7. EQUILIBRIUM

Ramanujan School of Mathematics and Physics

Theory + NCERT MCQs + Topic wise Question MCQs+NEET PYQs

EQUILIBRIUM

Ionic Product of Water

$2H_2O \rightleftharpoons H_3O^+ + OH^-$

$K_w = [H_3O^+][OH^-] = 1 \times 10^{-14} M^2$

$\therefore [OH^-] = [H^+] = 10^{-7} M$ @ 298 K

$pK_w = pK_a + pK_b = 14$

Solubility Product (KSP)

$aA \rightleftharpoons cC + dD$

$K_{sp} = [C]^c[D]^d$

Ostwald's Dilution Law

Applicable for weak K electrolytes

$\therefore K_c = C\alpha^2$ or $\alpha = \sqrt{\left(\frac{K_c}{C}\right)}$

So, $\alpha = \frac{1}{\sqrt{C}}$ or $\alpha \propto \sqrt{v}$

where V is the volume of solution at infinite dilution

Law of Chemical Equilibrium Equilibrium Law

$aA + bB \rightleftharpoons cC + dD$

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

Here K_c is equilibrium constant

Relation between equilibrium constant K_p & K_c

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

Gibb's Energy

$\boxed{\Delta G = RT \ln K}$

ΔG = -ve. Spontaneous reaction

Reaction proceeds forward.

ΔG = +ve. Non spontaneous reation

Reaction proceeds backward

ΔG = zero. equilibrium achieved

PH Concept

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

$pH = -\log[H_3O^+]$

for weak acid $\rightarrow pH = \frac{1}{2}(C_pK_a - \log c)$

Hydrolysis of Salts

Salts of strong base and strong acid Neutral solution and does not undergo hydrolysis. eg. NaCl, KCl

Salt of weak base and Strong Acid

$K_h = \frac{K_w}{K_b}; pH = \frac{1}{2}(pK_a - pK_b - \log c)$

Salt of weak Acid and weak base

$K_h = \frac{K_w}{K_a \times K_b}; pH = \frac{1}{2}(pK_w - pK_a - pK_b)$

Acids and Base

Acids: Liberates H_2 on reacting with metals Turns blue litumus into red

Base: Taste bitter and feel soapy Turns red litmus into blue

Acidic $\Rightarrow [H_3O^+] > [OH^-]$

Basic $\Rightarrow [H_3O^+] < [OH^-]$

Neutral $\Rightarrow [H_3O^+] = [OH^-]$

Direction of reaction

$Q_c < K_c$

Reaction goes from left to right

$Q_c < K_c$

Reaction goes from right to left

$Q_c = K_c$

No Net reaction occurs

Factor's of reaction

Le Chatlier's Principle

Effect of concentration change concentration $\rightarrow$, equilibrium shift forward.

Effect of pressure change equilibrium will shift in the direction having smaller number of moles.

Effect of temperature change

For exothermic $\rightarrow$ low temperature favors formation of reactants.

For Endothermic $\rightarrow$ High temperature favors formation of products.

Effect of inert gas $\rightarrow$ No change

Effect of catalyst $\rightarrow$ No change

Definition

Chemical reaction reach a state of dynamic equilibrium in which the rate of forward reaction and reaction are same and there is no net charge in composition

Homogeneous

Reactant and Product are in same phase

Hetrogeneous

Reactant and Product are in different phase

Equilibrium set up in a physical process like evaporation of water etc.

$(S) \rightleftharpoons (L)$

$(L) \rightleftharpoons (g)$

$(g) \rightleftharpoons (S)$

Chemical

Equilibrium attained in a chemical reaction

$3H_2 + N_2 \rightleftharpoons 2NH_3$

Possible only in a closed system.

Both reaction occur at same rate

All measurable property remains constant

EQUILIBRIUM

Introduction

Equilibrium is the most important characteristic property of reversible reactions. These reactions for which the forward reaction occurs to a much greater extent are considered to be unidirectional in nature and whenever the rate of forward reaction is equal to rate of backward reaction, equilibrium is attained, not to forget that equilibrium exists only in closed system.

It is the state of system at which temperature, pressure, volume and composition have fixed value and does not vary with time. Chemical Reactions can be divided into two categories:

Irreversible Reactions: The reactions which proceed to completion and the products fail to recombine to give back reactants. For example:

$$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$$

Reversible Reactions: The reactions which never go to completion and products recombine to give back reactants. For example:

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

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Physical Equilibrium

We know that solid, liquid and gas are the three states of substance. Therefore, three types of physical equilibrium are possible. These are

Solid(s) ⇌ liquid (l)

Liquid(l) ⇌ gas(g)

Solid(s) ⇌ gas(g)

Here the sign double half arrows (⇌) pointing in the opposite directions is both for the reversible change as well as for the equilibrium state.

1. **Solid(s) – liquid(l) equilibrium:** At equilibrium two processes takes place at the same rate i.e.,

   Ice(s) ⇌ water(l)

$H_2O(s) \rightleftharpoons H_2O(l)$

At equilibrium,

Rate of melting of ice = Rate of freezing of water

The temperature at which the solid and liquid states of a pure substance are in equilibrium at the atmospheric.

pressure is called the normal freezing point or melting point of that substance.

2. Liquid(l) – gas(g) equilibrium:

   $H_2O(l) \rightleftharpoons H_2O(g)$

   In such type of equilibrium,

   Rate of vaporisation of water = Rate of condensation of water vapour

3. **Solid(s) – gas(g) equilibrium:** Such type of equilibrium is attained in case of volatile solids.

   Example: If solid iodine is placed in a closed vessel, violet vapours starts appearing in the vessel.

   The intensity of violet vapour increases with time and ultimately it becomes constant.

   $I_2(s) \rightleftharpoons I_2(g)$

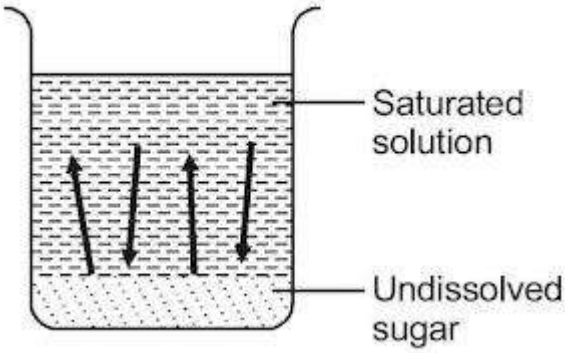
   In this equilibrium,

   Rate of sublimation = Rate of condensation

4. **Solids in liquids:** Suppose sugar is added continuously into a fixed volume of water at room temperature and stirred thoroughly with a glass rod. First the sugar will keep on dissolving but then a stage will come when no more sugar dissolves. Instead it settles down at the bottom. The solution is now said to be saturated and in a state of equilibrium. In this state

   Rate of dissolution = Rate of precipitation

   Sugar(s) ⇌ Sugar (in solution)



The amount of the solid in grams that dissolves in 100 g of the solvent to form a saturated solution at a particular temperature is called the solubility of that solid in the given solvent at that temperature.

Gases in liquids

Such type of equilibrium is present in soda water bottle in which CO_2 gas is dissolved in water under high pressure. There is a state of dynamic equilibrium between the CO_2 present in the solution and the vapours of the gas above the liquid surface at a given temperature.

$CO_2(g) \rightleftharpoons CO_2(aq)$

Henry's law

Periodic table may be defined as the tabular arrangement of elements in such a way that the elements having same properties are kept together.

$m \propto p$

$m = kp$

where k is Henry's constant and its value depends upon the nature of the gas, nature of liquid and temperature.

General Characteristics of Physical Equilibrium:

1. Equilibrium is possible only in a closed system at a given temperature.
2. Both the opposing processes occur at the same rate and there is a dynamic but stable condition.
3. All measurable properties of the system remain constant.
4. When equilibrium is attained for a physical process, it is characterised by constant value of one of its parameters at a given temperature.
5. The magnitude of such quantities at any stage indicates the extent to which

the physical process has proceeded before reaching equilibrium.

Chemical Equilibrium

Every reversible reaction consists of one pair of reaction, one is forward and other is backward reaction. At one stage during reversible reactions, forward and backward reaction proceed at the same time with the same rate, the reaction is then said to be in equilibrium. If the opposing processes involve chemical reactions, the equilibrium is called Chemical equilibrium.

1. **Law of Chemical Equilibrium:** This law states that the rate of an elementary reaction is proportional to the product of the concentration of the reactants. At a constant temperature, the rate of a chemical reaction is directly proportional to the product of the molar concentrations of the reactants each raised to a power equal to the corresponding stoichiometric coefficients as represented by the balanced chemical equation. Let us consider the reaction,

   $A + B \rightleftharpoons C + D$

   $r_f = K_f[A][B]$

   $r_b = K_b[C][D]$

   At equilibrium $r_f = r_b$.

   $K_f[A][B] = K_b[C][D]$

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

   Kc is called the equilibrium constant, [ ] ⟶ denotes active masses.

   For a general reversible reaction,

   $aA + bB \rightleftharpoons cC + dD$

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

2. **Equilibrium constant of reverse reaction:** Equilibrium constant for the reverse reaction is the inverse of the equilibrium constant for the reaction in the forward direction.

$$K'^{c} = \frac{1}{K_c}$$

Relation between K_p and K_c

For a general reversible reaction

aA + bB ⇌ cC+ dD

$$K_c = \frac{[C_C]^C[C_D]^d}{[C_A]^a[C_B]^b} \ldots (1)$$

$$K_p = \frac{[P_C]^C[P_D]^d}{[P_A]^a[P_B]^b}$$

$$K_p = \frac{[C_C]^C \times [RT]^C . [C_D]^d \times [RT]^d}{[C_A]^a \times [RT]^a . [C_B]^b \times [RT]^b}$$

$$K_p = \frac{[C_C]^C[C_D]^d(RT)^{c+b}}{[C_A]^a[C_B]^b(RT)^{a+b}}$$

From Eq. (1),

$$K_p = K_C(RT)^{(c+d)-(a+b)}$$

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

Where, Δn = difference of stoichiometric coefficients of gaseous products and reactants.

3. Characteristics of Equilibrium

i. At the state of equilibrium, certain available properties like pressure, concentration and density becomes constant.

ii. Chemical equilibrium can be established from either side.

iii. A catalyst can cause the state of equilibrium to be reached faster, but does not alter the state of equilibrium.

iv. Chemical equilibrium is dynamic in nature.

v. Any change in external stress (Pressure, temperature or concentration) causes disturbance in equilibrium state. The state of equilibrium being stable, is again reached by some adjustment.

vi. If temperature is changed, a new equilibrium is achieved with a new value for relative concentration of products and reactants.

vii. If temperature is kept constant, pressure and concentration of reactants / products is altered, system shifts in forward or backward direction in order to nullify the alteration (stress).

Factors Affecting Equilibria

i. **Change in Concentration:** When the concentration of any of the reactants or products in an equilibrium reaction is altered, the equilibrium mixture's composition changes in order to minimize the effect of the concentration change.

ii. **Change in Temperature:** According to Le-Chatelier's principle if the temperature of an equilibrium system is increased, the equilibrium will move in the direction of the added heat.

iii. **Change in Pressure:** The pressure has no effect on the equilibrium if the number of moles of gaseous reactants and products does not change. The change in pressure in both liquids and solids can be neglected in heterogeneous chemical equilibrium.

iv. **Change in Volume:** When the volume of a gaseous mixture at equilibrium is increased, the equilibrium moves in the direction of a larger number of gaseous molecules.

v. **Effect of a Catalyst:** The equilibrium is unaffected by the catalyst. This is due to the fact that the catalyst favours both forward and backward reactions equally.

Homogeneous Equilibria

When in an equilibrium reaction, all the reactants and the products are present in the same phase (i.e., gaseous or liquid) it is called a homogeneous equilibrium. For example,

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

Heterogeneous Equilibria

When in an equilibrium reaction, the reactants and the products are present in two or more than two phases, it is called a heterogeneous equilibrium.

The equilibrium between water vapour and liquid water in a closed container is an example of heterogeneous equilibrium.

$H_2O(I) \rightleftharpoons H_2O(g)$

Le Chatelier's Principle

It states that if a stress is applied to a system in equilibrium, the equilibrium for the time being gets disturbed. As a result system moves in a direction which tends to relieve the external stress and finally a new equilibrium is attained.

Ionic Equilibrium

Electrolyte: Electrolytes are the substances which conduct electricity in molten state or in solution. Example HCl, NaCl, KCl, CH_3COOH etc.

Arrhenius theory of Electrolytic dissociation: When an electrolyte is dissolved in a solvent it spontaneously dissociates into oppositely charged particles called ions, to a considerable extent. Electrolytic ionization or dissociation gives ions and unionized molecules in solution. For neutrality, the total charge on cations is equal to the total charge on the anions.

Degree of Dissociation (α): It is the fraction of one mole of the electrolyte that has dissociated under the given conditions. The value of α depends on temperature, dilution of electrolyte, nature of electrolyte and solvent.

α = No. of ionized moles/ Total mo. Moles

Ostwald's Law of Dilution:

According to this law, "The degree of ionization (or dissociation) of any weak electrolyte is inversely proportional to the square root of concentration."

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

Where, K = proportionality constant

Concepts of Acids and Bases

1. **Arrhenius Concept:**
   - **Acid:** Any substance when dissolved in water, increases the concentration of H+. e.g., HCl, H_2SO_4, HNO_3 etc.
   - **Base:** Any substance when dissolved in water, increases the concentration of OH-. e.g., NaOH, KOH etc.

2. **Bronsted - Lowry Concept:**

- **Acid:** Species (Molecule or ion) that donates a proton to another species.
- **Base:** Species (Molecule or ion) that accepts a proton from another species.
  $HCl(Acid) + NH_3(Base) \longrightarrow NH^+_4 + Cl^-$

3. **Lewis acids and bases:**

According to Lewis concept of acids and bases, a Lewis acid is an electron pair acceptor and a Lewis base is an electron pair donor.

Lewis acids: H^+, Ag^+, Fe^{2+}, $AlCl^3$, BF^3, BCl^3, $BeCl^2$ etc.

Lewis Bases: Cl^-, CN^-, OH^-, X^-, NH^{-2}, SH^- etc.

An acid base reaction is the sharing of an electron pair with an acid by a base. This process is simply defined as coordination or neutralisation.

a. A strong acid is an acid that ionizes completely in water.

$HCl(aq) + H_2O\ (l) \longrightarrow H_3O^+\ (aq) + Cl^-(aq)$

b. A weak acid is an acid that is only partially ionized in water

$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_2COO^-(aq) + H_3O^+(aq)$

c. A strong base is a base that ionizes completely in water.

$KOH(aq) \longrightarrow K^+(aq) + OH^-(aq)$

d. A weak base is a base which is partially ionized in water

$NH_3(aq) + H_2O(l) \rightleftharpoons NH{+}4(aq) + OH^-(aq)$

The pH Scale

pH of solution may be defined as negative logarithm of hydronium ion concentration.

$$\mathrm{pH} = -\log[\mathrm{H_3O^+}]$$

$$\mathrm{pH} = \log\frac{1}{[\mathrm{H_3O^+}]}$$

The pH range at 25°C is taken as 0 to 14.

pH = 7 Neutral

pH > 7 Basic

pH < 7 Acidic

Common Ion Effect

The suppression in the dissociation of a weak electrolyte by the addition of a strong electrolyte having a common ion is called common ion effect.

For example: Ionisation of acetic acid (CH_3COOH) and effect of addition of a small amount of acetate ion.

$CH_3COOH(aq) \longrightarrow CH_3COO^-(aq) + H^+(aq)$

$CH_3COONa(aq) \longrightarrow CH_3COO^-(aq) + Na^+(aq)$

Buffer Solution

A buffer solution is that solution which resists any change in its pH value on addition of small amount of acid or base. Although the pH of buffer changes on doing so, but the change in pH value will be less than the expected change. There are three types of buffer solution.

Acidic Buffer: This consists of solution of a weak acid and its salt with strong base. Example; CH_3COOH and CH_3COONa.

Basic Buffer: This consists of solution of weak base and its salt with strong acid. e.g., NH_4OH and NH_4Cl

Salt Buffer: It is a solution of salt which itself can act as a buffer. Such a salt is the salt of weak acid and weak base. For example,

$CH_3COONH_4 \rightleftharpoons NH_4^+ + CH_3COO^-$

When an acid is added, it reacts with CH_3COO^- to produce CH_3COOH and when a base is added, it react with NH^{4+} to produce NH_4OH.

Buffer Capacity

It is the number of moles of acid or base required by one litre of a buffer solution for changing its pH by one unit.

Buffer Capacity = No. of moles of acid or base adder per litre / Change in pH

Solubility and Solubility Product

The number of moles of solute in one litre of a saturated solution (mole / L) is defined as solubility. Let us calculate solubility of salt AgCl.

$AgCl \rightleftharpoons Ag^+ + Cl^-$

$Ksp = [Ag^+][Cl^-]$

Ksp is called solubility product.

In pure water,

$Ksp = [Ag^+][Cl^-]$

$Ksp = S_2 \quad \{\because S = [Ag^+] = [Cl^-]\}$

$S = \sqrt{K_{sp}}$

Summary

1. **Equilibrium:** It represents the state of a process in which the properties like temperature, pressure, concentration of the system do not show any change with the passage of time.
2. **Physical equilibrium:** It is a state of equilibrium in which the two opposing processes involve changes in physical state only.
3. **Chemical equilibrium:** It is a state of equilibrium in which the two opposing processes involve change of chemical species.
4. **Reversible reaction:** A reversible reaction is one which proceeds in both the forward and backward directions.
5. **Law of mass action:** This law states that at constant temperature, the rate of chemical reaction is directly proportional to the product of molar concentrations of the reacting substances.
6. **Equilibrium constant (K):** It is the ratio of the product of the molar concentrations of the substances formed to that of the reacting substances raised to the powers equal to their stoichiometric coefficients in the chemical equation at a particular temperature.
7. **Henry's law:** The mass of a gas dissolved in a given mass of a solvent at any temperature is directly proportional to the pressure of the gas above the solvent.
8. **Le Chatelier's Principle:** When a system in dynamic equilibrium is subjected to a stress such as a change in concentration, pressure or temperature, the equilibrium shifts in a direction that opposes or reduces the stress.
9. **Strong electrolytes:** Electrolytes which are ionized almost completely in aqueous solution under similar conditions of concentration and temperature

are called strong electrolytes.

10. **Weak electrolytes:** Electrolytes which are poorly ionized in aqueous solution under similar conditions of concentration and temperature are called weak electrolytes.

11. **Solubility product:** It is the product of concentration of ions in a saturated solution of a sparingly soluble salt at a given temperature.

12. **Arrhenius acid-base concept:** According to Arrhenious, an acid is a substance which gives hydrogen ions and base is a substance which gives hydroxyl ions in aqueous solutions.

13. **Bronsted-Lowry acid-base concept:** According to this concept, an acid is a proton donor and base is a proton acceptor.

14. **Lewis acid-bases concept:** According to this concept, an acid is an electron pair acceptor and base is an electron pair donor.

15. **pH value:** pH value of a solution is the negative logarithm of the hydrogen ion concentration (in moles per litre) present in it. Thus $pH = -\log[H^+]$

16. **Irreversible reaction:** If a reaction cannot take place in the reverse direction i.e., the products formed do not react to give back the reactants under the same conditions is called an irreversible reaction.

17. **Homogeneous equilibria:** When in an equilibrium reaction, all the reactants and the products are present in the same phase (i.e., gaseous or liquid), it is called a homogeneous equilibrium.

18. **Heterogeneous equilibrium:** When in an equilibrium reaction, the reactants and the products are present in two or more than two phases, it is called a heterogeneous equilibrium.

19. **Buffer solution:** It is defined as a solution which resists in its pH value even when small amounts of the acid or the base are added to it.

20. **Conjugate base:** A base formed by the loss of proton by an acid is called conjugate base of the acid.

NCERT LINE BY LINE QUESTIONS

(1.) **Assertion:** If in a saturated solution of $NaCl, HCl$ gas is passed, then sodium chloride is precipitated. Reason: Concentration of chloride ions decreases due to common ion effect of HCl. **[Page: 230]**

(a.) Both A and R are true and R is the correct explanation of A.
(b.) Both A and R are true but R is not the correct explanation of A.
(c.) A is true but R is false.
(d.) Both A and R are false.

(2.) At what dilution benzoic acid would be 10% dissociated? (Dissociation constant for benzoic acid $= 6.6\times10^{-5}$) **[Page: 219]**

(a.) 198
(b.) 158
(c.) 168
(d.) 178

(3.) Which of the following statements is correct? **[Page: 205]**
(I) Equilibrium constant does not depend on concentration as well as temperature.
(II) Expression of equilibrium constant is applicable only when concentrations of the reactants and products have attained constant value at equilibrium state.
(III) The equilibrium constant for the reverse reaction is equal to half of the equilibrium constant for the forward reaction.

(a.) I only
(b.) I and II only
(c.) II only
(d.) L II and III

(4.) Consider the following statements: **[Page: 228]**
(I) Solubility depends on lattice enthalpy of the salt and solution enthalpy of the ions.
(II) For a salt to be able to dissolve in a particular solvent its solvation enthalpy must be lesser than its lattice enthalpy.
(II1) In case of non-polar solvents, solvation enthalpy is not small and it is sufficient to overcome the lattice enthalpy of the salt. Thus, salt is able to dissolve in non-polar solvents. Which of the following is incorrect?

(a.) I only
(b.) II and III only
(c.) I and II only
(d.) III only

(5.) Which of the following statements is correct? **[Page: 2l6]**

(I) Bronsted acids are species which give H^+ ions only in aqueous solution.

(II) H_3O^+ is an acid according to Arrhenius and Bronsted but not according to Lewis.

(III) NH_3 can act as base according to both Bronsted and Lewis.

(a.) I and II only

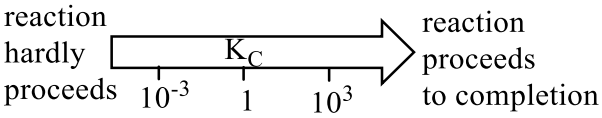
(b.) II and III only

(c.) III only

(d.) None of these

(6.) **Assertion:** For a given reaction,

$A(g)+B(g) \rightleftharpoons C(g)+D(g)$



Reason: If the value of K_c is very large, the reaction proceeds nearly to completion and if $K_c < 10^{-3}$, then reactants predominate over products. **[Page: 206]**

(a.) Both A and R are true and R is the correct explanation of A.

(b.) Both A and R are true but R is not the correct explanation of A.

(c.) A is true but R is false.

(d.) Both A and R are false.

(7.) Read the given reactions,

[NCERT Exemplar Modified, Page: 20l]

$H_2S \rightleftharpoons H^+ + HS^- \quad ---K_1$

$HS^- \rightleftharpoons H^+ + S^{2-} \quad -K_2$

$H_2S \rightleftharpoons 2H^+ + S^{2-} \quad -K_3$

The relation between K_1, K_2 and K_3 is

(a.) $K_3 = K_1 + K_2$

(b.) $K_3 = K_2 \times K_1$

(c.) $K_3 = K_2 / K_1$

(d.) $K_3 = K_1 / K_2$

(8.) For the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$ in a vessel, after the addition of equal number of moles of N_2 and H_2, equilibrium state is formed. Which of the following is correct?. **[Page: 198]**

(a.) $[H_2] = [N_2]$

(b.) $[H_2] < [N_2]$

(c.) $[H_2] > [N_2]$

(d.) $[N_2] = [H_2] = [NH_3]$

(9.) **Assertion:** For any chemical reaction at a particular temperature, the equilibrium constant is fixed and is a characteristic property.

Reason: Equilibrium constant is independent of temperature. **[Page: 209]**

(a.) Both A and R are true and R is the correct explanation of A.

(b.) Both A and R are true but R is nor the correct explanation of A.

(c.) A is true but R is false.

(d.) Both A and R are false.

(10.) Which of the following conditions will favour maximum formation of the product in the reaction? **[NEET-2018, Page: 211]**

$A_2(g)+B_2(g) \rightleftharpoons X_2(g) \quad \Delta_r H = -X\ kJ$

(a.) High temperature and low pressure (b.) Low temperature and low pressure
(c.) Low temperature and high pressure (d.) High temperature and low pressure

(11.) The pH of a solution containing $2.0\,g$ of NaOH per litre of water is

(a.) 12 (b.) 14
(c.) 12.7 (d.) 13.3

(12.) The concentration of $\left[Ag^{+}\right]=10^{-6}$ in a solution. Then the concentration of $\left[Cl^{-}\right]$ required for the precipitation of $AgCl\left(K_{sp}=1\times10^{-13}\right)$ is **[Page: 230]**

(a.) 10^{-6} (b.) 10^{-10}
(c.) 10^{-8} (d.) 10^{-9}

(13.) Accepting the definition that an acid is a proton donor, the acid in the following reaction is:
$NH_3+H_20\rightarrow NH_4^{+}+OH^{-}$

(a.) NH_3
(b.) H_2O
(c.) NH_4^{+}
(d.) OH^{-}

(14.) If the equilibrium constant for **[AIPMT-2O15, Page: 20l]**
$N_2(g)+O_2(g)\rightleftharpoons 2NO(g)$ is K, the equilibrium constant for
$1/2N_2(g)+1/2O_2(g)\rightleftharpoons NO(g)$ will be

(a.) $\frac{1}{2}K$ (b.) K
(c.) K^2 (d.) $K^{1/2}$

(15.) The reaction $N_2(g)+3H_2(g)\rightleftharpoons 2NH_3(g)$ is exothermic. Which of the following sets correctly states the changes in the equilibrium yield of NH_3 when the reaction conditions are changed? **[Page: 211]**

	Increase in pressure	**Increase in temp.**	**Use of catalyst**
(a.)	More NH_3	Less NH_3	No change
(b.)	Less NH_3	More NH_3	No change
(c.)	More NH_3	Less NH_3	More NH_3
(d.)	Less NH_3	More NH_3	More NH_3

(16.) What is the pH of the resulting solution when equal volumes of $0.1M$ NaOH and $0.01M$ HCl are Mixed? **[AIPNIT-2015, Page: 218]**

(a.) 7.0 (b.) 1.04
(c.) 12.35 (d.) 2.0

(17.) Which of the following pairs of substances cannot exist together in solution?

(a.) Na_2CO_3+NaOH (b.) $NaHCO_3+Na_2CO_3$

(c.) $NaHCO_3 + NaOH$ (d.) $NaOH + NaCl$

(18.) The concentration of HCN and NaCN in a solution is 0.0l M each. The concentration of hydroxyl ions is $\left(K_a \text{ for HCN is } 7.2\times10^{-10}\right)$

(a.) 7.2×10^{-8} (b.) 1.39×10^{-5}

(c.) 3.97×10^{-6} (d.) 1.49×10^{-7}

(19.) 0.2 moles of A, 0.3 moles of B and 0.5 moles of C are taken in a *2* L flask to obtain the following equilibrium: **[Page: 206]**

$A(g)+B(g) \rightleftharpoons C(g)$

If the equilibrium constant for the reaction is 3.0×10^{-1}, then predict the direction of the reaction.

(a.) Forward direction (b.) Backward direction

(c.) At equilibrium (d.) Can't be predicted.

(20.) For the reaction,

$ZnCO_3(s) \rightleftharpoons ZnO(s)+CO_2(g)$ **[Page: 202]**

The correct expression for K_p is

(a.) $K_p = \frac{[ZnO][CO_2]}{[ZnCO_3]}$ (b.) $K_p = \frac{P_{ZnCO_3}\times P_{CO_2}}{P_{ZnCO_3}}$

(c.) $K_P = P_{CO_2}$ (d.) $K_p = P_{ZnO}^2 \times P_{co_2}$

(21.) 3. Which of the following will nor act as a buffer **[Page: 227]**

(a.) $NaHCO_3 + H_2CO_3$ (b.) $CH_3COOH + CH_3COONa$

(c.) $NaOH + NaCl$ (d.) $NH_4OH + NH_4CL$

(22.) Solution of $0.1\,N NH_4OH$ and $0.1\,N NH_4Cl$ has pH 9.25. Then find out pK_b of NH_4OH.

(a.) 9.25 (b.) 4.75

(c.) 3.75 (d.) 8.25

(23.) Which of the following cannot act both as Bronsted acid and as Bronsted base **[Odisha NEET-2019, Page: 214]**

(a.) HSO_4^- (b.) HCO_3^-

(c.) NH_3 (d.) HCl

(24.) In which of the following reactions, the yield of the products decreases by increasing the pressure. **[Page: 210]**

(a.) $N_2(g)+3H_2(g) \rightleftharpoons 2NH_3(g)$ (b.) $H_2(g)+I_2(g) \rightleftharpoons 2HI(g)$

(c.) $C(s)+O_2(g) \rightleftharpoons CO_2(g)$ (d.) $PCl_5(g) \rightleftharpoons PCl_3(g)+Cl_2(g)$

(25.) If at 25°C only 25.5g AgBr salt is soluble in 2L water, then K_{sp} for the salt at 25°C will be (in mol^2L^{-2}) **[Page: 228]**

(a.) 4.61×10^{-3} (b.) 6.79×10^{-2}
(c.) 144.89 (d.) 6.85×10^{-3}

(26.) $CaCO_3 \rightleftharpoons CaO + CO_2$ reaction in a lime kiln goes to completion because **[Page: 193]**

(a.) CaO does not react to CO_2 to give $CaCO_3$ (b.) Backward reaction is very slow
(c.) CO_2 formed escapes out (d.) None of these

(27.) The value of ΔG for an isomerization reaction is 5.5 kJ / mol. Then the value of equilibrium constant K_c for the reaction at 298 K is **[Page: 208]**

(a.) 1.5 (b.) 0.596
(c.) 0.1086 (d.) 10.59

(28.) For the given reaction, equilibrium can be reached
$A + B \rightleftharpoons C + D$ **[Page: 196]**

(a.) When only A and B are initially present (b.) When only C and D are initially present
(c.) Any of these (d.) None of these

(29.) **Assertion:** If K_c for the reaction, $2HI \rightleftharpoons H_2 + I_2$ is 5, then K_c for the reaction $H_2 + I_2 \rightleftharpoons 2HI$ is 0.2.
Reason: Equilibrium constant for the reverse reaction is the inverse of the equilibrium constant for the reaction in the forward direction. **[Page: 200]**

(a.) Both A and R are true and R is the correct explanation of A. (b.) Both A and R are true but R is not the correct explanation of A.
(c.) A is true but R is false. (d.) Both A and R are false.

(30.) In the reaction $A + B \rightleftharpoons 2C$, the initial concentrations of A and B are 1.0 and 2.0 mol L^{-1} respectively, If at equilibrium the concentration of *C* is 0.4 mol L^{-1}, the value of equilibrium constant is

(a.) 0.028 (b.) 0.111
(c.) 0.166 (d.) 0.200

(31.) Which of the following statements is correct about acids? **[Page: 213]**

(a.) The lining of our stomach secretes approx. 2.l to 5.l L of hydrochloric acid per day. (b.) Lemon and tamarind contain ascorbic and citric acids.
(c.) Acids turn blue litmus red. (d.) Acids liberate hydrogen gas on reacting with non-metals.

(32.) Which of the given statements is/are correct? **[NCERT Exemplar Modified, Page: 210]**
(I) On addition of catalyst, the equilibrium constant value is not affected.
(II) The intensity of red colour increases when oxalic acid is added to a solution containing iron nitrate and potassium thiocyanate.

(a.) I only (b.) II only
(c.) Both of these (d.) None of these

(33.) The molar solubility of $CaF_2\left(K_{sp}=5.3\times10^{-11}\right)$ in 0.1M Solution of NaF will be **[Odisha NEET-2019, Page: 22S]**

(a.) $5.3\times10^{-10}\,mol\,L^{-1}$ (b.) $5.3\times10^{-11}\,mol\,L^{-1}$

(c.) $5.3\times10^{-8}\,mol\,L^{-1}$ (d.) $5.3\times10^{-9}\,mol\,L^{-1}$

(34.) Which one of the following is the correct order of basic strength?

(a.) $NH_3 > CH_3NH_2 > NF_3$ (b.) $CH_3NH_2 > NH_3 > NF_3$

(c.) $N\Gamma_3 > CH_3NH_2 > NH_3$ (d.) $NH_3 > NP_3 > CH_3NH_2$

(35.) The solubility product of $Al(OH)_3$ is 2.7×10^{-11}. Then its solubility will be **[Page: 228]**

(a.) 10^{-4} (b.) 10^{-6}

(c.) 10^{-3} (d.) 10^{-2}

(36.) What will be the value of pH of $0.01\,mol/dm^3$ HCOOH? $\left(K_a \text{ for } HCOOH = 1.8\times10^{-4}\right)$ **[NCERT Exemplar Modified, Page: 22O]**

(a.) 1.127 (b.) 3.412

(c.) 1.89 (d.) 2.87

(37.) The pH of an aqueous solution of 1 M ammonium formate is if pK_a of HCOOH is 3.8 and pK_b of ammonia is 4.8 (assuming complete dissociation)

(a.) 5.5 (b.) 6.5

(c.) 7.5 (d.) 4.5

(38.) Acidity of $AlCl_3$ can be explained on the basis of which of the following concepts? **[NCERT Exemplar Modified, Page: 216]**

(a.) Arrhenius concept (b.) Bronsted-Lowry concept

(c.) Lewis concept (d.) Bronsted-Lowry as well as Lewis concept

(39.) At 927K and 1 atm pressure, a gaseous mixture of CO and CO_2 an equilibrium with solid carbon has 80.15% CO by mass

$C(s)+CO_2(g) \square \; 2CO(g)$

K_c for the reaction at 927K will be

(a.) 0.189 (b.) 0.072l

(c.) 0.0893 (d.) 0.285

(40.) Match the following: **[Page: 228]**

Salt	K_{sp}
(P) $BaSO_4$	(1) $4s^3$
(Q) Ag_2Cr0_4	(2) $27s^4$
(R) $Zr_3(PO_4)_4$	(3) s^2
(S) $Al(OH)_3$	(4) $6912s^7$

(a.) P-2, Q- 1, R-4, S-3 (b.) P- 1, Q-4, R-3, S-2
(c.) P-3, Q- 1, R-4, S-2 (d.) P- 3, Q- 1, R- 2, S-4

(41.) The aqueous solution of which of the following salts will have the lowest pH?

(a.) $NaClO_3$ (b.) $NaClO$
(c.) $NaClO_4$ (d.) $NaClO_2$

(42.) Consider the following statements: **[Page: 228]**
(I) Ionic product should be greater than solubility product for precipitation.
(II) If solubility is less than 0.0l M, then the salt is known as a sparingly soluble salt.
(1II) Solubility for AB_2 type electrolyte is given by $\left(K_{sp}/4\right)^{1/3}$.
Then the incorrect statement(s) is/are

(a.) I only (b.) II only
(c.) Both II and III (d.) None of these

(43.) **Assertion:** Increasing order of acidity of hydrogen halides is $HF > HCl > HBr > HI$
Reason: Polar nature of bond is the most important factor compared to H-A bond strength to determine acidity. **[Page: 224]**

(a.) Both A and R are true and R is the correct explanation of A. (b.) Both A and R are true but R is not the correct explanation of A.
(c.) A is true but R is false. (d.) Both A and R are false.

(44.) The percentage of pyridine $\left(C_5H_5N\right)$ that forms pyridinium ion $\left(C_5H_5NH\right)$ in a 0.10 M aqueous pyridine solution $\left(K_b \text{ for } C_5H_5N = 1.7\times10^{-9}\right)$ is **[NEET-2016, Phase-II, Page: 22l]**

(a.) 0.0060% (b.) 0.013%
(c.) 0.77% (d.) 1.6%

(45.) The two Bronsted bases in the reaction,

$$\underset{\text{I}}{HC_2O_4} + \underset{\text{II}}{PO_4^{3-}} \rightarrow \underset{\text{III}}{HPO_4^{2-}} + \underset{\text{IV}}{C_2O_4^{2-}}$$

are **[Page: 215]**

(a.) I and II (b.) II and III
(c.) II and IV (d.) I and III

(46.) Boiling point of a liquid can be increased by **[Page: 194]**

(a.) Increasing atmospheric pressure (b.) Decreasing atmospheric pressure
(c.) Going at high altitude (d.) Boiling point is fixed for a substance, it cannot be increased.

(47.) Consider the following equilibrium:
$C(s) + 2H_2(g) \square CH_4(g)$ **[Page: 210]** increasing pressure will shift the equilibrium

(a.) towards left (b.) towards right

(c.) first left then right

(d.) equilibrium state will not change.

(48.) The equilibrium constant for the reaction: **[Page: 199]**
$2A + 3B \rightarrow 2C$ will be

(a.) $\frac{[2A][3B]}{[2C]}$

(b.) $\frac{[A]^2[3B]}{[C]^2}$

(c.) $\frac{[2C]}{[2A][3C]}$

(d.) $\frac{[C]^2}{[A]^2[B]^2}$

(49.) Which one of the following pairs of solution is not an acidic buffer? **[AIPNIT-2015, Page: 227]**

(a.) H_2CO_3 and Na_2CO_3

(b.) H_3PO_4 and Na_3PO_4

(c.) $HClO_4$ and $NaClO_4$

(d.) CH_3COOH and CH_3COONa

(50.) The following solutions were prepared by mixing different volumes of NaOH and HCl of different concentrations: **[NEET-2018, Page: 218]**

(I) $60mL\,\frac{M}{10}HCl + 40mL\,\frac{M}{10}NaOH$

(II) $55mL\,\frac{M}{10}HCl + 45mL\,\frac{M}{10}NaOH$

(III) $75mL\,\frac{M}{5}HCl + 25mL\,\frac{M}{5}NaOH$

(IV) $100mL\,\frac{M}{10}HCl + 100mL\,\frac{M}{10}NaOH$

pH of which one of them will be equal to l?

(a.) IV

(b.) I

(c.) II

(d.) III

TOPIC WISE PRACTICE QUESTIONS

TOPIC 1: Law of Mass Action, Equilibrium Constant (K_c and K_p) and its Applications

1. A cylinder fitted with a movable piston contains liquid water in equilibrium with water vapour at 25 °C. Which operation result in a decrease in the equilibrium vapour pressure?
   1) Moving the piston downward a short distance
   2) Removing a small amount of vapour
   3) Removing a small amount of the liquid water
   4) Dissolving salt in the water

2. The reaction $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ goes to completion in lime kiln because
   1) Of the high temperature.
   2) CaO is more stable than $CaCO_3$.
   3) CaO is not dissociated.
   4) CO_2 escapes continuously.

3. Which of the following conditions represents an equilibrium?
   1) Freezing of ice in an open vessel, temperature of ice is constant.
   2) Few drops of water is present along with air in a balloon, temperature of balloon is constant.
   3) Water is boiling in an open vessel over stove, temperature of water is constant.
   4) All the statements 1), 2) and 3) are correct for the equilibrium.

4. Which of the following statement(s) is/are correct regarding chemical equilibrium?
(i) Equilibrium is maintained rapidly.
(ii) The concentration of reactants and products become same at equilibrium.
(iii) The concentration of reactants and products are constant but different.
(iv) Both forward and backward reactions occur at all times with same speed.
1) (i) and (iii) 2) (i), (ii) and (iii) 3) (iii) and (iv) 4) (ii), (iii) and (iv)

5. At the triple point of water there exists an equilibrium between
1) Ice and water 2) ice and vapours
3) Ice, water and vapours 4) none of the above.

6. Two moles of PCl_5 were heated in a closed vessel of 2 L. At equilibrium 40% of PCl_5 is dissociated into PCl_3 and Cl_2. The value of equilibrium constant is
1) 0.53 2) 0.267 3) 2.63 4) 5.3

7. For the following reaction in gaseous phase $CO(g)+\frac{1}{2}O_2(g)\rightarrow CO_2(g), K_p/K_c$ is
1) $(RT)^{1/2}$ 2) $(RT)^{-1/2}$ 3) (RT) 4) $(RT)^{-1}$

8. At a given temperature the equilibrium constant for the reaction of PCl_5 PCl_3+Cl_2 is 2.4 × 10–3. At the same temperature, the equilibrium constant for the reaction $PCl_3(g)+Cl_2(g)$ $PCl_5(g)$ is
1) 2.4×10^{-3} 2) -2.4×10^{-3} 3) 4.2×10^{2} 4) 4.8×10^{-2}

9. Which of the following is correct for the reaction?
$N_2(g)+3H_2(g)$ $2NH_3(g)$
1) $K_p=K_c$ 2) $K_p<K_c$ 3) $K_p>K_c$ 4) Pressure is required to predict the correlation

10. Boiling point of the liquid depends on the atmospheric pressure. It depends on the altitude of the place; at high altitude the boiling point……..
1) increases 2) decreases 3) either decreases or increases 4) remains same

11. 5 mole of NH_4HS (s) start to decompose at a particular temperature in a closed vessel. If pressure of NH_3(g) in the vessel is 2 atm, then K_p for the reaction,
$NH_4HS(s)$ $NH_3(g)+H_2S(g)$, will be
1) 2 2)4 3)0.4 4)0.8

12. A reaction is said to be in equilibrium when
1) the rate of transformation of reactant to products is equal to the rate of transformation of products to the reactants.
2) 50% of the reactants are converted to products.
3) the reaction is near completion and all the reactants are converted to products.
4) the volume of reactants is just equal to the volume of the products.

13. If the K_p for the equilibrium, $M.5H_2O(s)$ $M.3H_2O(s)+2H_2O(g)$ is 1×10^{-4}. Then $M.5H_2O(s)$ will show efflorescence when it is exposed to an atmosphere where vapour pressure of water is
1) more than 10^{-2} atm 2) below 10^{-2} atm 3) more than 10^{-4} atm 4) below 10^{-4} atm

14. For the reaction : $2NO_2(g)$ $2NO(g)+O_2(g)$, $(K_c=1.8\times10^{-6}$ at 184^0C and $R=0.0831$kJ/mol K$)$
When K_p and K_c are compared at 184^0C, it is found that
1)Whether K_p is greater than, less than or equal to K_c depends upon the total gas pressure
2) $K_p = K_c$ 3) K_p is less than K_c 4) K_p is greater than K_c

15. The equilibrium constant for the reversible reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$ is K and for reaction $\frac{1}{2}N_2 + \frac{3}{2}H_2 \rightleftharpoons NH_3$, the equilibrium constant is K'

The K and K' will be related as:

1) $K \times K' = 1$ 2) $K = K'$ 3) $K' = \sqrt{K}$ 4) $K = \sqrt{K'}$

16. For the reaction $CO + 2H_2 \rightleftharpoons CH_3OH$ at 427^oC, the partial pressure of CH_3OH, CO and H_2 at equilibrium are 2,1 and 0.1atm. respectively. The value of K_p for the decomposition of CH_3OH to CO and H_2 is

1) $5 \times 10^{-3} atm^{-2}$ 2) $2 \times 10^{-2} atm^{-2}$ 3) $5 \times 10^{-2} atm^2$ 4) $2 \times 10^{-1} atm^2$

17. In $A + B \rightleftharpoons C$. The unit of equilibrium constant is

1) Litre mol^{-1} 2) Mol litre 3) Mol $litre^{-1}$ 4) No unit

18. A reaction is A + B $\rightleftharpoons$ C + D. Initially we start with equal concentrations of A and B. At equilibrium we find that the moles of C is two times of A. What is the equilibrium constant of the reaction?

1) 1 2) 3 3) 4 4) 2

19. The decomposition of N_2O_4 to NO_2 is carried out at 280 K in chloroform. When equilibrium has been established, 0.2 mole of N_2O_4 and 2×10^{-3} mole of NO_2 are present in a 2L solution. The equilibrium constant for the reaction, $N_2O_4 \rightleftharpoons 2NO_2$ is

1) 1×10^{-2} 2) 2×10^{-3} 3) 1×10^{-5} 4) 2×10^{-5}

20. For the reaction $H_2(g) + I_2(g) \rightleftharpoons 2HI\ (g)$ at 721 K, the value of equilibrium constant is 50, when equilibrium concentration of both is 5M. Value of K_p under the same conditions will be

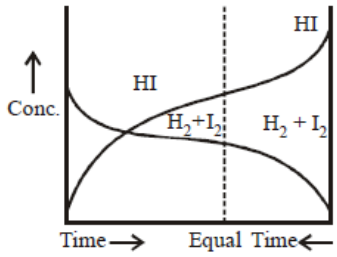
1) 0.02 2) 0.2 3) 50 4) 50 RT

21. The rate constant for forward and backward reaction of hydrolysis of ester are 1.1×10^{-2} and 1.5×10^{-3} per minute respectively. Equilibrium constant for the reaction

$CH_3COOC_2H_5 + H^+ \rightleftharpoons CH_3COOH + C_2H_5OH$ is

1) 4.33 2) 5.33 3) 6.33 4) 7.33

22. Consider the following graph and mark the correct statement.



1) Chemical equilibrium in the reaction, $H_2 + I_2 \rightleftharpoons 2HI$ can be attained from either directions.

2) Equilibrium can be detained when H_2 and I_2 are mixed in an open vessel.

3) The concentrations of H_2 and I_2 keep decreasing while concentration of HI keeps increasing with time.

4) We can find out equilibrium concentration of H_2 and I_2 from the given graph.

23. Steam reacts with iron at high temperature to give hydrogen gas and Fe_3O_4 (s). The correct expression for the equilibrium constant is

1) $\frac{P_{H_2}^2}{P_{H_2O}^2}$ 2) $\frac{(P_{H_2})^4}{(P_{H_2O})4}$ 3) $\frac{(P_{H_2})^4[Fe_3O_4]}{(P_{H_2O})4[Fe]}$ 4) $\frac{[Fe_3O_4]}{[Fe]}$

24. For a reaction the free energy change. $\Delta G = -RT \ln K_p + RT \ln Q_p$ where K_p = equilibrium constant. Q_p = reaction quotient. For the reaction to be in equilibrium state

1) $\frac{Q_p}{K_p} > 1$ 2) $\frac{Q_p}{K_p} < 1$ 3) $\frac{Q_p}{K_p} = 1$ 4) $Q_p K_p = 1$

TOPIC 2: Relation between K, Q, G and Factors Affecting Equilibrium

25. What is the effect of halving the pressure by doubling the volume on the following system at 500 °C?

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

1) Shift to reactant side 2) Shift to product side
3) Liquefaction of HI 4) No effect

26. The equilibrium constant for a reaction, $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ is 4×10^{-4} at 2000 K. In the presence of catalyst, the equilibrium is attained 10 times faster. The equilibrium constant in presence of catalyst at 2000 K is:

1) 10×10^{-4} 2) 4×10^{-2} 3) 4×10^{-4} 4) 40×10^{-4}

27. The formation of SO_3 takes place according to the following reaction:

$2SO_2 + O_2 \rightleftharpoons 2SO_3; \Delta H = -45.2 kcal$

The formation of SO_3 is favoured by

1) Increase in temperature 2) removal of oxygen
3) Increase of volume 4) increase of pressure

28. According to Le-chatelier's principle, adding heat to a solid $\rightleftharpoons$ *f* liquid equilibrium will cause the

1) Temperature to increase 2) temperature to decrease
3) Amount of liquid to decrease 4) amount of solid to decrease

29. Consider the reaction

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ for which $K_c = 278\ M^{-1}$. 0.001 mole of each of the reagents $SO_2(g)$, $O_2(g)$ and $SO_3(g)$ are mixed in a 1.0 L flask. Determine the reaction quotient of the system and the spontaneous direction of the system:

1) $Q_c = 1000$; the equilibrium shifts to the right
2) $Q_c = 1000$; the equilibrium shifts to the left
3) $Q_c = 0.001$; the equilibrium shifts to the left
4) $Q_c = 0.001$; the equilibrium shifts to the right

30. For the reaction $XCO_3(s) \rightleftharpoons XO(s) + CO_2(g)$, $K_p = 1.642$ atm at 727°C. If 4 moles of $XCO_3(s)$ was put into a 50 litre container and heated to 727°C. What mole percent of the XCO_3 remains unreacted at equilibrium?

1) 20 2) 25 3) 50 4) None of these

31. Which one of the following information can be obtained on the basis of Le Chatelier principle?

1) Dissociation constant of a weak acid 2) Entropy change in a reaction
3) Equilibrium constant of a chemical reaction
4) Shift in equilibrium position on changing value of a constraint

32. The equilibrium constant K_p for a homogeneous gaseous reaction is 10^{-8}. The standard Gibbs free energy change $\Delta G°$ for the reaction (using R = 2 cal K^{-1} mol^{-1}) is

1) 10.98 kcal 2) –1.8 kcal 3) –4.1454 kcal 4) +4.1454 kcal

33. For the reaction $NH_4HS(g) \rightleftharpoons NH_3(g) + H_2S(g)$ in a closed flask, the equilibrium pressure is *P* atm. The standard free energy of the reaction would be:
1) – RT ln p 2) – RT (ln p – ln 2) 3) – 2 RT ln p 4) – 2 RT (ln p – ln 2)

TOPIC 3: Theories of Acids and Bases, Ionic Product of Water and pH Scale

34. In HS–, I–, RNH2 and NH3, order of proton accepting tendency will be
1) $I^- > NH_3 > RNH_2 > HS^-$ 2) $HS^- > RNH_2 > NH_3 > I^-$
3) $RNH_2 > NH_3 > HS^- > I^-$ 4) $NH_3 > RNH_2 > HS^- > I^-$

35. Which one of the following compounds is **not** a protonic acid?
1) $SO_2(OH)_2$ 2) $B(OH)_3$ 3) $PO(OH)_3$ 4) $SO(OH)_2$

36. Which of the following molecules acts as a Lewis acid?
1) $(CH_3)_2O$ 2) $(CH_3)_3P$ 3) $(CH_3)_3N$ 4) $(CH_3)_3B$

37. Which one of the following is the correct statement?
1) HCO_3^- is the conjugate base of CO_3^{2-} 2) NH_2^- is the conjugate acid of NH_3.
3) H_2SO_4 is the conjugate acid of HSO_4^- 4) NH_3 is the conjugate base of NH_2^-

38. Which of the following has highest pH?
1) $\frac{M}{4}KOH$ 2) $\frac{M}{4}NaOH$ 3) $\frac{M}{4}NH_4OH$ 4) $\frac{M}{4}Ca(OH)_2$

39. The pH of a 10^{-3} M HCl solution at 25 °C if it is diluted 1000 times, will be –
1) 3 2) zero 3) 5.98 4) 6.02

40. What will be the pH of a solution formed by mixing 50 mL of 0.5 M HCl solution and 150 mL of 0.5 M NaOH solution and 300 mL H_2O?
1) 13 2) 12.7 3) 7 4) 11

41. Water is well known amphoprotic solvent. In which chemical reaction water is behaving as a base?
1) $H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$ 2) $H_2O + H_2O \rightarrow H_3O^+ + OH^-$
3) $H_2O + NH_2^- \rightarrow NH_3 + OH^-$ 4) $H_2O + NH_3 \rightarrow NH_4^+ + OH^-$

42. The value of ionic product of water at 393 K is
1) Less than 1×10^{-14} 2) greater than 1×10^{-14} 3) Equal to 1×10^{-14} 4) equal to 1×10^{-7}

43. Three reactions involving $H_2PO_4^-$ are given below:
1) $H_3PO_4 + H_2O \rightarrow H_3O^+ + H_2PO_4^-$ 2) $H_2PO_4^- + H_2O \rightarrow HPO_4^{2-} + H_3O^+$
3) $H_2PO_4^- + OH^- \rightarrow H_3PO_4 + O^{2-}$
In which of the above does $H_2PO_4^-$ act as an acid?
1) (ii) only 2) (i) and (ii) 3) (iii) only 4) (i) only

44. The value of the ionic product of water
1) depends on volume of water 2) depends on temperature
3) changes by adding acid or alkali 4) always remains constant

45. Which of the following increasing order of pH of 0.1 *M* solution of the compounds 1) $HCOONH_4$, 2) CH_3COONH_4, 3) CH_3COONa and 4) NH_4Cl is correct?
1) A < D < B < C 2) D < A < C < B 3) A < D < C < B 4) D < A < B < C

46. On increasing the temperature of pure water
1) both pH and pOH increase 2) both pH and pOH decrease
3) pH increases and pOH decreases 4) pH decreases and pOH increases

TOPIC 4: Ionization of Weak Acids and Bases and Relation between K_a and K_b

47. The dissociation constant of two acids HA_1 and HA_2 are 3.14×10^{-4} and 1.96×10^{-5} respectively. The relative strength of the acids will be approximately
1) 1 : 4 2) 4 : 1 3) 1 : 16 4) 16 : 1

48. At 298*K* a 0.1 M CH_3COOH solution is 1.34% ionized. The ionization constant *Ka* for acetic acid will be
1) 1.82×10^{-5} 2) 18.2×10^{-5} 3) 0.182×10^{-5} 4) none of these

49. K_{a1} , K_{a2} and K_{a3} are the respective ionisation constants for the following reactions.

$H_2S \rightleftharpoons H^+ + HS^-$

$HS^- \rightleftharpoons H^+ + S^{2-}$

$H_2S \rightleftharpoons 2H^+ + S^{2-}$

The correct relationship between K_{a1}, K_{a2} and K_{a3} is
1) $K_{a_3} = K_{a_1} \times K_{a_2}$ 2) $K_{a_3} = K_{a_1} + K_{a_2}$ 3) $K_{a_3} = K_{a_1} - K_{a_2}$ 4) $K_{a_3} = K_{a_1} / K_{a_2}$

50. A monobasic weak acid solution has a molarity of 0.005 and pH of 5. What is the percentage ionization in this solution?
1) 2.0 2) 0.2 3) 0.5 4) 0.25

51. Which of the following will occur if 0.1 M solution of a weak acid is diluted to 0.01 M at constant temperature?
1) $[H^+]$ will decrease to 0.01 M 2) pH will decrease
3) Percentage ionization will increase 4) K_a will increase

52. At a certain temperature the dissociation constants of formic acid and acetic acid are 1.8×10^{-4} and $1.8 \times 10_{-6}$ respectively. The concentration of acetic acid solution in which the hydrogen ion has the same concentration as in 0.001 M formic acid solution is equal to
1) 0.001 M 2) 0.01 M 3) 0.1 M 4) 0.0001 M

53. The degree of dissociation of acetic acid in a 0.1 M solution is 1.32×10^{-2}. Find out the dissociation constant of the acid.
1) 1.50×10^{-4} 2) 1.80×10^{-16} 3) 1.76×10^{-5} 4) 1.2×10^{-3}

TOPIC 5: Common Ion Effect, Salt Hydrolysis, Buffer Solutions and Solubility Product

54. Solubility of salt A_2B_3 is 1×10^{-4}, its solubility product is
1) 1.08×10^{20} 2) 1.08×10^{18} 3) 2.6×10^{-18} 4) 1.08×10^{-18}

55. Which of the following metal sulphides has maximum solubility in water?
1) CdS ($K_{sp} = 36 \times 10^{-30}$) 2) FeS ($K_{sp} = 11 \times 10^{-20}$)
3) HgS ($K_{sp} = 32 \times 10^{-54}$) 4) ZnS ($K_{sp} = 11 \times 10^{-22}$)

56. The K_{sp} for $Cr(OH)_3$ is 1.6×10^{-30}. The solubility of this compound in water is :
1) $\sqrt[4]{1.6 \times 10^{-30}}$ 2) $\sqrt[4]{1.6 \times 10^{-30} / 27}$ 3) $1.6 \times 10^{-30/27}$ 4) $\sqrt{1.6 \times 10^{-30}}$

57. Consider the following equilibrium

$AgCl\downarrow + 2NH_3 \rightleftharpoons [Ag(NH_3)_2]$

White precipitate of AgCl appears on adding which of the following?
1) NH_3 2) aqueous NaCl 3) aqueous HNO_3 4) aqueous NH_4Cl

58. A saturated solution of Ag_2SO_4 is 2.5×10^{-2} M, the value of its solubility product is :
1) 62.5×10^{-6} 2) 6.25×10^{-4} 3) 15.625×10^{-6} 4) 3.125×10^{-6}

59. The solubility of AgI in NaI solution is less than that in pure water because

1) the temperature of the solution decreases.
2) solubility product of AgI is less than that of NaI.
3) of common ion effect. 4) AgI forms complex with NaI.

60. A litre of solution is saturated with AgCl. To this solution if 1.0×10^{-4} mole of solid NaCl is added, what will be the $[Ag^+]$, assuming no volume change?
1) More 2) Less 3) Equal 4) Zero

61. The solubility product of a sparingly soluble salt BA_2 is 4×10^{-12}. The solubility of BA_2 is
1) 4×10^{-4} 2) 4×10^{-12} 3) 4×10^{-3} 4) 1×10^{-4}

62. The pH of an acidic buffer mixture is
1) > 7 2) = 7 3) < 7 4) depends upon K_a of the acid

63. The pH of a weak mono acidic base, neutralized up to 80% with a strong acid in a dilute solution, is 7.40. The ionization constant of the base is
1) 1.0×10^{-5} 2) 1.6×10^{-7} 3) 1.0×10^{-6} 4) None of these

64. When a buffer solution, sodium acetate and acetic acid is diluted with water :
1) Acetate ion concentration increases 2) H^+ ion concentration increases
3) OH^- ion conc. Increases 4) H^+ ion concentration remains unaltered

65. A buffer solution is prepared by mixing 0.1 M ammonia and 1.0 M ammonium chloride. At 298 K, the pK_b of NH_4OH is 5.0. The pH of the buffer is
1) 10.0 2) 9.0 3) 6.0 4) 8.0

66. The product of ionic concentration in a saturated solution of an electrolyte at a given temperature is constant and is known as
1) Ionic product of the electrolyte 2) Solubility product
3) Ionization constant 4) Dissociation constant

67. What is the molar solubility of $Fe(OH)_3$ if $K_{sp} = 1.0 \times 10^{-38}$?
1) 3.16×10^{-10} 2) 1.386×10^{-10} 3) 1.45×10^{-9} 4) 1.12×10^{-11}

68. The correct order of increasing solubility of AgCl in 1) water, 2) 0.1 M NaCl, 3) 0.1 M, $BaCl_2$, 4) 0.1 M NH_3 is
1) D > A > B > C 2) D > C > B > A 3) B > A > D > C 4) A > D > B > C

69 . If s and S are respectively solubility and solubility product of a sparingly soluble binary electrolyte then
1) s = S 2) $s = S^2$ 3) $s = S^{1/2}$ 4) s 1/2 S

70. Solubility product of $M(OH)_2$ is 10^{-14}. What should be the concentration of M^{2+} in 0.1 M solution of NH_4OH, if NH_4OH gets 10% ionised ?
1) 10^{-10} 2) 10^{-5} 3) 10^{-12} 4) 10^{-4}

NEET PREVIOUS YEARS QUESTIONS

1. Which one of the following conditions will favour maximum formation of the product in the reaction,
$A_2(g) + B_2(g) \rightleftharpoons X_2(g)$ $\Delta_r H = -X$ kJ : **[2018]**
1) Low temperature and high pressure 2) Low temperature and low pressure
3) High temperature and low pressure 4) High temperature and high pressure

2. The solubility of $BaSO_4$ in water is 2.42×10^{-3} gL^{-1} at 298 K. The value of its solubility product (K_{sp}) will be (Given molar mass of $BaSO_4$ = 233 g mol^{-1})
[2018]
1) 1.08×10^{-10} mol^2L^{-2} 2) 1.08×10^{-12} mol^2L^{-2}

3) 1.08×10^{-8} mol²L^{-2} 4) 1.08×10^{-14} mol²L^{-2}

3. Following solutions were prepared by mixing different volumes of NaOH and HCl of different concentrations:

[2018]

1. 60mL $\frac{M}{10}$ HCl + 40mL $\frac{M}{10}$ NaOH 2) 55mL $\frac{M}{10}$ HCl + 45mL $\frac{M}{10}$ NaOH

3) 75mL $\frac{M}{5}$ HCl + 25mL $\frac{M}{5}$ NaOH 4) 100mL $\frac{M}{10}$ HCl + 100mL $\frac{M}{10}$ NaOH

pH of which one of them will be equal to 1?

1) 2 2) 1 3) 3 4) 4

4. The equilibrium constant of the following are : **[2017]**

$N_2 + 3H_2 \rightleftharpoons 2NH_3 \quad K_1$

$N_2 + O_2 \rightleftharpoons 2NO \quad K_2$

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O \quad K_3$

The equilibrium constant (K) of the reaction:

$2NH_3 + \frac{5}{2}O_2 \overset{K}{\rightleftharpoons} 2NO + 3H_2O$, will be;

1) $K_2K_3^3/K_1$ 2) K_2K_3/K_1 3) $K_2^3K_3/K_1$ 4) $K_1K_3^3/K_2$

5. Concentration of the Ag^+ ions in a saturated solution of $Ag_2C_2O_4$ is 2.2×10^{-4} mol L^{-1}. Solubility product of $Ag_2C_2O_4$ is

[2017]

1) 2.66×10^{-12} 2) 4.5×10^{-11} 3) 5.3×10^{-12} 4) 2.42×10^{-8}

6. Which one of the following statements is not correct? **[2017]**

1) The value of equilibrium constant is changed in the presence of a catalyst in the reaction at equilibrium

2) Enzymes catalyse mainly bio-chemical reactions

3) Coenzymes increase the catalytic activity of enzyme

4) Catalyst does not initiate any reaction

7. A 20 litre container at 400 K contains CO_2(g) at pressure 0.4 atm and an excess of SrO (neglect the volume of solid SrO). The volume of the container is now decreased by moving the movable piston fitted in the container. The maximum volume of the container, when pressure of CO_2 attains its maximum value, will be : **[2017]**

(Given that : $SrCO_3(s) \rightleftharpoons SrO(s) + CO_2(g), K_p = 1.6atm$)

1) 10 litre 2) 4 litre 3) 2 litre 4) 5 litre

8. MY and NY_3, two nearly insoluble salts, have the same K_{sp} values of 6.2×10^{-13} at room temperature. Which statement would be true in regard to MY and NY_3?

[2016]

1) The molar solubilities of MY and NY_3 in water are identical.

2) The molar solubility of MY in water is less than that of NY_3

3) The salts MY and NY_3 are more soluble in 0.5 M KY than in pure water.

4) The addition of the salt of KY to solution of MY and NY_3 will have no effect on their solubilities.

9. Consider the nitration of benzene using mixed conc of H_2SO_4 and HNO_3. If a large amount of $KHSO_4$ is added to the mixture, the rate of nitration will be **[2016]**

1) faster 2) slower 3) unchanged 4) doubled

10. The Ksp of Ag_2CrO_4, AgCl, AgBr and AgI are respectively, 1.1×10^{-12}, 1.8×10^{-10}, 5.0×10^{-13}, 8.3×10^{-17}. Which one of the following salts will precipitate last if $AgNO_3$ solution is added to the solution containing equal moles of NaCl, NaBr, NaI and Na_2CrO_4? **[2015]**
1) AgCl 2) AgBr 3) Ag_2CrO_4 4) AgI

11. If the value of an equilibrium constant for a particular reaction is 1.6 × 1012, then at equilibrium the system will contain :- **[2015]**
1) mostly reactants 2) mostly products
3) similar amounts of reactants and products 4) all reactants

12. Which of the following statements is correct for a reversible process in a state of equilibrium? **[2015]**
1) $\Delta G = 2.30\ RT \log K$ 2) $\Delta G^o = -2.30\ RT \log K$
3) $\Delta G^o = 2.30\ RT \log K$ 4) $\Delta G = -2.30\ RT \log K$

13. Which of the following salts will give highest pH in water? **[2014]**
1) KCl 2) NaCl 3) Na_2CO_3 4) $CuSO_4$

14. Using the Gibbs energy change, $\Delta G° = +63.3kJ$, for the following reaction, **[2014]**
$Ag_2CO_3 \rightarrow 2Ag^+(aq) + CO_3^{2-}(aq)$ the K_{sp} of $Ag_2CO_3(s)$ in water at 25^0C is: $\left(R = 8.314JK^{-1}mol^{-1}\right)$
1) 3.2×10^{-26} 2) 8.0×10^{-12} 3) 2.9×10^{-3} 4) 7.9×10^{-2}

15. For the reversible reaction, $N_2(g) + 3H_2(g) \square 2NH_3(g) + heat$
The equilibrium shifts in forward direction **[2014]**
1) By increasing the concentration of $NH_3(g)$
2) By decreasing the pressure
3) By decreasing concentration of $N_2(g)$ and $H_2(g)$
4) By increasing pressure and decreasing temperature.

16. For a given exothermic reaction, K_p and $K_p^{'}$ are the equilibrium constants at temperatures T_1 and T_2, respectively. Assuming that heat of reaction is constant in temperature range between T_1 and T_2, it is readily observed that: **[2014]**
1) $K_p > K_p^{'}$ 2) AgBr 3) Ag_2CrO_4 4) AgI

17. pH of a saturated solution of $Ca(OH)_2$ is 9. The solubility product (Ksp) of $Ca(OH)_2$ is **[NEET-2019]**
(1) 0.5×10^{-15} (2) 0.25×10^{-10} (3) 0.125×10^{-15} (4) 0.5×10^{-10}

18. Conjugate base for Bronsted acids H_2O and HF are:- **[NEET-2019]**
(1) OH^- and H_2F^+ respectively (2) H_3O^+ and F^-, respectively
(3) OH^- and F^-, respectively (4) H_3O^+ and H_2F^+, respectively

19. Which will make basic buffer? **[NEET-2019]**
(1) 50 mL of 0.1 M NaOH + 25 mL of 0.1 M CH_3COOH
(2) 100 mL of 0.1 M CH_3COOH + 100 mL of 0.1M NaOH
(3) 100 mL of 0.1 M HCl + 200 mL of 0.1 M NH_4OH
(4) 100 mL of 0.1 M HCl + 100 mL of 0.1 M NaOH

20. The pH of 0.01 M NaOH (aq) solution will be **[NEET-2019 ODISSA]**
(1) 7.01 (2) 2 (3) 12 (4) 9

21. Which of the following cannot act both as Bronsted acid and as Bronsted base? **[NEET-2019 ODISSA]**
(1) HCO_3^- (2) NH_3 (3) HCl (4) HSO_4^-

22. The molar solubility of CaF_2 ($K_{sp} = 5.3 \times 10^{-11}$) in 0.1 M solution of NaF will be **[NEET-2019 ODISSA]**
(1) 5.3×10^{-11} mol L^{-1} (2) 5.3×10^{-8} mol L^{-1} (3) 5.3×10^{-9} mol L^{-1} (4) 5.3×10^{-10} mol L^{-1}

23. Which among the following salt solutions is basic in nature ? **[NEET-2020 COVID]**
(1) Ammonium chloride (2) Ammonium sulphate
(3) Ammonium nitrate (4) Sodium acetate

24. The solubility product for a salt of the type AB is 4×10^{-8}. What is the molarity of its standard solution? **[NEET-2020 COVID]**
(1) 2×10^{-4}mol/L (2) 16×10^{-16}mol/L (3) 2×10^{-16}mol/L (4) 4×10^{-4}mol/L

25. Hydrolysis of sucrose is given by the following reaction **[NEET-2020]**
Sucrose + H_2O $\rightarrow$ $Glucose + Fructose$
If the equilibrium constant (K_C) is 2×10^{13} at 300K, the value of $\Delta_r G^O$ at the same temperature will be
1. $-8.314J\ mol^{-1}K^{-1}\times300K\times\ln\left(4\times10^{13}\right)$ 2. $-8.314J\ mol^{-1}K^{-1}\times300K\times\ln\left(2\times10^{13}\right)$
3. $8.314J\ mol^{-1}K^{-1}\times300K\times\ln\left(2\times10^{13}\right)$ 4. $8.314J\ mol^{-1}K^{-1}\times300K\times\ln\left(3\times10^{13}\right)$

26. Find out the solubility of $Ni(OH)_2$ in $0.1M\ NaOH$. Given that the ionic product of $Ni(OH)_2$ is 2×10^{-15}. **[NEET-2020]**
1) $1\times10^{8}M$ 2) $2\times10^{-13}M$ 3) $2\times10^{-8}M$ 4) $1\times10^{-13}M$

27 The pK_b of dimethylamine and pK_a of acetic acid are 3.27 and 4.77 respectively at T (K). The correct option for the pH of dimethylammonium acetate solution is : **[NEET-2021]**
1) 5.50 2) 7.75 3) 6.25 4) 8.50

28. $3O_2(g) \rightarrow 2O_3(g)$ **[NEET-2022]**
For the above reaction at 298 K, K_c is found to be 3.0×10^{-59}. If the concentration of O_2 at equilibrium is 0.040M then concentration of O_3 in M is
1) 4.38×10^{-32} 2) 1.9×10^{-63} 3) 2.4×10^{31} 4) 1.2×10^{21}

29. The pH of the solution containing 50mL each of 0.10 M sodium acetate and 0.01M acetic acid is [Given pK_a of CH_3COOH = 4.57] **[NEET-2022]**
1) 5.57 2) 3.57 3)4.57 4) 2.57

NCERT LINE BY LINE QUESTIONS – ANSWERS

(1.)	c	**(2.)**	c	**(3.)**	c	**(4.)**	b	**(5.)**	c
(6.)	a	**(7.)**	b	**(8.)**	b	**(9.)**	c	**(10.)**	c
(11.)	c	**(12.)**	a	**(13.)**	b	**(14.)**	d	**(15.)**	a
(16.)	c	**(17.)**	c	**(18.)**	b	**(19.)**	b	**(20.)**	c
(21.)	c	**(22.)**	b	**(23.)**	d	**(24.)**	d	**(25.)**	a

(26.)	c	**(27.)**	c	**(28.)**	c	**(29.)**	a	**(30.)**	b
(31.)	c	**(32.)**	a	**(33.)**	a	**(34.)**	b	**(35.)**	c
(36.)	d	**(37.)**	b	**(38.)**	c	**(39.)**	b	**(40.)**	c
(41.)	c	**(42.)**	d	**(43.)**	d	**(44.)**	b	**(45.)**	c
(46.)	a	**(47.)**	b	**(48.)**	d	**(49.)**	c	**(50.)**	d

TOPIC WISE PRACTICE QUESTIONS - ANSWERS

1) 4	**2)** 4	**3)** 2	**4)** 3	**5)** 3	**6)** 2	**7)** 2	**8)** 3	**9)** 2	**10)** 2
11) 2	**12)** 1	**13)** 2	**14)** 4	**15)** 3	**16)** 1	**17)** 1	**18)** 3	**19)** 3	**20)** 3
21) 4	**22)** 1	**23)** 2	**24)** 3	**25)** 4	**26)** 3	**27)** 4	**28)** 4	**29)** 2	**30)** 4
31) 4	**32)** 1	**33)** 4	**34)** 3	**35)** 2	**36)** 4	**37)** 3	**38)** 4	**39)** 3	**40)** 1
41) 1	**42)** 2	**43)** 1	**44)** 2	**45)** 4	**46)** 2	**47)** 2	**48)** 1	**49)** 1	**50)** 2
51) 3	**52)** 2	**53)** 3	**54)** 4	**55)** 2	**56)** 2	**57)** 3	**58)** 1	**59)** 3	**60)** 2
61) 4	**62)** 4	**63)** 3	**64)** 4	**65)** 1	**66)** 2	**67)** 2	**68)** 1	**69)** 3	**70)** 1

NEET PREVIOUS YEARS QUESTIONS-ANSWERS

1) 1	**2)** 1	**3)** 3	**4)** 1	**5)** 3	**6)** 1	**7)** 4	**8)** 2	**9)** 2	**10)** 3
11) 2	**12)** 2	**13)** 3	**14)** 2	**15)** 4	**16)** 1	**17)** 1	**18)** 3	**19)** 3	**20)** 3
21) 3	**22)** 3	**23)** 4	**24)** 1	**25)** 2	**26)** 2	**27)** 2	**28)** 1	**29)** 1	

NCERT LINE BY LINE QUESTIONS – SOLUTIONS

(1.) (c) On passing HCl gas in saturated $NaCl$ solution, sodium chloride is precipitated due to increased concentration of chloride ion available from the dissociation of HCl.

(2.) (c) $K = \dfrac{\alpha^2}{V(1-\alpha)}$ $\alpha = \dfrac{10}{100} = 0.1$

$$V = \frac{\alpha^2}{K(1-\alpha)} = \frac{(0.1)^2}{6.6\times10^{-5}(1-0.1)}$$

$$= \frac{0.1\times0.1}{6.6\times9\times10^{-6}} = \frac{10,000}{6.6\times9} = 168.35L$$

(3.) (c) Equilibrium constant depends only on temperature and equilibrium constant for the reverse reaction is equal to the inverse of the equilibrium constant for the forward reaction.

(4.) (b) For a salt to be able to dissolve, its solvation enthalpy should be greater than its lattice enthalpy. As in non-polar solvents, solvation energy is small, it does not overcome the lattice enthalpy, so the salt is unable to dissolve in non-polar solvents.

(5.) (c) The species which can donate H^+ are acids according to Bronsted.

(6.) (a)

(7.) (b) On adding equations I and II we can get equation III. Thus, $K_3 = K_1 \times K_2$

(8.) (b) To attain equilibrium 1 mol of N_2 reacts with 3 mol of H_2

	N_2	+	$3H_2$		$2NH_3$
Initially	1		1		0
At equilibrium	$1-x$		$1-3x$		$2x$

Thus, N_2 is more than H_2 at equilibrium.

(9.) (c) Equilibrium constant is a characteristic property but it depends on temperature.

(10.) (c) In an exothermic process, according to Le Chatelier's principle, raising the temperature shifts the equilibrium to the left. And on increasing pressure the equilibrium shifts in a direction where the

number of moles of gas decreases (*i.e*, pressure decreases). So for maximum yield we require low temperature and high pressure.

(12.) (a) $AgCl \rightleftharpoons Ag^{+} + Cl^{-}$

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

$\left[Cl^{-}\right] = K_{sp} / \left[Ag^{+}\right] \frac{1\times10^{-13}}{10^{-6}} = 10^{-7}$

For precipitation $Q_{sp} > K_{sp}$ thus $\left[Cl^{-}\right] > 10^{-7}$

(14.) (d) For dividing equation by two

$K' = \sqrt{K}$ or $K^{1/2}$

(15.) (a) As the reaction is exothermic and has less number of moles at the product side, so decrease in temperature and increase is pressure are favourable conditions.

(16.) (c) Normality of resulting solution

$N_3V_3 = N_1V_1 - N_2V_2$

$N_3 \times 2 = 0.1\times1 - 0.01\times1$

$N_3 \times 2 = 0.09 \Rightarrow N_3 = 0.09/2$

As $NaOH$ is higher in concentration, thus

$\left[OH^{-}\right] = 0.09/4 = 0.0225$

$pOH = -\log 0.0225 = 1.65$

$pH = 14 - 1.65 = 12.35$

(19.) **(b)** For $\mathbf{A(g) + B(g) \rightleftharpoons C(g)}$

$$\mathbf{Q_c = \frac{[C]}{[A][B]} = \frac{0.5}{0.3\times0.2} = 8.33}$$

$\mathbf{K_c = 0.3}$

Since, $\mathbf{Q_c > K_c}$, so reaction will proceed in the backward direction.

(20.) (c) The pressure for solid substances is taken as unity. Thus, $K_p = P(CO_2)$

(21.) (c) It is a mixture of a strong base and its salt with strong acid, so it is not a buffer solution.

(23.) (d) Bronsted acid is capable of donating a proton and Bronsted base is 'capable of' accepting a proton. Hense HSO_4^-, HCO_3^- and NH_3 can act as both Bronsted acid and base but HCl can act only as Bronsted acid.

(24.) (d) According to Le Chatelier's principle, on increasing the pressure equilibrium shifts to left since the volume of the gaseous products is increasing in the given reaction.

(25.) (a) Solubility of $AgBr = 25.5/2 = 12.75$ g/L

Solubility in moles/L $= \frac{12.75}{187.8} = 6.79\times10^{-2}$ mol / L

For $AgBr$ $K_{sp} = s^2 = \left(6.79\times10^{-2}\right)^2$

$K_{sp} = 4.61\times10^{-3} mol^2L^{-2}$

(26.) (c) As CO_2 is gas it escapes out and thus reverse reaction is not possible.

(27.) (c) $\mathbf{\Delta G = -RT\ln K_c}$

$5.5\times10^3 = -8.314\times298\times\mathbf{\ln K_c}$

$$\mathbf{\ln K_c} = -\frac{5.5\times10^3}{8.314\times298} = -2.22$$

$\mathbf{K_c} = \boldsymbol{e}^{-222} = 0.1086$

(28.) (c) Equilibrium can be achieved from either direction.

(29.) (a) If $2HI \rightleftharpoons H_2 + I_2$ $K_c = 5$

Then for reaction

$H_2 + I_2 \rightleftharpoons 2HI$ $K_c^{'} = \frac{1}{K_c} = \frac{1}{5} = 0.2$

(31.) (c) The lining of our stomach secretes approx. 1.2-1.51L/day and tamarind contains tartaric acid not ascorbic and citric acid.

(32.) (a) The intensity of red colour decreases when oxalic acid is added to a solution containing iron (III) nitrate and potassium thiocyanate as concentration of $\left[\mathbf{Fe(SCN)}\right]^{2+}$ decreases.

(33.) (a) $CaF_2 \geq Ca^{2+} + 2F^-$

$s(2s + 0.1)$ (0.1 from NaF)

$K_{sp} = s(2s + 0.1) = 2s^2 + 0.1s$ 0. $1s >>> 2s^2$ thus $K_{sp} = 0.1s$

$s = K_{sp} / 0.1 = \frac{5.3\times10^{-11}}{0.1} = 5.3\times10^{-10}\,mol/L$

$Al(OH)_3 \rightleftharpoons Al^{3+} + 3OH^-$

(35.) (c) s 3s

$K_{sp} = s\times(3s)^3 = 27s^4$

$27s^4 = 2.7\times10^{-11} \Rightarrow s^4 = 10^{-12}$

$s = 10^{-3}$

$HCOOH + H_2O \rightleftharpoons H_3O^+ + HCOO^-$

(36.) (d)

Initial	0.01		0	0
At equilibrium	0.01 – x		x	x

$K_a = \frac{X^2}{(0.01 - x)}$

$x << 0.01$ thus $0.01 - x \approx 0.01$

$\frac{X^2}{0.01} = 1.8\times10^{-4} \Rightarrow X^2 = 1.8\times10^{-6}$

$x = 1.34\times10^{-3}$

$pH = -\log H^+ = -\log(1.34\times10^{-3}) = 3 - \log 1.34 = 3\text{-}0.127 = 2.873$

(38.) (c) According to Lewis concept, the molecules or compound which can accept an electron act as acids; thus $AlCl_3$ due to lack of electrons in valence shell (6) can accept electron and act as an acid.

(40.) (c) $BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO^{2-}(aq)$

s s

$K_{sp} = s^2$

$Ag_2CrO_4 \rightleftharpoons 2Ag^+ + CrO_4^{2-}$

2s s

$K_{sp} = (2s)^2 s = 4s^3$

$Al(OH)_3 \rightleftharpoons Al^{3+} + 3OH^-$

s 3s

$K_{sp} = s \times (3s)^3 = 27s^4$

$Zr_3(PO_4)_4 \rightleftharpoons 3Zr^{4+} + 4PO_4^{3-}$

$3s \quad 4s$

$K_{sp} = (3s)^3 (4s)^4 = 6912s^7$

(42.) (d) All the given statements are correct.

(43.) (d) Increasing order of acidic strength is $HI > HBr > HCl > HF$ as bond strength of $(H\text{-}A)$ bond is a more important factor in determining the acidity of an acid.

(44.) (b) $C_5H_5N + H_2O \rightleftharpoons C_5H_5NH^+ + OH$

$$\alpha = \sqrt{\frac{K_b}{C}} = \sqrt{\frac{17 \times 10^{-9}}{01}} = 1.30 \times 10^{-4}$$

Percentage of ionization $= 1.3 \times 10^{-4} \times 100 = 0.013$

(45.) (c) Bronsted base is one that accepts H^+ion , so here PO_4^{3-} and $C_2O_4^{2-}$ are Bronsted bases.

(46.) (a) Boiling point of a liquid increases with increase in atmospheric pressure. Boiling point also depends on the altitude of the place; at high altitude boiling point decreases.

(47.) (b) As the number of gaseous molecules decreases in the forward direction, increasing pressure will shift the equilibrium in the forward direction.

(48.) (d)

(49.) (c) Acidic buffer is a mixture of weak acid and its sal[with strong base. But $HClO_4$ is a strong acid.

(50.) (d) pH = 1 i.*e.* $H^+ = 10^{-1}$

Por acid, base mixing: $N_3V_3 = N_1V_1 - N_2V_2$

For I $N_3V_3 = 60 \times \frac{1}{10} - 40 \times \frac{1}{10} \Rightarrow N_3 \times 100 = 2 \Rightarrow N_3 = 2 \times 10^{-2}$

For II $N_3V_3 = 55 \times \frac{1}{10} - 45 \times \frac{1}{10} \Rightarrow N_3 \times 100 = 1 \Rightarrow N_3 \times 10^{-2}$

For III $N_3V_3 = 75 \times \frac{1}{5} - 25 \times \frac{1}{5} \Rightarrow N_3 \times 100 = 10 \Rightarrow N_3 = 10^{-1}$

For IV $N_3V_3 = 100 \times \frac{1}{10} - 100 \times \frac{1}{10} \Rightarrow N_3V_3 = 0$

Means neutral solution pH = 7

TOPIC WISE PRACTICE QUESTIONS – SOLUTIONS

1. (4) Dissolution of salt lowers the V.P. It is also effected by temperature.
2. (4) In lime kiln CO_2 escapes continuously so reaction proceeds in forward direction.
3. (2) Equilibrium can be achieved only in closed system.
4. (3)
5. (3) Triple point of a substance is the temperature and pressure at which the three phases (gas, liquid and solid) of that substance coexist in thermodynamic equilibrium.
6. 2) $\underset{a(1-x)}{PCl_5} \rightleftharpoons \underset{ax}{PCl_3} + \underset{ax}{Cl_2}$

a = 2, x = 0.4, V = 2L

$$\therefore [PCl_5] = \frac{2(1-0.4)}{2} = 0.6 \text{molL}^{-1}$$

$$[PCl_3] = [Cl_2] = \frac{2 \times 0.4}{2} = 0.4 \text{molL}^{-1}$$

$\therefore K_c = \frac{0.4 \times 0.4}{0.6} = 0.267$

7. 2)For a gaseous phase reaction K_p and K_c are related as $K_p = K_c (RT)^{\Delta n_g}$ For the given reaction

$CO(g) + \frac{1}{2} O_2(g) \rightarrow CO_2(g)$

$\Delta n_g = 1 - (1 + 0.5) = -0.5 \text{or} -\frac{1}{2}$

$\therefore K_p = K_c (RT)^{-\frac{1}{2}} \text{ or } \frac{K_p}{K_c} = (RT)^{-\frac{1}{2}}$

8. 3) $K' = \frac{1}{K}; \therefore K' = \frac{1}{2.4 \times 10^{-3}} = 4.2 \times 10^2$

9. 2) $N_2(g) + 3H_2(g) \square \quad 2NH_3(g)$

$\Delta_n = n_{\text{products}} - n_{\text{readants}} = 2 - 4 = -2$

$\therefore K_p = K_c (RT)^{-2} \text{ or } K_p = \frac{K_c}{(RT)^2}$

Thus $K_p < K_c$

10. 2)

11. 2) $P_{NH_3} = P_{H_2S} = 2\text{atm}$

$\therefore K_p = P_{NH_3} \times PH_{2S} = 4$

12. 1) A reaction is said to be in equilibrium when rate of forward reaction is equal to the rate of backward reaction.

13. 2) $K_p = (P_{H_2O})^2$

P_{H_2O} (at equilibrium) = $\sqrt{K_p} = 10^{-2}$

Forward reaction will occur if value of P_{H_2O} is less than 10^{-2} atm.

14. 4)For the reaction ; $2NO_2(g) \square \quad 2NO(g) + O_2(g)$

Given $K_c = 1.8 \times 10^{-6} \text{at} 184^0 C$

$R = 0.0831 KJ / \text{mol K}$

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

$K_p = 1.8 \times 10^{-6} \times 0.0831 \times 457 = 6.836 \times 10^{-5}$

$\left[\because 184^0 C = (273 + 184) = 457K, \Delta n = (2 + 1, -1) = 1\right]$

Hence it is clear that $K_p > K_c$

15. 3) $N_2 + 3H_2 \square \quad 2NH_3$

$\therefore K = [NH_3]^2 / [N_2][H_2]^3 \ldots\ldots\ldots\ldots\ldots\ldots\text{(i)}$

$\frac{1}{2} N_2 + \frac{3}{2} H_2 \square \quad NH_3$

$\therefore K' = \frac{[NH_3]}{[N_2]^{1/2}[H_2]^{3/2}} \ldots\ldots\ldots\ldots\ldots\text{(ii)}$

Dividing equation (i) by equation (ii), we get

$K' = \sqrt{K}$

16. 1) $K_p = \frac{P_{CH_3OH}}{P_{CO} \times p_{H_2}^2}; K_p = \frac{2}{1 \times (0.1)^2} = 200 \text{ atm}^{-2}$

For reverse reaction, $K_p = \frac{1}{200} = 5 \times 10^{-3} \text{atm}^{-2}$

17. 1) For $A + B \rightleftharpoons C$, $\Delta n = 1 - 2 = -1$

Unit of $K_c = \left[\frac{\text{mol}}{\text{litre}}\right]^{\Delta n} = \left[\frac{\text{mol}}{\text{litre}}\right]^{-1} = \text{litre mol}^{-1}$

18. 3)

19. 3)

20. 3) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

$K_p = K_c (RT)^{\Delta n}$; $\Delta n = 2 - 2 = 0; \therefore K_p = K_c$

21. 4) Rate constant of forward reaction (K_f) = 1.1×10^{-2} and rate constant of backward reaction (K_b) = 1.5×10^{-3} per minute.

Equilibrium constant (K_c) = $\frac{K_f}{K_b} = \frac{1.1 \times 10^{-2}}{1.5 \times 10^{-3}} = 7.33$

22. (1) Equilibrium can be attained by either side of the reactions of equilibrium.

23. 2) $3Fe(s) + 4H_2O(\text{steam}) \rightleftharpoons Fe_3O_4(s) + 4H_2(g)$

$K_p = \frac{(p_{H2})^4}{(p_{H_2O})^4}$ only gaseous products and reactants

24. 3)

25. 4) $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

Since the reaction does not involve any change in number of moles, the equilibrium state is not affected by change of pressure.

26. 3) Catalyst only changes time to achieve equilibrium, it does not affect the value of equilibrium constant.

27. 4)

28. 4) Solid $\rightleftharpoons$ Liquid

It is an endothermic process. So when temperature is raised, more liquid is formed. Hence adding heat will shift the equilbrium in the forward direction.

29. 2) $Q_C = \frac{[SO_3(g)]^2}{[SO_2(g)]^2 [O_2(g)]^1} = 1000 > K_c$ i.e., reaction moves backward also if volume increases

pressure decreases $\Rightarrow Q_c \uparrow$

30. 4)Moles of CO_2 present at equilibrium = $= \frac{1.642 \times 50}{0.0821 \times 1000} = 1$

Mole % of XCO_3decomposed = $\frac{1}{4} \times 100 = 25\%$

Hence, 75% remains undecomposed

31. 4) According to Le-chatelier's principle whenever a constraint is applied to a system in equilibrium, the system tends to readjust so as to nullify the effect of the constraint.

32. 1) $\Delta G^0 = -2.303\ RT \log K = -2.303 \times 2 \times 298 \log 10^{-8} = -2.303 \times 2 \times 298 \times (-8) \text{cal} = 10980 \text{cal} = 10.98 \text{kcal}$

33. 4) $P_{NH_3} = P_{H_2S} = \frac{P}{2} \text{atm}$

$$K_p = P_{NH_3} P_{H_2S} = \left(\frac{P}{2}\right)^2 = \frac{P^2}{4}$$

$$\Delta G = -RT \ln K_p = -RT \ln \left(\frac{p}{2}\right)^2 = -2RT\left[\ln p - \ln 2\right]$$

34. 3) Strong base has higher tendency to accept the proton. Increasing order of base and hence the order of accepting tendency of proton is $I^- < HS^- < NH_3 < RNH_2$

35. 2) $B(OH)_3$ does not provide H^+ ions in water instead it accepts OH^- ion and hence it is Lewis acid.

$$B(OH)_3 + H_2O \square \left[B(OH)_4\right]^- + H^+$$

36. 4) $(CH_3)_3 B-$ is an electron deficient, thus behave as a Lewis acid.

37. 3) HSO_4^- accepts a proton to form H_2SO_4.

Thus H_2SO_4 sis the conjugate acid of HSO_4^-

$$\underset{\text{base}}{HSO_4^-} \xrightarrow{+H^+} \underset{\text{conjugate acid of } HSO_4^-}{H_2SO_4}$$

38. 4) Among M/4 KOH, M/4 NaOH, M/4 NH_4OH and M/4 $Ca(OH)_2$, $Ca(OH)_2$ furnishes highest number of OH^- ions $\left(\because Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-\right)$ so pH of $M/4\,Ca(OH)_2$ is highest.

39. 3)on dilution $\left[H^+\right] = 10^{-6} M = 10^{-6}$ mol

Now dissociation of water cannot be neglected, total $\left[H^+\right] = 10^{-6} + 10^{-7} = 11 \times 10^{-7}$

$$pH = -\log\left[H^+\right] = -\log\left(11 \times 10^{-7}\right) = 5.98$$

40. 1) Acid and base will neutralise each other. Since volume of NaOH is more (150 mL) in solution.

$$\therefore \left[OH^-\right] = \frac{M_2V_2 - M_1V_1}{V_1 + V_2 + V_3} = \frac{150 \times 0.5 - 50 \times 0.5}{50 + 150 + 300} = \frac{1}{10} \text{ Now}$$

$$\left[OH^+\right]\left[OH^-\right] = 10^{-14} \Rightarrow \left[H^+\right] = 10^{-13}; pH = 13$$

41. 1) Bronsted base is a substance which accepts proton. In option 1), H_2O is accepting proton, i.e., acting as a base.

42. (2) K_w increases with temperature.

43. 1)

(i) $\underset{\text{acid}_1}{H_3PO_4} + \underset{\text{base}_2}{H_2O} \longrightarrow \underset{\text{acid}_2}{H_3O^+} + \underset{\text{base}_1}{H_2PO_4^-}$

(ii) $\underset{\text{acid}_1}{H_2PO_4^-} + \underset{\text{base}_2}{H_2O} \longrightarrow \underset{\text{base}_1}{HPO_4^{2-}} + \underset{\text{acid}_2}{H_3O^+}$

(iii) $\underset{\text{base}_1}{H_2PO_4^-} + \underset{\text{acid}_2}{OH^-} \longrightarrow \underset{\text{acid}_1}{H_3PO_4} + \underset{\text{base}_2}{O^{2-}}$

Hence only in (ii) reaction $H_2PO_4^-$ is acting as an acid.

44. (2) The value of ionic product of water changes with the temperature.

45. (4) All the salts undergo hydrolysis in aqueous solution. The acids formed have their strength in the order: HCl > HCOOH > CH_3COOH
The strength of the bases formed follow the order : NaOH > NH_4OH.

46. 2) On increasing T, Kw increases and so also $\left[H_3O^+\right]$ and $\left[OH^-\right]$

47. 2) $\frac{\alpha_1}{\alpha_2} = \sqrt{\frac{K_{al}}{K_{a_2}}} = \sqrt{\frac{3.14 \times 10^{-4}}{1.96 \times 10^{-5}}} = 4:1$

48. 1) $K = c\alpha^2 = 0.1\times\left(\frac{1.34}{100}\right)^2 = 1.8\times10^{-5}$

49. 1)

50. 2) $HA \rightarrow H^+ + A^-$

$K_a = \frac{[H^+][A^-]}{[HA]}, \because [H^+] = 10^{-pH} \therefore [H^+] = 10^{-5}$; and at equilibrium $[H^+] = [A^-]$

$\therefore K_a = \frac{10^{-5}\times10^{-5}}{0.0015} = 2\times10^{-8}$

$\alpha = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{2\times10^{-8}}{005}} = \sqrt{4\times10^{-6}} = 2\times10^{-3}$

Percentage ionization = 0.2

51. 3) Ostwald's dilution law, dilution increases ionisation.

52. 2) $[H^+] = \sqrt{C\times K_a} = \sqrt{0.001\times1.8\times10^{-4}}$ for formic acid

53. 3)

54. 4)

55. 2)Higher the K_{sp}, more soluble is that compound in H_2O.

$\therefore K_{sp}$ of $FeS\left(11\times10^{-20}\right)$ is highest

So it is more soluble and has maximum solubility in H_2O

56. 2) $Cr(OH_3)(s) \square \underset{s}{Cr^{3+}}(aq) + \underset{3s}{3OH^-}(aq)$

$(s)(3s)^3 = K_{sp}$

$27s^4 = K_{sp}$

$s = \left(\frac{K_{sp}}{27}\right)^{1/4} = \left(\frac{1.6\times10^{-30}}{27}\right)^{1/4}$

57. 3) $2HNO_3(aq) + \left[Ag(NH_3)_2\right]^+ + Cl^- \rightarrow AgCl(s)\downarrow + 2NH_4^+ + 2NO_3^-$

58. 1) $Ag_2SO_4 \rightarrow \underset{2x}{2Ag^+} + \underset{x\ 4}{SO^{2-}}$ (moles after diss)

$K_{sp} = [Ag^+]^2[SO_4^{2-}] = (2x)^2\times x = 4x^3$

Given solubility = $x = 2.5\times10^{-2}$ mole/litre

$\therefore\ K_{sp} = 4\times\left(2.5\times10^{-2}\right)^3 = 4\times15.6\times10^{-6} = 62.5\times10^{-6}$

59. (3) Solubility of weak electrolyte decreases in solvent having common ion. So solubility of AgI in NaI solution is less than in pure water because of common ion effect.

60. (2) The dissociation of a weak electrolyte (AgCl) is suppressed on adding the strong electrolyte having a common ion (Cl^-), hence concentration of Ag^+ ion will be less.

61. 4) $BA_2 \rightarrow \underset{x}{B} + \underset{2x}{2A}$

Solubility product = $[x][2x]^2 = 4x^3$

$4\times10^{-12} = 4x^3$ or $x = \sqrt[3]{\frac{4\times10^{-12}}{4}}$

$\therefore x = 10^{-4}$

62. 4) pH of acidic buffer = $-\log K_a + \log\left[\frac{salt}{acid}\right]$

$\therefore$ pH of acidic buffer depends upon value of K_a

63. 3) At 80% neutralization, $\left[\frac{\text{salt}}{\text{base}}\right]=\frac{8}{2}=\frac{4}{1}$

$\therefore \text{pOH}=\text{pK}_b+\log\frac{[\text{salt}]}{[\text{base}]}$

14-7.4 = $\text{pK}_b+\log 4$; $\text{pK}_b=6.6-0.6=;\text{K}_b=10^{-6}$

64. 4) pH or $\left[\text{H}^+\right]$ of a buffer does not change with dilution.

65. 1) The pOH of a buffer consisting of NH_3 (i.e. NH_4OH) and salt NH_4Cl (salt) is given by the equation

$\text{pOH}=\text{pK}_b+\log\square=5.0+\log\frac{[0.1]}{[1.0]}=5.0-\log 10=5-1=4.0$

$\text{pH}+\text{pOH}=14$; $\text{pH}=14-4\rightarrow 10$

66. 2) Solubility product is the product of ionic concentration in a saturated solution of an electrolyte at a given temperature.

67. 2) $\text{K}_{sp}=\left[\text{Fe}^{3+}\right].\left[3\text{OH}^-\right]$

So molar solubility of $\text{Fe}^{3+}=\text{s}$ and $\left[3\text{OH}^-\right]=3\text{s}$

$\text{Fe}(\text{OH})_3\square\ \underset{[s]}{\text{Fe}^{3+}}+\underset{[3s]}{3\text{OH}^-}$

$1.0\times10^{-38}=[\text{s}][3\text{s}]^3=\text{s}^4\times27$; $\text{s}^4=\frac{1.0\times10^{-38}}{27}=3.703\times10^{-40}$; $\text{s}=\left(3.703\times10^{-40}\right)^{1/4}=1.386\times10^{-10}$

68. 1) The solubility decreases by common ion effect

$\text{A};\left[\text{Cl}^-\right]=0,\text{B};\left[\text{Cl}^-\right]=0.1$

$\text{C};\left[\text{Cl}^-\right]=0.2,\text{D};\left[\text{Cl}^-\right]=0$

[soluble compelx is formed $[Ag(NH_3)_2Cl]$

$\underset{s}{\text{AB}}\square\ \underset{s}{\text{A}^+}+\underset{s}{\text{B}^-}$

Hence, solubility product of AB

$\text{K}_{sp}=\left[\text{A}^+\right]\left[\text{B}^-\right]$; $\text{S}=[\text{s}][\text{s}]\Rightarrow\text{s}=\text{S}^{1/2}$

70. 1) $\text{K}_{sp}=\left[\text{M}^{+2}\right]\left[\text{OH}^-\right]^2$; $\left[\text{OH}^-\right]=\frac{10}{100}\times0.1=0.01$; $[\text{M}^{+2}]=\frac{\text{K}_{sp}}{\left[\text{OH}^-\right]^2}=\frac{10^{-14}}{[0.01]^2}=10^{-10}$

NEET PREVIOUS YEARS QUESTIONS-EXPLANATIONS

1. 1) $\text{A}_2(\text{g})+\text{B}_2(\text{g})\square\ \text{X}_2(\text{g});\Delta\text{H}=-\text{xkJ}$

On increasing pressure equilibrium shifts in a direction where number of moles decreases i.e. forward direction. On decreasing temperature, equilibrium shifts in exothermic direction i.e., forward direction. So, high pressure and low temperature favours maximum formation of product.

2. 1)Solubility of $\text{BaSO}_4=2.42\times10^{-3}\text{gL}^{-1}$

$\therefore\text{s}=\frac{2.42\times10^{-3}}{233}=1.038\times10^{-5}\text{mol L}^{-1}$

$\text{K}_{sp}=\text{s}^2=\left(1.038\times10^{-5}\right)^2=1.08\times10^{-10}\text{mol}^2\text{L}^{-2}$

3. 3)Meq. Of HCI = $75\times\frac{1}{5}\times1=15$

Meq. of NaOH = $25\times\frac{1}{5}\times1=5$

Meq. of HCl in resulting solution = 10

Molarity of $[H^+]$ in resulting mixture = $\frac{10}{100} = \frac{1}{10}$

$pH = -\log\left[H^+\right] = -\log\left[\frac{1}{10}\right] = 1.0$

4. 1)

(i) $N_2 + 3H_2 \rightleftharpoons 2NH_3; K_1 = \frac{[NH_3]^2}{[N_2][H_2]^3}$

(ii) $N_2 + O_2 \rightleftharpoons 2NO; K_2 = \frac{[NO]^2}{[N_2][O_2]}$

(iii) $H_2 + \frac{1}{2}O_2 \longrightarrow H_2O; K_3 = \frac{[H_2O]}{[H_2][O_2]^{1/2}}$

Applying $(II + 3 \times III - I)$ we will get

$2NH_3 + \frac{5}{2}O_2 \overset{K}{\rightleftharpoons} 2NO + 3H_2O;$

$K = \frac{[NO]^2}{[N_2][O_2]} \times \frac{[H_2O]^3}{[H_2]^3 \times [O_2]^{3/2}} \Big/ \frac{[NH_3]^2}{[N_2][H_2]^3}$

5. 3) $Ag_2C_2O_4(s) \rightleftharpoons 2Ag^+(aq) + C_2O_4^{2-}(aq)$

$K_{sp} = \left[Ag^+\right]^2 [C_2O_4^{2-}]$

$\left[Ag^+\right] = 2.2 \times 10^{-4} M$

Given that

$\therefore$ Concentration of $C_2O_4^{2-}$ ions

$[C_2O_4^{2-}] = \frac{2.2 \times 10^{-4}}{2} M = 1.1 \times 10^{-4} M$

$\therefore K_{sp} = \left(2.2 \times 10^{-4}\right)^2 \left(1.1 \times 10^{-4}\right) = 5.324 \times 10^{-12}$

$\therefore K = K_2 \times K_3^3 / K_1$

6. (1) A catalyst speeds up both forward and backward reaction with the same rate. So, equilibrium constant is not affected by the presence of a catalyst at any given temperature.

7. (4) Max. pressure of CO_2 = Pressure of CO_2 at equilibrium For reaction, $SrCO_3(s) \rightleftharpoons SrO(s) + CO_2$

$K_p = P_{CO_2} = 1.6$ atm = maximum pressure of CO_2

Volume of container at this stage

$V = nRT/P$ .....(i)

Since container is sealed and reaction was not earlier at equilibrium

$\therefore n$ = constant

$n = \frac{PV}{RT} = \frac{0.4 \times 20}{RT}$ .....(ii)

Put equation (ii) in equation (i)

$V = \left[\frac{0.4 \times 20}{RT}\right] \frac{RT}{1.6} = 5L$

8. 2)

$MY \rightleftharpoons M^+ + Y^-$

$K_{SP} = s^2 = 6.2 \times 10^{-13}$

$s = \sqrt{6.2 \times 10^{-13}}$

$s = 7.87 \times 10^{-7} \text{ mol } L^{-1}$

$NY_3 \rightleftharpoons N^{3+} + 3Y^-$

$K_{SP} = s \times (3s)^3 = 27s^4 = 6.2 \times 10^{-13}$

$s = \left(\frac{6.2 \times 10^{-13}}{27} \right)^{1/4}$

$s = 3.89 \times 10^{-4} \text{ mol } L^{-1}$

$\therefore$ molar solubility of NY_3 is more than MY in water.

9. 2) The presence of large amount of $KHSO_4$ will decrease ionisation of H_2SO_4 that result in lesser ionisation of nitric acid and lesser formation of nitronium [NO_2^+]. Hence the rate of nitration will be slower.

10. 3) Ag_2CrO_4

$K_{sp} = \left[Ag^+ \right]^2 \left[Cr_2O_4^{-2} \right] = 1.1 \times 10^{-12}$

$\left[Ag^+ \right] = \sqrt{\frac{1.1 \times 10^{-12}}{[Cr_2O_4^{-2}}}$

AgCl

$K_{sp} = \left[Ag^+ \right] \left[Cl^- \right] = 1.8 \times 10^{-10}$

$\left[Ag^+ \right] = \frac{1.8 \times 10^{-10}}{\left[Cl^- \right]}$

AgBr

$K_{sp} = \left[Ag^+ \right] \left[Br^- \right] = 5.0 \times 10^{-13}$

$\left[Ag^+ \right] = \frac{5.3 \times 10^{-13}}{\left[Br^- \right]}$

AgI

$K_{sp} = \left[Ag^+ \right] \left[I^- \right] = 8.3 \times 10^{-17}$

$\left[Ag^+ \right] = \frac{8.3 \times 10^{-17}}{\left[I^- \right]}$

If we take $\left[Cr_2O_4^{-2} \right] = \left[Cl^- \right] = \left[Br^- \right] = \left[I^- \right] = 1$ then maximum $\left[Ag^+ \right]$ will be required in case of Ag_2 Cro_4

11. 2) Equilibrium constant for reaction:

$K = 1.6 \times 10^{12} = \frac{[\text{Product}]}{[\text{Reactant}]}$ The value of K is very high so the system will contain mostly products at equilibrium

12. 2) $\Delta G^0 = -2.30RT \log K$

Because at equilibrium $\Delta G = 0$

13. 3) Na_2CO_3 is a salt of strong base (NaOH) and weak acid (H_2CO_3). On hydrolysis this salt will produce strongly basic solution. i.e. pH will be highest (pH > 7) for this sotluion. Others are combination of

KCl = Strong acid + Strong base $\rightarrow$ neutral solution ($pH \approx 7$)
$NaCl$ = Strong acid + Strong base $\rightarrow$ neutral solution ($pH \approx 7$)
$CuSO_4$ = Strong acid + weak base $\rightarrow$ Acidic solution ($pH \approx 7$)

14. 2) $\Delta G = -2.303 \log K$ here

$$K = \left[Ag^+\right]^2 \left[CO_3^{--}\right] = K_{sp}$$

$$\therefore 63.3 \times 10^{-3} = -2.303 \times 8.314 \times 298 \log K_{sp}$$

$$\therefore \log K_{sp} = -\frac{63.3 \times 10^{-3}}{5705.8} = -11.09$$

$$\therefore K_{sp} = \text{Anti}\log(-11.09) = 8 \times 10^{-12}$$

15. (4) Given reaction is exothermic reaction. Hence according to Le-Chatelier's principle low temperature favours the forward reaction and on increasing pressure equilibrium will shift, towards lesser number of moles i.e. forward direction.

16. (1) In exothermic reactions on increasing temperature value of K_p decreases
$\therefore K_p > K_p^{'}$ (Assuming $T_1 < T_2$)

17.

$$Ca(OH)_2(s) \rightleftharpoons Ca^{+2}(aq) + 2OH^-(aq)$$
$$\quad\quad\quad\quad\quad\quad S \quad\quad\quad 2S$$

$pH = 9$; $pOH = 5$; $[OH^-] = 10^{-5} = 2S$

$$S = \frac{10^{-5}}{2}$$

$$K_{sp} = [Ca^{+2}]\,[OH^-]^2$$
$$K_{sp} = S \times (2S)^2$$
$$K_{sp} = 4S^3$$
$$K_{sp} = 4 \times \left(\frac{10^{-5}}{2}\right)^3$$
$$K_{sp} = 0.5 \times 10^{-15}$$

18. Conjugate base of H_2O is OH^-
Conjugate base of HF is F^-

19.. Basic buffer is mixture of weak base and salt of weak base with strong acid
milli mole of HCl = 100 × 0.1 = 10 milli mole
milli mole of NH_4OH = 200 × 0.1 = 20 milli mole

$$HCl + NH_4OH \rightarrow NH_4Xl + H_2O$$

10 20 – –

– 10 10

20.

$NaOH(aq)$ is strong base solution

So, $[OH^-] = N = 10^{-2}N$

$$pOH = -\log[OH^-] = -\log 10^{-2} = 2$$

$$pH = 14 - pOH = 14 - 2$$

$$pH = 12$$

21.

HCl cannot act both as Bronsted acid and Bronsted base because HCl can only donate proton.

22

$CaF_2(s) \rightleftharpoons Ca^{+2}(aq) + 2F^-(aq)$

(a –s')	s'	2s'

$NaF(aq) \rightarrow Na^+(aq) + F^-(aq)$

C	0	0
0	C	C

In solution- $[F^-] = (2s' + C)$

$[F^-] \approx C$ (due to common ion effect)

$K_{sp(CaF_2)} = [Ca^{+2}].[F^-]^2$

$K_{sp(CaF_2)} = s'.C^2$

$$s' = \frac{5.3\times10^{-11}}{(10^{-1})^2}$$

$s' = 5.3 \times 10^{-9}$ mol L^{-1}

23. $CH_3COONa \Rightarrow$ Salf of $CH_3COOH(WA) + NaOH(SB)$

Solution of CH_3COONa shows basic nature

24. K_{sp} for $AB = s^2 = 4\times10^{-8}$

$\therefore$ Molarity of solution = solubility = $\sqrt{K_{sp}}$

$= \sqrt{4\times10^{-8}} = 2\times10^{-4}$ mol / L

25. $\Delta G^0 = -2.303RT\log_{10}K = -RT\ln k = -8.314J.mol^{-1}k^{-1}\times300k\times\ln\left(2\times10^{13}\right)$

26. 2)

$$Ni(OH)_2 \rightleftharpoons N_1^{2+} + 2OH^-$$

$S \qquad S \qquad 2S$

$(2S + 0.1)$

in presence NaOH of 0.1M

$K_{sp} = \left[Ni^{2+}\right]\left[OH^-\right]^2$

$2\times10^{-15} = (S)(2S + 0.1)^2$

$2\times10^{-15} = (S)(0.1)^2 \quad (\because 2S << 0.1)$

$\therefore S = 2\times10^{-13}$

27. $P^H = 7 + \frac{1}{2}\left(P^{ka} - P^{kb}\right)$;

$P^H = 7 + \frac{1}{2}(4.77 - 3.27) = 7 + \frac{1}{2}(1.5) = 7.75$

28 $3O_2 \rightleftharpoons 2O_3$

$K_c = 3.0\times10^{-5.9}$

$[O_2] = 0.040M$

$[O_3] = ?$

$$3.0\times10^{-59} = \frac{[O_3]^2}{[O_2]^3}$$

$[O_3]^2 = 3.0\times10^{-59}\times0.040\times0.040\times0.040 = 0.000192$

$=1.92\times10^{-4} = \sqrt{1.92\times10^{-4}\times10^{-59}} = \sqrt{1.92\times10^{-63}} = \sqrt{19.2\times10^{-64}}$

4.38×10^{-32}

29 $P^H = P^{ka}$ + log (N.V) salt/(NV) acid = 4.57 + log(50x0.10/50x0.01) = 5.57