpoons XeO_3\ F_2(g) + H_2O(g)$$

a) $K_1/(K_2)^2$ b) K_1/K_2 c) K_1/K_2 **d) K_2/K_1**

Solution : -

$$XeF_6(\mathrm{~g}) + H_2O(\mathrm{g}) \rightleftharpoons XeOF_4(\mathrm{~g}) + 2HF(\mathrm{g})$$

$$K_1 = \frac{[XeOF_4][HF]^2}{[XeF_6][H_2O]} \quad ....(i)$$

$$XeO_4(\mathrm{~g}) + XeOF_6(\mathrm{~g}) \rightleftharpoons XeOF_4(\mathrm{~g}) + XeO_3\ F_2(\mathrm{~g})$$

$$K_2 = \frac{[XeOF_4][XeO_3\ F_2]}{[XeO_4][XeF_2]} \ldots\ldots\ldots.(ii)$$

for the reaction,

$$XeO_4(\mathrm{~g}) + 2HF(\mathrm{g}) \rightleftharpoons XeO_3\ F_2(\mathrm{~g}) + H_2O(\mathrm{g})$$

$$K = \frac{[XeO_3\ F_2][H_2O]}{[XeO_4][HF]^2} \quad ....(iii)$$

By dividing eq.(ii) by (i) we get,

$$K = \frac{K_2}{K_1}$$

5. According to Lewis concept, an acid is a/an

a) proton donor b) electron pair donor c) proton acceptor **d) electron pair acceptor.**

Solution : -

Lewis gave the concept of acids and bases on the basis of the ability of a species to donate or accept an electron.

Lewis acids are those which can accept an electron pair. These include species having vacant orbital or positive charges. They should either be electron deficient or have an expandable octet.

Examples include K^+, Mg^{2+}, etc.

Lewis bases are those which can donate a pair of electrons. These include species having L.P. or negative charge. They should have an excess of electrons.

Examples include OH^-, F^-, etc.

6. Acidic character of BF_3 can be explained on the basis of which of the following concepts?

a) Arrhenius concept b) Bronsted-Lowry concept **c) Lewis concept**

d) Bronsted-Lowry as well as Lewis concept

Solution : -

Property of acidic nature of BF_3 can be explained using the Lewis concept.

According to Lewis concept if a molecular/ion can accept a lone pair of electrons then it is called an acid.

$BF_3 + H_2O \rightarrow BF_3(OH)^- + H^+$

7. 0.6 mole of PCl_5 , 0.3 mole of PCl_3 and 0.5 mole of Cl_2 are taken in a 1 L flask to obtain the following equilibrium:

$PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$

If the equilibrium constant K_c for the reaction is 0.2. Predict the direction of the reaction.

a) Forward direction **b) Backward direction** c) Direction of the reaction cannot be predicted

d) Reaction does not move in any direction

Solution : -

$PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$

$Q_c = \frac{0.5\times0.3}{0.6} = 0.25$

$K_c = 0.2$, Since, $Q_c > K_c$ reaction will proceed in backward direction.

8. In which of the following equilibrium K_c and K_p are not equal

a) $2NO_{(g)} \rightleftharpoons N_{2(g)} + O_{2(g)}$ b) $SO_{2(g)} + NO_{2(g)} \rightleftharpoons SO_{3(g)} + NO_{(g)}$ c) $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

d) $2C_{(s)} + O_{2(g)} \rightleftharpoons 2CO_{2(g)}$

Solution : -

According to option,

$2C_{(s)} + O_{2(s)} \rightleftharpoons 2CO_{2(g)}$

$\Delta n = 2 - 1 = +1$

Hence, K_c and K_p are not equal.

9. At 100^{o}c the K_w of water is 55 times its value at 25^{o}C. What will be the pH of neutral solution? (log 55 = 1.74).

a) 6.13 b) 7.00 c) 7.87 d) 5.13

Solution : -

Given condition Kw 25^{o}C = 1 x 10^{-14}

At 25^{o}C

$K_w = [H^+][H^-] = 10^{-14}$

At 100^{o}C(given)

$K_w = [H^+][H^-] = 55 \times 10^{-14}$

for a neutral solution

$[H^+] = [OH^-]$

$\therefore [H^+]^2 = 55 \times 10^{-14}$

$\text{or } [H^+] = (55 \times 10^{-14})^{1/2}$

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

Taking log on both side in (i)

$-\log[H^+] = -\log(55 \times 10^{-14})^{1/2}$

$pH = \frac{1}{2} - \log 55 + 14 \log 10$

= 6.13

10. Solution of a monobasic acid has a pH = 5. If one mL of it is diluted to 1 litre, what will be the pH of be the resulting solution?

a) 3.45 **b) 6.96** c) 8.58 d) 10.25

Solution : -

pH = 5, $[H^+] = 10^{-5}$ M

After dilution = $\frac{10^{-5}}{1000}$

= 10^{-8} M

Total $[H^+] = 10^{-8} + 10^{-7} = 1.1 \times 10^{-7}$

pH = $-\log[H^+] = -\log(1.1 \times 10^{-7})$ = 6.96

11. Which of the following is conjugate acid of SO_4^{2-}?

a) HSO_4^- b) H^+ c) H_2SO_4 d) SO_4^{2-}

Solution : -

conjugate acid of SO_4^{2-} is HSO_4^-

12. The strongest conjugate base is

a) NO_3^- b) Cl^- c) $SO_4{}^{2-}$ **d)** CH_3COO^-

Solution : -

The strength of acid is measured by the ability to lose a proton. An acid is called as strong if it completely ionizes in water while the weak acids do not ionize completely in water, rather they partially dissociates and forms equilibrium with both the acid and conjugate base in the solution.

The ethnic acid (acetic acid) is considered as a weak acid because it does not releases all of its hydrogen in water; rather it dissociates partially and establishes equilibrium with its conjugate base.

$H_3COOH \rightleftharpoons CH_3COO^- + H^+$

Strong conjugate base has a weak conjugate acid. Since CH_3COOH is weakest acid, therefore its conjugate base.

13. Equimolar solution of the following substances were prepared separately. Which one of these will record the highest pH value?

a) $BaCl_2$ b) $AlCl_3$ c) LiCl d) $BeCl_2$

Solution : -

All of the given salts have same anion i.e., Cl which on hydrolysis gives HCl which is a strong acid. Now, among the salts which have cation that gives a strongest base on hydrolysis of salt have the highest pH value. As Ba form $Ba(OH)_2$ which is a stronger base thus, it results in the highest pH value.

14. What is pOH of an aqueous solution with hydrogen ion concentration equal to 3×10^{-5} mol L^{-1}?

a) 9.47 b) 4.52 c) 12.69 d) 11.69

Solution : -

pH = $-\log[H^+] = -\log(3 \times 10^{-5}) = 4.5229$

pOH = 14 - pH, 14 - 4.5229 = 9.47

15. The correct relationship between free energy change in a reaction and the corresponding equilibrium constant, K is

a) $-\Delta G = RT \ln K$ b) $\Delta G^o = RT \ln K$ **c) $\Delta G = -RT \ln K$** d) $-\Delta G^o = RT \ln K$

Solution : -

The correct relationship between free energy change in a reaction and the corresponding equilibrium constant K is $-\Delta G = RT \ln K$ or $\Delta G = -RT \ln K$

16. Which one of the following molecular hydrides acts as a Lewis acid?

a) NH_3 b) H_2O c) B_2H_6 **d) CH_4**

Solution : -

Electron-deficient molecules behave as Lewis acid. Among the given molecules, only diborane is the electron deficient, it does not have the complete octet. Hence it behaves like a Lewis acid.

17. Which one of the following compounds is not a protonic acid?

a) $SO_2(OH_2)$ **b) $B(OH)_3$** c) $PO(OH)_3$ d) $SO(OH_2)$

Solution : -

The acid which gives H^+ when placed in aqueous solution is called protonic acid.

Ortho boric acid does not donate proton like most of the acids but rather it accepts OH^- therefore it is a lewis acid.

$B(OH)_2 + 2H_2O \leftrightharpoons H_3O^+ + [B(OH)_4]^-$

18. A solution which is 10^{-3} M each in Mn^{2+}, Fe^{2+}, Zn^{2+} and Hg^{2+} is treated with 10^{-16} M sulphide ion. If K_{sp} of MnS, FeS, ZnS and HgS are 10^{-15}, 10^{-25}, 10^{-20} and 10^{-54} respectively, which one will precipitate first?

a) FeS b) MnS **c) HgS** d) ZnS

Solution : -

Ionic product in the solution = $10^{-3} \times 10^{-16} = 10^{-19}$. The metal sulphide having the lowest solubility will precipitate first provided the ionic product is higher than the K_{sp}. Here, all salts are of the same valence type. So, the sulphide having the lowest K_{sp} value will precitate first provided $K_{sp} < 10^{-19}$. hgS has the lowest K_{sp} value (10^{-54}), so it will precipitate first.

19. For dibasic acid correct order is

a) K_{a1}a2 **b) $K_{a1} > K_{a2}$** c) $K_{a1} = K_{a2}$ d) not certain

Solution : -

In polyprotic acids the loss of second proton occurs much less readily than the first. Generally the K_a values for successive loss of protons from these acids differ by at least a factor of 10^{-3} i.e.,

$$K_{a_1} < K_{a_2}$$
$$H_2X \rightleftharpoons H^+ + HX^- \ (K_{a_1})$$
$$HX^- \rightleftharpoons H^+ + H^{2-} \ (K_{a_2})$$

20. Which of the following is not true about a reversible reaction?
a) The reaction does not proceed to completion. b) It cannot be influenced by a catalyst
c) Number of moles of reactants and products is always equal.
d) It can be attained only in a closed container

Solution : -
In a reversible reaction, number of moles of reactants and products is not always equal.
It is attained only if the system is closed.
A catalyst does not alter the equilibrium point. It alters the rate of the reaction.
The equilibrium can be shifted either to left or right by altering the volume, pressure, temperature etc.

21. Identify the correct order of solubility in aqueous medium:
a) $ZnS > Na_2S > CuS$ **b) $Na_2S>CuS>ZnS$** c) $Na_2S > ZnS > CuS$ d) $CUS>ZnS>Na_2S$

Solution : -
The order or strength of solubility in the aqueous medium depends upon the lattice energy, hybridization and size of the cation.
The anions here are sulphide ions and the size of cations decrease across the period: $Na^+ > Zn^+ > Cu^+$
solubility will be more when there is the maximum distance between the atoms. $Na_2S > ZnS > Cus$

22. If the value of an equilibrium constant for a particular reaction is 16×10^{12}, then at equilibrium the system will contain:
a) mostly reactants **b) mostly products** c) similar amounts of reactants and products d) all reactants

Solution : -
Given equilibrium constant for reaction
$$\mathrm{K} = 1.6 \times 10^{12} = \frac{[\,\text{Product}\,]}{[\,\text{Reactant}\,]}$$
Then the value of K is very high so the system will contain max. products at equilibrium.

23. For the reaction, $2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$. What is K_c when the equilibrium concentration of $[SO_2] = 0.60$ M, $[O_2] = 0.82$ M and $[SO_3] = 1.90$ M?
a) 12.229 L mol^{-1} b) 24.5 L mol^{-1} c) 36.0 L mol^{-1} d) 2.67×10^3 L mol^{-1}

Solution : -
The formula for calculating equilibrium constant $K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$

$$K_c = \frac{(1.9^2)}{[(0.6)^2] \times [0.82]}$$
$$K_c = \frac{3.61}{0.36 \times 0.82}$$
$$K_c = \frac{3.61}{0.2952}$$
$\therefore K_c = 12.229$L/mol

24. For the reversible reaction:
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$+ heat
The equilibrium shifts in forward direction:
a) By increasing the concentration of $NH_{3(g)}$ b) By decreasing the pressure
c) By decreasing the concentration of $N_{2(g)}$ and $H_{2(g)}$
d) By increasing pressure and decreasing temperature.

Solution : -
For the reversible reaction:
$N_2(g)+3H_2(g) \rightleftharpoons 2NH_3\,(g)$+ heat The equilibrium shifts in forward direction by increasing pressure and decreasing temperature.
According to Le-Chatalier principle when equilibrium is disturbed by a change, the system will try to nullify the effect of change to restore the equilibrium.

Thus, when pressure is increased, the equilibrium will shift in a direction in which there is a decrease in the number of moles of gaseous substances. This will nullify the effect of increased pressure. This happens in the forward direction.

Similarly for an exothermic reaction, when the temperature is decreased, the equilibrium shifts in the forward direction.

25. When sulphur is heated at 900 K, S_8 is converted to S_2 . What will be the equilibrium constant for the reaction if initial pressure of 1 atm falls by 25% at equilibrium?

a) 0.75 atm^3 b) 2.55 atm^3 c) 25.0 atm^3 **d) 1.33 atm^3**

Solution : -

The given reaction is:-

$S_8 \rightleftharpoons 4S_2$

Initial pressure: 1 atm 0 (given)

At equilibrium: 1- 0.25 4×0.25 = 1 atm (At eqm P of S_8 falls by 25%)

= 0.75 atm

So, $K_P = \frac{(P_{S_2})^4}{P_{S_8}}$

$= \frac{(1)^4}{0.75} = \frac{4}{3} = 1.33 atm^3$

26. The hydrogen ion concentration of a 10^{-8} M HCl aqueous solution at 298 K (Kw = 10^{-14}) is

a) 11 x 10^{-8}M b) 9.525 x 10^{-8}M c) 10 x 10^{-8}M d) 10 x 10^{-6} M

Solution : -

Given solution of 10^{-8}M HCl $[H^+] = 10^{-8}$

$[H^+]$ of water = 10^{-7}

$\text{Total } [H^+] = 10^{-7} + 10^{-8} = 10 \times 10^{-8} \times 10^{-8}$

$10^{-8}(10+1) = 11 \times 10^{-8}$

27. The solubility product of $BaCl_2$ is 3.2 x 10^{-9}. What will be its solubility in mol L^{-1}?

a) 4 x 10^{-3} b) 3.2 x 10^{-9} **c) 1 x 10^{-3}** d) 1 x 10^{-9}

Solution : -

$BaCl_2 \rightleftharpoons Ba^{2+} + 2Cl^-$

$K_{sp} = [Ba^{2+}][Cl^-]^2 = x \times (2x)^2 = 4x^3$

$4x^3$ = 3.2 x 10^{-9}

$\Rightarrow x$ = 9.28 x 10^{-4} = 0.928 x 10^{-3} = 1 x 10^{-3}

28. Equimolar solutions of the following substances were prepared separately. Which of these will record the highest pH value

a) $BaCl_2$ b) $AlCl_3$ c) LiCl d) $BeCl_2$

Solution : -

($AlCl_3$, LiCl and $BeCl_2$) all these solutions are acidic because of cationic hydrolysis, whereas BaClr, is salt of strong base ($Ba(OH)_2$) and strong acid (HCl), thus it will have maximum PH.

29. The pH value of a 10M solution of HCl is :

a) Less than 0 b) Equal to 0 c) Equal to 1 d) Equal to 2

Solution : -

HCl(aq) → H^+(aq) + Cl^-(aq) [S=$\sqrt{K_{sp}}$]

[HCl] = 10 M

$\Rightarrow [H^+]$ = 10 mol / L

pH = - loh$[H^+]$ = -log 10

= -1, so the pH is less than zero.

30. pH of a saturated solution of $Ba(OH)_2$ is 12. The value of solubility product K_{sp} of $Ba(OH)_2$ is :

a) 3.3×10^{-7} **b) 5.0×10^{-7}** c) 4.0×10^{-6} d) 5.0×10^{-6}

Solution : -

$Ba(OH)_2 \rightleftharpoons Ba^{2+} + 2OH$

At equilibrium x 2x

$12 = -\log [H^+] \Rightarrow [H^+] = 10^{-12}$

As, $[H^+][OH^-]$- $K_w = 10^{-14}$

$10^{-12}[OH^-] = 10^{-14} \Rightarrow [OH^-] = 10^{-2}$

If $[OH^-] = 2x = 10^{-2}$ then $x = 5.0 \times 10^{-3}$

Now, $K_{sp} = [Ba^{2+}][OH^-]2$

$K_{sp} = (5 \times 10^{-3})(10^{-2})^2 = 5.0 \times 10^{-7}$

31. The compound whose aqueous solution hai the highest pH is

a) NaCl b) $NaHCO_3$ **c) Na_2CO_3** d) NH_4Cl

Solution : -

NaCl is a salt of strong acid and strong base hence its aqueous solution will be neutral ie pH=7.$NaHCO_3$ is an acidic salt hence pH<7. Na_2 CO_3 is a salt of weak acid and strong base. Hence its aqueous solution will be strongly basic ie. pH>7.NH_4Cl is salt of weak base and strong acid, hence its aqueous solution will be strongly acidic i.e. pH<7.

32. For which of the following reactions, $K_p = K_c$?

a) $PCl_{3(g)} + Cl_{2(g)} \rightleftharpoons PCl_{5(g)}$ **b) $H_{2(g)} + Cl_{2(g)} \rightleftharpoons 2HCl_{(g)}$** c) $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

d) $CaCO_{3(g)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$

Solution : -

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

If $\Delta n=0$, $K_p = K_c$

If $\Delta n > 0$, then $K_p > K_c$

If $\Delta n < 0$, then $K_p < K_c$

For the reaction: $H_2(g) + Cl_2(g) \rightleftharpoons 2HCl_{(g)}$

Δn is the number of moles of gaseous products - number of moles of gaseous reactants in a balanced equation.

$\Delta n = 2 - (1 + 1) = 0$

$\therefore K_p = K_c$

33. What will be the pH of 1×10^{-4} M H_2SO_4 solution?

a) 10.4 **b) 3.70** c) 3 d) 13

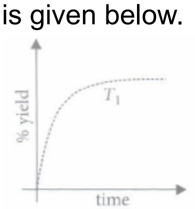
Solution : -

$H_2SO_4 \rightleftharpoons 2H^+ + SO_4^{2-}$

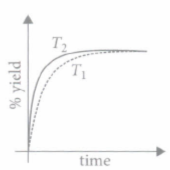
$[H^+] = 2 \times 1 \times 10^{-4}$ M

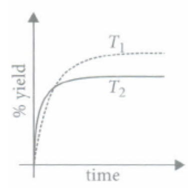
$pH = -\log (2 \times 10^{-4}) = 3.70$

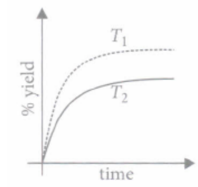
34. The % yield of ammonia as a function of time in the reaction $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}, \Delta H < 0$ at (P,T_1) is given below.

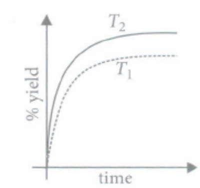


If this reaction is conducted at (P, T_2), with $T_2 > T_1$, the % yield of ammonia as a function of time is represented by

a) 

b) 

c) 

d) 

Solution : -

$N_{2(g)} + 3H_{2(g)} \underset{endo}{\overset{exo}{\rightleftharpoons}} 2NH_{3(g)}, \Delta H < 0$

Initially, with increase in temperature ($T_2 > T_1$) % yield increases.Afterwards, equilibrium is reached and if the temperature is increased, i.e., heat is supplied to the system, then according to Le Chatelier's principle, the equilibrium will shift in the backward direction, where the heat is absorbed. Hence, the % yield decreases.

35. The pH value of blood does not change appreciably by a small addition of an acid or base, because the blood
a) is a body fluid b) can be easily coagulated c) contains iron as a part of the molecule
d) contains serum protein that acts as buffer

Solution : -

The buffer system present in serum is H_2CO_3 + $NaHCO_3$ and as we know that a buffer solution resist the change in pH therefore pH value of blood does not change by a small addition of an acid or a base.

36. The dissociation equilibrium of a gas AB_2 can be represented as $2AB_{2(g)} \rightleftharpoons 2AB_{(g)} + B_{2(g)}$ The degree of dissociation is x and is small compared to 1. The expression relating the degree of dissociation (x) with equilibrium constant K_p and total pressure P is :
a) $(2K_P/P)$ **b) $(2K_P/P)^{1/3}$** c) $(2K_p/P)^{1/2}$ d) (K_p/P)

Solution : -

$2AB_{2(g)} \rightleftharpoons 2AB_{(g)} + B_{2(g)}$
2 0 0 (Initial)
2(1 - x) 2x x (at equilibrium) Amount of moles at equilibrium
= 2(1 - x) + 2x + x = 2 + X

$K_p = \frac{[p_{AB}]^2[p_{b_2}]}{[p_{AB_2}]^2}$

$K_p = \frac{\left(\frac{2x}{2+x}P\right)^2 \times \left(\frac{x}{2+x} \times P\right)}{\left(\frac{2(1-x)}{2+x} \times P\right)} = \frac{\frac{4x^3}{2+x} \times P}{4(1-x)^2}$

$K_p = \frac{4x^3 \times P}{2} \times \frac{1}{4}$ ($\because$1 - x≈ 1 and 2 + x ≈2)

$x = \left(\frac{8K_p}{4P}\right)^{1/3} \Rightarrow x = \left(\frac{2K_p}{P}\right)^{1/3}$

37. 5 moles of SO_2 and 5 moles of O_2 react in a dosed vessel. At equilibrium 60% of the SO_2 is consumed. The total number of gaseous moles (SO_2 , O_2 and SO_3) in the vessel is
a) 5.1 b) 3.9 c) 10.5 **d) 8.5**

Solution : -

$2SO_2 + O_2 \rightleftharpoons 2SO_3$
Initial moles 5 5 0
At equilibrium 5-3 $5 - \frac{3}{2} 3 (5 \times \frac{60}{100} = 3)$
Total number of moles in the vessel = 2 + 3.5 + 3 = 8.5

38. The rate of reaction depends upon the :
a) Volume b) Force c) Pressure **d) Concentration of reactants**

Solution : -

According to law of mass action, the rate of a reaction depends upon the active masses or molar concentration of reactants. Rate of reaction ∝ Concentration of reactants

39. Which of these is least likely to act as Lewis base?

a) CO b) F^- **c) BF_3** d) PF_3

Solution : -

CO, F^- and PF_3 can either donate a lone pair of electron or pi electrons. Hence, they act as Lewis bases.
On the other hand, BF_3 has vacant orbital in which it can accent a lone pair of electrons or pi electrons. Hence, it acts as a Lewis acid. Thus, it has least tendency to act as a Lewis base.

40. The solubility product of sparingly soluble salt AX_2 is 3.2×10^{-11}. Its solubility (in mol/L) is :

a) 5.6×10^{-6} b) 3.1×10^{-4} **c) 2×10^{-4}** d) 4×10^{-4}

Solution : -

$AX_2 \rightleftharpoons A^{2+} + 2X^-$

s mol L^{-1} s mol L^{-1} 2s mol L^{-1}

$K_{sp} = [A^{2+}] [X^-]^2 = s \times (2s)^2 = 4s^3$

$3.2 \times 10^{-11} = 4s^3$

$\Rightarrow s^3 = 8 \times 10^{-12}$

$s = 2 \times 10^{-4}$ mol L^{-1}

41. For the reaction

$\mathrm{CH_4(g) + 2O_2(g) \rightleftharpoons CO_2(g) + 2H_2O(l)}$

$\triangle H_r = 170.8$ kJmol^{-1}

Which of the following statement is not true?

a) The equilibrium constant for the reaction is given by $\mathrm{K_C} = \frac{[\mathrm{CO_2}]}{[\mathrm{CO_4}][\mathrm{O_2}]}$

b) Addition of $CH_{4(g)}$, or $O_{2(g)}$ at equilibrium will cause a shift to the right. c) The reaction is exothermic

d) At equilibrium, the concentrations of $CO_2(g)$ and $H_2O(l)$ are not equal

Solution : -

The given reaction is:-

$CH_4(g) + 2O_2(g) \rightleftharpoons CO_2(g) + 2H_2O$ (g)

Now, $K_C = \frac{[CO_2][H_2O]^2}{[CH_4][O_2]^2}$

Now, H_2O is pure liquid, so, $[H_2O]=1$

$\Rightarrow K_C = \frac{[CO_2]}{[CH_4][O_2]^2}$

$\because \triangle Hr = -170.8$KJ/mol is negative, so reaction is exothermic by adding O_2 (g) or CH_4(g) at equilibrium, by Le Chatelier's principle, the equilibrium shift towards right side.

42. A buffer solution is prepared in which the concentration of NH_3 is 0.30 M and the concentration of NH_4^+ is 0.20 M. If the equilibrium constant, K_b for NH_3 equals 1.8×10^{-5}, what is the pH of this solution? (log 2.7 = 0.43)

a) 9.43 b) 11.72 c) 8.73 d) 9.08

Solution : -

$pOH = pK_b + \log \frac{[Salt]}{[base]}$

$= -\log Kb + \log \frac{[Salt]}{[base]}$

$= -\log 1.8 \times 10\text{-}5 + \log \frac{0.20}{0.30}$

= 5 - 0.25 + (-0.176) = 4.57

Now, pH = 14 - pOH = 14 - 4.57 = 9.43

43. Consider the following liquid-vapour equilibrium

$\mathrm{Liquid} \rightleftharpoons \mathrm{Vapour}$

Which of the following relations is correct?

a) $\frac{d\ell n\mathrm{G}}{\mathrm{dT}^2} = \frac{\Delta \mathrm{H}_v}{\mathrm{RT}^2}$ b) $\frac{d\ell n\mathrm{P}}{\mathrm{dT}} = \frac{\Delta \mathrm{H}_v}{\mathrm{RT}}$ c) $\frac{d\ell n\mathrm{P}}{\mathrm{dT}^2} = \frac{-\Delta \mathrm{H}_v}{\mathrm{T}^2}$ **d)** $\frac{d\ell n\mathrm{P}}{d\,\mathrm{T}} = \frac{\Delta \mathrm{H}_v}{\mathrm{RT}^2}$

Solution : -

According to Clausius - Clapeyron's equation

$\frac{dlnP}{dT} = \frac{\triangle_{Hv}}{RT^2}$

According to this equation, the rate at which the natural logarithm of the vapor pressure of a liquid changes with temperature is determined by the molar enthalpy of vaporization of the liquid, the ideal gas constant, and the temperature of the system.

44. The pH of neutral water at 25°C is 7.0. As the temperature increases, ionisation of water increases, however, the concentration of H^+ ions and OH^- ions are equal. What will be the pH of pure water at 60°C?

a) Equal to 7.0 b) Greater than 7.0 **c) Less than 7.0** d) Equal to zero

Solution : -

The pH of neutral water at 25°C is 7.0

$\therefore [H^+] = [OH^-] = 10^{-7}$ ($\because$ pH = -log $[H^+]$)

Now, $K_w = [H^+][OH^-] = (l \times 10^{-7})^2 = 1 \times 10^{-14}$

As the temperature increases, ionisation of water increases, thus [H+] and [OH-] increases equally. Now , $K_w = [H^+][OH^-] > 1 \times 10^{-14}$ ($\because [H^+]=[OH^-]$)

or $[H^+]2 > 1 \times 10^{-14}$

$\therefore [H^+] > 1 \times 10^{-7}$ and pH < 7

45. Which of the following is most soluble?

a) $\mathrm{Bi_2\ S_3}\ (\mathrm{K_{sp}} = 1 \times 10^{-70})$ **b)** $\mathrm{MnS}\ (\mathrm{K_{sp}} = 7 \times 10^{-16})$ c) $\mathrm{CuS}(\mathrm{K_{sp}} = 8 \times 10^{-37})$

d) $\mathrm{Ag_2\ S}\ (\mathrm{K_{cn}} = 6 \times 10^{-51})$

Solution : -

Solubility of an electrolyte increases with increase in their ionic product i.e. solubility product. Now,

K_{sp} (MnS) > K_{sp} (CuS) > K_{sp} (Ag_2S) > K_{sp}(Bi_2S_3)

Therefore MnS will be most soluble.

46. Given the reaction between 2 gases represented by A_2 and B_2 to give the compound AB(g).

$\mathrm{A_2(g) + B_2(g) \rightleftharpoons 2AB(g)}$

At equilibrium, the concentration

of $A_2 = 3.0 \times 10^{-3}$ M

of $B_2 = 4.2 \times 10^{-3}$M

of AB = 2.8×10^{-3}M

If the reaction takes place in a sealed vessel at 527°C, then the value of K. will be:

a) 2.0 b) 1.9 **c) 0.62** d) 4.5

Solution : -

$A_2 + B_2 \rightleftharpoons 2AB; \quad K_c = \frac{[AB]^2}{[A_2][B_2]}$

$K_c = \frac{(2.8 \times 10^{-3})^2}{3 \times 10^{-3} \times 4.2 \times 10^{-3}}$

$= \frac{(2.8)^2}{3 \times 4.2} = 0.62$

47. In the following question, a statement of assertion is followed by a statement of reason. Mark the correct choice as :

Assertion: For the reaction:

$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$, $K_p = K_c$

Reason: Concentration of gaseous reactants and products is taken as unity.

a) If both assertion and reason are true and reason is the correct explanation of assertion.

b) If both assertion and reason are true but reason is not the correct explanation of assertion

c) If assertion is true but reason is false. **d) If both assertion and reason are false.**

Solution : -

(I) The given reaction is:-

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$; $K_p = K_C$

we know that, $K_P = K_C(RT)^{\Delta n}$

Δn = no. of moles of gaseous products - No. of moles of gaseous reactants.

= 2 - (3 + 1) = -2

$\Rightarrow K_P = K_C(RT)^{-2}$.

(II) the concentrations of pure solids and liquids are taken as unity.

48. For the reaction a + b $\rightleftharpoons$ c + d, initially concentrations of a and b are equal and at equilibrium the concentration of d will be twice of that of a. What will be the equilibrium constant for the reaction?

a) 2 b) 9 **c) 4** d) 3

Solution : -

a + b $\rightleftharpoons$ c + d

Initial concentrations of 'a' and 'b' are equal let it be 'x'.

At equilibrium, the concentration will be twice that of 'a' i.e. 2x.

Since concentration of c and d are equal thus the concentration of c at equilibrium will be 2x.

$\therefore$ The equilibrium constant for the reaction $K_c = \frac{2x \times 2x}{x \times x} = \frac{x^2}{x^2}$

$\therefore K_c = 4$.

49. For a reversible reaction at 298 K the equilibrium constant K is 200. What is the value of ΔG^0 at 298 K?

a) -13.13 kcal b) -0.13 kcal **c) -3.158 kcal** d) -0.413 kcal

Solution : -

Applying ΔG^0 = - 2.303 RT x 10gK

= - 2.303 x 2 x 298 x log 200

= - 3158.4 cal = - 3.158 kcal

50. In the system X + 2 Y $\rightleftharpoons$ Z, the equilibrium concentrations are, [X] = 0.06 mol L^{-1}, [y] = 0.12 mol L^{-1}, [Z] = 0.216 mol L^{-1}. Find the equilibrium constant of the reaction.

a) 250 b) 500 c) 125 d) 273

Solution : -

X + 2Y $\rightleftharpoons$ Z

0.06 0.12 0.216

$K_c = \frac{[Z]}{[X][Y]^2} = \frac{0.216}{0.06 \times 0.12 \times 0.12} = 250$