

# Chapter 7. Equilibrium

## Question-1

Why is vapour pressure independent of initial amount?

### Solution:

Vapour pressure of liquid is defined as pressure exerted by liquid at equilibrium at a given temperature. Different amounts of liquids exert same equilibrium pressure at a given temperature.

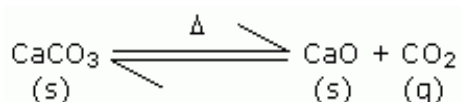
At equilibrium, rate of evaporation = rate of condensation

Different amounts of liquid will give same ratio of amount of evaporation to amount of condensation.

## Question-2

Write  $K_p$  for  $\text{CaCO}_3$  decomposition.

### Solution:



$K_p = p_{\text{CO}_2}$  where  $p_{\text{CO}_2}$  - Partial vapor pressure.

## Question-3

At 500 k, 1.04 g of hydrogen, and 12,060 g of  $\text{I}_2$  are in equilibrium with 5.058 g of hydrogen iodide in a vessel of capacity 1 lit. Calculate the equilibrium constant.

### Solution:

$$[\text{H}_2] = \frac{1.04}{2} \text{ g/mole} \quad [\text{H I}] = \frac{5.058}{128} \text{ g/mole} \quad (\text{H}_2 = 2; \text{I}_2 = 254; \text{HI} = 128)$$



$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{\left(\frac{5.058}{128}\right)^2}{\left(\frac{1.04}{2}\right)\left(\frac{12060}{254}\right)} = 6.48$$

$$K_c = 6.48.$$

### Question-4

If the value of  $K_c$  be 50.21, for the reaction  $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$  how much HI would be present at the equilibrium, if we start with 5.30 mol of  $\text{I}_2$  and 7.94 mol of  $\text{H}_2$ ?

#### Solution:

At equilibrium of  $2x$  mol HI is formed,  $x$  mol  $\text{H}_2$  and  $x$  mole of  $\text{I}_2$  should be consumed.

$$\text{At equilibrium } [\text{HI}] = 2x; [\text{H}_2] = [7.94 - x]$$

$$[\text{I}_2] = [5.30 - x]$$

$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]}$$

$$50.21 = \frac{[2x]^2}{[7.94 - x][5.30 - x]}$$

$$4x^2 = 2112.93 - 664.78x + 50.21 x^2$$

$$46.21 x^2 - 664.78 x + 2112.93 = 0;$$

$$x^2 - 14.3 x + 46 = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = \frac{14.3 \pm \sqrt{205 - 184}}{2}$$

$$x = \frac{14.3 \pm 4.58}{2}$$

$$2x = 9.72 \text{ moles}$$

At equilibrium  $[\text{HI}] = 9.72$  moles.

### Question-5

$K_c$  for  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$  is 61. What is  $K_c$  for  $2\text{NH}_3 \rightleftharpoons \text{N}_2(\text{g}) + 3\text{H}_2(\text{g})$ ?

#### Solution:

Equilibrium constant for the reverse reaction is the inverse of the equilibrium constant for the reaction in the forward reaction.

$$K'_c \text{ for } 2\text{NH}_3 \rightleftharpoons \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \text{ is } \frac{1}{61}.$$

### Question-6

For the equilibrium  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$   $K_p$  is  $6.8 \times 10^5$  at 298 K. What is the  $K_c$  value for the same equilibrium?

#### Solution:

For the reaction  $\Delta n = 2 - (1 + 3) = -2$

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

$$6.8 \times 10^5 = K_c \times (0.0831 \times 298)^{-2}$$

$$6.8 \times 10^5 = K_c \times \frac{1}{(0.0831 \times 298)^2}$$

$$\text{(i.e.) } K_c = 6.8 \times 10^5 \times (0.0831 \times 298)^2 = 4170.1 \times.$$

### Question-7

If the reaction between iron and steam proceeds as:  $3\text{Fe} + 4\text{H}_2\text{O} \rightleftharpoons \text{Fe}_3\text{O}_4 + 4\text{H}_2$  and the partial pressure of steam be 50 mm and of hydrogen 940 mm at  $250^\circ$ . Calculate the pressure of steam at equilibrium when the partial pressure of hydrogen is 1800 mm.

#### Solution:



$$K_p = \frac{(p_{\text{H}_2})^4}{(p_{\text{H}_2\text{O}})^4} = \frac{(940)^4}{(50)^4}$$

When  $p_{\text{H}_2} = 1800$  mm

$$K_p = \frac{(940)^4}{(50)^4} = \frac{(1800)^4}{x^4} \text{ where } x \text{ is } p_{\text{H}_2\text{O}}$$

$$x = \frac{50 \times 1800}{940} = 95.7.$$

∴ Partial pressure of steam is 95.7 mm.

### Question-8

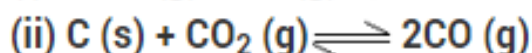
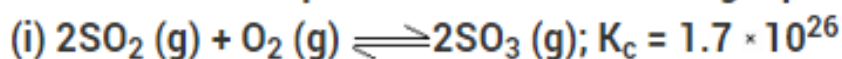
Why is NaCl precipitated by passing HCl gas in a NaCl saturated solution?

#### Solution:

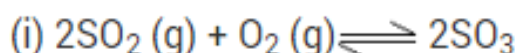
$\text{NaCl} \rightleftharpoons \text{Na}^+ \text{Cl}^-$ , when HCl is passed,  $[\text{Cl}^-]$  is more. To decrease the concentration of  $\text{Cl}^-$  backward reaction takes place. Hence NaCl is precipitated. This could also be attributed to common ion effect.

### Question-9

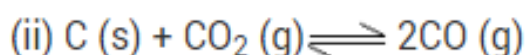
What is effect of pressure for the following equilibrium?



#### Solution:



When equilibrium pressure is increased, to annul the effect, the reaction in which pressure is consumed will take place. Pressure is used to decrease the volume (i.e.) forward reaction will take place where 3 moles are reduced to 2 moles. The equilibrium is shifted to forward direction.



When pressure increased the reaction in which number of moles reduced will take place. (i.e.) equilibrium is shifted to reverse direction.

### Question-10

$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) ; \Delta H^\circ = -92.38 \text{ kJ mole}^{-1}$ , what happens temp is increased at equilibrium?

#### Solution:

When temperature is increased at equilibrium, the reaction in which heat is consumed will take place to annul the effect of increase in temperature. Backward reaction is a heat consuming reaction. Hence, decomposition of ammonia takes decomposition when temp is increased.

### Question-11

Explain the dynamic equilibrium between ice and water.

#### Solution:

The dynamic equilibrium consists of a forward reaction in which ice melts to give water and a reverse reaction in which water solidifies to ice.



At 273K, both ice and water are in equilibrium. Both forward and backward reactions occur at the same rate. At  $T > 273\text{K}$ , more water is present and  $\Delta G < 0$ . At  $T < 273\text{K}$ , more ice is present and  $\Delta G > 0$ .

### Question-12

Explain why the equilibrium constant for a gaseous reaction can be written in terms of partial pressures instead of concentrations.

#### Solution:

The molar concentration of a gas is directly proportional to its partial pressure at a fixed temperature. Therefore, it is convenient to express the composition of a gaseous reaction mixture in terms of the partial pressures of the components rather than in terms of their molar concentrations.

### Question-13

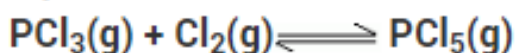
The equilibrium constant of a reaction decrease with increase in temperature. Is the reaction exothermic or endothermic?

#### Solution:

Exothermic because with increase in temperature, the equilibrium shifts in the backward direction.

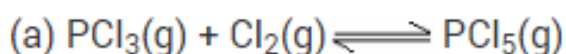
### Question-14

(a) Predict the direction of reaction when chlorine gas is added to an equilibrium mixture of  $\text{PCl}_3$ ,  $\text{PCl}_5$  and  $\text{Cl}_2$ . The reaction is

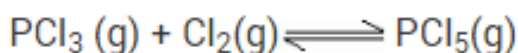


(b) What is the direction of reaction when chlorine gas is removed from an equilibrium mixture of these gases?

#### Solution:



(a) When  $\text{Cl}_2$  is added to the reaction mixture, increasing its concentration, the reaction goes in the forward direction (more  $\text{PCl}_5$  is formed).

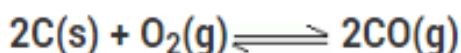


(b) When  $\text{Cl}_2$  is removed from the reaction mixture, lowering its concentration, the reaction goes in the reverse direction (more  $\text{PCl}_5$  dissociates to  $\text{PCl}_3$  and  $\text{Cl}_2$ ) to partially restore the  $\text{Cl}_2$  that was removed.



### Question-15

What is the effect of reducing the volume on the system described below?



#### Solution:

The forward reaction is accompanied by increase in volume. Hence according to Le-Chatelier's principle, reducing the volume will shift the equilibrium in the forward direction.

### Question-16

Consider the reaction:



Indicate the direction in which the equilibrium will shift when

(i) a catalyst is added

(ii) pressure is decreased

#### Solution:

(i) **A catalyst is added:** The equilibrium will not shift in any direction because both the forward and backward rates are increased to the same extent.

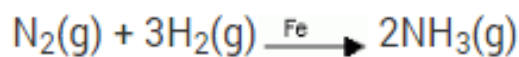
(ii) **Pressure is decreased:** The equilibrium will shift to the left to produce more number of moles of the reactants.

### Question-17

What is the effect of catalyst on the equilibrium? Give an example.

#### Solution:

A catalyst is a substance that increases the rate of a reaction but is not consumed by it. However, the presence of a catalyst speeds the reverse reaction as well as the forward reaction, leaving the equilibrium constant unchanged. For example,



### Question-18

Classify conjugate pairs among the following:

Q.  $\text{HCl}$ ,  $\text{HS}^-$ ,  $\text{HSO}_4^{2-}$ ,  $\text{H}_2\text{O}$ ,  $\text{HCO}_3^-$ ,  $\text{H}_2\text{S}$ ,  $\text{HPO}_4^{2-}$ ,  $\text{NH}_3$ ,  $\text{Cl}^-$ ,  $\text{H}_3\text{O}^+$ ,  $\text{NH}_4^+$ ,  $\text{SO}_4^{2-}$ .

#### Solution:

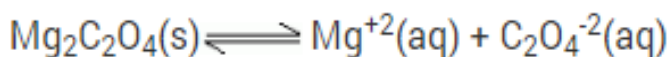
$\text{HCl}$  and  $\text{Cl}^-$ ,  $\text{H}_2\text{S}$  and  $\text{HS}^-$ ,  $\text{HSO}_4^{2-}$  &  $\text{SO}_4^{2-}$ ,  $\text{H}_3\text{O}^+$  and  $\text{H}_2\text{O}$ .

### Question-19

The solubility of magnesium oxalate,  $\text{MgC}_2\text{O}_4$ , in water is  $0.0093 \text{ mol/L}$ .

Calculate  $K_{\text{sp}}$ .

#### Solution:



Starting                      0                      0

Equilibrium                 $9.3 \times 10^{-3}$      $9.3 \times 10^{-3}$

$$[\text{Mg}^{+2}] = 9.3 \times 10^{-3} \text{ mol L}^{-3}. [\text{C}_2\text{O}_4^{-2}] = 9.3 \times 10^{-3} \text{ mol L}^{-1}$$

$$K_{\text{sp}} = [\text{Mg}^{+2}][\text{C}_2\text{O}_4^{-2}] = (9.3 \times 10^{-3})(9.3 \times 10^{-3}) = 8.65 \times 10^{-5} \text{ M}^2.$$

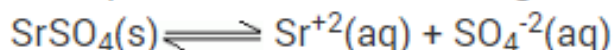


### Question-20

What is the solubility of strontium sulphate,  $\text{SrSO}_4$ , in 0.15 M sodium sulphate,  $\text{Na}_2\text{SO}_4$ ?  $K_{\text{sp}}$  of  $\text{SrSO}_4 = 2.5 \times 10^{-7}$

#### Solution:

The equation for the dissolving  $\text{SrSO}_4$  is



In this solution there are two sources of  $\text{SO}_4^{-2}(\text{aq})$ . The soluble  $\text{Na}_2\text{SO}_4$  is completely dissociated into  $\text{Na}^+$  ions and  $\text{SO}_4^{-2}(\text{aq})$  ions, and the small amount of  $\text{SrSO}_4$  that dissolves also provides  $\text{SO}_4^{-2}(\text{aq})$  ions. The  $\text{SO}_4^{-2}(\text{aq})$  ion is common to both the insoluble electrolyte  $\text{SrSO}_4$  and the soluble salt  $\text{Na}_2\text{SO}_4$ . There is, however, only one source of  $\text{Sr}^{+2}$  ions, namely the dissolution of solid  $\text{SrSO}_4$

Let

$S$  = molar solubility of  $\text{SrSO}_4$  in 0.15 M  $\text{Na}_2\text{SO}_4$

Then,  $[\text{Sr}^{+2}] = S$  and  $[\text{SO}_4^{-2}] = 0.15 + S$

$$K_{\text{sp}}(\text{SrSO}_4) = [\text{Sr}^{+2}] [\text{SO}_4^{-2}] = S(0.15 + S)$$

$$2.5 \times 10^{-7} = 0.15S + S^2$$

Assuming  $S$  to be small compared to 0.15 and that

$$\therefore 2.5 \times 10^{-7} = 0.15S$$

$$\text{or } S = 1.67 \times 10^{-6} \text{ M.}$$

### Question-21

The solubility product of  $\text{Fe}(\text{OH})_3$  is  $1 \times 10^{-36}$ . What is the minimum concentration of  $\text{OH}^-$  ions required to precipitate  $\text{Fe}(\text{OH})_3$  from a 0.001 M solution of  $\text{FeCl}_3$ ?

#### Solution:

$$K_w = [\text{Fe}^{3+}] [\text{OH}^-]^3$$

Precipitation will occur when ionic product,  $[\text{Fe}^{3+}] [\text{OH}^-]^3$  becomes greater than  $K_s$

$$[\text{Fe}^{3+}] = [\text{FeCl}_3] = 0.001 \text{ M}$$

The concentration of  $\text{OH}^-$  ions required to start the precipitation is

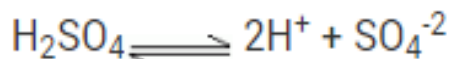
$$[\text{OH}^-]^3 = \frac{K_{\text{sp}}}{[\text{Fe}^{3+}]} = \frac{1 \times 10^{-36}}{0.001} = 1 \times 10^{-11}$$

$$\therefore [\text{OH}^-] = (1 \times 10^{-11})^{1/3} = 1 \times 10^{-11} \text{ molL}^{-1}.$$

### Question-22

Calculate the pH of 0.5 molar solution of sulphuric acid.

#### Solution:



In a 0.5 m  $\text{H}_2\text{SO}_4$  solution

$$[\text{H}^+] = 2 \times 0.5 \text{ mol L}^{-1} = 1 \text{ mol L}^{-1}$$

$$\therefore \text{pH} = -\log [\text{H}^+] = -\log (1) = 0$$

### Question-23

The pH of 0.05M aqueous solution of diethyl amine is 12.0 calculate its  $K_b$ .

#### Solution:

$$\text{pH} = 12.0; \text{pOH} = 14 - 12.00 = 2.00$$

$$[\text{OH}^-] = 10^{-2} \text{M}; K_b = \frac{x^2}{c}$$

$$k_b = \frac{(10^{-2})(10^{-2})}{5 \times 10^{-2}} = 0.2 \times 10^{-2} = 2 \times 10^{-3}.$$

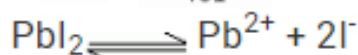
### Question-24

The solubility of lead iodide in water is 0.63 gm/litre. Calculate the solubility product of lead iodide (At mass of Pb = 207; I = 127).

#### Solution:

Solubility of lead iodide in moles / litre

$$= \frac{0.63}{461} = \frac{630 \times 10^{-3}}{461} = 1.36 \times 10^{-3} \text{ moles / litres}$$



$$K_{\text{SP}} = [\text{Pb}^{2+}][\text{I}^-]^2$$

$$= [1.36 \times 10^{-3}][2 \times 1.36 \times 10^{-3}]^2$$

$$= 4 \times (1.36 \times 10^{-3})^3$$

$$= 4 \times (10^{-3})^3 (1.36)^3$$

$$= 4 \times 10^{-9} \times 1.4 \times 1.4 \times 1.4$$

$$= 4 \times 10^{-9} \times 1.4 \times 1.4 \times 1.4$$

$$= 10.9 \times 10^{-9}$$

$$K_{\text{SP}} = 1.09 \times 10^{-8} \text{ mol}^{-3} \text{ L}^{-3}.$$

### Question-25

How many moles of AgBr ( $K_{SP} = 5 \times 10^{-13} \text{ mol}^{-2} \text{ L}^{-2}$ ) will dissolve in 0.01M NaBr solution?



$K_{SP} = [x] [x + 0.01]$  where  $x = \text{solubility of Ag}^+ \text{ or Br}^-$ .

#### Solution:

$5 \times 10^{-13} \approx (x) (10^{-2})$  (compare to 0.01,  $x$  is small)

(i.e.)  $x = \frac{5 \times 10^{-3}}{10^{-2}} = 5 \times 10^{-1} \text{ mol L}^{-1}$ .

### Question-26

Two moles of HI when heated at  $44.4^\circ\text{C}$  until equilibrium is reached, were found to be 22% dissociated. Calculate the equilibrium constant for the reaction.

#### Solution:



$x = 22\%$  (i.e.) that  $\frac{22}{100} = 0.22$



0.22 dissociate to give 0.11 moles of  $\text{H}_2$  and 0.11 mole of  $\text{I}_2$

Equilibrium concentration =  $(1 - 0.22) = 0.78$  moles / litre

Equilibrium concentration of HI = 0.78 moles / litre.

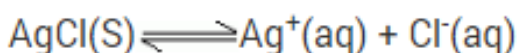
$$K_C = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = \frac{0.11 \times 0.11}{(0.78)^2}$$

$$K_C = 0.0198.$$

### Question-27

How does the solubility of AgCl change in the presence of NaCl solution?  
( $S = 1.3 \times 10^{-5} \text{M}$ )

#### Solution:



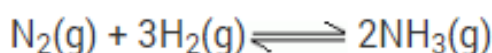
$$K_{\text{SP}} = [\text{Ag}^+][\text{Cl}^-]$$

Addition of NaCl to saturated solution of AgCl increases the concentration of  $\text{Cl}^-$  ions. According to Le Chatelier's  $\text{AgCl}(s)$ , there by reducing the  $\text{Ag}^+$  concentration by precipitating AgCl. Thus solubility of AgCl in NaCl solution is lower than water, due to common ion effect.

### Question-28

A 50.0 L reaction vessel contains 1.00 mol  $\text{N}_2$ , 3.00 mol  $\text{H}_2$ , and 0.055 mol  $\text{NH}_3$ . Will more ammonia, ( $\text{NH}_3$ ), be formed or will it dissociate when the mixture goes to equilibrium at  $400^\circ\text{C}$ ? The equation is  $\text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g)$   $K_c$  is 0.500 at  $400^\circ\text{C}$ .

#### Solution:



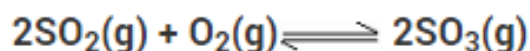
$$\text{Concentrations } \frac{1.00}{50.0} \frac{3.00}{50.0} \frac{0.055}{50.0} = 0.0200 \text{ M}, 0.0600 \text{ M}, 0.0011 \text{ M}$$

$$\therefore Q_c = \frac{[\text{NH}_3]^2}{[\text{H}_2]^3[\text{N}_2]} = \frac{(0.0011)^2}{(0.0600)^3(0.0200)} = 0.2801$$

Because  $Q_c = 23.1$  is greater than  $K_c = 0.500$ , the reaction will go to the left as it approaches equilibrium. Therefore, ammonia will dissociate.

### Question-29

Consider the exothermic formation of sulphur trioxide from sulphur dioxide and oxygen in the gas-phase.



At 900 K,  $K_p$  for this reaction is  $40.5 \text{ atm}^{-1}$  and  $\Delta H = -198 \text{ kJ}$

(a) Write the expression for the equilibrium constant for this reaction.

(b) Will the equilibrium constant for this reaction at room temperature (300 K) be greater than, less than, or equal to the equilibrium constant at 900 K? Explain your answer.

(c) How will the equilibrium be affected if the volume of the vessel containing the three gases is reduced, keeping the temperature constant: what happens?

(d) What is the effect of adding one mole of He(g) to a flask containing  $\text{SO}_2$ ,  $\text{O}_2$  and  $\text{SO}_3$  at equilibrium at constant temperature?

#### Solution:

$$(a) K_p = \frac{P_{\text{SO}_3}^2}{P_{\text{SO}_2}^2 P_{\text{O}_2}}$$

(b) The equilibrium constant at 300 K will be greater than the equilibrium constant at 900 K. This is an exothermic reaction. If we decrease the temperature from 900 to 300 K, the equilibrium will be displaced to the right, releasing heat, more  $\text{SO}_3$  is produced and  $\text{SO}_2$  and  $\text{O}_2$  are used up. The equilibrium constant increases.

(c) Decreasing the volume of the vessel will increase the partial pressure of each gas and therefore increase the total pressure. The system will shift to the side with fewer numbers of moles of gas. Since there are two moles of gas on the right, but three on the left, the system will shift to the right. More  $\text{SO}_3$  will be produced, and some  $\text{SO}_2$  and  $\text{O}_2$  will be used up.

(d) Adding He(g) has no effect at all. The partial pressures of  $\text{SO}_2$  and  $\text{O}_2$  and  $\text{SO}_3$  are unchanged by the addition of helium. The total pressure in the container increases, but as the partial pressure of the gases involved in the equilibrium are unaffected, the equilibrium does not shift.

### Question-30

Suppose we mix  $25.0 \text{ cm}^3$  of  $0.001 \text{ M AgNO}_3(\text{aq})$  with  $75.0 \text{ cm}^3$  of  $0.001 \text{ M Na}_2\text{CO}_3(\text{aq})$ . Does a precipitate of  $\text{Ag}_2\text{CO}_3$  form? The  $K_{\text{sp}}$  of  $\text{Ag}_2\text{CO}_3$  is  $6.2 \times 10^{-12}$  at  $298 \text{ K}$ .

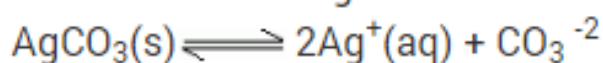
#### Solution:

At the instant of mixing,

$$\text{Conc. Of } [\text{Ag}^+] = \frac{(0.025\text{L})(0.001\text{M})}{(0.100\text{L})} = 2.5 \times 10^{-4} \text{ M}$$

$$\text{Conc. Of } [\text{CO}_3^{2-}] = \frac{(0.075\text{L})(0.001\text{M})}{(0.100\text{L})} = 7.5 \times 10^{-4} \text{ M}$$

We are considering the reaction



$$K_i = [\text{Ag}^+]^2 [\text{CO}_3^{2-}] = (2.5 \times 10^{-4})^2 (7.5 \times 10^{-4})$$

$$K_i = 4.69 \times 10^{-11}$$

Since,  $K_i > K_{\text{sp}}$ , the solution is saturated and  $\text{Ag}_2\text{CO}_3$  will form.

### Question-31

Which of the following are Lewis bases?

$\text{H}_2\text{O}$ ,  $\text{BF}_3$ ,  $\text{H}^+$ ,  $\text{Ag}^+$ ,  $\text{NH}_3$  &  $\text{CO}$ .

#### Solution:

$\text{H}_2\text{O}$ ,  $\text{NH}_3$  and  $\text{CO}$  are Lewis bases as they have lone pairs of electrons to donate to Lewis acids.

### Question-32

Which of the following act both as Bronsted acids and bases?

$\text{HSO}_4^-$ ,  $\text{NH}_3$ ,  $\text{CO}$ ,  $\text{H}_2\text{O}$ ,  $\text{SO}_4^{2-}$ .

#### Solution:

$\text{HSO}_4^-$ ,  $\text{NH}_3$ ,  $\text{H}_2\text{O}$  can donate as well as accept a proton. Hence these can act as Bronsted acids and bases.

$\text{CO}$ ,  $\text{SO}_4^{2-}$  can only accept proton, hence they are Bronsted's bases.

### Question-33

A solution has a hydroxide concentration of  $1.0 \times 10^{-5}$  M at  $25^\circ$  C. Is the solution acidic, natural or basic?

#### Solution:

$$[\text{OH}^-] = 1.0 \times 10^{-5} \text{ M};$$

$$[\text{H}^+][\text{OH}^-] = 1 \times 10^{-14};$$

$$[\text{H}^+] = \frac{1 \times 10^{-14}}{1 \times 10^{-5}} = 1.0 \times 10^{-9} \text{ M}$$

as  $[\text{H}^+]$  is less than  $1 \times 10^{-7}$  M, the solution is basic.

### Question-34

Calculate the pH value of mixture containing 50cc M-HCl and 30cc M-NaOH solution assuming both to be completely ionised.

#### Solution:

$$\text{Total volume after mixing} = 50 + 30 = 80 \text{ cc}$$

$$\text{Molarity of HCl after mixing} = \frac{50}{80} \text{ M}$$

$$\text{Molarity of NaOH after mixing} = \frac{30}{80} \text{ M}$$

$$\text{Net molarity of HCl after mixing} = \frac{50}{80} - \frac{30}{80} = 0.25 \text{ M}$$

$$[\text{H}^+] = 0.25 = 2.5 \times 10^{-1}$$

$$\text{pH} = -\log(2.5 \times 10^{-1}) = 0.6021.$$

### Question-35

The dissociation constant of an acid HA is  $1.6 \times 10^{-5}$ . Calculate  $\text{H}_3\text{O}^+$  ion concentration of its 0.01 M solution.

#### Solution:

$$K_a = \frac{C\alpha^2}{1-\alpha}; K_a = C \alpha^2$$

$$\alpha = \sqrt{\frac{K_a}{C}}$$

$$\alpha = \sqrt{\frac{1.6 \times 10^{-5}}{1 \times 10^{-2}}} = \sqrt{1.6 \times 10^{-3}} = \sqrt{10^{-4} \times 16} = 4 \times 10^{-2}$$

$$[\text{H}^+] = \alpha \times C = (4 \times 10^{-2}) \times 10^{-2} = 4 \times 10^{-4} \text{ mol. L}^{-1}.$$

### Question-36

Calculate the concentration of  $\text{H}_3\text{O}^+$  ion in a mixture of 0.02 m acetic acid and 0.2 m sodium acetate. (Given  $K_a$  for acetic acid is  $1.8 \times 10^{-5}$ )

**Solution:**



$$\therefore K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]}$$

$$[\text{CH}_3\text{COO}^-] = 0.2 \text{ M} + x \approx 0.2 \text{ M}$$

$$[\text{CH}_3\text{COOH}] = 0.02 \text{ M} - x \approx 0.02 \text{ M}$$

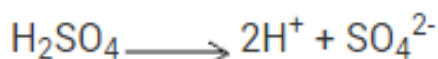
$$1.8 \times 10^{-5} = \frac{[\text{H}_3\text{O}^+][0.2]}{[0.02]}$$

$$\text{(i.e.) } \text{H}_3\text{O}^+ = \frac{1.8 \times 10^{-5} \times 0.02}{0.2} = 1.8 \times 10^{-6} \text{ mol. L}^{-1}.$$

### Question-37

Calculate pH of features  $10^{-2} \text{ M H}_2\text{SO}_4$ .

**Solution:**



One mole of  $\text{H}_2\text{SO}_4$  gives 2 moles of  $\text{H}^+_{(\text{aq})}$  ions.

$$[\text{H}^+] = 2[\text{H}_2\text{SO}_4]$$

$$[\text{H}^+] = 2[10^{-2} \text{ M}]$$

$$\text{pH} = -\log [2 \times 10^{-2}]$$

$$= 2 - \log 2 = 2 - 0.3010$$

$$\text{pH} = 1.699.$$



### Question-38

Calculate pH of 0.001 M Ba (OH)<sub>2</sub>.

**Solution:**



Each mole of Ba(OH)<sub>2</sub> gives 2[OH<sup>-</sup>] moles.

∴ 0.001 m Ba(OH)<sub>2</sub> gives 2 × [0.001 m] OH<sup>-</sup> ions

$$[\text{H}^+] = \frac{10^{-14}}{2 \times 10^{-3}} = \frac{10^{-11}}{2} = 0.5 \times 10^{-11}$$

$$[\text{H}^+] = 5 \times 10^{-12}$$

$$\text{pH} = -\log_{10} [5 \times 10^{-12}]$$

$$\text{pH} = -[\log_{10} 5 + \log 10^{-12}]$$

$$= -(-12) \log 10 - \log 5$$

$$= 12 \times 1 - 0.6990$$

$$\text{pH} = 11.301.$$

### Question-39

How many grams of NaOH must be dissolved in one litre of solution to give it a pH of 12?

**Solution:**

$$\text{pH} = 12; -\log [\text{H}^+] = 12; [\text{H}^+] = \text{antilog of } -12 = 1 \times 10^{-12}.$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}^+]} = \frac{1 \times 10^{-14}}{1 \times 10^{-12}} = 1 \times 10^{-2} = 0.01 \text{ moles/litre}$$

We know, weight per litre  $\equiv$  molarity  $\times$  m.wt.

$$= 0.01 \times 40 = 0.4 \text{ gms/litre.}$$

∴ Weight of NaOH present = 0.4 gms/litre.

### Question-40

A buffer solution is prepared by mixing 6 g. of acetic acid and 13.6 g of sodium acetate  $\text{CH}_3\text{COONa} \cdot 3\text{H}_2\text{O}$  and making the total volume 250 ml. Calculate the

(i) pH of the solution and

(ii) pH change on addition of 1 ml of 1M HCl to it.  $K_a = 1.8 \times 10^{-5}$ .

#### Solution:

$$[\text{CH}_3\text{COOH}] = \frac{\text{wt}}{\text{mwt}} = \frac{6}{60} = 0.1 \text{ moles ; } 0.1/250 = 0.004$$

$$[\text{CH}_3\text{COONa}] = \frac{\text{wt}}{\text{mwt}} = \frac{13.6}{136} = 0.1 \text{ moles} = 0.004$$

$$\text{pH} = \text{pK}_a + \log_{10} \frac{[\text{Salt}]}{[\text{Acid}]}$$

$$\text{pK}_a = -\log_{10} K_a = -\log 1.8 \times 10^{-5} \cong 5 - 0.2553 = 4.74$$

$$\text{pH} = 4.74 + \log [0.004]/[0.004]$$

$$\text{pH} = 4.74 + \log_{10} 1 = 4.74 + 0 = 4.74.$$

### Question-41

Calculate the amount of  $(\text{NH}_4)_2 \text{SO}_4$  in gms which must be added to 500 ml of 0.200 M  $\text{NH}_3$  to yield a solution with pH 9.35 ( $K_b$  for  $\text{NH}_3 = 1.78 \times 10^{-5}$ )

#### Solution:

$$\text{pOH} = \text{pK}_b + \log \frac{[\text{Salt}]}{[\text{Base}]}$$

$$\text{pH} = 9.35$$

$$\text{pOH} = 14 - 9.35 = 4.65; K_b = 1.78 \times 10^{-5}$$

$$\text{pK}_b = 4.7447$$

Number of moles of  $\text{NH}_3$  added: one litre of 0.2 M  $\text{NH}_3$  contain 0.2 mole  
500 ml contains  $0.2 \times \frac{1}{2} = 0.1$  mole

$$\text{Moles of } (\text{NH}_4)_2 \text{SO}_4 = x$$

$$\therefore 4.65 = 4.7447 + \log \left( \frac{x}{0.1} \right)$$

$$\log \left( \frac{x}{0.1} \right) = -0.0947 \text{ or } \log \left( \frac{0.1}{x} \right) = 0.0947$$

$$\frac{0.1}{x} = 1.2436; x = \frac{0.1}{1.2436} = 0.08 \text{ mole}$$

M. wt. of  $(\text{NH}_4)_2 \text{SO}_4 = \text{mole} \times \text{m.wt} = 0.08 \times 132 = 10.56 \text{ grams.}$

### Question-43

What would be the pH of a solution that contains 100 ml of 0.1 N HCl and 9.9 ml of 1 N NaOH solution?

#### Solution:

9.9 ml of 1 N solution of NaOH should be converted into 0.1 N NaOH

$$9.9 \times 1 \text{ N} \equiv x \times 0.1; \text{ (i.e.) } x = \frac{9.9 \times 1}{0.1} = 99 \text{ ml.}$$

9.9 ml of 1.0 N NaOH = 99 ml of 0.1 N, NaOH

Volume of HCl left unneutralised  $\equiv (100 - 99 \text{ ml})$  of 0.1 N  $\equiv 1 \text{ ml}$  of 0.1 N

Total volume of solution after mixing  $\equiv 100 + 9.9 = 109.9 \text{ ml} \equiv 110 \text{ ml}$

(app)

Conc. HCl after mixing  $\equiv \frac{0.1}{110} = 9.09 \times 10^{-4} \text{ m}$

$$[\text{H}^+] = 9.09 \times 10^{-4}$$

$$\text{pH} = -\log [9.09 \times 10^{-4}] = 4 - 0.9546 = 3.0454.$$

### Question-44

Calculate the degree of hydrolysis of 0.1M solution of sodium acetate at 25° C ;  $K_a = 1.8 \times 10^{-5}$ .

#### Solution:

Since sodium acetate is a salt of strong base and weak acid

$$K_b = \frac{k_w}{k_a} = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.5 \times 10^{-10}$$

$$\begin{aligned} \text{Since solution is 0.1, } h &= \sqrt{\frac{k_b}{c}} \\ &= \sqrt{\frac{5.5 \times 10^{-10}}{0.1}} = 7.452 \times 10^{-5}. \end{aligned}$$

### Question-45

$K_a$  for butyric acid is  $2.0 \times 10^{-5}$ . Calculate pH and hydroxyl ion concentration of 0.2 M aqueous solution of sodium butyrate.

#### Solution:

For salt of a weak acid with strong base

$$\begin{aligned} \text{pH} &= \frac{1}{2}\text{pK}_w + \frac{1}{2}\text{pK}_a + \frac{1}{2}\log C \\ &= \frac{1}{2} \times 14 + \frac{1}{2}\log \left( \frac{1}{2 \times 10^{-5}} \right) + \frac{1}{2}\log 0.2 \\ &= 7 + 2.3494 - 0.3494 \\ &= 9 \end{aligned}$$

$$\text{pOH} = 14 - 9 = 5$$

$$[\text{OH}^-] = 10^{-5} \text{ M.}$$

### Question-46

The pH of blood stream is maintained by a proper balance of  $\text{H}_2\text{CO}_3$  and  $\text{NaHCO}_3$  concentrations. What volume of 5 M  $\text{NaHCO}_3$  solution should be mixed with a 10 ml sample of blood which is 2 M in  $\text{H}_2\text{CO}_3$  in order to maintain a pH of 7.4?  $K_a$  for  $\text{H}_2\text{CO}_3$  in blood is  $7.8 \times 10^{-7}$ .

#### Solution:

Let the amount of  $\text{NaHCO}_3$  mixed = x ml

$$\text{Number of moles of NaHCO}_3 \text{ in x ml of 5 m NaHCO}_3 = \frac{5 \times x}{1000} = 0.005x$$

$$\text{Number of moles of H}_2\text{CO}_3 \text{ in 10 ml of 2 m H}_2\text{CO}_3 = \frac{2 \times 10}{1000} = 0.02$$

$$\text{pH} = -\log K_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$$

$$7.4 = -\log 7.8 \times 10^{-7} + \log \left( \frac{0.005x}{0.02} \right)$$

$$x = 78.36 \text{ ml.}$$

### Question-47

To a buffer solution containing 0.02 mole of propionic acid and 0.015 mole of sodium propionate per litre, 0.01 mole / litre of HCl is added. What is the pH of the solution?  $K_a$  propionic acid =  $1.34 \times 10^{-5}$ .

#### Solution:

$$[\text{propionic acid}] = (0.02 + 0.01)\text{M}$$

$$[\text{propionate}] = (0.015 + 0.010)\text{M}$$

$$\text{pH} = \text{p}K_a + \log \frac{[\text{salt}]}{[\text{Acid}]}$$

$$\text{pH} = -\log (1.34 \times 10^{-5}) + \log \frac{0.025}{0.025} = 4.9528.$$

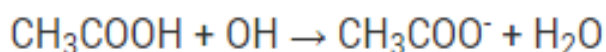
### Question-48

A buffer solution of acetic acid and sodium acetate of 1 M concentration was prepared. Calculate pH of the system when 0.01 mole of NaOH is added per litre of the above solution.  $K_a = 1.8 \times 10^{-5}$ .

#### Solution:

NaOH ionises fully: Hence  $\text{OH}^-$  ions added will react with  $\text{CH}_3\text{COOH}$  of the buffer system.

Molar conc. before adding NaOH



Molar conc. after adding 0.01 mole of  $(1 - 0.01) = 0.99$

$$\text{PH} = \text{p}K_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$$

$$= -\log 1.8 \times 10^{-5} + \log \frac{[1.01]}{[0.99]}$$

$$= 4.74 + 0.01 = 4.75.$$

### Question-49

The solubility product of AgCl in water is  $1.5 \times 10^{-10}$ . Calculate its solubility in 0.01 M NaCl aqueous.

#### Solution:

Let the solubility of AgCl in 0.01 M NaCl solution =  $x$  g mole  $L^{-1}$

$[Ag^+]$  in NaCl solution =  $x$  mol  $L^{-1}$

$[Cl^-]$  in 0.01 M NaCl solution  $\equiv x + 0.01$  M  $L^{-1}$

$\therefore [Ag^+][Cl^-] = K_{sp}$

$(x)(x + 0.01) = 1.5 \times 10^{-10}$

$0.01x + x^2 = 1.5 \times 10^{-10}$

Since value of  $x$  is small  $x^2$  is neglected.

$0.01x = 1.5 \times 10^{-10}$

$x = 1.5 \times 10^{-8}$  mol  $L^{-1}$ .

## CBSE Class 11 Chemistry

### Important Questions

#### Chapter 7

#### Equilibrium

#### 1 Marks Questions

##### 1. Define dynamic equilibrium.

**Ans.** When the reactants in a closed vessel at a particular temperature react to give products, the concentrations of the reactants keep on decreasing, while those of products keep on increasing for sometime after which there is no change in the concentrations of either the reactants or products. This stage of the system is the dynamic equilibrium.

##### 2. What is physical equilibrium? Give an example.

**Ans.** Physical equilibrium is an equilibrium between two different physical states of same substance e.g.  $\text{H}_2\text{O}(s) \rightleftharpoons \text{H}_2\text{O}(l)$

##### 3. What is meant by the statement 'Equilibrium is dynamic in nature'?

**Ans.** At equilibrium, reaction does not stop rather it still continues, the equilibrium is dynamic in nature. It appears to stop because rate of forward reaction is equal to the rate of backward reaction.

##### 4. How does dilution with water affect the pH of a buffer solution?

**Ans.** Dilution with water has no effect on the pH of any buffer. This is because pH of a buffer depends on the ratio of the salt, acid or salt base and dilution does not affect this ratio.

##### 5. On what factor does the boiling point of the liquid depend?

**Ans.** Boiling point depends on the atmospheric pressure.

**6..State Henry's law.**

**Ans .**The mass of a gas dissolved in a given mass of a solvent at any temperature is proportional to the gas above the solvent.

**7.What happens to the boiling point of water at high altitude?**

**Ans .**Boiling point of water depends on the altitude of the place. At high altitude atmoosp here pressure thetore is less boiling point decreases.

**8.On which factor does the concentration of solute in a saturated solution depends?**

**Ans .** The concentration of solute in a saturated solution depends upon the temperature.

Sugar (soln.)  $\rightleftharpoons$  sugar (solid).

**9.What conclusion is drawn from the following –**

**Solid  $\rightleftharpoons$  Liquid**

**H<sub>2</sub>O(s)  $\rightleftharpoons$  H<sub>2</sub>O (l)**

**Ans .** Melting point is fixed at constant pressure.

**10.State the law of chemical equilibrium.**

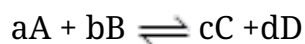
**Ans.** At a given temperature, the product of concentrations of the reaction products raised to the respective stoichiometric coefficient in the balanced chemical equation divided by the product of concentrations of the reactants raised to their individual stoichiometric coefficients has a constant value. This is known as the equilibrium law or law of chemical equilibrium.

**11. Write the equilibrium constant for the following equation :**

**aA +bB  $\rightleftharpoons$  cC + dD**



**Ans .The** equilibrium constant for a general reaction



is expressed as

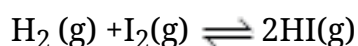
$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Where [A], [B], [C] and [D] are the equilibrium concentrations of the reactants and products.

**12. Write the chemical equation for the following chemical constant.**

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

**Ans .The** chemical equation is given by



**13. Write the expression for equilibrium constant K<sub>p</sub> for the reaction**



$$\text{Ans . } K_p = \frac{(P_{H_2})^4}{(P_{H_2O})^4} = \frac{P_{H_2}}{P_{H_2O}}.$$

**14..The equilibrium constant for the reaction  $H_2O + CO \rightleftharpoons H_2 + CO_2$**

**Is 0.44 at 1260k. What will be the value of the equilibrium constant for the reaction :**



**Ans .The** reaction is reversed and also doubled,

$$\therefore K_c = \left( \frac{1}{0.44} \right)^2 = \underline{\underline{5.16}}$$

**15. Define reaction quotient.**

**Ans .** The reaction quotient, Q is same as equilibrium constant  $K_c$ , except that the concentrations in

**Qc are not necessarily equilibrium values.**

**16. If  $Q_c > K_c$ , what would be the type of reaction?**

**Ans .** If  $Q_c > K_c$ , the reaction will proceed in the direction of the reactants (reverse reactions)

**17. What inference you get when  $Q_c = K_c$ ?**

**Ans .** If  $Q_c = K_c$ , the reaction mixture is already at equilibrium.

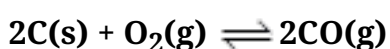
**18. State Le chatelier's principle.**

**Ans .** It states that a change in any of the factors that determine the equilibrium conditions of a system will cause the system to change in such a manner so as to reduce or to counteract the effect **of the change**.

**19. Can a catalyst change the position of equilibrium in a reaction?**

**Ans .** No, a catalyst cannot change the position of equilibrium in a chemical reaction. A catalyst, **however, affects the rate of reaction**.

**20. What is the effect of reducing the volume on the system described below?**



**Ans.** The forward reaction is accompanied by increase in volume. Hence according to Chatelier's principle, reducing the volume will shift the equilibrium in the forward direction.

**21..What happens when temperature increases for a reaction?**

**Ans .**The equilibrium constant for an exothermic reaction ( $\Delta H - ve$ ) decreases as the temperature increases.

**22.Can a catalyst change the position of equilibrium in a reaction?**

**Ans .**No, a catalyst cannot change the position of equilibrium in a chemical reaction. A catalyst affects the rate of reaction.

**23.If  $Q_c < K_c$ , when we continuously remove the product, what would be the direction of the reaction?**

**Ans.**Continuous removal of a product maintains  $Q_c$  at a value less than  $K_c$  and reaction continues to move in the forward direction.

**24.Define strong and weak electrolyte.**

**Ans .** Those electrolytes which dissociate almost completely into ions in aqueous solutions are

Known as strong electrolytes while those which show poor dissociation into ions in aqueous solutions are called weak electrolytes.

**25.Write the conjugate acids for the following Bronsted bases :  $NH_2$ ,  $NH_3$  and  $HCOO^-$ .**

**Ans.**

Species	Conjugate acids
$NH_2$	$NH_3^+$
$NH_3$	$NH_4^+$
$HCOO^-$	$HCOOH$ .

**26. Which conjugate base is stronger  $\text{CN}^-$  or  $\text{F}^-$  ?**

**Ans .**  $\text{F}^- < \text{CN}^-$  basic character.

**27. What is the difference between a conjugate acid and a conjugate base?**

**Ans.** A conjugate acid and base differ by a proton.

**28. Select Lewis acid and Lewis base from the following :**

$\text{Cu}^{2+}$  ,  $\text{H}_2\text{O}$ ,  $\text{BF}_3$   $\text{OH}^-$

**Ans .** Lewis acids :  $\text{Cu}^{2+}$ ,  $\text{BF}_3$

Lewis bases :  $\text{H}_2\text{O}$ ,  $\text{OH}^-$

**29. The dimethyl ammonium ion,  $(\text{CH}_3)_2 \text{NH}_2^+$ , is a weak acid and ionizes to a slight degree in water what is its conjugate base?**

**Ans .**  $(\text{CH}_3)_2 \text{NH}$

**30. If  $\text{p}^{\text{H}}$  of a solution is 7, calculate its  $\text{p}^{\text{OH}}$  value.**

**Ans .**  $\text{p}^{\text{H}} + \text{p}^{\text{OH}} = 14$

$\therefore \text{p}^{\text{OH}} = 14 - \text{p}^{\text{H}}$

$= 14 - 7$

$= 7.$

**31. What happens to the  $p^H$  if a few drops of acid are added to  $CH_3COONH_4$  solution?**

**Ans .** pH will almost remain constant.

**32. What is the concentration of  $H_3O^+$  and  $OH^-$  ions in water at 298K?**

**Ans .**  $[H_3O^+] = [OH^-] = 1 \times 10^{-7} \text{ mol}^{-1}$

**33. The  $p^{ka}$  of acetic acid and  $p^{kb}$  of ammonium hydroxide are 4.76 and 4.75 respectively. Calculate the  $p^H$  of ammonium acetate solution.**

**Ans .**  $p^H = 7 + \frac{1}{2} [p^{ka} - p^{kb}]$

$$= 7 + \frac{4.76 - 4.75}{2}$$

$$= 7 + \frac{1}{2} [0.01]$$

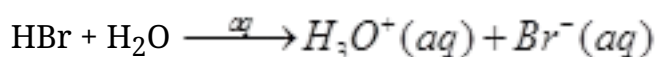
$$= 7 + 0.005$$

$$= 7.005$$

**34. Calculate the pH of the solution**

**0.002 M HBr**

**Ans .**  $p^H$  value of 0.002M HBr.



0.002M

$$pH = -(\log H_3O^+) = -\log (2 \times 10^{-3})$$

$$= (3 - \log 2) = 3 - 0.3010 = \underline{\underline{2.7}}$$

**35. Define Buffer solution.**

**Ans .** The solutions which resist change in pH on dilution or with the addition of small amounts of acid or alkali are called Buffer solutions.

**36. When is a solution called unsaturated?**

**Ans .** When the ionic product is less than the solubility product the solution is unsaturated.

**37. Give an example of acidic buffer?**

**Ans.**  $\text{CH}_3\text{COOH} + \text{CH}_3\text{COONa}$ .

**38. Calculate the solubility of AgCl (s) in pure water.**

**Ans .** Let the solubility of AgCl in water be  $S \text{ mol L}^{-1}$



$$[\text{Ag}^+] = S; [\text{Cl}^-] = S$$

$$K_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-]$$

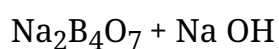
$$2.8 \times 10^{-10} = s \times s$$

$$\text{Or } S = \sqrt{2.8 \times 10^{-10}}$$

$$= \underline{\underline{1.673 \times 10^{-5} \text{ mol}^{-1}}}$$

**39. Name a basic buffer having pH around 10.**

**Ans .** Basic buffer



Borax sodium hydroxide.

**CBSE Class 12 Chemistry**  
**Important Questions**  
**Chapter 7**  
**Equilibrium**

**2 Marks Questions**

**1. Mention the general characteristics of equilibria involving physical processes.**

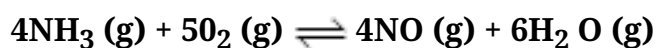
**Ans .(a)** For solid  $\rightleftharpoons$  liquid equilibrium, there is only one temperature at 1 atm at which two phases can co-exist. If there is no exchange of heat with the surroundings, the mass of the two phases remain constant.

**(b)** For liquid  $\rightleftharpoons$  vapors equilibrium, the vapors pressure is constant at a given temperature.

**(c)** For dissolution of solids in liquids, the solubility is constant at a given temperature.

**(d)** For dissolution of gases in liquids, the concentration of a gas in liquid is proportional to pressure of the gas over the liquid.

**2. Write the expression for the equilibrium constant for the reaction :**

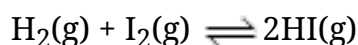


**Ans .**The equilibrium constant is given by

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

**3. When the total number of moles of product and reactants are equal, K has no unit. Give reason.**

**Ans .**When the total number of moles of products is equal to the total number of moles of reactants the equilibrium constant k has no unit for eg.



$$K = \frac{[\text{HI}(\text{g})]^2}{[\text{H}_2(\text{g})][\text{I}_2(\text{g})]}$$

$$\text{Units of } K = \frac{\text{mol/L} \times \text{mol/L}}{\text{mol/L} \times \text{mol/L}} = \text{No units} .$$

4. What is the unit of equilibrium for the reaction  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ .

Ans .

$$K = \frac{[\text{NH}_3(\text{g})]^2}{[\text{N}_2(\text{g})][\text{H}_2(\text{g})]^3}$$

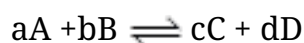
$$\text{units of } K = \frac{(\text{mol/L})^2}{(\text{mol/L})(\text{mol/L})^3} = \frac{1}{(\text{mol/L})^2}$$

$$(\text{mol/L})^{-2}$$

$$= \underline{\underline{L^2 \text{mol}^{-2}}}$$

5. Give the relation  $K_p = K_c (RT)^{\Delta n}$ .

Ans . Let us consider a reaction



$$K_c = \frac{[\text{C}]^c [\text{D}]^d}{[\text{A}]^a [\text{B}]^b} \dots (i) < K_p = \frac{p_C^c p_D^d}{p_A^a p_B^b}$$

Assuming the gaseous components to behave ideally,

$$p_i V_i = n_i RT \dots$$

$$\text{Or, } p_i = \frac{n_i}{V_i} RT = C_i RT = [c] RT \dots (iv).$$



Where [i] is the molar concentration of the species i

Then,

$$K_p = \frac{p_C^c p_D^d}{p_A^a p_B^b} = \frac{([C]RT)^c \times ([D]RT)^d}{([A]RT)^a \times ([B]RT)^b}$$
$$= \frac{[C]^c [D]^d}{[A]^a [B]^b} \times (RT)^{(\overline{c+d} - \overline{a+b})} \dots (v)$$

$$\Delta n = (\overline{c+d} - \overline{a+b})$$

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

#### 6. The value of $K_c$ for the reaction

$2A \rightleftharpoons B + C$  is  $2 \times 10^{-3}$ . At a given time, the composition of the reaction mixture is  $[A] = [B] = [C] = 3 \times 10^{-4}$  M. In which direction the reaction will proceed?

Ans . For the reaction the reaction  $Q_c$  is given by

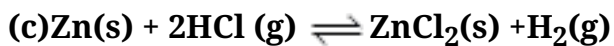
$$Q_c = \frac{[B][C]}{[A]^2}$$

As  $[A] = [B] = [C] = 3 \times 10^{-4}$  M

$$Q_c = \frac{(3 \times 10^{-4})(3 \times 10^{-4})}{(3 \times 10^{-4})^2} = 1$$

As  $Q_c > K_c$  so the reaction will proceed in the reverse direction.

7. Write the equilibrium constant expression for each of the following reactions. In each case, indicate which of the reaction is homogeneous or heterogeneous.



$$\text{Ans. (a) } K_c = \frac{[CO_2]^2}{[CO]^2 [O_2]}$$

$$(b) K_c = \frac{[NO_2][NO_3]}{[N_2O_5]}$$

$$(c) K_c = \frac{[H_2]}{[HCl]^2} \quad (d) K_c = [O_2]$$

Homogeneous : a, b

Heterogeneous : c, d

**8. The dissociation of HI is independent of pressure, while dissociation of  $PCl_5$  depends upon the pressure applied. Why?**

Ans. For  $2HI \rightleftharpoons H_2 + I_2$

$$K_c = \frac{x^2}{4(1-x)^2}$$

Where x is degree of dissociation

For  $PCl_5 \rightleftharpoons PCl_3 + Cl_2$

$$K_c = \frac{x^2}{v(1-x)}$$

Where x is degree of dissociation

Since  $K_c$  for HI does not have volume terms and thus dissociation of HI is independent of pressure. On the other hand  $K_c$  for  $\text{PCl}_5$  has volume in denominator and thus an increase in pressure reduces volume. And to have  $K_c$  constant, x decrease.

**9. On what factors does the value of the equilibrium constant of a reaction depend?**

**Ans.** The equilibrium constant of a reaction depends upon

**(i)** Temperature

**(ii)** Pressure, &

**(iii)** Stoichiometry of the reaction

**10. Why the addition of inert gas does not change the equilibrium?**

**Ans.** It is because the addition of an inert gas at constant volume does not change the partial pressures or the molar concentrations of the substance involved in the reaction.

**11. The equilibrium constant of a reaction increases with rise in temperature. Is the reaction exo – or endothermic?**

**Ans.** The equilibrium constant increases with a rise in temperature. Therefore, the reaction is endothermic.

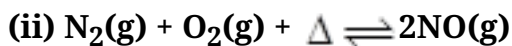
**12. Using Le – chatelier principle, predict the effect of**

**(a)** decreasing the temperature

**(b)** increasing the temperature

**in each of the following equilibrium systems:**

**(i)**  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) + \Delta$



**Ans .(i)** For an exothermic reaction increase in temperature shifts the equilibrium to the left and decrease in temperature shifts it to the left.

**(ii)** For an endothermic reaction increase in temperature shifts the equilibrium to the right and decrease in temperature shifts it to the right.

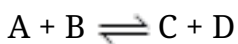
**13.(i) In the reaction equilibrium**



**What will happen to the concentrations of A, B and D if concentration of C is increased.**

**(ii) what will happen if concentration of A is increased?**

**Ans .(i)** For an equilibrium reaction



$$K_c = \frac{[\text{C}][\text{D}]}{[\text{A}][\text{B}]}$$

If the concentration of a product is increased, the concentration of other components changes in such a way that the conc of C decreases and vice – versa.

If the conc of C is increased the conc of D will decrease and those of A and B will increase simultaneously so that the numerical value of  $K_c$  is the same and vice – versa. The equilibrium shifts to the left.

**(ii)** If the conc of A is increase, conc of B will decrease and those of C and D will increase simultaneously so that the numerical value of  $K_c$  is the same and vice – versa. The equilibrium shifts to the right

**14. Give two examples of actions which can act as Lewis acids.**

**Ans .**  $\text{Ag}^+, \text{H}^+.$

**15. Justify the statement that water behaves like an acid and also like a base on the basis of protonic concept**

**Ans .** Water ionizes as  $\text{H}_2\text{O} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$

With strong acid water behaves as a base and accept the proton given by the acid e.g.  $\text{HCl} + \text{H}_2\text{O} \rightleftharpoons \text{Cl}^- + \text{H}_3\text{O}^+$

While with strong base, water behaves as an acid by liberating a proton e.g. :

$\text{H}_2\text{O} + \text{NH}_3 \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$ .

**16. The degree of dissociation of  $\text{N}_2\text{O}_4$ ,**

$\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$ , at temperature T and total pressure is  $\alpha$ . Find the expression for the equilibrium constant of this reaction at this temperature and pressure?

**Ans .**  $\text{N}_2\text{O}_4 \rightleftharpoons 2\text{NO}_2$

At eq 1- $\alpha$ 2 $\alpha$

If p is the total pressure then

$$P_{\text{N}_2\text{O}_4} = P_{\text{N}_2\text{O}_4} = \frac{(1-\alpha)}{(1+\alpha)} P$$

$$P_{\text{NO}_2} = P_{\text{NO}_2} = \frac{2\alpha}{1+\alpha} P$$

$$\text{Then } K_p = \frac{p_{\text{NO}_2}^2}{P_{\text{N}_2\text{O}_4}} = \frac{[2\alpha p / (1+\alpha)]^2}{[(1-\alpha) p / (1+\alpha)]} = \frac{4\alpha^2 p}{(1-\alpha^2)}$$

A solution give the following colors with different indicators. Methyl orange – yellow, methyl red – yellow, and bromothymol blue Orange . what is the  $p^{\text{H}}$  of the solution?

**17. Show that, in aqueous solutions**

$$p^H + p^{OH} = p^{kw}$$

**What is the value of  $p^H + p^{OH}$  at 25°C?**

**Ans .(i)** The colors in methyl orange indicates that  $p^H > 4.5$

**(ii)** Colors in methyl red indicates that  $p^H > 6.0$  and

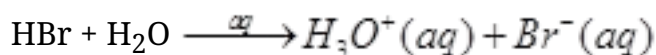
**(iii)** colors in bromothymol blue indicates that  $p^H < 6.3$ .

Therefore, the pH of the solution is between 6.0 to 6.3.

**18. Calculate the pH of the solution**

**0.002 M HBr**

**Ans .**  $p^H$  value of 0.002M HBr.



0.002M

$$pH = -(\log H_3O^+) = -\log (2 \times 10^{-3})$$

$$= (3 - \log 2) = 3 - 0.3010 = \underline{\underline{2.7}}$$

**19. The concentration of  $H^+$  in a soft drink is  $3.8 \times 10^{-3}$  M. what is its  $p^H$ ?**

**Ans .**  $[H^+] = 3.8 \times 10^{-3}$  M

$$pH = -\log [H^+] = \log (3.8 \times 10^{-3})$$

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

$$= -0.5798 + 3 = \underline{\underline{2.42}}$$

**20..Define solubility product.**

**Ans .**The solubility product of a salt at a given temperature is equal to the product of the concentration of its ions in the saturated solution, with each concentration term raised to the power equal to the number of ions produced on dissociation of one mole of the substance.

**21.Ksp for Hg SO<sub>4</sub> is 6.4 x 10<sup>-5</sup>. What is the solubility of the salt?**

**Ans .** $S = (k_{sp})^{1/2}$

$= (6.4 \times 10^{-5})^{1/2}$

$= (64 \times 10^{-6})^{1/2}$

$= \underline{\underline{8 \times 10^{-3}}}$

**22.Calculate the pH of a buffer solution containing 0.1 mole of acetic acid and 0.15 mole of sodium acetate. Ionisation constant for acetic acid is 1.75 × 10<sup>-5</sup>.**

**Ans .** $p^H = p^{ka} + \log \frac{\textit{salt}}{\textit{acid}}$

$pH = -\log 1.75 \times 10^{-5} + \log \frac{0.15}{0.10}$

or,  $p^H = -\log 1.75 \times 10^{-5} + \log 1.5 = \underline{\underline{4.9}}$

**CBSE Class 12 Chemistry**  
**Important Questions**  
**Chapter 7**  
**Equilibrium**

**3 Marks Questions**

**1. Name the three groups into which chemical equilibrium can be classified.**

**Ans.** Chemical equilibrium can be classified into three groups

**(i)** The reaction that proceeds nearly to completion and only negligible concentrations of the reactants are left.

**(ii)** The reactions in which only small amounts of products are formed and most of the reactants remain unchanged at equilibrium stage.

The reactions in which the concentrations of the reactants and products are comparable, when the system is in equilibrium.

**2. Give the generalizations concerning the composition of equilibrium mixtures.**

**Ans. (i)** If  $K_c > 10^3$ , products predominate over reactants i.e; if  $K_c$  is very large, the reaction proceeds nearly to completion.

**(ii)** If  $K_c < 10^{-3}$ , reactants predominate over products i.e; if  $K_c$  is very small, the reaction proceeds rarely.

**(iii)** If  $K_c$  is in the range of  $10^{-3}$  to  $10^3$ , appreciable concentration of both reactants and products are present.

**1. Predict if the solutions of the following salts are neutral, acidic or basic:**

**NaCl, KBr, NaCN, NaOH, H<sub>2</sub>SO<sub>4</sub>, NaNO<sub>2</sub>, NH<sub>4</sub>NO<sub>3</sub>, KF**

**Ans.**



NaCl - Neutral

KBr - Neutral

NaCN - Basic

NaOH - Basic

H<sub>2</sub>SO<sub>4</sub> - Acidic

NaNO<sub>2</sub> - Basic

NH<sub>4</sub>NO<sub>3</sub> - Acidic

KF - Basic

**CBSE Class 12 Chemistry**  
**Important Questions**  
**Chapter 8**  
**Equilibrium**

**5 Marks Questions**

**1. Find the oxidation state of sulphur in the following compounds :**

**$\text{H}_2\text{S}$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{S}_2\text{O}_4^{2-}$ ,  $\text{S}_2\text{O}_8^{2-}$  and  $\text{HSO}_3^-$ .**

**Ans.** In  $\text{H}_2\text{S}$

$$2 + x = 0$$

$$x = -2$$

In  $\text{HSO}_3^-$

$$+1 + x - 6 = -1$$

$$\text{or } x - 5 = -1$$

$$\text{or } x = +4$$

In  $\text{H}_2\text{SO}_4$

$$+2 + x - 8 = 0$$

$$\text{Or } x = +6$$

In  $\text{S}_2\text{O}_5^{2-}$

There is peroxide linkage, thus

oxidation state of S is 6

In  $\text{S}_2\text{O}_4^{2-}$

$$2x - 8 = -2$$

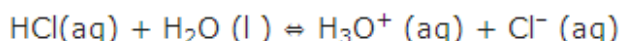
$$2x = 6$$

$$x = +3$$

## Equilibrium

### Short Answer Type Questions

1. The ionisation of hydrochloric in water is given below:



Label two conjugate acid-base pairs in this ionisation.

2. The aqueous solution of sugar does not conduct electricity. However, when sodium chloride is added to water, it conducts electricity. How will you explain this statement on the basis of ionisation and how is it affected by concentration of sodium chloride?
3.  $\text{BF}_3$  does not have proton but still acts as an acid and reacts with  $\text{NH}_3$ . Why is it so? What type of bond is formed between the two?
4. Ionisation constant of a weak base MOH, is given by the expression

$$K_b = \frac{[\text{M}^+][\text{OH}^-]}{[\text{MOH}]}$$

Values of ionisation constant of some weak bases at a particular temperature are given below:

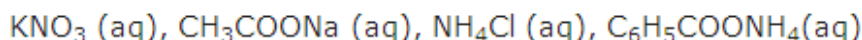
Base	Dimethylamine	Urea	Pyridine	Ammonia
$K_b$	$5.4 \times 10^{-4}$	$1.3 \times 10^{-14}$	$1.77 \times 10^{-9}$	$1.77 \times 10^{-5}$

Arrange the bases in decreasing order of the extent of their ionisation at equilibrium. Which of the above base is the strongest?

5. Conjugate acid of a weak base is always stronger. What will be the decreasing order of basic strength of the following conjugate bases?

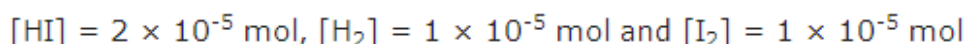


6. Arrange the following in increasing order of pH.



7. The value of  $K_c$  for the reaction  $2\text{HI} \text{ (g)} \rightleftharpoons \text{H}_2 \text{ (g)} + \text{I}_2 \text{ (g)}$  is  $1 \times 10^{-4}$

At a given time, the composition of reaction mixture is



In which direction will the reaction proceed?

8. On the basis of the equation  $\text{pH} = -\log [\text{H}^+]$ , the pH of  $10^{-8} \text{ mol dm}^{-3}$  solution of HCl should be 8. However, it is observed to be less than 7.0. Explain the reason.
9. pH of a solution of a strong acid is 5.0. What will be the pH of the solution obtained after diluting the given solution a 100 times?
10. A sparingly soluble salt gets precipitated only when the product of concentration of its ions in the solution ( $Q_{\text{sp}}$ ) becomes greater than its solubility product. If the solubility of  $\text{BaSO}_4$  in water is  $8 \times 10^{-4} \text{ mol dm}^{-3}$ . Calculate its solubility in  $0.01 \text{ mol dm}^{-3}$  of  $\text{H}_2\text{SO}_4$ .
11. pH of  $0.08 \text{ mol dm}^{-3}$  HOCl solution is 2.85. Calculate its ionisation constant.
12. Calculate the pH of a solution formed by mixing equal volumes of two solutions A and B of a strong acid having pH = 6 and pH = 4 respectively.
13. The solubility product of  $\text{Al}(\text{OH})_3$  is  $2.7 \times 10^{-11}$ . Calculate its solubility in  $\text{g L}^{-1}$  and also find out pH of this solution. (Atomic mass of Al = 27 u).
14. Calculate the volume of water required to dissolve 0.1 g lead (II) chloride to get a saturated solution. ( $K_{\text{sp}}$  of  $\text{PbCl}_2 = 3.2 \times 10^{-8}$ , atomic mass of Pb = 207 u).
15. A reaction between ammonia and boron trifluoride is given below:  
:  $\text{NH}_3 + \text{BF}_3 \rightarrow \text{H}_3\text{N} : \text{BF}_3$   
Identify the acid and base in this reaction. Which theory explains it? What is the hybridisation of B and N in the reactants?
16. Following data is given for the reaction:  $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$   
 $\Delta_f H^\ominus [\text{CaO}(\text{s})] = - 635.1 \text{ kJ mol}^{-1}$   
 $\Delta_f H^\ominus [\text{CO}_2(\text{g})] = - 393.5 \text{ kJ mol}^{-1}$   
 $\Delta_f H^\ominus [\text{CaCO}_3(\text{s})] = - 1206.9 \text{ kJ mol}^{-1}$   
Predict the effect of temperature on the equilibrium constant of the above reaction.

## Long Answer Type Questions

1. How can you predict the following stages of a reaction by comparing the value of  $K_c$  and  $Q_c$  ?
  - (i) Net reaction proceeds in the forward direction.
  - (ii) Net reaction proceeds in the backward direction.
  - (iii) No net reaction occurs.
2. On the basis of Le Chatelier principle explain how temperature and pressure can be adjusted to increase the yield of ammonia in the following reaction.  
$$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) \quad \Delta H = -92.38 \text{ kJ mol}^{-1}$$

What will be the effect of addition of argon to the above reaction mixture at constant volume?
3. A sparingly soluble salt having general formula  $\text{A}^{p+}_x \text{B}^{q-}_y$  and molar solubility  $S$  is in equilibrium with its saturated solution. Derive a relationship between the solubility and solubility product for such salt.
4. Write a relation between  $\Delta G$  and  $Q$  and define the meaning of each term and answer the following :
  - (a) Why a reaction proceeds forward when  $Q < K$  and no net reaction occurs when  $Q = K$ .
  - (b) Explain the effect of increase in pressure in terms of reaction quotient  $Q$ . for the reaction :  $\text{CO}(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons \text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g})$