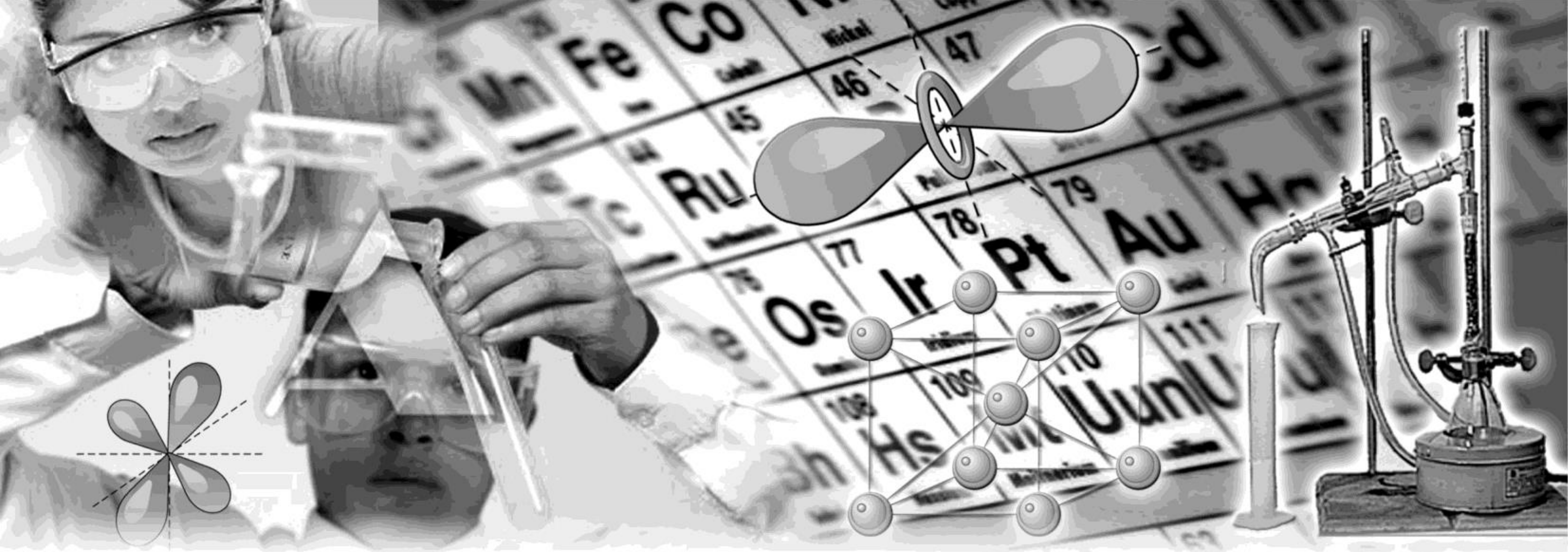


Chapter
9
Ionic Equilibrium

In chemical equilibrium we studied reaction involving molecules only but in ionic equilibrium we will study reversible reactions involving formation of ions in water. When solute is polar covalent compound then it reacts with water to form ions.

Electrical conductors

Substances, which allow electric current to pass through them, are known as conductors or electrical conductors. Conductors can be divided into two types,

(1) Conductors which conduct electricity without undergoing any chemical change are known as metallic or electronic conductors.

(2) Conductors which undergo decomposition (a chemical change) when an electric current is passed through them are known as electrolytic conductors or electrolytes.

Electrolytes are further divided into two types on the basis of their strengths,

(i) Substances which almost completely ionize into ions in their aqueous solution are called strong electrolytes. *Degree of ionization for this type of electrolyte is one i.e.,* $\alpha \approx 1$.

For example : $HCl, H_2SO_4, NaCl, HNO_3, KOH, \ NaOH,$ $HNO_3, AgNO_3, CuSO_4$ etc. means all strong acids, bases and all types of salts.

(ii) Substances which ionize to a small extent in their aqueous solution are known as weak electrolytes. *Degree of ionization for this types of electrolytes is* $\alpha <<< 1$.

For example : $H_2O, CH_3COOH, NH_4OH, \ HCN,$ $Liq. SO_2$, $HCOOH$ etc. means all weak acids and bases.

Arrhenius theory of electrolytic dissociation

(1) **Postulates of Arrhenius theory**

(i) In aqueous solution, the molecules of an electrolyte undergo spontaneous dissociation to form positive and negative ions.

(ii) Degree of ionization (α)

$$= \frac{\text{Number of dissociated molecules}}{\text{Total number of molecules of electrolyte before dissociation}}$$

(iii) At moderate concentrations, there exists an equilibrium between the ions and undissociated molecules, such as, $NaOH \rightleftharpoons Na^+ + OH^-$; $KCl \rightleftharpoons K^+ + Cl^-$

This equilibrium state is called *ionic equilibrium.*

(iv) Each ion behaves osmotically as a molecule.

(2) **Factors affecting degree of ionisation**

(i) At normal dilution, value of α is nearly 1 for strong electrolytes, while it is very less than 1 for weak electrolytes.

(ii) *Higher the dielectric constant of a solvent more is its ionising power.* Water is the most powerful ionising solvent as its dielectric constant is highest.

(iii) $\alpha \propto \frac{1}{\text{Con. of solution}} \propto \frac{1}{\text{wt. of solution}}$

$\propto$ Dilution of solution $\propto$ Amount of solvent

(iv) Degree of ionisation of an electrolyte in solution increases with rise in temperature.

(v) ***Presence of common ion*** **:** The degree of ionisation of an electrolyte decreases in the presence of a strong electrolyte having a common ion.

Ostwald's dilution law

The strength of an acid or a bas is experimentally measured by determining its dissociation or ionisation constant.

When acetic acid (a weak electrolyte) is dissolved in water, it dissociates partially into H^+ or H_3O^+ and CH_3COO^- ions and the following equilibrium is obtained,

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

Applying law of chemical equilibrium,

$$K = \frac{[CH_3COO^-] \times [H_3O^+]}{[CH_3COOH] \times [H_2O]}$$

In dilute solution, $[H_2O]$ is constant. The product of K and constant $[H_2O]$ is denoted as K_a, *the ionization constant or dissociation constant of the acid* is,

$$K_a = \frac{[CH_3COO^-] \times [H_3O^+]}{[CH_3COOH]} \qquad .....(i)$$

The fraction of total number of molecules of an electrolyte which ionise into ions is known as degree of dissociation/ionisation α.

If 'C' represents the initial concentration of the acid in moles L^{-1} and α the degree of dissociation, then equilibrium concentration of the ions (CH_3COO^- and H_3O^+) is equal to $C\alpha$ and that of the undissociated acetic acid $= C(1-\alpha)$ *i.e.,* we have

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

Initial conc	C	0	0
Conc. at eqb.	$C(1-\alpha)$	$C\alpha$	$C\alpha$

Substituting the values of the equilibrium concentrations in equation (i), we get

$$K_a = \frac{C\alpha . C\alpha}{C(1-\alpha)} = \frac{C^2\alpha^2}{C(1-\alpha)} = \frac{C\alpha^2}{1-\alpha} \qquad .....(ii)$$

In case of weak electrolytes, the value of α is very small and can be neglected in comparison to 1 *i.e.,* $1-\alpha = 1$.

Hence, we get

$$K_a = C\alpha^2 \text{ or } \alpha = \sqrt{\frac{K_a}{C}} \qquad .....(iii)$$

The degree of dissociation, α can therefore be calcualted at a given concentration, C if K_a is known. Furher, if V is the volume of the solution in litres containing 1 mole of the electrolyte, $C = 1/V$. Hence we have

$$\alpha = \sqrt{K_a V} \qquad .....(iv)$$

Similarly, for a weak base like NH_4OH, we have

$$\alpha = \sqrt{K_b / C} = \sqrt{K_b V} \qquad .....(v)$$

The above equations lead to the following result

"For a weak electrolyte, the degree of ionisation is inversely proportional to the square root of molar concentration or directly proportional to the square root of volume containing one mole of the solute."

This is called ***Ostwald's dilution law***.

Dissociation constants of acids and Bases

(1) **Dissociation constant for weak acid :** Consider an acid HA which, when dissolved in water ionizes as,

$$HA \rightleftharpoons H^+ + A^-$$

Applying the law of mass action, $K_a = \frac{[H^+][A^-]}{[HA]}$

Where, K_a is the dissociation constant of the acid, HA. It has constant value at definite temperature and does not change with the change of concentration.

Dissociation Constant for polybasic acid : Polybasic acids ionise stepwise as, for example, orthophosphoric acid ionises in three steps and each step has its own ionisation constant.

$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$ (I step)

$H_2PO_4^- \rightleftharpoons H^+ + HPO_4^{-2}$ (II step)

$HPO_4^{-2} \rightleftharpoons H^+ + PO_4^{-3}$ (III step)

Let K_1, K_2 and K_3 be the ionization constants of first, second and third steps respectively. Thus,

$$K_1 = \frac{[H^+][H_2PO_4^-]}{[H_3PO_4]} ; K_2 = \frac{[H^+][HPO_4^{-2}]}{[H_2PO_4^-]} ; K_3 = \frac{[H^+][PO_4^{-3}]}{[HPO_4^{-2}]}$$

In general, $K_1 > K_2 > K_3$

The overall dissociation constant (K) is given by the relation,

$$K = K_1 \times K_2 \times K_3$$

(2) **Dissociation constant for weak base :** The equilibrium of NH_4OH (a weak base) can be represented as,

$$NH_4OH \rightleftharpoons NH_4^+ + OH^-$$

Applying the law of mass action, $K_b = \frac{[NH_4^+][OH^-]}{[NH_4OH]}$

K_b is constant at a definite temperature and does not change with the change of concentration.

Common ion effect

The degree of dissociation of an electrolyte (weak) is suppressed by the addition of another electrolyte (strong) containing a common ion, this is termed as *common ion effect*. Acetic acid is a weak electrolyte and its ionisation is suppressed in presence of a strong acid (H^+ ion as common ion) or a strong salt like sodium acetate (acetate ion is a common ion). Similarly, the addition of NH_4Cl or $NaOH$ to NH_4OH solution will suppress the dissociation of NH_4OH due to common ion either NH_4^+ or OH^-.

$$CH_3COOH \rightleftharpoons \boxed{CH_3COO^-} + H^+ \qquad NH_4OH \rightleftharpoons \boxed{NH_4^+} + OH^-$$

$$CH_3COONa \rightarrow \underbrace{\boxed{CH_3COO^-}}_{\text{Common ion}} + Na^+ \qquad NH_4Cl \rightarrow \underbrace{\boxed{NH_4^+}}_{\text{Common ion}} + Cl^-$$

As a result of common ion effect, the concentration of the ion of weak electrolyte which is not common in two electrolytes, is decreased. The use of this phenomenon is made in qualitative analysis to adjust concentration of S^{2-} ions in second group and OH^- ion concentration in third group.

Isohydric solution

If the concentration of the common ions in the solution of two electrolytes , for example H^+ ion concentration in HCl and HNO_3 or OH^- ion concentration in $Ca(OH)_2$ and $Ba(OH)_2$ is same, then on mixing them there is no change in the degree of dissociation of either of the electrolytes. Such solutions are called *isohydric solutions*.

Consider two isohydric solutions of acids HA_1 and HA_2. Let V_1 and V_2 be their dilutions and α_1 and α_2 be their degree of dissociation at the respective dilution. Then,

$$\frac{\alpha_1}{V_1} = \frac{\alpha_2}{V_2}$$

Above equation is useful for calculating the relative dilution of two acids at which they would be isohydric.

Solubility product

In a saturated solution of sparingly soluble electrolyte two equilibria exist and can be represented as, $\underbrace{AB}_{\text{Solid}} \rightleftharpoons$

$$\underset{\substack{\text{Unionised}\\\text{(Dissolved)}}}{AB} \rightleftharpoons \underset{\text{ions}}{A^+ + B^-}$$

Applying the law of mass action, $\frac{[A^+][B^-]}{[AB]} = K$

Since the solution is saturated, the concentration of unionised molecules of the electrolyte is constant at a particular temperature, *i.e.,* $[AB] = K' =$ constant.

Hence, $[A^+][B^-] = K[AB] = KK' = K_{sp}$ (constant)

K_{sp} is termed as the solubility product. It is defined *as the product of the concentration of ions in a saturated solution of an electrolyte at a given temperature.*

Consider, in general, the electrolyte of the type A_xB_y which dissociates as, $A_xB_y \rightleftharpoons xA^{y+} + yB^{x-}$

Applying law of mass action, $\frac{[A^{y+}]^x[B^{x-}]^y}{[A_xB_y]} = K$

When the solution is saturated, $[A_xB_y] = K'$ (constant) or

$[A^{y+}]^x[B^{x-}]^y = K[A_xB_y] = KK' = K_{sp}$ (constant)

Thus, solubility product is defined as the product of concentrations of the ions raised to a power equal to the number of times the ions occur in the equation representing the dissociation of the electrolyte at a given temperature when the solution is saturated.

(1) **Difference between solubility product and ionic product** : Both ionic product and solubility product represent the product of the concentrations of the ions in the solution. The term ionic product has a broad meaning since, it is applicable to all types of solutions, either unsaturated or saturated and varies accordingly.

On the other hand, the term solubility product is applied only to a saturated solution in which there exists a dynamic equilibrium between the undissolved salt and the ions present in solution. Thus the solubility product is in fact the ionic product for a saturated solution at a constant temperature.

(2) **Different expression for solubility products**

(i) Electrolyte of type AB (1 : 1 type salt) *e.g.,* $AgCl, BaSO_4$

$$AgCl \rightleftharpoons \underset{x}{Ag^+} + \underset{x}{Cl^-}$$

$$K_{sp} = [Ag^+][Cl^-] \ ; \ K_{sp} = x^2 \ ; \ x = \sqrt{K_{sp}}$$

(ii) Electrolytes of type AB_2 (1:2 type salt) *e.g.,* $PbCl_2, CaF_2$

$$PbCl_2 \rightleftharpoons \underset{x}{Pb^{2+}} + \underset{2x}{2Cl^-}$$

$$K_{sp} = [Pb^{2+}][Cl^-]^2 \ ; K_{sp} = [x][2x]^2 \ ; K_{sp} = 4x^3$$

$$x = 3\sqrt{K_{sp}/4}$$

(iii) Electrolyte of type A_2B (2 : 1 type salt) *e.g.,* Ag_2CrO_4, H_2S

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

$$K_{sp} = [Ag^+]^2\,[CrO_4^{2-}]\,;\; K_{sp} = [2x]^2\,[x]\,;\; K_{sp} = 4x^3$$

$$x = \sqrt[3]{\frac{K_{sp}}{4}}$$

(iv) Electrolyte of type A_2B_3 (2 : 3 type salt)

e.g., As_2S_3, Sb_2S_3

$$As_2S_3 \rightleftharpoons \underset{2x}{2As^{3+}} + \underset{3x}{3S^{2-}}$$

$$K_{sp} = [As^{3+}]^2[S^{2-}]^3 \;\;;\;\; K_{sp} = [2x]^2[3x]^3\,;\;\; K_{sp} = 4x^2 \times 27x^3$$

$$K_{sp} = 108\,x^5 \;\;;\;\; x = \sqrt[5]{\frac{K_{sp}}{108}}$$

(v) Electrolyte of type AB_3 (1 : 3 type salt)

e.g., $AlCl_3, Fe(OH)_3$

$$AlCl_3 \rightleftharpoons \underset{x}{Al^{+++}} + \underset{3x}{3Cl^-}$$

$$K_{sp} = [Al^{+3}][3Cl^-] \;\;;\;\; K_{sp} = [x]\,[3x]^3$$

$$K_{sp} = 27x^4\,;\;\; x = \sqrt[4]{\frac{K_{sp}}{27}}\;.$$

(3) **Criteria of precipitation of an electrolyte :** When Ionic product of an electrolyte is greater than its solubility product, precipitation occurs.

(4) **Applications of solubility product**

(i) ***In predicting the formation of a precipitate***

Case I : When $K_{ip} < K_{sp}$, then solution is unsaturated in which more solute can be dissolved. *i.e.,* no precipitation.

Case II : When $K_{ip} = K_{sp}$, then solution is saturated in which no more solute can be dissolved but no ppt. is fomed.

Case III : When $K_{ip} > K_{sp}$, then solution is supersaturated and precipitation takes place.

When the ionic product exceeds the solubility product, the equilibrium shifts towards left-hand side, *i.e.,* increasing the concentration of undissociated molecules of the electrolyte. As the solvent can hold a fixed amount of electrolyte at a definite temperature, the excess of the electrolyte is thrown out from the solutions as precipitate.

(ii) *In predicting the solubility of sparingly soluble salts* Knowing the solubility product of a sparingly soluble salt at any given temperature, we can predict its solubility.

(iii) ***Purification of common salt*** **:** HCl gas is circulated through the saturated solution of common salt. HCl and $NaCl$ dissociate into their respective ions as,

$$NaCl \rightleftharpoons Na^+ + Cl^- \,;\; HCl \rightleftharpoons H^+ + Cl^-$$

The concentration of Cl^- ions increases considerably in solution due to ionisation of HCl and due to common ion effect, dissociation of $NaCl$ is decreased. Hence, the ionic product $[Na^+][Cl^-]$ exceeds the solubility product of $NaCl$ and therefore pure $NaCl$ precipitates out from the solution.

(iv) ***Salting out of soap*** **:** From the solution, soap is precipitated by the addition of concentrated solution of $NaCl$.

$$\underset{\text{Soap}}{RCOONa} \rightleftharpoons RCOO^- + Na^+ \,;\; NaCl \rightleftharpoons Na^+ + Cl^-$$

Hence, the ionic product $[RCOO^-]$ $[Na^+]$ exceeds the solubility product of soap and therefore, soap precipitates out from the solution.

(v) ***In qualitative analysis*** **:** The separation and identification of various basic radicals into different groups is based upon solubility product principle and common ion effect.

(a) *Precipitation of group first radicals* (Pb^{+2}, Ag^+ , Hg^{+2}) The group reagent is dilute HCl. $[Ag^+][Cl^-] > K_{sp}$ for $AgCl$.

(b) *Precipitation of group second radicals* (Hg^{+2}, Pb^{+2}, Bi^{+3}, Cu^{+2}, Cd^{+2}, As^{+3}, Sb^{+3} and Sn^{+2}) : The group reagent is H_2S in presence of dilute HCl. $[Pb^{+2}][S^{-2}] > K_{sp}$ for PbS .

(c) *Precipitation of group third radicals* (Fe^{+3}, Al^{+3} and Cr^{+3}) The group reagent is NH_4OH in presence of NH_4Cl .

$$[Fe^{+3}][OH^-]^3 > K_{sp}$$

(d) *Precipitation of group fourth radicals* (Co^{+2}, Ni^{+2}, Mn^{+2} and Zn^{+2}) : The group reagent is H_2S in presence of NH_4OH .

$$[Co^{+2}][S^{-2}] > K_{sp}$$

(e) *Precipitation of group fifth radicals* (Ba^{+2}, Sr^{+2}, Ca^{+2}) The group reagent is ammonium carbonate in presence of NH_4Cl and NH_4OH . $[Ba^{+2}][CO_3^{-2}] > K_{sp}$

(vi) ***Calculation of remaining concentration after precipitation*** : Sometimes an ion remains after precipitation if it is in excess. Remaining concentration can be determined, *Example* : $[A^+]_{left} = \frac{K_{sp}[AB]}{[B^-]}$;

$$[Ca^{2+}]_{left} = \frac{K_{sp}[Ca(OH)_2]}{[OH^-]^2}$$

In general $[A^{n+}]^m_{left} = \frac{K_{sp}[A_m B_n]}{[B^{m-}]^n}$

% precipitation of ion

$$= \left[\frac{\text{Initial conc. - Remaining conc.}}{\text{Initial conc.}}\right] \times 100$$

(vii) ***Calculation of simultaneous solubility*** : Solubility of two electrolytes having common ion; when they are dissolved in the same solution, is called *simultaneous solubility*.

Calculation of simultaneous solubility is divided into two cases.

Case I : When the two electrolytes are almost equally strong (having close solubility product).

e.g., $AgBr\ (K_{sp} = 5 \times 10^{-13})$; $AgSCN\ (K_{sp} = 10^{-12})$

Here, charge balancing concept is applied.

Charge of Ag^+ = Charge of Br^- + Charge of SCN^-

$$[Ag^+] \quad = \quad [Br^-] \quad + \quad [SCN^-]$$

$$(a+b) \quad = \quad a \qquad\qquad b$$

Case II **:** When solubility products of two electrolytes are not close, i.e., they are not equally strong.

e.g., $CaF_2 (K_{sp} = 3.4 \times 10^{-11})$; $SrF_2\ (K_{sp} = 2.9 \times 10^{-9})$

Most of fluoride ions come of stronger electrolyte.

Acid and Bases

(1) **Arrhenius concept** : According to Arrhenius concept *all substances which give H^+ ions when dissolved in water are called acids while those which ionise in water to furnish OH^- ions are called bases.*

$$\underset{(Acid)}{HCl} \xrightleftharpoons{H_2O} \underset{(aq.)}{H^+} + \underset{(aq.)}{Cl^-} \ ; \quad \underset{(Base)}{NaOH} \xrightleftharpoons{H_2O} \underset{(aq.)}{Na^+} + \underset{(aq)}{OH^-}$$

Some acids and bases ionise almost completely in solutions and are called *strong acids and bases*. Others are dissociated to a limited extent in solutions and are termed *weak acids and bases*. $HCl, HNO_3, H_2SO_4, HClO_4$, etc., are examples of strong acids and $NaOH, KOH, (CH_3)_4NOH$ are strong bases. Every hydrogen compound cannot be regarded as an acid, e.g., CH_4 is not an acid. Similarly, CH_3OH, C_2H_5OH, etc., have OH groups but they are not bases.

(i) *Utility of Arrhenius concept* : The Arrhenius concept of acids and bases was able to explain a number of phenomenon like neutralization, salt hydrolysis, strength of acids and bases etc.

(ii) *Limitations of Arrhenius concept*

(a) For the acidic or basic properties, the presence of water is absolutely necessary. Dry HCl shall not act as an acid. HCl is regarded as an acid only when dissolved in water and not in any other solvent.

(b) The concept does not explain acidic and basic character of substances in non-aqueous solvents.

(c) The neutralisation process is limited to those reactions which can occur in aqueous solutions only, although reactions involving salt formation do occur in absence of solvent.

(d) It cannot explain the acidic character of certain salts such as $AlCl_3$ in aqueous solution.

(2) **Bronsted–Lowry concept** : According to this concept,

"An acid is defined as a substance which has the tendency to give a proton (H^+) and a base is defined as a substance which has a tendency to accept a proton. In other words, an acid is a proton donor whereas a base is a proton acceptor."

$$\underset{\text{Acid}}{HCl} + \underset{\text{Base}}{H_2O} \rightleftharpoons H_3O^+ + Cl^- \qquad .....(i)$$

$$\underset{\text{Acid}}{CH_3COOH} + \underset{\text{Base}}{H_2O} \rightleftharpoons H_3O^+ + CH_3COO^- \qquad .....(ii)$$

(i) HCl and CH_3COOH are acids because they donate a proton to H_2O .(ii) NH_3 and CO_3^{2-} are bases because they accept a proton from water.

In reaction (i), in the reverse process, H_3O^+ can give a proton and hence is an acid while Cl^- can accept the proton and hence is a base. Thus there are two acid-base pairs in reaction (i). These are HCl – Cl^- and H_3O^+– H_2O. These acid-base pairs are called *conjugate acid-base pairs*.

Conjugate acid $\rightleftharpoons$ *Conjugate base* $+ H^+$

Conjugate base of a strong acid is a weak base and vice a versa. Weak acid has a strong conjugate base and vice a versa.

Levelling effect and classification of solvents : In acid-base strength series, all acids above H_3O^+ in aqueous solution fall to the strength of H_3O^+. Similarly the basic strength of bases above OH^- fall to the strength of OH^-in aqueous solution. This is known as *levelling effect*. Levelling effect of water is due to its high dielectric constant and strong proton accepting tendency.

On the basis of proton interaction, solvents are of four types,

(i) ***Protophilic solvents*** **:** Solvents which have greater tendency to accept protons, *i.e.,* water, alcohol, liquid ammonia, etc.

(ii) ***Protogenic solvents*** **:** Solvents which have the tendency to produce protons, *i.e.,* water, liquid hydrogen chloride, glacial acetic acid, etc.

(iii) ***Amphiprotic solvents*** **:** Solvents which act both as protophilic or protogenic, *e.g.,* water, ammonia, ethyl alcohol, etc.

(iv) ***Aprotic solvents*** **:** Solvents which neither donate nor accept protons, *e.g.,* benzene, carbon tetrachloride, carbon disulphide, etc.

HCl acts as acid in H_2O, stronger acid in NH_3, weak acid in CH_3COOH, neutral in C_6H_6 and a weak base in HF.

$$\underset{Base}{HCl}+\underset{Acid}{HF} \rightarrow \underset{Acid}{H_2Cl^+}+\underset{Base}{F^-}$$

Utility of Bronsted – Lowry concept

(i) Bronsted – Lowry concept is not limited to molecules but includes even the ionic species to act as acids or bases.

(ii) It can explain the basic character of the substances like Na_2CO_3, NH_3 etc.

(iii) It can explain the acid-base reactions in the non-aqueous medium or even in the absence of a solvent (e.g., between HCl and NH_2).

Limitations of Bronsted lowry concept

(i) The protonic definition cannot be used to explain the reactions occuring in non-protonic solvents such as $COCl_2$, SO_2, N_2O_4, etc.

(ii) It cannot explain the reactions between acidic oxides like CO_2, SO_2, SO_3 etc and the basic oxides like CaO, BaO, MgO etc which take place even in the absence of the solvent *e.g.,*

$$CaO+SO_3 \rightarrow CaSO_4$$

There is no proton transfer in the above example.

(iii) Substances like BF_3, $AlCl_3$ etc, do not have any hydrogen and hence cannot give a proton but are known to behave as acids.

Table: 9.1 Conjugate acid-base pairs

Acid			Conjugate base		
$HClO_4$	(Perchloric acid)	↑ Increasing order of acidic strength	ClO_4^-	(Perchlorate ion)	Increasing order of basic strength ↓
H_2SO_4	(Sulphuric acid)		HSO_4^-	(Hydrogen sulphate ion)	
HCl	(Hydrogen chloride)		Cl^-	(Chloride ion)	
HNO_3	(Nitric acid)		NO_3^-	(Nitrate ion)	
H_3O^+	(Hydronium ion)		H_2O	(Water)	
HSO_4^-	(Hydrogen sulphate ion)		SO_4^{2-}	(Sulphate ion)	
H_3PO_4	(Ortho phosphoric acid)		$H_2PO_4^-$	(Dihydrogen phosphate ion)	
CH_3COOH	(Acetic acid)		CH_3COO^-	(Acetate ion)	
H_2CO_3	(Carbonic acid)		HCO_3^-	(Hydrogen carbonate ion)	
H_2S	(Hydrogen sulphide)		HS^-	(Hydrogen sulphide ion)	
NH_4^+	(Ammonium ion)		NH_3	(Ammonia)	
HCN	(Hydrogen cyanide)		CN^-	(Cyanide ion)	
C_6H_5OH	(Phenol)		$C_6H_5O^-$	(Phenoxide ion)	
H_2O	(Water)		OH^-	(Hydroxide ion)	
C_2H_5OH	(Ethyl alcohol)		$C_2H_5O^-$	(Ethoxide ion)	
NH_3	(Ammonia)		NH_2^-	(Amide ion)	
CH_4	(Methane)		CH_3^-	(Methyl carbanion)	

(3) **Lewis concept** : This concept was proposed by G.N. Lewis, in 1939. According to this concept*, "a base is defined as a substance which can furnish a pair of electrons to form a coordinate bond whereas an acid is a substance which can accept a pair of electrons."* The acid is also known as electron pair acceptor or *electrophile* while the base is electron pair donor or *nucleophile.*

A simple example of an acid-base is the reaction of a proton with hydroxyl ion, $\underset{Acid}{H^+}+\underset{Base}{OH^-} \rightarrow HOH$

Lewis concept is more general than the Bronsted Lowry concept. *All Bronsted bases are also Lewis bases but all Bronsted acids are not Lewis acids.* [e.g., HCl, H_2SO_4 as they are not capable of accepting a pair of electrons]

(i) ***Types of Lewis acids*** **:** According to Lewis concept, the following species can act as Lewis acids.

(a) Molecules in which the central atom has incomplete octet BF_3, BCl_3, $AlCl_3, BeCl_2$, etc.

(b) All cations are expected to act as Lewis acids since they are deficient in electrons.

(c) *Molecules in which the central atom has empty d-orbitals*. e.g., $SiF_4, SnCl_4, PF_5$ etc.

(d) *Molecules having a multiple bond between atoms of dissimilar electronegativity e.g.,* CO_2, SO_2.

(ii) ***Types of Lewis bases*** **:** The following species can act as Lewis bases.

(a) Neutral species having at least one lone pair of electrons $:NH_3, -\ddot{N}H_2, R-\underset{\cdot\cdot}{\ddot{O}}-H$

(b) Negatively charged species or anions

(iii) ***Hard and Soft principle of acids and bases*** **:** Lewis acids and bases are classified as hard and soft acids and bases. Hardness is defined as the property of retaining valence electrons very strongly. Thus a hard acid is that in which electron-accepting atom is small, has a high positive charge and has no electron which can be easily polarised or removed *e.g.,* $Li^+, Na^+, Be^{2+}, Mg^{+2}, Al^{+3}$ BF_3, SO_3 *etc.*. On the contrary, a soft acid is that in which the acceptor atom is large, carries a low positive charge or it has electrons in orbitals which are easily polarised or distorted *e.g.,* $Pb^{+2}, Cd^{+2}, Pt^{+2}, Hg^{+2}, Ro^+, Rs^+, I_2$ etc..

A Lewis base which holds its electrons strongly is called hard base, e.g., $OH^-, F^-, H_2O, NH_3, CH_3OCH_3$, etc. on the other hand, a Lewis base in which the position of electrons is easily polarised or removed is called a soft base e.g., $I^-, CO, CH_3S^-, (CH_3)_3P$, etc.

In general, hard acids prefer to bind to hard bases and soft acids prefer to bind to soft bases. The bonding between hard acids and hard bases is chiefly ionic and that between soft bases and soft acids is mainly covalent.

(iv) ***Utility of Lewis concept*** **:** Lewis concept is the most general of all the concepts and can explain the acidic and basic nature of all those substances which could not be explained by the earlier concepts. Similarly, it can explain even those acid-base reactions which could not be explained by the other concepts.

(v) ***Limitations of lewis concept*** **:** It does not explain behaviour of well known protonic acids, as HCl, H_2SO_4 etc, as which do not form coordinate bonds with bases.

It does not explain relative strengths of acids and bases. Many lewis acids do not posses catalytic property.

Relative strength of acids and Bases

In practice K_a is used to define the strength only of those acids that are weaker than H_3O^+ and K_b is used to define the strength of only those bases that are weaker than OH^-. For two weak acids HA_1 and HA_2 of ionisation constant K_{a_1} and K_{a_2} respectively at the same concentration C, we have,

$$\frac{\text{Acid strength of } HA_1}{\text{Acid strength of } HA_2} = \sqrt{\frac{K_{a_1}}{K_{a_2}}}$$

Similarly, relative strengths of any two weak bases at the same concentration are given by the ratio of the square-roots of their dissociation constants. i.e.,

$$\frac{\text{Basic strength of } BOH_1}{\text{Basic strength of } BOH_2} = \sqrt{\frac{K_{b_1}}{K_{b_2}}}$$

(1) **Relative strength of Inorganic acids**

(i) ***Hydrides***

(a) The acidic strength increases with the increase in the electronegativity of the element directly attached with the hydrogen.

$H-F > H-OH > H-NH_2 > H-CH_3$

$HCl > H_2S > PH_3 > SiH_4$

(b) The acidic strength increases with the increase in atomic size,

$HF < HCl < HBr < HI$; $H_2O < H_2S < H_2Se < H_2Te$

(ii) ***Oxyacids***

(a) Among oxyacids of the same type formed by different elements, acidic nature increases with increasing electronegativity,

$HOI < HOBr < HOCl$; $HIO_4 < HBrO_4 < HClO_4$

(b) In oxyacids of the same element, acidic nature increases with its oxidation number

$\underset{+1}{HOCl} < \underset{+3}{HClO_2} < \underset{+5}{HClO_3} < \underset{+7}{HClO_4}$; $H_2SO_3 < H_2SO_4$

$HNO_2 < HNO_3$

(c) The strength of oxyacids increases from left to right across a period

$H_4SiO_4 < H_3PO_4 < H_2SO_4 < HClO_4$

(d) For the same oxidation state and configuration of the elements, acid strength decreases with increase in size of the atom.

$HNO_3 > HPO_3$; $H_3PO_4 > H_3AsO_4$

$HClO_4 > HBrO_4 > HIO_4$

(2) **Relative strength of organic acids**

(i) A compound is acidic in nature, if its conjugate base can stabilize through resonance. Thus phenol is acidic while ethanol is neutral because the conjugate base of phenol $(C_6H_5O^-)$ can be stabilized through resonance while that of alcohol $(C_2H_5O^-)$ can not.

(ii) Hydrogen atom attached to sp-hybridized carbon is more acidic than that on sp^2 hybridized carbon which in turn is more acidic than that on sp^3 hybridized carbon.

Thus, $\underset{sp}{HC \equiv CH} > \underset{sp^2}{CH_2 = CH_2} > \underset{sp^3}{CH_3 - CH_3}$

(3) **Relative strength of Inorganic bases**

(i) The basicity of a compound decreases with increase in electronegativity of the atom holding the electron pair, $\ddot{N}H_3 > H_2\ddot{O}: > H\ddot{\underset{\cdot\cdot}{F}}:$

(ii) The larger the size of the atom holding the unshared electrons, the lesser is the availability of electrons.

$F^- > Cl^- > Br^- > I^-$; $O^{2-} > S^{2-}$

(iii) Presence of negative charge on the atom holding the electron pair increases the basicity, while the presence of positive charge on the atom holding the electron pair decreases the basicity. $OH^- > H_2O > H_3O^+$

(iv) Among alkali and alkaline earth hydroxides (oxides) the basic nature increases with electropositivity

$LiOH < NaOH < KOH < RbOH < CsOH$;

$Be(OH)_2 < Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$

$CsOH$ is the strongest known base

(v) On going down the group; basic nature decreases with size of the central atom due to decrease in the ability to donate the lone pair. $NH_3 > PH_3 > AsH_3 > SbH_3 > BiH_3$

(4) **Relative strength of Organic bases**

(i) Higher the electron density on nitrogen, more is the basic character of amine.

(ii) A compound is basic in nature, if its conjugate acid can be stabilized through resonance. Thus guanidine $(NH_3 - \overset{NH_2}{\overset{|}{C}} = NH)$ is as strong alkali as metal hydroxides because its conjugate acid $(H_3N^+ - \overset{NH_2}{\overset{|}{C}} = NH)$ is very much stabilised through resonance.

The acid-base neutralisation and Salt

The reaction between an acid and a base to form salt and water is termed *neutralisation*

$$HCl_{(aq.)} + NaOH_{(aq.)} \rightleftharpoons \underset{\substack{\text{Sodium Chloride}\\ \text{Salt}}}{NaCl_{(aq.)}} + H_2O_{(l)}$$

The process of neutralisation does not produce the resulting solution always neutral; no doubt it involves the interaction of H^+ and OH^- ions. The nature of the resulting solution depends on the particular acid and the particular base involved in the reaction.

Salts : Salts are regarded as compounds made up of positive and negative ions. *The positive part comes from a base while negative part from an acid*. Salts are ionic compounds.The salts can be classified into following classes,

(1) **Simple salts :** The salt formed by the interaction between acid and base, is termed as simple salt. These are of three types,

(i) ***Normal salts*** **:** the salts formed by the loss of all possible protons (replaceable hydrogen atoms as H^+) are called normal salts. Such a salt does not contain either replacable hydrogen or a hydroxyl group.

Examples : $NaCl, NaNO_3, K_2SO_4, Ca_3(PO_4)_2, Na_3BO_3$, Na_2HPO_3 (one H atom is not replaceable as H_3PO_2 is a dibasic acid) NaH_2PO_2 (both H atoms are not replaceable as H_3PO_2 is a monobasic acid) etc.

(ii) ***Acidic salts*** **:** Salts formed by incomplete neutralisation of poly-basic acids are called acidic salts. Such salts still contain one or more replaceable hydrogen atoms. These salts when neutralised by bases form normal salts.

Examples : $NaHCO_3, NaHSO_4, NaH_2PO_4, Na_2HPO_4$, etc.

(iii) ***Basic salts*** **:** Salts formed by incomplete neutralisation of poly acidic bases are called basic salts. Such salts still contain one or more hydroxyl groups. These salts when neutralised by acids form normal salts.

Examples: $Zn(OH)Cl, Mg(OH)Cl, Fe(OH)_2Cl, Bi(OH)_2Cl$

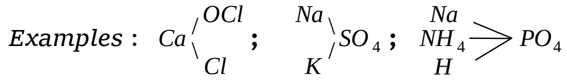
(2) **Double salts :** The addition compounds formed by the combination of two simple salts are termed double salts. Such salts are stable in solid state only.

Examples : Ferrous ammonium sulphate, Potash alum and other alums.

(3) **Complex salts :** These are formed by combination of simple salts or molecular compounds. These are stable in solid state as well as in solutions.

$$\underset{\text{Simple salts}}{FeSO_4 + 6KCN} \rightarrow \underset{\text{Complex salt}}{K_4[Fe(CN)_6]} + K_2SO_4$$

(4) **Mixed salts :** The salt which furnishes more than one cation or more than one anion when dissolved in water is called a mixed salt.



Examples : $Ca\begin{matrix} / OCl \\ \backslash Cl \end{matrix}$; $\begin{matrix} Na \backslash \\ K / \end{matrix} SO_4$; $\begin{matrix} Na \\ NH_4 \\ H \end{matrix} > PO_4$

Ionic product of water

Water is a weak electrolyte and undergoes selfionistion to a small extent.

"The product of concentrations of H^+ and OH^- ions in water at a particular temperature is known as ionic product of water." It is designated as K_w.

$$H_2O \rightleftharpoons H^+ + OH^- ; \ \Delta H = +57.3 \ kJM^{-1}$$

$$K = \frac{[H^+][OH^-]}{[H_2O]} ; K[H_2O] = [H^+][OH^-] ; K_w = [H^+][OH^-]$$

The value of K_w increases with the increase of temperature, *i.e., the concentration H^+ and OH^- ions increases with increase in temperature*.

The value of K_w at 25^oC is 1×10^{-14} mole/litre. Since pure water is neutral in nature, H^+ ion concentration must be equal to OH^- ion concentration.

$[H^+]=[OH^-]=x$ or $[H^+][OH^-]=x^2=1\times10^{-14}$ or $x=1\times10^{-7}M$ or $[H^+]=[OH^-]=1\times10^{-7}$ $mole\ litre^{-1}$

This shows that at 25^oC, in 1 *litre* only 10^{-7} *mole* of water is in ionic form out of a total of approximately 55.5 *moles*.

Thus when, $[H^+]=[OH^-]$; the solution is neutral

$[H^+]>[OH^-]$; the solution is acidic

$[H^+]<[OH^-]$; the solution is basic

Hydrogen ion concentration – *pH* scale

Sorensen, a *Danish biochemist* developed a scale to measure the acidity in terms of concentrations of H^+ in a solution. As defined by him, *"pH of a solution is the negative logarithm to the base 10 of the concentration of H^+ ions which it contains."*

$pH=-\log[H^+]$ or $pH=\log\frac{1}{[H^+]}$

Just as *pH* indicates the hydrogen ion concentration, the *pOH* represents the hydroxyl ion concentration, *i.e.,*

$pOH=-\log[OH^-]$

Considering the relationship, $[H^+][OH^-]=K_w=1\times10^{-14}$

Taking log on both sides, we have

$\log[H^+]+\log[OH^-]=\log K_w=\log(1\times10^{-14})$ or

$-\log[H^+]-\log[OH^-]=-\log K_w=-\log(1\times10^{-14})$

or $pH+pOH=pK_w=14$

	$[H^+]$	$[OH^-]$	*pH*	*pOH*
Acidic solution	$> 10^{-7}$	$< 10^{-7}$	< 7	> 7
Neutral solution	10^{-7}	10^{-7}	7	7
Basic solution	$< 10^{-7}$	$> 10^{-7}$	> 7	< 7

***pH* of some materials**

Material	*pH*	Material	*pH*
Gastric juice	1.4	Rain water	6.5
Lemon juice	2.1	Pure water	7.0
Vinegar	2.9	Human saliva	7.0
Soft drinks	3.0	Blood plasma	7.4
Beer	4.5	Tears	7.4
Black coffee	5.0	Egg	7.8
Cow's milk	6.5	Household ammonia	11.9

Limitations of *pH* scale

(i) *pH* values of the solutions do not give us immediate idea of the relative strengths of the solutions. A solution of *pH* =1 has a hydrogen ion concentration 100 times that of a solution *pH* = 3 (not three times). A $4\times10^{-5}N\ HCl$ is twice concentrated of a $2\times10^{-5}N\ HCl$ solution, but the pH values of these solutions are 4.40 and 4.70 (not double).

(ii) *pH* value zero is obtained in $1N$ solution of strong acid. If the concentration is $2N, 3N, 10N$, etc. the respective *pH* values will be negative.

(iii) A solution of an acid having very low concentration, say $10^{-8}N$, can not have *pH* 8, as shown by *pH* formula but the actual *pH* value will be less than 7.

***pK* value :** *p* stands for negative logarithm. Just as H^+ and OH^- ion concentrations range over many negative powers of 10, it is convenient to express them as *pH* or *pOH*, the dissociation constant (*K*) values also range over many negative powers of 10 and it is convenient to write them as *pK*. Thus, *pK* is the negative logarithm of dissociation constant.

$pK_a=-\log K_a$ and $pK_b=-\log K_b$

Weak acids have higher pK_a values. Similarly weak bases have higher pK_b values

For any conjugate acid-base pair in aqueous solution, $K_a\times K_b=K_w$

$pK_a+pK_b=pK_w=14$ (at 298° *K*)

Calculation of the *pH* of $10^{-8}M\ HCl$ & $10^{-8}M\ NaOH$

If we use the relation $pH=-\log[H_3O^+]$ we get pH equal to 8, but this is not correct because an acidic solution connot have pH greater than 7. In this condition H^+ concentration of water cannot be neglected.

Therefore, $[H^+]_{total}=H^+_{Acid}+H^+_{water}$

Since HCl is strong acid and completely ionised,

$[H^+]_{HCl}=1\times10^{-8}$, $[H^+]_{H_2O}=10^{-7}$

$[H^+]_{total}=[H^+]_{HCl}+[H^+]_{H_2O}=10^{-8}+10^{-7}=10^{-8}[1+10]$ $=10^{-8}\times11$

$pH=-\log10^{-8}+\log11=6.958$

Similarly if $NaOH$ concentration is $10^{-8}M$

Then, $[OH^-]_{total}=[10^{-8}]_{NaOH}+[10^{-7}]_{H_2O}$

$[OH^-]=10^{-8}\times11$; $pOH=6.96$ $pH=7.04$

Buffer solutions

A solution whose pH is not altered to any great extent by the addition of small quantities of either sirong acid (H^+ ions) or a sirong base (OH^- ions) is called the buffer solution. It can also be defined as a solution of *reserve acidity* or *alkalinity* which resists change of pH upon the addition of small amount of acid or alkali.

(1) **Types of buffer solutions :** There are two types of buffer solutions,

(i) ***Solutions of single substances* :** The solution of the salt of a weak acid and a weak base.

Example : ammonium acetate (CH_3COONH_4), NH_4CN act as a buffer.

(ii) ***Solutions of Mixtures* :** These are further of two types,

(a) *Acidic buffer* : It is the solution of a mixture of a weak acid and a salt of this weak acid with a strong base.

Example : $CH_3COOH + CH_3COONa$

(b) *Basic buffer* : It is the solution of a mixture of a weak base and a salt of this weak base with a strong acid.

Example : $NH_4OH + NH_4Cl$

(2) **Buffer action** : *Buffer action is the mechanism by which added H^+ ions or OH^- ions are almost neutralised; so that pH practically remains constant.* Reserved base of buffer neutralises the added H^+ ions while the reserved acid of buffer neutralises the added OH^- ions.

(3) **Examples of buffer solutions**

(i) Phthalic acid + potassium hydrogen phthalate

(ii) Citric acid + sodium citrate.

(iii) Boric acid + borax (sodium tetraborate).

(iv) Carbonic acid (H_2CO_3)+ sodium hydrogen carbonate $(NaHCO_3)$. This system is found in blood and helps in maintaining pH of the blood close to 7.4 (pH value of human blood lies between 7.36 – 7.42; a change in pH by 0.2 units may cause death).

(v) $NaH_2PO_4 + Na_3PO_4$

(vi) $NaH_2PO_4 + Na_2HPO_4$

(vii) Glycerine + HCl

(viii) The pH value of gastric juice is maintained between 1.6 and 1.7 due to buffer system.

(4) **Henderson - Hasselbalch equation :** pH of an acidic or a basic buffer can be calculated by Henderson- Hasselbalch equation.

For acidic buffers, $pH = pK_a + \log\frac{[salt]}{[acid]}$

When $\frac{[salt]}{[acid]} = 10$, then, $pH = 1 + pK_a$ and

when $\frac{[salt]}{[acid]} = \frac{1}{10}$, then, $pH = pK_a - 1$

So weak acid may be used for preparing buffer solutions having pH values lying within the ranges $pK_a + 1$ and $pK_a - 1$. The acetic acid has a pK_a of about 4.8; it may, therefore, be used for making buffer solutions with pH values lying roughly within the range 3.8 to 5.8.

For basic Buffers, $pOH = pK_b + \log\frac{[salt]}{[base]}$

Knowing pOH , pH can be calculated by the application of formula, $pH + pOH = 14$

pH of a buffer solution does not change with dilution but it varies with temperature because value of K_w changes with temperature.

(5) **Buffer capacity :** The property of a buffer solution to resist alteration in its *pH* value is known as *buffer capacity*. It has been found that if the ratio $\frac{[salt]}{[acid]}$ or $\frac{[salt]}{[base]}$ is unity, the *pH* of a particular buffer does not change at all. Buffer capacity is defined quantitatively as number of moles of acid or base added in one litre of solution as to change the *pH* by unity, *i.e.,*

Buffer capacity

$$(\phi) = \frac{\text{Number of moles of acid or base added to 1 litre}}{\text{Change in pH}}$$

Thus greater the buffer capacity, the greater is its capacity to resist change in pH value. Buffer capacity is greatest when the concentration of salt and weak acid/base are equal, or when $pH = pK_a$ or $pOH = pK_b$.

(6) **Significance of buffer solutions**

(i) Buffer solutions are used for comparing colorimetrically the hydrogen ion concentration of unknown solutions.

(ii) Acetic acid-sodium acetate is used in the removal of phosphate radical during the qualitative analysis of the mixture.

(iii) NH_4Cl/NH_4OH buffer is used for the precipitation of hydroxides of third group of qualitative analysis.

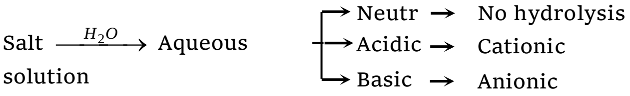
(iv) In industries, buffer solutions are used in the alcoholic fermentation (*pH* 5 to 6.5), tanning of leather, electroplating, manufacture of sugar, paper manufacturing etc.,

(v) In bacteriological research culture media are generally buffered to maintain the *pH* required for the growth of the bacteria being studied.

(vi) In biological systems buffer system of carbonic acid and sodium bicarbonate is found in our blood. It maintains the *pH* of blood to a constant value (about 7.4) inspite of various acid and base-producing reactions going on in our body.

Salt hydrolysis

It is the reaction of the cation or the anion or both the ions of the salt with water to produce either acidic or basic solution. Hydrolysis is the reverse of neutralization.



(1) **Hydrolysis constant :** The general equation for the hydrolysis of a salt (BA), $\underset{\text{salt}}{BA} + H_2O \rightleftharpoons \underset{\text{acid}}{HA} + \underset{\text{base}}{BOH}$

Applying the law of chemical equilibrium, we get

$\frac{[HA][BOH]}{[BA][H_2O]} = K$, where K is the equilibrium constant.

Since water is present in very large excess in the aqueous solution, its concentration $[H_2O]$ may be regarded as constant so,

$$\frac{[HA][BOH]}{[BA]} = K[H_2O] = K_h$$

where K_h is called the hydrolysis constant.

(2) **Degree of hydrolysis :** It is defined as the fraction (or percentage) of the total salt which is hydrolysed at equilibrium. For example, if 90% of a salt solution is hydrolysed, its degree of hydrolysis is 0.90 or as 90%. It is generally represented by 'h'.

$$h = \frac{\text{Number of moles of the salt hydrolysed}}{\text{Total number of moles of the salt taken}}$$

Types of salt	**Exp. for K_h**	**Exp. for *h***	**Exp. for *pH***
(i) Salt of weak acid and strong base	$K_h = K_w / K_a$	$h = \sqrt{\left(\frac{K_h}{C}\right)}$	$pH = -\frac{1}{2}$ [log K_w + log K_a – log C]
(ii) Salt of strong acid and weak base	$K_h = K_w / K_b$	$h = \sqrt{\left(\frac{K_h}{C}\right)}$	$pH = -\frac{1}{2}$ [log K_w – log K_b + log C]
(iii) Salt of weak acid and weak base	$K_h = \frac{K_w}{K_a K_b}$	$h = \sqrt{(K_h)}$	$pH = -\frac{1}{2}$ [log K_a + log K_w – log K_b]

(iv) Salts of strong acids and strong bases do not undergo hydrolysis (they undergo only ionization) hence the resulting aqueous solution is neutral.

Indicators

An indicator is a substance, which is used to determine the end point in a titration. In acid-base titrations, organic substance (weak acids or weak bases) are generally used as indicators. They change their colour within a certain *pH* range. The colour change and the *pH* range of some common indicators are tabulated below

Table : 9.2 Colour changes of indicators with *pH*

Indicator	pH range	Colour	
		Acid solution	Base solution
Cresol red	1.2 – 1.8	Red	Yellow
Thymol blue	1.2 – 2.8	Red	Yellow
Methyl yellow	2.9 – 4.0	Red	Yellow
Methyl orange	3.1 – 4.4	Pink	Yellow
Methyl red	4.2 – 6.3	Red	Yellow
Litmus	5.0 – 8.0	Red	Blue
Bromothymol blue	6.0 – 7.6	Yellow	Blue
Phenol red	6.4 – 8.2	Yellow	Red
Thymol blue (base)	8.1 – 9.6	Yellow	Blue
Phenolphthalein	8.3 – 10.0	Colourless	Pink
Thymolphthalein	8.3 – 10.5	Colourless	Blue
Alizarin yellow R	10.1 – 12.0	Blue	Yellow
Nitramine	10.8 – 13.0	Colourless	Orange, Brown

Two theories have been proposed to explain the change of colour of acid-base indicators with change in *pH*.

(i) *Ostwald's Theory* (ii) *Quinonoid theory*

(1) **Selection of suitable indicator or choice of indicator :** In order to choose a suitable indicator, it is necessary to understand the *pH* changes in the titrations. The change in *pH* in the vicinity of the equivalence point is most important for this purpose. The curve obtained by plotting *pH* as ordinate against the volume of alkali added as abscissa is known as neutralisation or titration curve. The suitable indicators for the following titrations are,

(i) ***Strong acid Vs strong base*** **:** Phenolphthalein (*pH* range 8.3 to 10.5), methyl red (pH range 4.4 – 6.5) and methyl orange (*pH* range 3.2 to 4.5).

(ii) ***Weak acid Vs strong base*** **:** Phenolphthalein.

(iii) ***Strong acid Vs weak base*** **:** Methyl red and methyl orange.

(iv) ***Weak acid vs. weak base*** **:** No suitable indicator can be used for such a titration.

Reason for use of different indicators for different systems : Indicators are either weak acids or weak bases and when dissolved in water their dissociated form acquires a colour different from that of the undissociated form. Consider a weak acid indicator of the general formula *HIn*, where in represents indicator. The equilibrium established in aqueous solution will be

$$\underset{\text{Red}}{HIn(aq.)} \rightleftharpoons H^+(aq.) + \underset{\text{Green}}{In^-(aq.)}$$

Let K_{In} be the equilibrium constant

$$K_{In} = \frac{[H^+][In^-]}{[HIn]} \quad \text{or} \quad \frac{[HIn]}{[In^-]} = \frac{[H^+]}{K_{In}}$$

The human eye can detect the change in colour if the ratio of the two forms of indicator ranges between 0.1 to 10.

If, $\frac{[HIn]}{[In^-]} = 1.0$, the colour visible will be yellow

$\frac{[HIn]}{[In^-]} = 10$, the colour visible will be red.

$\frac{[HIn]}{[In^-]} = 0.1$, the colour visible will be green.

In other words,

The colour visible will be red, when $pH = pK_{In} - 1$

The colour visible will be yellow, when $pH = pK_{In}$

The colour visible will be green, when $pH = pK_{In} + 1$

Thus, our imaginary indicator will be red at any pH which just falls below $pK_{In} - 1$ and green at any pH which just exceeds $pK_{In} + 1$. The indicator changes its colour in the narrow pH range $pK_{In} - 1$ to $pK_{In} + 1$ from red to (red-yellow, yellow, yellow-green) green. We can therefore use this indicator to locate this narrow pH range. In other words, in order to use the indicator effectively in this range, we should have a solution for which pH is very near to pK_{In} of the indicator. The colour change of an indicator can, therefore, be summarised as,

	First change of colour	**Mid point of change**	**Colour change complete**
$[H^+]$	$10\ K_{\text{In}}$	K_{In}	$0.1\ K_{\text{In}}$
pH	$PK_{\text{In}} - 1$	PK_{In}	$PK_{\text{In}} + 1$

It is for this reason that we use different indicators for different systems.

Tips & Tricks

- pH of boiling water is 6.5625. It does not mean that boiling water is not neutral. It is due to greater dissociation of H_2O into H^+ and OH^-.
- pH values of solutions do not give the exact idea of their relative strengths *e.g.,* (i) A solution with pH = 1 had $[H^+]$ 100 times than that with pH = 3 and not 3 times. (ii) A 4×10^{-5} M HCl solution is twice concentrated as compared to 2×10^{-5} M HCl solution but pH values of these solutions are 4.4 and 4.7 respectively and not double.
- pH can be zero in 1M HCl or it can be negative for more concentrated solutions like 2M, 3M, 10M etc.
- At the temperature of the human body which is nearly 37°C, pH of neutral solution is 6.8.
- Buffer solutions have reserve acidity and reserve alkalinity.
- The greater the buffer capacity, the greater is its capacity to resist change in pH value.
- Buffers cannot withstand the addition of large amounts of acids or alkalies. The addition of 0.1*mol* per litre of $[H^+]$ or OH^- is about the maximum that any buffer can be expected to withstand.

Ordinary Thinking

Objective Questions

Electrical conductors, Arrhenius theory and Ostwald's dilution law

1. Which of the following is non-electrolyte **[CPMT 2001]**

(a) $NaCl$ (b) $CaCl_2$

(c) $C_{12}H_{22}O_{11}$ (d) CH_3COOH

2. Ammonium hydroxide is a **[CPMT 1977]**

(a) Strong electrolyte

(b) Weak electrolyte

(c) Both under different conditions

(d) Non-electrolyte

3. Ammonium hydroxide is a weak base because **[MP PET 2000]**

(a) It has low vapour pressure

(b) It is only slightly ionized

(c) It is not a hydroxide of any metal

(d) It has low density

4. Electrolytes when dissolved in water dissociate into their constituent ions. The degree of dissociation of an electrolyte increases with **[CPMT 1974]**

(a) Increasing concentration of the electrolyte

(b) Decreasing concentration of the electrolyte

(c) Decreasing temperature

(d) Presence of a substance yielding a common ion

5. An electrolyte **[MP PMT/PET 1988; CPMT 1974]**

(a) Gives complex ions in solution

(b) Dissolves in water to give ions

(c) Is ionized in the solid state

(d) Generates ions on passing electric current

6. A monoprotic acid in 1.00 M solution is 0.01% ionised. The dissociation constant of this acid is **[BVP 2003]**

(a) 1×10^{-8} (b) 1×10^{-4}

(c) 1×10^{-6} (d) 10^{-5}

7. Molten sodium chloride conducts electricity due to the presence of **[BHU 2001]**

(a) Free electrons

(b) Free ions

(c) Free molecules

(d) Atoms of sodium and chlorine

8. An example for a strong electrolyte is **[KCET 2002]**

(a) Urea (b) Ammonium hydroxide

(c) Sugar (d) Sodium acetate

9. Which one is strongest electrolyte in the following **[CPMT 1990]**

(a) $NaCl$ (b) CH_3COOH

(c) NH_4OH (d) $C_6H_{12}O_6$

10. The equivalent conductance at infinite dilution of a weak acid such as HF **[Pb. PMT 1998]**

(a) Can be determined by measurement of very dilute HF solution

(b) Can be determined by extrapolation of measurements on dilute solutions of HCl, HBr and HI

(c) Can best be determined from measurements on dilute solutions of NaF, $NaCl$ and HCl

(d) Is an undefined quantity

11. If α is the degree of ionization, C the concentration of a weak electrolyte and K_a the acid ionization constant, then the correct relationship between α, C and K_a is **[CET Pune 1998; Pb. PMT 1998; RPMT 2002]**

(a) $\alpha^2 = \sqrt{\frac{K_a}{C}}$ (b) $\alpha^2 = \sqrt{\frac{C}{K_a}}$

(c) $\alpha = \sqrt{\frac{K_a}{C}}$ (d) $\alpha = \sqrt{\frac{C}{K_a}}$

12. Theory of ionization was given by **[AMU 1983; DPMT 1985]**

(a) Rutherford (b) Graham

(c) Faraday (d) Arrhenius

13. An ionizing solvent has

(a) Low value of dielectric constant

(b) High value of dielectric constant

(c) A dielectric constant equal to 1

(d) Has a high melting point

14. The extent of ionization increases **[MNR 1982]**

(a) With the increase in concentration of solute

(b) On addition of excess water to solution

(c) On decreasing the temperature of solution

(d) On stirring the solution vigorously

15. Which is generally true about ionic compounds **[Pb. PMT 2002]**

(a) Have low boiling point

(b) Have low melting point

(c) Soluble in non polar solvents

(d) Conduct electricity in the fused state

16. At infinite dilution, the percentage ionisation for both strong and weak electrolytes is **[CPMT 1999]**

(a) 1% (b) 20%

(c) 50% (d) 100%

17. The degree of ionization of a compound depends on **[MNR 1980]**

(a) Size of solute molecules

(b) Nature of solute molecules
(c) Nature of vessel used
(d) Quantity of electricity passed

18. For a weak acid $HA,$ Ostwald's dilution law is represented by the equation

(a) $K_a = \frac{\alpha c}{1-\alpha^2}$ (b) $K_a = \frac{\alpha^2 c}{1-\alpha}$

(c) $\alpha = \frac{K_a c}{1-c}$ (d) $K_a = \frac{\alpha^2 c}{1-\alpha^2}$

19. Acetic acid is a weak electrolyte because **[CPMT 1974]**

(a) Its molecular weight is high
(b) It is covalent compound
(c) It does not dissociate much or its ionization is very less
(d) It is highly unstable

20. In which of the following dissociation of NH_4OH will be minimum **[MP PET 2000]**

(a) $NaOH$ (b) H_2O
(c) NH_4Cl (d) $NaCl$

21. Vant hoff factor of $BaCl_2$ of conc. $0.01M$ is 1.98. Percentage dissociation of $BaCl_2$ on this conc. Will be **[Kerala CET 2005]**

(a) 49 (b) 69
(c) 89 (d) 98
(e) 100

22. In which of the following solutions, ions are present **[NCERT 1981]**

(a) Sucrose in water (b) Sulphur in CS_2
(c) Caesium nitrate in water (d) Ethanol in water

23. The following equilibrium exists inaqueous solution, $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$ if dil HCl is added, without change in temperature, the

(a) Concentration of CH_3COO^- will increase
(b) Concentration of CH_3COO^- will decrease
(c) The equilibrium constant will increase
(d) The equilibrium constant will decrease

24. Which will not affect the degree of ionisation **[MP PMT 1994]**

(a) Temperature (b) Concentration
(c) Type of solvent (d) Current

25. The addition of a polar solvent to a solid electrolyte results in **[NCERT 1973]**

(a) Polarization (b) Association
(c) Ionization (d) Electron transfer

26. The degree of dissociation of $0.1 M\, HCN$ solution is 0.01%. Its ionisation constant would be **[RPMT 1999]**

(a) 10^{-3} (b) 10^{-5}
(c) 10^{-7} (d) 10^{-9}

27. The hydrogen ion concentration in weak acid of dissociation constant K_a and concentration c is nearly equal to **[CBSE PMT 1989; RPMT 2000]**

(a) $\sqrt{K_a / c}$ (b) c / K_a
(c) $K_a c$ (d) $\sqrt{K_a c}$

28. Degree of dissociation of $0.1 N\; CH_3COOH$ is (Dissociation constant $= 1 \times 10^{-5}$) **[MP PET 1997]**

(a) 10^{-5} (b) 10^{-4}
(c) 10^{-3} (d) 10^{-2}

29. Which of the following substance is an electrolyte **[MADT Bihar 1980]**

(a) Chloroform
(b) Benzene
(c) Toluene
(d) Magnesium chloride

30. In weak electrolytic solution, degree of ionization

(a) Will be proportional to dilution
(b) Will be proportional to concentration of electrolyte
(c) Will be proportional to the square root of dilution
(d) Will be reciprocal to the dilution

31. 0.2 molar solution of formic acid is ionized 3.2%. Its ionization constant is **[MP PMT 1991]**

(a) 9.6×10^{-3} (b) 2.1×10^{-4}
(c) 1.25×10^{-6} (d) 4.8×10^{-5}

32. The best conductor of electricity is a 1.0 M solution of **[NCERT 1973]**

(a) Boric acid (b) Acetic acid
[UPSEAT 2000, 02]
(c) Sulphuric acid (d) Phosphoric acid

33. The colour of an electrolyte solution depends on **[DPMT 1985]**

(a) The nature of the anion
(b) The nature of the cation
(c) The nature of both the ions
(d) The nature of the solvent

34. Ionisation depends upon **[CPMT 2004]**

(a) Pressure (b) Volume
(c) Dilution (d) None of these

35. The values of dissociation constants of some acids (at $25^\circ C$) are as follows. Indicate which is the strongest acid in water **[MP PMT 1991]**

(a) 1.4×10^{-2} (b) 1.6×10^{-4}
(c) 4.4×10^{-10} (d) 4.3×10^{-7}

36. Concentration CN^- in $0.1\,M\,HCN$ is $[K_a = 4\times 10^{-10}]$ **[RPET 2000]**

(a) $2.5\times 10^{-6}M$ (b) $4.5\times 10^{-6}M$

(c) $6.3\times 10^{-6}M$ (d) $9.2\times 10^{-6}M$

Acids and Bases

1. Which of the following is not a Lewis acid **[MP PET 2002]**

(a) CO (b) $SiCl_4$

(c) SO_3 (d) Zn^{2+}

2. Review the equilibrium and choose the correct statement $HClO_4 + H_2O \rightleftarrows H_3O^+ + ClO_4^-$ **[RPMT 2000]**

(a) $HClO_4$ is the conjugate acid of H_2O

(b) H_3O^+ is the conjugate base of H_2O

(c) H_2O is the conjugate acid of H_3O^+

(d) ClO_4^- is the conjugate base of $HClO_4$

3. A solution of $FeCl_3$ in water acts as acidic due to **[BVP 2003]**

(a) Hydrolysis of Fe^{3+} (b) Acidic impurities

(c) Dissociation (d) Ionisation

4. A white substance having alkaline nature in solution is **[BVP 2003]**

(a) $NaNO_3$ (b) NH_4Cl

(c) Na_2CO_3 (d) Fe_2O_3

5. Which of the following can act both as Bronsted acid and Bronsted base **[MP PET 1995; MP PET/PMT 1998]**

(a) Cl^- (b) HCO_3^-

(c) H_3O^+ (d) OH^-

6. Lewis acid **[MP PMT 1987]**

(a) Presence of H atom is necessary

(b) Is a electron pair donor

(c) Always a proton donor

(d) Is a electron pair acceptor

7. For two acids A and B, $pK_a = 1.2$, $pK_b = 2.8$ respectively in value, then which is true **[Bihar MEE 1998]**

(a) A and B both are equally acidic

(b) A is stronger than B

(c) B is stronger than A

(d) Neither A nor B is strong

(e) None of these

8. Aq. solution of sodium cyanide is **[BHU 1981]**

(a) Acidic (b) Amphoteric

(c) Basic (d) Netural

9. Which of the following is the strongest conjugate base **[MADT Bihar 1983; CBSE PMT 1999; KCET (Med.) 2001]**

(a) Cl^- (b) CH_3COO^-

(c) SO_4^{--} (d) NO_2^-

10. $NaOH$ is a strong base because **[AIIMS 2001]**

(a) It gives OH^- ion (b) It can be oxidised

(c) It can be easily ionised (d) Both (a) and (c)

11. Which one of the following can be classified as a Bronsted base **[KCET 2001]**

(a) NO_3^- (b) H_3O^+

(c) NH_4^+ (d) CH_3COOH

12. Which one of the following substance has the highest proton affinity **[AIEEE 2003]**

(a) H_2O (b) H_2S

(c) NH_3 (d) PH_3

13. Which of the following is the strongest Lewis acid **[EAMCET 1998]**

(a) BI_3 (b) BBr_3

(c) BCl_3 (d BF_3

14. An aqueous solution of ammonia consists of **[MP PET 2001]**

(a) H^+ (b) OH^-

(c) NH_4^+ (d) NH_4^+ and OH^-

15. Which of the following is not a Lewis acid **[CBSE PMT 1996]**

(a) BF_3 (b) $FeCl_3$

(c) SiF_4 (d) C_2H_4

16. The conjugate base of NH_2^- is **[EAMCET 1998]**

(a) NH_3 (b) NH^{2-}

(c) NH_4^+ (d) N_3^-

17. The strength of an acid depends on its tendency to **[MP PET 1996]**

(a) Accept protons (b) Donate protons

(c) Accept electrons (d) Donate electrons

18. Which is not a electrophile **[RPET 1999]**

(a) $AlCl_3$ (b) BF_3

(c) $(CH_3)_3C^+$ (d) NH_3

19. Ammonia gas dissolves in water to give NH_4OH. In this reaction water acts as **[CPMT 1990; MP PMT 1990]**

(a) An acid (b) A base

(c) A salt (d) A conjugate base

20. In the equilibrium

$CH_3COOH + HF \rightleftarrows CH_3COOH_2^+ + F^-$ **[BHU 1987]**

(a) F^- is the conjugate acid of CH_3COOH

(b) F^- is the conjugate base of HF

(c) CH_3COOH is the conjugate acid of $CH_3COOH_2^+$

(d) $CH_3COOH_2^+$ is the conjugate base of CH_3COOH

21. The compound that is not a Lewis acid is **[IIT 1985]**

(a) BF_3 (b) $AlCl_3$

(c) $BeCl_2$ (d) NH_3

22. Which of the following dissolves in water to give a neutral solution **[Bihar MADT 1980]**

(a) $(NH_4)_2SO_4$ (b) $Ba(NO_3)_2$

(c) $CrCl_3$ (d) $CuSO_4$

23. Which of the following is the strongest acid **[AMU 1982]**

(a) H_3PO_4 (b) H_2SO_4

(c) HNO_2 (d) CH_3COOH

24. An example of a Lewis acid is **[MADT Bihar 1982; AMU 1982; MNR 1994; RPMT 1997; KCET 2000; Orissa JEE 2005]**

(a) $NaCl$ (b) $MgCl_2$

(c) $AlCl_3$ (d) $SnCl_4$

25. In the equilibrium $HClO_4 + H_2O \rightleftharpoons H_2O^+ + ClO_4^-$ **[BHU 1981, 86]**

(a) $HClO_4$ is the conjugate acid of H_2O

(b) H_2O is the conjugate acid of H_3O^+

(c) H_3O^+ is the conjugate base of H_2O

(d) ClO_4^- is the conjugate base of $HClO_4$

26. Which of the following would be expected to form ionic solution in water **[CPMT 1976; Kurukshetra CEE 1998]**

(a) CCl_4 (b) O_2

(c) $NaBr$ (d) $CHBr_3$

27. A solution of sodium bicarbonate in water turns **[NCERT 1971]**

(a) Phenolphthalein pink (b)Methyl orange yellow

(c) Methyl orange red (d) Blue litmus red

28. Accepting the definition that an acid is a proton donor, the acid in the following reaction $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$ is **[Kerala (Med.) 2003]**

(a) NH_3 (b) H^+

(c) NH_4^+ (d) H_2O

29. With reference to protonic acids, which of the following statements is correct **[CPMT 1990]**

(a) PH_3 is more basic than NH_3

(b) PH_3 is less basic than NH_3

(c) PH_3 is equally basic as NH_3

(d) PH_3 is amphoteric while NH_3 is basic

30. NH_4OH is weak base because **[CPMT 1979]**

(a) It has low vapour pressure

(b) It is only slightly ionized

(c) It is not a hydroxide of metal

(d) It has low density

31. HNO_3 in liquid hydrogen fluoride behaves

(a) As an acid

(b) As a base

(c) Neither as a base nor as an acid

(d) As a base and as an acid

32. Aqueous solution of $CuSO_4.5H_2O$ changes blue litmus paper to red due to **[CPMT 1979; MP PMT 1989]**

(a) Presence of Cu^{++} ions

(b) Presence of SO_4^{--} ions

(c) Hydrolysis taking place

(d) Reduction taking place

33. In the following reaction

$$HC_2O_4^- + PO_4^{---} \rightleftarrows HPO_4^{--} + C_2O_4^{--}$$

Which are the two Bronsted bases **[MP PMT 1991]**

(a) $HC_2O_4^-$ and PO_4^{---} (b) HPO_4^{--} and $C_2O_4^{--}$

(c) $HC_2O_4^-$ and HPO_4^{--} (d) PO_4^{---} and $C_2O_4^{--}$

34. Which of the following is the weakest acid **[CPMT 2001]**

(a) HF (b) HCl

(c) HBr (d) HI

35. The degree of dissociation in a weak electrolyte increases **[CBSE PMT 1989; MP PMT 1997]**

(a) On increasing dilution (b)On increasing pressure

(c) On decreasing dilution (d) None of these

36. H^+ is a **[MADT Bihar 1983]**

(a) Lewis acid (b) Lewis base

(c) Bronsted-Lowry base (d) None of the above

37. Dissociation of H_3PO_4 takes place in following steps **[CPMT 1976; NCERT 1987]**

(a) 1 (b) 2

(c) 3 (d) 4

38. The aqueous solution of disodium hydrogen phosphate is **[MADT Bihar 1982]**

(a) Acidic (b) Neutral

(c) Basic (d) None

39. Which of the following is a conjugated acid-base pair **[MADT Bihar 1984; DPMT 2001]**

(a) $HCl, NaOH$ (b) NH_4Cl, NH_4OH

(c) H_2SO_4, HSO_4^- (d) KCN, HCN

40. The solution of strong acid and weak base $(FeCl_3)$ is **[MADT Bihar 1981; CPMT 1979, 83, 84]**

(a) Acidic (b) Basic

(c) Neutral (d) None of the above

41. The conjugate acid of HPO_3^{2-} is **[EAMCET 1989]**

(a) H_3PO_4 (b) H_3PO_3

(c) $H_2PO_3^-$ (d) PO_4^{3-}

42. What name is given to the reaction between hydrogen ion and hydroxyl ion **[MP PMT 1990]**

(a) Hydrogenation (b) Hydroxylation

(c) Hydrolysis (d) Neutralization

43. Among the following, the weakest Lewis base is **[NCERT 1981]**

(a) H^- (b) OH^-

(c) Cl^- (d) HCO_3^-

44. The pKa for acid A is greater than pKa for acid B. The strong acid is **[DPMT 2000]**

(a) Acid B (b) Acid A

(c) Both A and B (d) Neither A nor B

45. The conjugate acid of NH_2^- is**[IIT 1985; Roorkee 1995; EAMCET 1997;CBSE PMT PMT 2000; MP PET 1996,2000]**

(a) NH_3 (b) NH_4^+

(c) NH_2OH (d) N_2H_4

46. Correct statement is **[CPMT 1985]**

(a) NH_4Cl gives alkaline solution in water

(b) CH_3COONa gives acidic solution in water

(c) CH_3COOH is a weak acid

(d) NH_4OH is a strong base

47. pK_a of a weak acid is defined as **[JIPMER 1999]**

(a) $\log_{10} K_a$ (b) $\frac{1}{\log_{10} K_a}$

(c) $\log_{10} \frac{1}{K_a}$ (d) $-\log_{10} \frac{1}{K_a}$

48. A salt $'X'$ is dissolved in water $(pH = 7)$, the resulting solution becomes alkaline in nature. The salt is made of **[CPMT 1983]**

(a) A strong acid and strong base

(b) A strong acid and weak base

(c) A weak acid and weak base

(d) A weak acid and strong base

49. Which one is not an acid salt **[MNR 1984]**

(a) NaH_2PO_2 (b) NaH_2PO_3

(c) NaH_2PO_4 (d) None

50. A white substance was alkaline in solution. Which of the following substances could it be **[CPMT 1989]**

(a) Fe_2O_3 (b) Na_2CO_3

(c) NH_4Cl (d) $NaNO_3$

51. An aqueous solution of ammonium carbonate is **[MP PMT 1989]**

(a) Weakly acidic (b) Weakly basic

(c) Strongly acidic (d) Neither acidic nor basic

52. $100\ ml$ of $0.2\ M\ H_2SO_4$ is added to $100\ ml$ of $0.2\ M\ NaOH$. The resulting solution will be**[BHU 1996]**

(a) Acidic (b) Basic

(c) Neutral (d) Slightly basic

53. H_3BO_3 is **[IIT Screening 2003]**

(a) Monobasic and weak Lewis acid

(b) Monobasic and weak Bronsted acid

(c) Monobasic and strong Lewis acid

(d) Tribasic and weak Bronsted acid

54. In the reaction $SnCl_2 + 2Cl^- \rightarrow SnCl_4$, Lewis acid is **[EAMCET 1987]**

(a) $SnCl_2$ (b) Cl^-

(c) $SnCl_4$ (d) None of these

55. Lewis base is **[RPMT 1997]**

(a) CO_2 (b) SO_3

(c) SO_2 (d) ROH

56. $10\ ml$ of $1\ M\ H_2SO_4$ will completely neutralise

(a) $10\ ml$ of $1\ M\ NaOH$ solution

(b) $10\ ml$ of $2\ M\ NaOH$ solution

(c) $5\ ml$ of $2\ M\ KOH$ solution

(d) $5\ ml$ of $1\ M\ Na_2CO_3$ solution

57. Which of the following compounds are diprotic **[Pb. PMT 2000]**

(a) H_2PO_5 (b) H_2S

(c) $HClO_3$ (d) H_3PO_3

58. When 100 ml of 1M $NaOH$ solution is mixed with 10 ml of 10 M H_2SO_4, the resulting mixture will be **[AMU 2002]**

(a) Acidic (b) Alkaline

(c) Neutral (d) Strongly alkaline

59. The pH indicators are **[KCET 1996]**

(a) Salts of strong acids and strong bases

(b) Salts of weak acids and weak bases

(c) Either weak acids or weak bases

(d) Either strong acids or strong bases

60. Which of the following is not Lewis acid**[BHU 1997]**

(a) BF_3 (b) $AlCl_3$

(c) $FeCl_3$ (d) PH_3

61. A solution of sodium acetate in water will**[MNR 1979]**

(a) Turn red litmus blue (b) Turn blue litmus red

(c) Not effect litmus (d) Decolourises litmus

62. Cl^- is the conjugate base of **[NCERT 1979; CPMT 1976; MP PET/PMT 1988]**

(a) $HClO_4$ (b) HCl

(c) $HOCl$ (d) $HClO_3$

63. Which of the following behaves as both Lewis and Bronsted base **[MP PMT 2003]**

(a) BF_3 (b) Cl^-

(c) CO (d) None of these

64. The conjugate acid of a strong base is a**[EAMCET 1979]**

(a) Strong acid (b) Weak acid

(c) Strong base (d) Weak base

65. Which one is the weakest acid **[DPMT 2002]**

(a) HNO_3 (b) $HClO_4$

(c) H_2SO_4 (d) HBr

66. Conjugate base of HPO_4^{2-} is **[MP PMT 1995]**

(a) PO_4^{3-} (b) $H_2PO_4^-$

(c) H_3PO_4 (d) H_4PO_3

67. Which of the following is not Lewis acid **[RPET/PMT 2002]**

(a) $FeCl_3$ (b) $AlCl_3$

(c) BCl_3 (d) NH_3

68. (i) A strong acid has a weak conjugate base

(ii) An acid is an electron pair acceptor

The above statements (i) and (ii)

(a) Correct

(b) Wrong

(c) (i) Correct and (ii) Wrong

(d) (i) Wrong and (ii) Correct

69. The *pH* is less than 7, of the solution of **[Pb. PMT 2002; MP PMT 2003]**

(a) $FeCl_3$ (b) $NaCN$

(c) $NaOH$ (d) $NaCl$

70. In the reaction $I_2 + I^- \rightarrow I_3^-$, the Lewis base is **[CPMT 1997; RPMT 2002; BCECE 2005]**

(a) I_2 (b) I^-

(c) I_3^- (d) None of these

71. The strength of an acid depends on its tendency to **[UPSEAT 2001]**

(a) Accept protons (b) Donate protons

(c) Accept electrons (d) Donate electrons

72. In Lewis acid-base theory, neutralization reaction may be considered as **[NCERT 1977]**

(a) Formation of salt and water

(b) Competition for protons by acid and base

(c) Oxidation reduction

(d) Coordinate covalent bond formation

73. The salt that forms neutral solution in water is **[EAMCET 1981]**

(a) NH_4Cl (b) $NaCl$

(c) Na_2CO_3 (d) K_3BO_3

74. Which of the following cannot act as a Lewis or Bronsted acid **[DCE 2001]**

(a) BF_3 (b) $AlCl_3$

(c) $SnCl_4$ (d) CCl_4

75. Which one of the following salts gives an acidic solution in water **[MP PET 2002]**

(a) CH_3COONa (b) NH_4Cl

(c) $NaCl$ (d) CH_3COONH_4

76. Which of the anhydrous salts when come in contact with water turns blue **[AMU 1981, 82]**

(a) Ferrous sulphate (b) Copper sulphate

(c) Zinc sulphate (d) Cobalt sulphate

77. The species among the following, which can act as an acid and a base is **[AIEEE 2002; KCET 2005]**

(a) HSO_4^- (b) SO_4^{2-}

(c) H_3O^+ (d) Cl^-

78. The strongest base from the following species is **[KCET 1996]**

(a) NH^{2-} (b) OH^-

(c) O^{2-} (d) S^{2-}

79. Which one is Lewis acid **[CPMT 1997]**

(a) Cl^- (b) Ag^+

(c) C_2H_5OH (d) S^{2-}

80. The conjugate acid of $H_2PO_4^-$ is **[MP PET 1990]**

(a) H_3PO_4 (b) $H_2PO_4^-$

(c) PO_4^{3-} (d) H_3O^+

81. The conjugate acid of $S_2O_8^{2-}$ is **[EAMCET 1984]**

(a) $H_2S_2O_8$ (b) H_2SO_4

(c) HSO_4^- (d) $HS_2O_8^-$

82. In the reaction $BCl_3 + PH_3 \rightarrow Cl_3B - PH_3$, Lewis base is **[EAMCET 1986]**

(a) BCl_3 (b) PH_3

(c) $Cl_3B - PH_3$ (d) None of these

83. Which of the following statement is true **[MP PMT 1996]**

(a) The conjugate base of a strong acid is a strong base

(b) The conjugate base of a weak acid is a strong base

(c) The conjugate base of a weak acid is a weak base

(d) The base and its conjugate acid react to form a neutral solution

84. What is the conjugate base of OH^- **[AIEEE 2005]**

(a) O_2 (b) H_2O

(c) O^- (d) O^{2-}

85. Which of the following is a Lewis base **[BHU 1995]**

(a) CH_4 (b) C_2H_5OH

(c) Acetone (d) Secondary amine

86. The correct order of acid strength is **[CBSE PMT 2005]**

(a) $HClO < HClO_2 < HClO_3 < HClO_4$

(b) $HClO_4 < HClO < HClO_2 < HClO_3$

(c) $HClO_2 < HClO_3 < HClO_4 < HClO$

(d) $HClO_4 < HClO_3 < HClO_2 < HClO$

87. The strongest acid is **[DPMT 2000]**

(a) H_3AsO_4 (b) H_3AsO_3

(c) H_3PO_3 (d) H_3PO_4

88. Which of the following is the strongest base **[Roorkee 2000]**

(a) $C_2H_5^-$ (b) $C_2H_5COO^-$

(c) $C_2H_5O^-$ (d) OH^-

89. The aqueous solution of which one of the following is basic

[MP PMT 2001]

(a) $HOCl$ (b) $NaHSO_4$

(c) NH_4NO_3 (d) $NaOCl$

90. Which of the following is the weakest base **[IIT 1980]**

(a) $NaOH$ (b) $Ca(OH)_2$

(c) NH_4OH (d) KOH

91. The suitable indicator for strong acid and weak base is

[RPMT 1997; UPSEAT 2002]

(a) Methyl orange (b) Methyl red

(c) Phenol red (d) Phenolphthalein

92. Among the following acids, the weakest one is

[NCERT 1984]

(a) HF (b) HCl

(c) HBr (d) HI

93. The compound *HCl* behaves as .... in the reaction, $HCl + HF \rightarrow H_2^+Cl + F^-$ **[JIPMER 2001]**

(a) Weak base (b) Weak acid

(c) Strong base (d) Strong acid

94. The conjugate base of a strong acid is a **[EAMCET 1978]**

(a) Strong base (b) Strong acid

(c) Weak acid (d) Weak base

95. Which among the following is strongest acid **[BHU 2005]**

(a) $H(ClO)O_2$ (b) $H(ClO)O_3$

(c) $H(ClO)O$ (d) $H(ClO)$

96. In the reaction $2H_2O \rightleftarrows H_3O^+ + OH^-$, water is

[MP PET 1989]

(a) A weak base

(b) A weak acid

(c) Both a weak acid and a weak base

(d) Neither an acid nor a base

97. In the reaction $HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$

[NCERT 1978; EAMCET 1982, 89]

(a) H_2O is the conjugate base of HCl acid

(b) Cl^- is the conjugate base of HCl acid

(c) Cl^- is the conjugate acid of H_2O base

(d) H_3O^+ is the conjugate base of HCl

98. According to Lewis concept, an acid is a substance which

[EAMCET 1981; NCERT 1981; CPMT 1986; MP PMT 1987]

(a) Accepts protons

(b) Donates protons

(c) Accepts a lone pair of electrons

(d) Donates a lone pair of electrons

99. Water is a **[KCET 2002]**

(a) Amphoteric acid (b) Aprotic solvent

(c) Protophobic solvent (d) None of these

100. Conjugate base of NH_3 is **[RPMT 2002]**

(a) $NH_4^{\oplus}$ (b) $NH_2^{\oplus}$

(c) $NH_2^{\ominus}$ (d) N_2

101. Which of the following species is an acid and also a conjugate base of another acid **[NCERT 1981]**

(a) HSO_4^- (b) H_2SO_4

(c) OH^- (d) H_3O^+

102. Which of the following is Lewis acid

[NCERT 1978; EAMCET 1987; MP PET 1994; CPMT 1990; AFMC 1997; RPMT 1999]

(a) BF_3 (b) Cl^-

(c) H_2O (d) NH_3

103. According to Bronsted-Lowry concept, base is a substance which is **[NCERT 1982; RPMT 1997 AFMC 1999]**

(a) A proton donor

(b) An electron pair acceptor

(c) A proton acceptor

(d) An electron pair donor

104. Which of the following is known as hydronium ion

[NCERT 1976]

(a) H^+ (b) H_2O^+

(c) H_3O^+ (d) $H_2O_2^+$

105. An aqueous solution of aluminium sulphate would show

[NCERT 1977]

(a) An acidic reaction

(b) A neutral reaction

(c) A basic reaction

(d) Both acidic and basic reaction

106. The aqueous solution of aluminium chloride is acidic due to

[MNR 1986, 88]

(a) Cation hydrolysis

(b) Anion hydrolysis

(c) Hydrolysis of both anion and cation

(d) Dissociation

107. HSO_4^- is the conjugate base of **[NCERT 1977]**

(a) H^+ (b) H_2SO_3

(c) SO_4^{2-} (d) H_2SO_4

108. An acid is a compound which furnishes (Bronsted-Lowry concept) **[EAMCET 1975]**

(a) An electron (b) A proton

(c) An electron and a proton (d) None of the above

109. The conjugate base of sulphuric acid is

[EAMCET 1974]

(a) Sodium hydroxide (b) Hydrochloric acid
(c) Bisulphate ion (d) Barium hydroxide

110. Which is strongest Lewis base **[CPMT 1994]**
(a) SbH_3 (b) AsH_3
(c) PH_3 (d) NH_3

111. According to Bronsted principle, an aqueous solution of HNO_3 will contain **[MP PMT/PET 1988]**
(a) NO_2^- (b) NO_3^-
(c) NO_2^+ (d) NO^+

112. Aqueous solution of an acid is characterised by the presence of **[NCERT 1977]**
(a) H^+ ions (b) H_2^+ ions
(c) H_3O^+ ions (d) H_4O^+ ions

113. Ammonium ion is **[RPMT 1999; KCET 2002]**
(a) Neither an acid nor base (b)Both an acid and a base
(c) A conjugate acid (d) A conjugate base

114. Aqueous solution of $AlCl_3$ is **[RPMT 2002]**
(a) Acidic (b) Basic
(c) Amphoteric (d) None of these

115. The species which acts as a Lewis acid but not a Bronsted acid is **[MP PMT 1999; Kurukshetra CET 2002]**
(a) NH_2^- (b) O^{2-}
(c) BF_3 (d) OH^-

116. Among the following, the weakest base is**[MP PMT 2002]**
(a) H^- (b) CH_3^-
(c) CH_3O^- (d) Cl^-

117. Which of the following is not lewis base **[EAMCET 1975; RPMT 2002]**
(a) NH_3 (b) PH_3
(c) $(CH_3)_3N$ (d) HN_3

118. pK_a value of the strongest acid among the following is **[MP PMT 1990]**
(a) 3.0 (b) 4.5
(c) 1.0 (d) 2.0

119. The most acidic compound in water is **[CBSE PMT 2001]**
(a) $AlCl_3$ (b) $BeCl_2$
(c) $FeCl_3$ (d) None of these

120. BF_3 is used as a catalyst in several industrial processes due to its **[Kerala (Med.) 2002]**
(a) Strong reducing agent
(b) Weak reducing agent
(c) Strong Lewis acid nature
(d) Weak Lewis acid character

121. The aqueous solution of $AlCl_3$ is acidic due to the hydrolysis of **[UPSEAT 2001, 02]**
(a) Aluminium ion
(b) Chloride ion
(c) Both aluminium and chloride ion
(d) None of these

122. Which shows weak ionisation in water **[MH CET 2001]**
(a) H_2SO_4 (b) $NaCl$
(c) HNO_3 (d) NH_3

123. An organic dye, cosine used to detect end point of precipitation titration by adsorption is called **[MH CET 1999]**
(a) Absorption indicator (b) Adsorption indicator
(c) Chemical indicator (d) None of these

124. The indicator used in the titration of iodine against sodium thiosulphate is **[AFMC 2002]**
(a) Starch (b) $K_3Fe(CN)_6$
(c) K_2CrO_4 (d) Potassium

125. Phenolphthalein does not act as an indicator for the titration between **[Pb. PMT 2002]**
(a) $NaOH$ and CH_3COOH
(b) $H_2C_2O_4$ and $KMnO_4$
(c) $Ba(OH)_2$ and HCl
(d) KOH and H_2SO_4

126. Which is not example of Bronsted Lowry theory **[AIEEE 2003]**
(a) $AlCl_3$ (b) H_2SO_4
(c) SO_2 (d) HNO_3

127. An aqueous solution of sodium carbonate is alkaline because sodium carbonate is a salt of **[MP PET 20**
(a) Weak acid and weak base
(b) Strong acid and weak base
(c) Weak acid and strong base
(d) Strong acid and strong base

128. The acid that results when a base accepts a proton is called **[Kerala (Med.) 2002]**
(a) Conjugate base of the acid
(b) Conjugate protonated base
(c) Lewis base
(d) Conjugate acid of the base
(e) None of these

129. Ammonia gas dissolves in water to form NH_4OH. In this reaction water acts as**[KCET (Engg./Med.) 1999]**
(a) A conjugate base (b) A non-polar solvent
(c) An acid (d) A base

130. The conjugate base in the following reaction

$H_2SO_4 + H_2O \rightleftarrows H_3O^+ + HSO_4^-$ **[DCE 1999]**

(a) H_2O (b) HSO_4^-

(c) H_3O^+ (d) SO_2

131. An aqueous solution of aluminium sulphate shows **[DPMT 2001]**

(a) A basic nature

(b) An acidic nature

(c) A neutral nature

(d) Both acidic and basic nature

132. Neutralization of an acid with a base invariably results in the production of

(a) H_3O^+ (b) OH^-

(c) H_2O (d) H^+ and OH^-

133. The conjugate acid of HPO_4^{2-} is **[MP PMT 1987, 90, 91; EAMCET 1993]**

(a) $H_2PO_4^-$ (b) PO_4^{3-}

(c) H_3PO_4 (d) H_3PO_3

134. Which of the following is not used as a Lewis acid **[MP PET 2000]**

(a) $SnCl_4$ (b) $FeCl_3$

(c) KCl (d) BF_3

135. Orthoboric acid in aqueous medium is **[AMU 2000]**

(a) Monobasic (b) Dibasic

(c) Tribasic (d) All are correct

136. According to Lewis concept which one of the following is not a base**[MP PET/PMT 1988; Pb. CET 2003]**

(a) OH^- (b) H_2O

(c) Ag^+ (d) NH_3

137. The aqueous solution of ammonium chloride is **[CPMT 1987]**

(a) Neutral (b) Basic

(c) Acidic (d) Amphoteric

138. In the process $BCl_3 + PH_3 \to BCl_3 : PH_3$ The Lewis acid is **[RPMT 2000]**

(a) PH_3 (b) BCl_3

(c) Both (d) None

139. The conjugate acid of NH_3 is **[BHU Varanasi 1999; Pb. PMT 2004]**

(a) NH_3 (b) NH_4^+

(c) N_2H_4 (d) NH_2OH

140. Which halide of nitrogen is least basic **[DPMT 2001]**

(a) NBr_3 (b) NI_3

(c) NCl_3 (d) NF_3

141. When $FeCl_3$ gets soluble in water, then its solution represents which of the characteristics **[MP PET/PMT 1988]**

(a) Amphoteric (b) Acidic

(c) Basic (d) Neutral

142. Lewis acid are those substances **[MP PMT 1987]**

(a) Which accept electron pair

(b) Which provide H^+ ion in the solution

(c) Which give electron pair

(d) Which accept OH^- ion

143. The conjugate base of HCO_3^- is

(a) H_2CO_3 (b) CO_3^{2-}

(c) CO_2 (d) H_2O

144. In the reaction $NH_3 + BF_3 \rightleftharpoons NH_3 \to BF_3$, BF_3 is **[CPMT 1983]** **[MP PMT 1989; MHCET 2001]**

(a) Lewis acid

(b) Lewis base

(c) Neither Lewis acid nor Lewis base

(d) Lewis acid and Lewis base both

145. The strongest Lewis base in the following **[MP PET/PMT 1988]**

(a) CH_3^- (b) F^-

(c) NH_2^- (d) OH^-

146. The aqueous solution of $CuSO_4$ is **[CPMT 1985]**

(a) Acidic (b) Basic

(c) Neutral (d) Amphoteric

147. The acid having the highest pK_a value among the following is **[JIPMER 2002]**

(a) $HCOOH$ (b) CH_3COOH

(c) $ClCH_2COOH$ (d) FCH_2COOH

148. The indicator used in the titration of sodium carbonate with sulphuric acid is **[DPMT 2001]**

(a) Phenolphthalein

(b) Methyl orange

(c) Potassium ferrocynide

(d) Potassium ferricynide

149. According to Bronsted law, water is a/an **[MP PET/PMT 1988]**

(a) Base (b) Acid

(c) Acid and base both (d) Salt

150. Which of the following can give base OH^- **[MP PET/PMT 1988]**

(a) H_2O (b) H_3O^+

(c) H_2 (d) HCl

151. Conjugate base of HBr is **[MP PET/PMT 1988]**

(a) H_2Br^+ (b) H^+

(c) Br^- (d) Br^+

152. Molar heat of neutralization of $NaOH$ with HCl in comparison to that of KOH with HNO_3 is **[MP PMT 1989**

(a) Less (b) More

(c) Equal (d) Depends on pressure

153. Which of the following is not a Lewis acid **[J & K 2005]**

(a) BF_3 (b) $AlCl_3$

(c) HCl (d) $LiAlH_4$

154. The solvent which neither accepts proton nor donates proton is called **[RPMT 2000]**

(a) Amphoteric (b) Neutral

(c) Aprotic (d) Amphiprotic

155. For the reaction in aqueous solution $Zn^{2+} + X^- \rightleftarrows ZnX^+$, the K_{eq} is greatest when X is **[Pb. PMT 1998]**

(a) F^- (b) NO_3^-

(c) ClO_4^- (d) I^-

156. Why are strong acids generally used as standard solutions in acid-base titrations **[Pb. PMT 1998]**

(a) The pH at the equivalence point will always be 7

(b) They can be used to titrate both strong and weak bases

(c) Strong acids form more stable solutions than weak acids

(d) The salts of strong acids do not hydrolysed

157. For an aqueous solution, the characteristic species of acid is

[RPMT 1999]

(a) H^+ ion (b) H_3O^+ ion

(c) H_2^+ ion (d) H_4O^+ ion

158. Which is a Lewis base **[CPMT 1988; JEE Orissa 2004]**

(a) B_2H_6 (b) $LiAlH_4$

(c) AlH_3 (d) NH_3

159. For a weak acid, the incorrect statement is **[Pb. PMT 2004]**

(a) Its dissociation constant is low

(b) Its pK_a is very low

(c) It is partially dissociated

(d) Solution of its sodium salt is alkaline in water

160. Boron halides behave as Lewis acids, because of their

[CBSE PMT 1996; BHU 2004]

(a) Ionic nature (b) Acidic nature

(c) Covalent nature (d) Electron deficient nature

161. Would gaseous HCl be considered as an Arrhenius acid

[UPSEAT 2004]

(a) Yes

(b) No

(c) Not known

(d) Gaseous HCl does not exist

162. Which one of the following is called amphoteric solvent

[UPSEAT 2004]

(a) Ammonium hydroxide (b) Chloroform

(c) Benzene (d) Water

163. Strongest conjugate base is **[DPMT 2004]**

(a) Cl^- (b) Br^-

(c) F^- (d) I^-

164. The conjugate base of $H_2PO_4^-$ is **[AIEEE 2004]**

(a) H_3PO_4 (b) P_2O_5

(c) PO_4^{3-} (d) HPO_4^{2-}

165. Conjugate base of HSO_4^- is **[MH CET 2004]**

(a) SO_4^{2-} (b) H_2SO_4

(c) $H_3SO_4^+$ (d) None of these

Common ion effect, Isohydric solutions, Solubility product, Ionic product of water and Salt hydrolysis

1. The expression for the solubility product of $Al_2(SO_4)_3$ is

(a) $K_{sp} = [Al^{3+}](SO_4^{2-}]$ (b) $K_{sp} = [Al^{3+}]^2\ (SO_4^{2-}]^3$

(c) $K_{sp} = [Al^{3+}]^3\ (SO_4^{2-}]^2$ (d) $K_{sp} = [Al^{3+}]^2\ (SO_4^{2-}]^2$

2. On addition of ammonium chloride to a solution of ammonium hydroxide **[CPMT 1976, 80, 81, 99; NCERT 1976, 77; MP PMT 1989, 99; DPMT 1983]**

(a) Dissociation of NH_4OH increases

(b) Concentration of OH^- increases

(c) Concentration of OH^- decreases

(d) Concentration of NH_4^+ and OH^- increases

3. The solubility product of a salt having general formula MX_2, in water is : 4×10^{-12}. The concentration of M^{2+} ions in the aqueous solution of the salt is **[AIEEE 2005]**

(a) $2.0 \times 10^{-6} M$ (b) $1.0 \times 10^{-4} M$

(c) $1.6 \times 10^{-4} M$ (d) $4.0 \times 10^{-10} M$

4. In a saturated solution of electrolyte, the ionic product of their concentration are constant at constant temperature and this constant for electrolyte is known as **[CPMT 1983]**

(a) Ionic product (b) Solubility product

(c) Ionization constant (d) Dissociation constant

5. If the solubility product K_{sp} of a sparingly soluble salt MX_2 at $25°C$ is 1.0×10^{-11}, the solubility of the salt in *mole litre*$^{-1}$ at this temperature will be **[RPMT 2**

(a) 2.46×10^{14} (b) 1.36×10^{-4}

(c) 2.60×10^{-7} (d) 1.20×10^{-10}

6. The unit of ionic product of water K_w are

[UPSEAT 2001, 02]

(a) $Mol^{-1}L^{-1}$ (b) $Mol^{-2}L^{-2}$

(c) $Mol^{-2}L^{-1}$ (d) $Mol^{2}L^{-2}$

7. A solution which is $10^{-3} M$ each in $Mn^{2+}, Fe^{2+}, Zn^{2+}$ and Hg^{2+} is treated with $10^{-16} M$ sulphide ion. If

K_{sp} of MnS, FeS, ZnS and HgS are $10^{-15}, 10^{-23}, 10^{-20}$ and 10^{-54} respectively, which one will precipitate first

[IIT Screening 2003]

(a) FeS (b) MgS

(c) HgS (d) ZnS

8. Let the solubility of an aqueous solution of $Mg(OH)_2$ be *x* then its k_{sp} is **[AIEEE 2002]**

(a) $4x^3$ (b) $108\,x^5$

(c) $27x^4$ (d) $9x$

9. The solubility product of $BaSO_4$ at $25°C$ is 1.0×10^{-9}. What would be the concentration of H_2SO_4 necessary to precipitate $BaSO_4$ from a solution of $0.01\,M\,Ba^{2+}$ ions

[RPMT 1999]

(a) 10^{-9} (b) 10^{-8}

(c) 10^{-7} (d) 10^{-6}

10. The solubility in water of a sparingly soluble salt AB_2 is $1.0 \times 10^{-5} mol\,l^{-1}$. Its solubility product number will be

[AIEEE 2003]

(a) 4×10^{-15} (b) 4×10^{-10}

(c) 1×10^{-15} (d) 1×10^{-10}

11. The solubility of CaF_2 is *a moles/litre*. Then its solubility product is ..... **[Orissa JEE 2002]**

(a) s^2 (b) $4s^3$

(c) $3s^2$ (d) s^3

12. On passing a current of HCl gas in a saturated solution of $NaCl$, the solubility of $NaCl$

[CPMT 1989; CBSE PMT 1989]

(a) Increases (b) Decreases

(c) Remains unchanged (d) $NaCl$ decomposes

13. Which is the correct representation of the solubility product constant of Ag_2CrO_4 **[NCERT 1974, 75]**

(a) $[Ag^+]^2\,[CrO_4^{-2}]$ (b) $[Ag^+]\,[CrO_4^{-2}]$

(c) $[2Ag^+]\,[CrO_4^{-2}]$ (d) $[2Ag^+]^2\,[CrO_4^{-2}]$

14. The solubility of CaF_2 is 2×10^{-4} *moles / litre*. Its solubilily product (K_{sp}) is**[NCERT 1981; BHU 1983, 86;**

MP PET 1992; CBSE PMT 1999]

(a) 2.0×10^{-4} (b) 4.0×10^{-3}

(c) 8.0×10^{-12} (d) 3.2×10^{-11}

15. Solubility product of a sulphide *MS* is 3×10^{-25} and that of another sulphide *NS* is 4×10^{-40}. In ammoniacal solution

[NCERT 1981]

(a) Only *NS* gets precipitated

(b) Only *MS* gets precipitated

(c) No sulphide precipitates

(d) Both sulphides precipitate

16. Which of the following salts when dissolved in water will get hydrolysed

[MNR 1985; CPMT 1989; CBSE PMT 1989; MP PET 1999]

(a) $NaCl$ (b) NH_4Cl

(c) KCl (d) Na_2SO_4

17. The aqueous solution of $FeCl_3$ is acidic due to

[CPMT 1972, 79, 83, 84;
MP PET/PMT 1988; RPMT 2000]

(a) Acidic impurities (b) Ionisation

(c) Hydrolysis (d) Dissociation

18. A precipitate of $AgCl$ is formed when equal volumes of the following are mixed. [K_{sp} for $AgCl = 10^{-10}$] **[KCET 2005]**

(a) $10^{-4}\,M\,AgNO_3$ and $10^{-7}\,M\,HCl$

(b) $10^{-5}\,M\,AgNO_3$ and $10^{-6}\,M\,HCl$

(c) $10^{-5}\,M\,AgNO_3$ and $10^{-4}\,M\,HCl$

(d) $10^{-6}\,M\,AgNO_3$ and $10^{-6}\,M\,HCl$

19. The solubility of silver chromate in 0.01 *M* K_2CrO_4 is $2 \times 10^{-8}\,mol\,dm^{-3}$. The solubility product of silver chromate will be

(a) 8×10^{-24} (b) 16×10^{-24}

(c) 1.6×10^{-18} (d) 16×10^{-18}

20. Some salts although containing two different metallic elements give test for only one of them in solution. Such salts are **[MNR 1979]**

(a) Double salts (b) Normal salts

(c) Complex salts (d) Basic salts

21. What is the pH value of $\frac{N}{1000} KOH$ solution

[CPMT 1975; MNR 1986, 91; Pb. CET 2004]

(a) 10^{-11} (b) 3

(c) 2 (d) 11

22. Mohr's salt is a **[MNR 1986]**

(a) Normal salt (b) Acid salt

(c) Basic salt (d) Double salt

23. Aqueous solution of sodium acetate is

[MNR 1978; CPMT 1971, 80, 81;
MADT Bihar 1982; MP PMT 1985;]

(a) Neutral (b) Weakly acidic

(c) Strongly acidic (d) Alkaline

24. Which is the correct alternate for hydrolysis constant of NH_4CN **[CBSE PMT 1989]**

(a) $\sqrt{\frac{K_w}{K_a}}$ (b) $\frac{K_w}{K_a \times K_b}$

(c) $\sqrt{\frac{K_b}{c}}$ (d) $\frac{K_a}{K_b}$

25. Which of the following salts undergoes hydrolysis

[CPMT 1972, 74, 78; DPMT 1985]

(a) CH_3COONa (b) KNO_3

(c) $NaCl$ (d) K_2SO_4

26. What will happen if CCl_4 is treated with $AgNO_3$ **[DPMT 1983]**

(a) A white ppt. of $AgCl$ will form

(b) NO_2 will be evolved

(c) CCl_4 will dissolve in $AgNO_3$

(d) Nothing will happen

27. The correct representation for solubility product of SnS_2 is **[CPMT 1977; MP PET 1999;RPMT 2000]**

(a) $[Sn^{4+}][S^{2-}]^2$ (b) $[Sn^{2+}][S^{2-}]^2$

(c) $[Sn^{2+}][2S^{2-}]$ (d) $[Sn^{4+}][2S^{2-}]^2$

28. A precipitate of calcium oxalate will not dissolve in **[CPMT 1971, 89; IIT 1986]**

(a) HCl (b) HNO_3

(c) Aquaregia (d) CH_3COOH

29. Baking soda is **[RPMT 2000]**

(a) Basic salt (b) Acidic salt

(c) Complex salt (d) Double salt

30. Which one of the following substances will be a mixed salt **[DPMT 1982; CPMT 1972]**

(a) $NaHCO_3$

(b) $Ca(OCl)Cl$

(c) $K_2SO_4\, Al_2(SO_4)_3.24H_2O$

(d) $Mg(OH)Br$

31. Solubility product of $BaCl_2$ is 4×10^{-9}. Its solubility in moles/litre would be **[AFMC 1982; Roorkee 1990; BHU 2000]**

(a) 1×10^{-3} (b) 1×10^{-9}

(c) 4×10^{-27} (d) 1×10^{-27}

32. Which hydroxide will have lowest value of solubility product at normal temperature (25^oC) **[IIT 1990; RPMT 1997]**

(a) $Mg(OH)_2$ (b) $Ca(OH)_2$

(c) $Ba(OH)_2$ (d) $Be(OH)_2$

33. Which will not be hydrolysed **[MP PMT 1989]**

(a) Potassium nitrate

(b) Potassium cyanide

(c) Potassium succinate

(d) Potassium carbonate

34. Which pair will show common ion effect **[MP PMT 1990, 99; Pb. PMT 2001]**

(a) $BaCl_2 + Ba(NO_3)_2$ (b) $NaCl + HCl$

(c) $NH_4OH + NH_4Cl$ (d) $AgCN + KCN$

35. Which is least soluble in water **[UPSEAT 1999]**

(a) $AgCl$ (b) AgF

(c) AgI (d) Ag_2S

36. A white salt is readily soluble in water and gives a colourless solution with a pH of about 9. The salt would be **[Pb. PMT 1998]**

(a) NH_4NO_3 (b) CH_3COONa

(c) CH_3COONH_4 (d) $CaCO_3$

37. If acetic acid mixed with sodium acetate, then H^+ ion concentration will be **[Roorkee 1995]**

(a) Increased (b) Decreased

(c) Remains unchanged (d) pH decreased

38. Solubility of $AgCl$ will be minimum in **[CBSE PMT 1995]**

(a) $0.001\, M\, AgNO_3$ (b) Pure water

(c) $0.01\, M\, CaCl_2$ (d) $0.01\, M\, NaCl$

39. In absence of formation of complex ions by the addition of a common ion, the solubility of a given salt is **[BHU 1979]**

(a) Increased

(b) Decreased

(c) Unaffected

(d) First increased and then decreased

40. At 298 K, the solubility product of $PbCl_2$ is 1.0×10^{-6}. What will be the solubility of $PbCl_2$ in moles/litre **[MP PMT 1990; CPMT 1985, 96]**

(a) 6.3×10^{-3} (b) 1.0×10^{-3}

(c) 3.0×10^{-3} (d) 4.6×10^{-14}

41. Solubility product is **[CET Pune 1998]**

(a) The ionic product of an electrolyte in its saturated solution

(b) The product of the solubilities of the ions of the electrolyte

(c) The product of solubilities of the salts

(d) The product of the concentration of the ions

42. Ionic product of water increases, if **[AMU 1983; MP PET 1986; MP PET/PMT 1988; JIPMER 2002]**

(a) Pressure is reduced (b) H^+ is added

(c) OH^- is added (d) Temperature increases

43. Which one is a mixed salt **[DPMT 1985]**

(a) $NaHSO_4$ (b) $NaKSO_4$

(c) $K_4Fe(CN)_6$ (d) $Mg(OH)Cl$

44. If K_{sp} for $HgSO_4$ is 6.4×10^{-5}, then solubility of the salt is **[AFMC 1997; KCET 2000; CPMT 2000; JIPMER 2001]**

(a) 8×10^{-3} (b) 8×10^{-6}

(c) 6.4×10^{-5} (d) 6.4×10^{-3}

45. The solubility of $BaSO_4$ in water is $2.33\times10^{-3}\, gm/litre$. Its solubility product will be (molecular weight of $BaSO_4 = 233$) **[AIIMS 1998]**

(a) 1×10^{-5} (b) 1×10^{-10}
(c) 1×10^{-15} (d) 1×10^{-20}

46. The solubility of $AgCl$ in $0.2\ M\ NaCl$ solution (K_{sp} for $AgCl = 1.20\times10^{-10}$) is **[MP PET 1996]**
(a) $0.2\ M$ (b) $1.2\times10^{-10}\ M$
(c) $0.2\times10^{-10}\ M$ (d) $6.0\times10^{-10}\ M$

47. The solubility of AgI in NaI solution is less than that in pure water because **[UPSEAT 2001]**
(a) AgI forms complex with NaI
(b) Of common ion effect
(c) Solubility product of AgI is less than that of NaI
(d) The temperature of the solution decreases

48. The solubility product of $BaSO_4$ is 1.5×10^{-9}. The precipitation in a 0.01 $M\ Ba^{2+}$ solution will start, on adding H_2SO_4 of concentration **[CPMT 1988]**
(a) $10^{-9}M$ (b) $10^{-8}M$
(c) $10^{-7}M$ (d) $10^{-6}M$

49. At 20^oC, the Ag^+ ion concentration in a saturated solution of Ag_2CrO_4 is 1.5×10^{-4} *mole / litre*. At 20^oC, the solubility product of Ag_2CrO_4 would be **[MP PET 1997; MP PMT 1999]**
(a) 3.3750×10^{-12} (b) 1.6875×10^{-10}
(c) 1.6875×10^{-12} (d) 1.6875×10^{-11}

50. The solubility of $PbCl_2$ is **[MP PMT 1995; DCE 1999]**
(a) $\sqrt{K_{sp}}$ (b) $\sqrt[3]{K_{sp}}$
(c) $\sqrt[3]{\frac{K_{sp}}{4}}$ (d) $\sqrt{8K_{sp}}$

51. The solubility product of $AgCl$ is 1.44×10^{-4} at 100^oC. The solubility of silver chloride in boiling water may be **[MP PMT 1994; Bihar MEE 1998]**
(a) $0.72\times10^{-4}\ M$ (b) $1.20\times10^{-2}\ M$
(c) $0.72\times10^{-2}\ M$ (d) $1.20\times10^{-4}\ M$

52. If the solubility of a sparingly soluble salt of the type BA_2 (giving three ions on dissociation of a molecule) is x moles per litre, then its solubility product is given by **[BHU 1987]**
(a) x^2 (b) $2x^3$
(c) $4x^2$ (d) $4x^3$

53. The solubility product of Ag_2CrO_4 is 32×10^{-12}. What is the concentration of CrO_4^- ions in that solution **[BHU 1997; DPMT 2004]**
(a) $2\times10^{-4}\ m/s$ (b) $16\times10^{-4}\ m/s$
(c) $8\times10^{-4}\ m/s$ (d) $8\times10^{-8}\ m/s$

54. The addition of HCl will not suppress the ionization of **[MP PET 1993]**
(a) Acetic acid (b) Benzoic acid
(c) H_2S (d) Sulphuric acid

55. On the addition of a solution containing CrO_4^{2-} ions to the solution of Ba^{2+}, Sr^{2+} and Ca^{2+} ions, the precipitate obtained first will be of **[MP PMT 1993]**
(a) $CaCrO_4$ (b) $SrCrO_4$
(c) $BaCrO_4$ (d) Mixture of (a), (b), (c)

56. The solubility product of a sparingly soluble salt AB at room temperature is 1.21×10^{-6}. Its molar solubility is **[CPMT 1987; MP PET 2001]**
(a) 1.21×10^{-6} (b) 1.21×10^{-3}
(c) 1.1×10^{-4} (d) 1.1×10^{-3}

57. The precipitation occurs if ionic concentration is **[AFMC 1995; J & K 2005]**
(a) Less than solubility product
(b) More than solubility product
(c) Equal to solubility product
(d) None of these

58. If S and K_{sp} are respectively solubility and solubility product of a sparingly soluble binary electrolyte, then **[CPMT 1988; MP PMT 1999]**
(a) $S = K_{sp}$ (b) $S = K_{sp}^2$
(c) $S = \sqrt{K_{sp}}$ (d) $S = \frac{1}{2}K_{sp}$

59. Any precipitate is formed when **[AIIMS 1982; DPMT 1985; KCET 1999; MP PMT 2004]**
(a) Solution becomes saturated
(b) The value of ionic product is less that than the value of solubility product
(c) The value of ionic product is equal than the value of solubility product
(d) The value of ionic product is greater than the value of solubility product

60. The solubility product of $AgCl$ is 4.0×10^{-10} at 298 K. The solubility of $AgCl$ in $0.04\,m\ CaCl_2$ will be **[KCET 1996]**
(a) $2.0\times10^{-5}\ m$ (b) $1.0\times10^{-4}\ m$
(c) $5.0\times10^{-9}\ m$ (d) $2.2\times10^{-4}\ m$

61. Hydrolysis of sodium acetate will give **[MNR 1978]**
(a) Acidic solution (b) Basic solution
(c) Neutral solution (d) Normal solution

62. If the solubility product of $BaSO_4$ is 1.5×10^{-9} in water, its solubility in moles per litre, is

[BHU 1995; MP PET 1995; UPSEAT 204]

(a) 1.5×10^{-9} (b) 3.9×10^{-5}

(c) 7.5×10^{-5} (d) 1.5×10^{-5}

63. On passing H_2S gas through a highly acidic solution containing Cd^{2+} ions, CdS is not precipitated because

(a) Of common ion effect

(b) The solubility of CdS is low

(c) Cd^{2+} ions do not form complex with H_2S

(d) The solubility product of CdS is low

64. Which of the following will occur if a 0.1 M solution of a weak acid is diluted to $0.01M$ at constant temperature

[UPSEAT 2001, 02]

(a) $[H^+]$ will decrease to 0.01 M

(b) pH will decrease

(c) Percentage ionization will increase

(d) K_a will increase

65. If solubility of calcium hydroxide is $\sqrt{3}$, then its solubility product will be **[MH CET 2002]**

(a) 27 (b) 3

(c) 9 (d) $12\sqrt{3}$

66. When NH_4Cl is added to NH_4OH solution, the dissociation of ammonium hydroxide is reduced. It is due to

[MP PMT 1993]

(a) Common ion effect (b) Hydrolysis

(c) Oxidation (d) Reduction

67. At 298 K, the solubility of $PbCl_2$ is $2 \times 10^{-2} mol/lit$, then k_{sp} = **[RPMT 2002]**

(a) 1×10^{-7} (b) 3.2×10^{-7}

(c) 1×10^{-5} (d) 3.2×10^{-5}

68. The solubility product of silver sulphide is 3.2×10^{-11}. Its solubility at the experimental temperature is

(a) 2×10^{-4} *moles per litre*

(b) 6×10^{-6} *moles per litre*

(c) 1.2×10^{-5} *moles per litre*

(d) 8×10^{-4} *moles per litre*

69. The solubility of $CaCO_3$ in water is $3.05 \times 10^{-4} moles\ /\ litre$. Its solubility product will be **[MP PMT 1997]**

(a) 3.05×10^{-4} (b) 10

(c) 6.1×10^{-4} (d) 9.3×10^{-8}

70. Solubility of BaF_2 in a solution $Ba(NO_3)_2$ will be represents by the concentration term

[UPSEAT 2001, 02; CPMT 2002]

(a) $[Ba^{++}]$ (b) $[F^-]$

(c) $\frac{1}{2}[F^-]$ (d) $2[NO_3^-]$

71. The solubility of $PbCl_2$ at 25^oC is 6.3×10^{-3} *mole/litre*. Its solubility product at that temperature is

[NCERT 1979; CPMT 1985]

(a) $(6.3 \times 10^{-3}) \times (6.3 \times 10^{-3})$

(b) $(6.3 \times 10^{-3}) \times (12.6 \times 10^{-3})$

(c) $(6.3 \times 10^{-3}) \times (12.6 \times 10^{-3})^2$

(d) $(12.6 \times 10^{-3}) \times (12.6 \times 10^{-3})$

72. Which of the following cannot be hydrolysed

[MP PMT 1996]

(a) A salt of weak acid and strong base

(b) A salt of strong acid and weak base

(c) A salt of weak acid and weak base

(d) A salt of strong acid and strong base

73. pH of water is 7. When a substance Y is dissolved in water, the pH becomes 13. The substance Y is a salt of

[MP PMT 1997]

(a) Strong acid and strong base

(b) Weak acid and weak base

(c) Strong acid and weak base

(d) Weak acid and strong base

74. Which is a basic salt **[MP PMT 1985]**

(a) PbS (b) $PbCO_3$

(c) $PbSO_4$ (d) $2PbCO_3.Pb(OH)_2$

75. The saturated solution of Ag_2SO_4 is $2.5 \times 10^{-2}\ M$. Its solubility product (K_{sp}) is **[NCERT 1980]**

(a) 62.5×10^{-6} (b) 6.25×10^{-4}

(c) 15.625×10^{-6} (d) 3.125×10^{-6}

76. K_{sp} for sodium chloride is $36\ mol^2 / litre^2$. The solubility of sodium chloride is **[BHU 1981]**

(a) $\frac{1}{36}$ (b) $\frac{1}{6}$

(c) 6 (d) 3600

77. Sodium chloride is purified by passing hydrogen chloride gas in an impure solution of sodium chloride. It is based on

[MP PMT 1996]

(a) Buffer action (b) Common ion effect

(c) Association of salt (d) Hydrolysis of salt

78. If the concentration of lead iodide in its saturated solution at 25^oC be 2×10^{-3} moles per litre, then its solubility product is **[CPMT 1984]**

(a) 4×10^{-6} (b) 8×10^{-12}

(c) 6×10^{-9} (d) 32×10^{-9}

79. The precipitate of $CaF_2\ (K_{sp} = 1.7 \times 10^{-10})$ is obtained when equal volumes of the following are mixed

[IIT 1992; UPSEAT 2000]

(a) $10^{-4}M\,Ca^{2+}+10^{-4}M\,F^-$

(b) $10^{-2}M\,Ca^{2+}+10^{-3}M\,F^-$

(c) Both

(d) None of these

80. In the reaction: $H_2S \rightleftarrows 2H^+ + S^{--}$, when NH_4OH is added, then **[KCET (Med.) 1999; AFMC 2000]**

(a) S^{--} is precipitate

(b) No action takes places

(c) Concentration of S^{--} decreases

(d) Concentration of S^{--} increases

81. What is the minimum concentration of SO_4^{2-} required to precipitate $BaSO_4$ in a solution containing $1.0\times10^{-4}mol\ Ba^{2+}$? (K_{sp} for $BaSO_4$ is 4×10^{-10}) **[MP PMT 2000]**

(a) $4\times10^{-10}M$ (b) $2\times10^{-7}M$

(c) $4\times10^{-6}M$ (d) $2\times10^{-3}M$

82. Solubility product for salt AB_2 is 4×10^{-12}. Calculate solubility **[RPET 2003]**

(a) $1\times10^{-3}gm\ mol\,/\,litre$

(b) $1\times10^{-5}gm\ mol\,/\,litre$

(c) $1\times10^{-4}gm\ mol\,/\,litre$

(d) $1\times10^{-2}gm\ mol\,/\,litre$

83. Solubility product of a salt AB is 1×10^{-8} in a solution in which concentration of A is $10^{-3}M$. The salt will precipitate when the concentration of B becomes more than **[MP PET 1990; KCET 2003]**

(a) $10^{-4}M$ (b) $10^{-7}M$

(c) $10^{-6}M$ (d) $10^{-5}M$

84. At equilibrium, if to a saturated solution of $NaCl, HCl$ is passed, $NaCl$ gets precipitated because **[RPMT 1999]**

(a) HCl is a strong acid

(b) Solubility of $NaCl$ decreases

(c) Ionic product of $NaCl$ becomes greater than its K_{sp}

(d) HCl is a weak acid

85. The solubility product of $BaSO_4$ is 1.3×10^{-9}. The solubility of this salt in pure water will be **[MP PET 2002]**

(a) $1.69\times10^{-9}mol\ litre^{-1}$ (b) $1.69\times10^{-18}mol\ litre^{-1}$

(c) $3.6\times10^{-18}mol\ litre^{-1}$ (d) $3.6\times10^{-5}mol\ litre^{-1}$

86. The solubility product of $AgCl$ under standard conditions of temperature is given by **[Kerala (Med.) 2003]**

(a) 1.6×10^{-5} (b) 1.5×10^{-8}

(c) 3.2×10^{-10} (d) 1.5×10^{-10}

87. An aqueous solution of CH_3COONa will be **[MP PET 2001]**

(a) Acidic (b) Alkaline

(c) Neutral (d) None of these

88. In which of the following salt hydrolysis takes place **[CPMT 1974, 78]**

(a) KCl (b) $NaNO_3$

(c) CH_3COOK (d) K_2SO_4

89. At 90^oC pure water has $[H_3O^+]=10^{-6}M$, the value of K_w at this temperature will be **[IIT 1981; MNR 1990; CBSE PMT 1993; UPSEAT 1999]**

(a) 10^{-6} (b) 10^{-12}

(c) 10^{-14} (d) 10^{-8}

90. Solubility of MX_2 type electrolyte is $0.5\times10^{-4}mole/litre$. The value of K_{sp} of the electrolyte is **[CBSE PMT 2002]**

(a) 5×10^{-13} (b) 25×10^{-10}

(c) 1.25×10^{-13} (d) 5×10^{12}

91. According to the reaction $PbCl_2 = Pb^2 + 2Cl^-$, the solubility coefficient of $PbCl_2$ is **[MP PET/PMT 1988]**

(a) $[Pb^{2+}][Cl^-]^2$ (b) $[Pb^{2+}][Cl^-]$

(c) $[Pb^{2+}]^2[Cl^-]$ (d) None of these

92. K_{sp} value of $Al(OH)_3$ and $Zn(OH)_2$ are 8.5×10^{-23} and 1.8×10^{-14} respectively. If NH_4OH is added in a solution of Al^{3+} and Zn^{2+}, which will precipitate earlier **[MP PMT 1989; CPMT 1989]**

(a) $Al(OH)_3$ (b) $Zn(OH)_2$

(c) Both together (d) None

93. Why pure $NaCl$ is precipitated when HCl gas is passed in a saturated solution of $NaCl$ **[NCERT 1977; MP PMT 1987; CPMT 1974, 78, 81]**

(a) Impurities dissolves in HCl

(b) The value of $[Na^+]$ and $[Cl^-]$ becomes smaller than K_{sp} of $NaCl$

(c) The value of $[Na^+]$ and $[Cl^-]$ becomes greater than K_{sp} of $NaCl$

(d) HCl dissolves in the water

94. Pure $NaCl$ is prepared by saturating a cold saturated solution of common salt in water with HCl gas. The principle used is

(a) Le Chatelier principle (b) Displacement law

(c) Common ion effect (d) Fractional distillation

95. What is the solubility of calcium fluoride in a saturated solution, if its solubility product is 3.2×10^{-11} **[CPMT 1997]**

(a) $2.0\times10^{-4}mole\,/\,litre$ (b) $12.0\times10^{-3}mole\,/\,litre$

(c) $0.2\times10^{-4}mole\,/\,litre$ (d) $2\times10^{-3}mole\,/\,litre$

96. The following equilibrium exists in an aqueous solution of hydrogen sulphide :

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

If dilute HCl is added to an aqueous solution of H_2S without any change in temperature **[NCERT 1989]**

(a) The equilibrium constant will change

(b) The concentration of HS^- will increase

(c) The concentration of undissociated H_2S will decrease

(d) The concentration of HS^- will decrease

97. Solubility of a salt M_2X_3 is $y\ mol\ dm^{-3}$. The solubility product of the salt will be **[IIT 1990, 97; AFMC 1991; RPMT 1999; MP PET 2001; MP PMT 2003; Orissa JEE 2005]**

(a) $6y^4$
(b) $64y^4$
(c) $36y^5$
(d) $108y^5$

98. Which one of the following is most soluble **[CBSE PMT 1994; RPMT 2000]**

(a) $CuS\ (K_{sp} = 8 \times 10^{-37})$
(b) $MnS\ (K_{sp} = 7 \times 10^{-16})$
(c) $Bi_2S_3\ (K_{sp} = 1 \times 10^{-70})$
(d) $Ag_2S\ (K_{sp} = 6 \times 10^{-51})$

99. The solubility product of $PbCl_2$ at 20^oC is 1.5×10^{-4}. Calculate the solubility **[Bihar CEE 1995; BHU 2002]**

(a) 3.75×10^{-4}
(b) 3.34×10^{-2}
(c) 3.34×10^{2}
(d) None of these

100. Which one of the following compounds is a Lewis acid **[EAMCET 1997]**

(a) PCl_3
(b) BCl_3
(c) NCl_3
(d) $CHCl_3$

101. Which one of the following salt is most acidic in water **[IIT 1995]**

(a) $NiCl_2$
(b) $BeCl_2$
(c) $FeCl_3$
(d) $AlCl_3$

102. Which of the following aqueous solution will have a pH less than 7.0 **[MP PMT 1991, 92]**

(a) KNO_3
(b) $NaOH$
(c) $FeCl_3$
(d) $NaCN$

103. Hydrolysis constant for a salt of weak acid and weak base would be **[RPMT 1999]**

(a) $K_h = \dfrac{K_w}{K_a}$
(b) $K_h = \dfrac{K_w}{K_b}$
(c) $K_h = \dfrac{K_w}{K_a K_b}$
(d) None of these

104. Which salt will give basic solution on hydrolysis **[RPMT 1997]**

(a) KCN
(b) KCl
(c) NH_4Cl
(d) CH_3COONH_4

105. Which of the following sulphides has the lowest solubility product **[KCET 1996]**

(a) FeS
(b) MnS
(c) PbS
(d) ZnS

106. The concentration of which ion is to be decreased, when NH_3 solution is added **[RPMT 1997]**

(a) OH^-
(b) NH_4^+
(c) H_3O^+
(d) O_2^-

107. The compound insoluble in acetic acid is **[IIT 1986]**

(a) Calcium oxide

(b) Calcium carbonate

(c) Calcium oxalate

(d) Calcium hydroxide

108. A saturated solution of Ag_2SO_4 is $2.5 \times 10^{-2}\ M$; The value of its solubility product is **[Pb.CET 2004]**

(a) 62.5×10^{-6}
(b) 6.25×10^{-4}
(c) 15.625×10^{-6}
(d) 3.125×10^{-6}

109. Solubility product of $AgCl$ is 1×10^{-6} at 298 K. Its solubility in mole $litre^{-1}$ would be

(a) $1 \times 10^{-6} mol / litre$

(b) $1 \times 10^{-3} mol / litre$

(c) $1 \times 10^{-12} mol / litre$

(d) None of these

110. 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 **[UPSEAT 2004]**

(a) More
(b) Less
(c) Equal
(d) Zero

111. The concentration of KI and KCl in certain solution containing both is 0.001M each. If 20 ml of this solution is added to 20 ml of a saturated solution of AgI in water? What will happen **[MP PMT 2004**

(a) $AgCl$ will be precipitated

(b) AgI will be precipitated

(c) Both $AgCl$ and AgI will be precipitated

(d) There will be no precipitated

112. The solubility product of a sparingly soluble salt AX_2 is 3.2×10^{-11}. Its solubility (in moles / litres) is **[CBSE PMT 2004]**

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

113. 0.5 M ammonium benzoate is hydrolysed to 0.25 percent, hence its hydrolysis constant is **[MH CET 2004]**

(a) 2.5×10^{-5}
(b) 1.5×10^{-4}
(c) 3.125×10^{-6}
(d) 6.25×10^{-4}

114. The solubility of Sb_2S_3 in water is 1.0×10^{-5} mol / litre at 298 K. What will be its solubility product **[CPMT 20**

(a) 108×10^{-25}
(b) 1.0×10^{-25}

(c) 144×10^{-25}
(d) 126×10^{-24}

115. The ionic product of water at 25^oC is 10^{-14}. The ionic product at 90^oC will be **[CBSE PMT 1996]**

(a) 1×10^{-20}
(b) 1×10^{-12}
(c) 1×10^{-14}
(d) 1×10^{-16}

116. In hydrolysis of a salt of weak acid and strong base, $A^- + H_2O \rightleftarrows HA + OH^-$, the hydrolysis constant (K_h) is equal to....

(a) $\frac{K_w}{K_a}$
(b) $\frac{K_w}{K_b}$
(c) $\sqrt{\frac{K_a}{C}}$
(d) $\frac{K_w}{K_a \times K_b}$

Hydrogen ion concentration- *p*H scale and Buffer solution

1. The pH of blood does not appreciably change by a small addition of an acid or a base because blood **[CBSE PMT 1995]**

(a) Contains serum protein which acts as buffer
(b) Contains iron as a part of the molecule
(c) Can be easily coagulated
(d) It is body fluid

2. The pH of a $0.001\,M\,NaOH$ will be **[MP PMT 1995; UPSEAT 2001]**

(a) 3
(b) 2
(c) 11
(d) 12

3. pH value of a solution, whose hydronium ion concentration is $6.2 \times 10^{-9} mol/l$, is **[AFMC 1999; AIIMS 2000]**

(a) 6.21
(b) 7.21
(c) 7.75
(d) 8.21

4. 0.1 mole of CH_3NH_2 ($K_b = 5 \times 10^{-4}$) is mixed with 0.08 mole of HCl and diluted to one litre. What will be the H^+ concentration in the solution? **[IIT 2005]**

(a) $8 \times 10^{-2}\, M$
(b) $8 \times 10^{-11}\, M$
(c) $1.6 \times 10^{-11}\, M$
(d) $8 \times 10^{-5}\, M$

5. What will be the sum of pH and pOH in an aqueous solution ? **[MP PET/PMT 1998]**

(a) 7
(b) pk_w
(c) Zero
(d) 1

6. Hydrogen ion concentration in mol/L in a solution of $pH = 5.4$ will be **[AIEEE 2005]**

(a) 3.98×10^8
(b) 3.88×10^6
(c) 3.68×10^{-6}
(d) 3.98×10^{-6}

7. When solid potassium cyanide is added in water then **[CPMT 2002; BHU 2002]**

(a) pH will increase
(b) pH will decrease
(c) pH will remain the same
(d) Electrical conductivity will not change

8. pH of a $10^{-3} M$ solution of hydrochloric acid will be **[MP PET 2000]**

(a) 1.3
(b) 2.0
(c) 3.0
(d) 4.5

9. The pH of water at 25^oC is nearly **[Orissa JEE 2002]** **[CPMT 1986, 89, 90, 93; CBSE PMT 1989; MADT Bihar 1995]**

(a) 2
(b) 7
(c) 10
(d) 12

10. pH of a solution is 5. Its hydroxyl ion concentration is **[JIPMER 1999]**

(a) 5
(b) 10
(c) 10^{-5}
(d) 10^{-9}

11. The pH of a solution in which the $[H^+] = 0.01$, is **[MADT Bihar 1980]**

(a) 2
(b) 1
(c) 4
(d) 3

12. At 25°C, the dissociation constant of a base BOH is 1.0×10^{-12}. The concentration of Hydroxyl ions in 0.01 M aqueous solution of the base would be**[CBSE PMT 2**

(a) $2.0 \times 10^{-6} mol\ L^{-1}$
(b) $1.0 \times 10^{-5} mol\ L^{-1}$
(c) $1.0 \times 10^{-6} mol\ L^{-1}$
(d) $1.0 \times 10^{-7} mol\ L^{-1}$

13. Aqueous solution of HCl has the $pH = 4$. Its molarity would be **[RPMT 1999]**

(a) 4 M
(b) 0.4 M
(c) 0.0001 M
(d) 10 M

14. Which is a buffer solution **[CPMT 1985, 88; AIIMS 1980; MP PMT 1994; AFMC 2004]**

(a) $CH_3COOH + CH_3COONa$
(b) $CH_3COOH + CH_3COONH_4$
(c) $CH_3COOH + NH_4Cl$
(d) $NaOH + NaCl$

15. The addition of solid sodium carbonate to pure water causes **[NCERT 1973]**

(a) An increase in hydronium ion concentration
(b) An increase in alkalinity
(c) No change in acidity
(d) A decrease in hydroxide ion concentration

16. The aqueous solution of which of the following salt has the lowest pH **[CBSE PMT 2002]**

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

17. The pH of a $10^{-10} M\,NaOH$ solution is nearest to

[UPSEAT 2001, 02]

(a) 10 (b) 7

(c) 4 (d) -10

18. Which will have maximum pH **[NCERT 1979]**

(a) Distilled water

(b) $1M\ NH_3$

(c) $1M\ NaOH$

(d) Water saturated by chlorine

19. pH of a solution is 9.5. The solution is **[MH CET 2000]**

(a) Neutral (b) Acidic

(c) Basic (d) Amphoteric

20. The pH of a $10^{-9}M$ solution of HCl in water is **[UPSEAT 2000, 02]**

(a) 8 (b) – 8

(c) Between 7 and 8 (d) Between 6 and 7

21. $pH + pOH$ equal to **[NCERT 1975]**

(a) Zero (b) Fourteen

(c) A negative number (d) Infinity

22. Which of the following 0.1*M* solution will contain the largest concentration of hydronium ions **[NCERT 1971, 73]**

(a) $NaHCO_3$ (b) NH_4Cl

(c) HCl (d) NH_3

23. Which one has pH 12 **[Roorkee 1995]**

(a) $0.01M\ KOH$ (b) $1N\ KOH\ ml$

(c) $1N\ NaOH\ ml$ (d) $1N\ Ca(OH)_2\ ml$

24. What is the correct relationship between the pHs of isomolar solutions of sodium oxide (pH_1), sodium sulphide (pH_2), sodium selenide (pH_3) and sodium telluride (pH_4)? **[CBSE PMT 2005]**

(a) $pH_1 > pH_2 = pH_3 > pH_4$

(b) $pH_1 < pH_2 < pH_3 < pH_4$

(c) $pH_1 < pH_2 < pH_3 = pH_4$

(d) $pH_1 > pH_2 > pH_3 > pH_4$

25. Given *pH* of a solution *A* is 3 and it is mixed with another solution *B* having *pH* 2. If both mixed then resultant *pH* of the solution will be **[BHU 2005]**

(a) 3.2 (b) 1.9

(c) 3.4 (d) 3.5

26. On adding solid potassium cyanide to water **[MP PMT 1989]**

(a) *pH* will increase

(b) *pH* will decrease

(c) *pH* will not change

(d) Electrical conductance will not change

27. *A* is an aqueous acid; *B* is an aqueous base. They are diluted separately, then **[KCET 2002]**

(a) pH of *A* increases and pH of *B* decreases

(b) pH of *A* increases and pH of *B* decreases till pH in each case is 7

(c) pH of *A* and *B* increase

(d) pH of *B* and *A* decrease

28. The compound whose 0.1 *M* solution is basic is **[IIT 1986; MP PMT 1991]**

(a) Ammonium acetate

(b) Calcium carbonate

(c) Ammonium sulphate

(d) Sodium acetate

29. The following reaction is known to occur in the body $CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$. If CO_2 escapes from the system **[NCERT 1973; RPMT 1997]**

(a) pH will decrease

(b) Hydrogen ion concentration will decrease

(c) H_2CO_3 concentration will be unaltered

(d) The forward reaction will be promoted

30. For preparing a buffer solution of pH 6 by mixing sodium acetate and acetic acid, the ratio of the concentration of salt and acid should be $(K_a = 10^{-5})$ **[MP PET**

(a) $1 : 10$ (b) $10 : 1$

(c) $100 : 1$ (d) $1 : 100$

31. Which is incorrect for buffer solution **[CPMT 1985]**

(a) It contains weak acid and its conjugate base

(b) It contains weak base and its conjugate acid

(c) In this there is very less change is pH value when very less amount of acid and base is mixed

(d) None of the above

32. pH values of HCl and $NaOH$ solutions each of strength $\frac{N}{100}$ will be respectively **[MP PMT 1999]**

(a) 2 and 2 (b) 2 and 12

(c) 12 and 2 (d) 2 and 10

33. When rain is accompained by a thunderstorm, the collected rain water will have a *pH* value **[AIEEE 2003]**

(a) Slightly lower than that of rain water without thunderstorm

(b) Slightly higher than that when the thunderstorm is not there

(c) Uninfluenced by occurrence of thunderstorm

(d) Which depends on the amount of dust in air

34. Which of the following is the buffer solution of strong acidic nature **[MP PET 2002]**

(a) $HCOOH + HCOO^-$

(b) $CH_3COOH + CH_3COO^-$

(c) $H_2C_2O_4 + C_2O_4^{2-}$

(d) $H_3BO_3 + BO_3^{3-}$

35. The dissociation constant of an acid *HA* is 1×10^{-5}. The pH of 0.1 molar solution of the acid will be **[KCET (Engg./Med.) 1999]**

(a) Five (b) Four

(c) Three (d) One

36. The pH value of $1.0 \times 10^{-8} M\,HCl$ solution is less than 8 because

(a) HCl is completely ionised at this concentration

(b) The ionization of water is negligible

(c) The ionization of water cannot be assumed to be negligible in comparison with this low concentration of HCl

(d) The pH cannot be calculated at such a low concentration of HCl

37. What is the pH for a neutral solutions at the normal temperature of the human body **[JIPMER 2000]**

(a) 7.2 (b) 14.0

(c) 6.8 (d) 6.0

38. 1 M $NaCl$ and 1 M HCl are present in an aqueous solution. The solution is **[AIEEE 2002]**

(a) Not a buffer solution with $pH < 7$

(b) Not a buffer solution with $pH > 7$

(c) A buffer solution with $pH < 7$

(d) A buffer solution with $pH > 7$

39. A solution has $pH = 5$, it is diluted 100 times, then it will become **[NCERT 1978; AFMC 2005]**

(a) Neutral (b) Basic

(c) Unaffected (d) More acidic

40. 0.02 M monobasic acid dissociates 2% hence, pH of the solution is **[MH CET 2000]**

(a) 0.3979 (b) 1.3979

(c) 1.699 (d) 3.3979

41. Components of buffer solution are $0.1 M\,HCN$ and $0.2 M\,NaCN$. What is the pH of the solution **[RPET 2000]**

(a) 9.61 (b) 6.15

(c) 2.0 (d) 4.2

42. pH of a solution of $10\,ml$. 1N sodium acetate and $50\,ml$ $2N$ acetic acid $(K_a = 1.8 \times 10^{-5})$, is approximately **[MP PMT 2003]**

(a) 4 (b) 5

(c) 6 (d) 7

43. By adding $20\,ml$ $0.1 N\,HCl$ to $20\,ml$ $0.001\,N$ KOH, the pH of the obtained solution will be **[KCET 2000]**

(a) 2 (b) 1.3

(c) 0 (d) 7

44. The pH of the solution containing $10\,ml$ of a $0.1 N\,NaOH$ and $10\,ml$ of $0.05\,N$ H_2SO_4 would be **[Pb. PMT 2002, 04]**

(a) 1 (b) 0

(c) 7 (d) > 7

45. The pH of $10^{-7} M\,NaOH$ is **[MP PMT 2001]**

(a) 7.01 (b) Between 7 and 8

(c) Between 9 and 10 (d) Greater than 10

46. The hydrogen ion concentration of $0.1 N$ solution of CH_3COOH, which is 30% dissociated, is **[JIPMER 2002]**

(a) 0.03 (b) 3.0

(c) 0.3 (d) 30.0

47. What is the pH of $0.1 M\,NH_3$ **[RPET 2000]**

(a) 11.27 (b) 11.13

(c) 12.0 (d) 9.13

48. By adding a strong acid to the buffer solution, the pH of the buffer solution **[DPMT 1996]**

(a) Remains constant (b) Increases

(c) Decreases (d) Becomes zero

49. The pH of $0.1 M\,NaOH$ is **[MP PET 2003]**

(a) 11 (b) 12

(c) 13 (d) 14

50. pH of human blood is 7.4. Then H^+ concentration will be **[RPMT 2002]**

(a) 4×10^{-8} (b) 2×10^{-8}

(c) 4×10^{-4} (d) 2×10^{-4}

51. Assuming complete ionisation, the pH of $0.1 M\,HCl$, is 1. The molarity of H_2SO_4 with the same pH is **[Pb. PMT 2002]**

(a) 0.1 (b) 0.2

(c) 0.05 (d) 2.0

52. Highest pH 14 is given by **[DCE 1999]**

(a) $0.1 M\,H_2SO_4$ (b) $0.1 M\,NaOH$

(c) $1N\,NaOH$ (d) $1N\,HCl$

53. What will be the pH of a $10^{-8}\,M\,HCl$ solution **[MP PET/PMT 1998; RPET 1999;MP PMT 2000]**

(a) 8.0 (b) 7.0

(c) 6.98 (d) 14.0

54. When 10 ml of 0.1 M acetic acid $(pK_a = 5.0)$ is titrated against 10 ml of 0.1M ammonia solution $(pK_b = 5.0)$, the equivalence point occurs at pH **[AIIMS 2005**

(a) 5.0 (b) 6.0

(c) 7.0 (d) 9.0

55. Which on reaction with water will have pH less than 7 **[MH CET 2001]**

(a) BaO (b) CaO

(c) Na_2O (d) P_2O_5

56. A solution of $MgCl_2$ in water has pH **[MP PMT 2002]**

(a) < 7 (b) > 7

(c) 7 (d) 14.2

57. pH of completely dissociated $0.005\,M\,H_2SO_4$ is **[RPET 2003]**

(a) 3 (b) 4

(c) 2 (d) 5

58. The pK_a of a weak acid is 4.8. What should be the ratio of [Acid]/[Salt] of a buffer if $pH = 5.8$ is required **[MP PET 2003]**

(a) 10
(b) 0.1
(c) 1
(d) 2

59. Which of the following salt is acidic **[CPMT 1979, 81; NCERT 1979, 81; MP PET 1990; JIPMER 2002]**

(a) Na_2SO_4
(b) $NaHSO_3$
(c) Na_2SO_3
(d) Na_2S

60. $20 ml$ of $0.5 N\ HCl$ and $35 ml$ of $0.1 N\ NaOH$ are mixed. The resulting solution will **[KCET 2005]**

(a) Be neutral
(b) Be basic
(c) Turn phenolphthalein solution pink
(d) Turn methyl orange red

61. The pH of a 0.02 M solution of hydrochloric acid is **[MP PMT 1993]**

(a) 2.0
(b) 1.7
(c) 0.3
(d) 2.2

62. A sample of $Na_2CO_3.H_2O$ weighing $0.62\ g$ is added to $100\ ml$ of $0.1 N\ (NH_4)_2SO_4$ solution. What will be the resulting solution **[BHU 1997]**

(a) Acidic
(b) Neutral
(c) Basic
(d) None of these

63. The *pH* of the solution is 4. The hydrogen ion concentration of the solution in *mol/litre* is **[UPSEAT 2000]**

(a) 9.5
(b) 10^{-4}
(c) 10^{4}
(d) 10^{-2}

64. $NaOH_{(aq)}, HCl_{(aq)}$ and $NaCl_{(aq)}$ concentration of each is $10^{-3} M$. Their *pH* will be respectively **[BHU 2003]**

(a) 10, 6, 2
(b) 11, 3, 7
(c) 10, 2, 6
(d) 3, 4, 7

65. The pH of $10^{-5} M$ aqueous solution of $NaOH$ is **[MP PET 1996]**

(a) 5
(b) 7
(c) 9
(d) 11

66. The pH of 0.05 M solution of dibasic acid is **[MH CET 2002]**

(a) $+1$
(b) -1
(c) $+2$
(d) -2

67. A buffer solution contains $0.1\ M$ of acetic acid and $0.1\ M$ of sodium acetate. What will be its pH, if pK_a of acetic acid is 4.75

(a) 4.00
(b) 4.75
(c) 5.00
(d) 5.25

68. To obtain a buffer which should be suitable for maintaining a pH of about $4-5$, we need to have in solution, a mixture of

(a) A strong base + its salt with a weak acid
(b) A weak base + its salt with a strong acid
(c) A strong acid + its salt with a weak base
(d) A weak acid + its salt with a strong base

69. The concentration of $NaOH$ solution is $10^{-8} M$. Find out the (OH^-) concentration **[CPMT 1993]**

(a) 10^{-8}
(b) Greater than 10^{-6}
(c) 10^{-6}
(d) Lies between 10^{-6} and 10^{-7}

70. The pH of $0.0001\ N$ solution of KOH will be **[BHU 1997; CET Pune 1998]**

(a) 4
(b) 6
(c) 10
(d) 12

71. Given that the dissociation constant for H_2O is $K_w = 1 \times 10^{-14}\ mole^2\ litre^{-2}$, what is the pH of a $0.001\ molar\ KOH$ solution **[MP PET 1995; MP PET/PMT 1998]**

(a) 10^{-11}
(b) 3
(c) 14
(d) 11

72. An acidic buffer solution can be prepared by mixing solution of **[MNR 1983]**

(a) Ammonium acetate and acetic acid
(b) Ammonium chloride and hydrochloric acid
(c) Sulphuric acid and sodium sulphate
(d) Acetic acid and sulphuric acid
(e) $NaCl$ and $NaOH$

73. Which of the following mixtures forms an acid buffer **[MP PMT 1993; IIT 1981; CPMT 1989; CBSE PMT 1989]**

(a) $NaOH + HCl$
(b) $CH_3COOH + CH_3COONa$
(c) $NH_4OH + NH_4Cl$
(d) $H_2CO_3 + (NH_4)_2CO_3$

74. A buffer solution has equal volumes of $0.2 M\ NH_4OH$ and $0.02 M\ NH_4Cl$. The pK_b of the base is 5. The *pH* is **[CBSE PMT 1989; KCET 2005]**

(a) 10
(b) 9
(c) 4
(d) 7

75. The pH of a simple sodium acetate buffer is given by $pH = pK_a + \log \frac{[Salt]}{[Acid]}$

K_a of acetic acid $= 1.8 \times 10^{-5}$

[MP PMT 1996]

If [Salt] = [Acid] = 0.1 *M*, the pH of the solution would be about **[BHU 1987]**

(a) 7
(b) 4.7
(c) 5.3
(d) 1.4

76. Amongst the following solutions, the buffer solution is **[MP PMT 1999]**

(a) $NH_4Cl + NH_4OH$ solution

(b) $NH_4Cl + NaOH$ solution

(c) $NH_4OH + HCl$ solution

(d) $NaOH + HCl$ solution

77. The pH of solution having $[OH^-] = 10^{-7}$ is **[AIIMS 1996]**

(a) 7 (b) 14

(c) Zero (d) –7

78. $50\ ml$ water is added to a $50\ ml$ solution of $Ba(OH)_2$ of strength $0.01\ M$. The pH value of the resulting solution will be **[MP PMT 1999]**

(a) 8 (b) 10

(c) 12 (d) 6

79. pH of a solution can be expressed as **[CPMT 1999; UPSEAT 2001]**

(a) $-\log_e(H^+)$ (b) $-\log_{10}(H^+)$

(c) $\log_e(H^+)$ (d) $\log_{10}(H^+)$

80. The solution of sodium carbonate has pH **[MP PET 2000]**

(a) Greater than 7 (b) Less than 7

(c) Equal to 7 (d) Equal to zero

81. The pH of $10^{-7}\,N\,HCl$ is **[RPMT 2000]**

(a) 6.0 (b) 6.97

(c) 8.0 (d) 10.0

82. If the pH of a solution is 2, its normality will be **[MADT Bihar 1982; MP PET 2000]**

(a) $2N$ (b) $\frac{1}{2}N$

(c) $0.01\ N$ (d) None of these

83. The buffer solution of 100 ml having a pH value 4 when added to 1 ml dilute HCl, then the pH of buffer solution **[NCERT 1976, 77]**

(a) Converts to 7 (b) Does not change

(c) Converts to 2 (d) Changes to 10

84. In a solution of acetic acid, sodium acetate is added, then its pH value **[NCERT 1977; DPMT 1985; MP PMT 1994]**

(a) Decreases

(b) Increases

(c) Remains unchanged

(d) (a) and (b) both are correct

85. If pOH of a solution is 6.0, then its pH will be **[MP PMT 1987]**

(a) 6 (b) 10

(c) 8 (d) 14

86. In a solution of $pH = 5$, more acid is added in order to reduce the $pH = 2$. The increase in hydrogen ion concentration is **[MP PET 1989; CPMT 1990]**

(a) 100 times (b) 1000 times

(c) 3 times (d) 5 times

87. Which solution contains maximum number of H^+ ion

(a) 0.1 M HCl (b) 0.1 M NH_4Cl

(c) 0.1 M $NaHCO_3$ (d) 0.1 M

88. A certain buffer solution contains equal concentration of X^- and HX. The K_b for X^- is 10^{-10}. The pH of the buffer is **[IIT 1984; RPMT 1997; CPMT 1996; DPMT 2004]**

(a) 4 (b) 7

(c) 10 (d) 14

89. The defination of pH is **[EAMCET 1980; UPSEAT 2001]**

(a) $pH = \log \frac{1}{[H^+]}$ (b) $pH = \log [H^+]$

(c) $pH = -\log \frac{1}{[H^+]}$ (d) $pH = -\log^{[H^+]}$

90. Which of the following does not make any change in pH when added to $10\ ml$ dilute HCl **[NCERT 1975]**

(a) 5 ml pure water (b) 20 ml pure water

(c) 10 ml HCl (d) Same 20 ml dilute HCl

91. A compound whose aqueous solution will have the highest pH **[CPMT 1974, 75, 78; MP PET 1996; DPMT 1982, 83]**

(a) $NaCl$ (b) Na_2CO_3

(c) NH_4Cl (d) $NaHCO_3$

92. At $80^\circ C$, distilled water has $[H_3O^+]$ concentration equal to 1×10^{-6} *mole / litre*. The value of K_w at this temperature will be **[CBSE PMT 1994; RPMT 2000; AFMC 2001; AIIMS 2002; BHU 2002]**

(a) 1×10^{-6} (b) 1×10^{-9}

(c) 1×10^{-12} (d) 1×10^{-15}

93. The pH value of $0.1\ M\ NaOH$ solution is (when there is a given reaction $[H^+][OH^-] = 10^{-15}$) **[CPMT 1997]**

(a) 13 (b) 12

(c) 11 (d) 2

94. Which oxychloride has maximum pH **[CPMT 1997]**

(a) $NaClO$ (b) $NaClO_2$

(c) $NaClO_3$ (d) $NaClO_4$

95. pH of $HCl(10^{-12}\,M)$ is **[CPMT 1997; Pb. PET/PMT 1999]**

(a) 12 (b) –12

(c) ≈ 7 (d) 14

96. Which one is buffer solution **[CPMT 1997]**

(a) $[PO_4^{--}]\,[[HPO_4^{--}]$ (b) $[PO_3^{3-}]\,[[H_2PO_4^{--}]$

(c) $[HPO_4^{--}]\,[[H_2PO_4^{--}]$ (d) All of these

97. When $100\ ml$ of $M/10\ NaOH$ solution and $50\ ml$ of $M/5\ HCl$ solution are mixed, the pH of resulting solution would be

(a) 0 (b) 7

(c) Less than 7 (d) More than 7

98. How many millilitres of $6.0\ M$ hydrochloric acid should be used to prepare $150\ ml$ of a solution which is $0.30\ M$ in hydrogen ion

(a) 3.0 (b) 7.5

(c) 9.3 (d) 30

99. The pH of 0.1 M acetic acid is 3, the dissociation constant of acid will be

(a) 1.0×10^{-4} (b) 1.0×10^{-5}

(c) 1.0×10^{-3} (d) 1.0×10^{-8}

100. The pH of a buffer solution containing $25\ ml$ of $1\ M\ CH_3COONa$ and $25\ ml$ of $1\ M\ CH_3COOH$ will be appreciably affected by $5\ ml$ of

(a) $1\ M\ CH_3COOH$ (b) $5\ M\ CH_3COOH$

(c) $5\ M\ HCl$ (d) $1\ M\ NH_4OH$

101. The pH value of decinormal solution of NH_4OH which is 20% ionised, is **[CBSE PMT 1998]**

(a) 13.30 (b) 14.70

(c) 12.30 (d) 12.95

102. A physician wishes to prepare a buffer solution at $pH = 3.58$ that efficiently resists changes in pH yet contains only small concentration of the buffering agents. Which of the following weak acids together with its sodium salt would be best to use **[CBSE PMT 1997]**

(a) $m-$chlorobenzoic acid $(pK_a = 3.98)$

(b) $p-$chlorocinnamic acid $(pK_a = 4.41)$

(c) $2, 5-$dihydroxy benzoic acid $(pK_a = 2.97)$

(d) Acetoacetic acid $(pK_a = 3.58)$

103. In a mixture of a weak acid and its salt, the ratio of concentration of acid to salt is increased ten-fold. The pH of the solution

(a) Decreases by one (b) Increases by one-tenth

(c) Increases by one (d) Increases ten-fold

104. When an acid or alkali is mixed with buffer solution, then pH of buffer solution **[CPMT 1997]**

(a) Not changes (b) Changes slightly

(c) Increases (d) Decreases

105. How much sodium acetate should be added to a $0.1\ m$ solution of CH_3COOH to give a solution of $pH = 5.5 (pK_a$ of $CH_3COOH = 4.5)$ **[KCET 1996]**

(a) $0.1\ m$ (b) $0.2\ m$

(c) $1.0\ m$ (d) $10.0\ m$

106. The pH of the aqueous solution containing $0.49\ gm$ of H_2SO_4 in one litre is **[EAMCET 1997]**

[RPMT 1997]

(a) 2 (b) 1

(c) 1.7 (d) 0.3

107. Which of the following solutions can act as buffer **[JIPMER 1997]**

(a) 0.1 molar aq. $NaCl$

(b) 0.1 molar aq. $CH_3COOH + 0.1$ molar $NaOH$

[Pb. PMT 1998]

(c) 0.1 molar aq. ammonium acetate

(d) None of the above

108. The pH of a solution obtained by mixing $50\ ml$ of $0.4\ N\ HCl$ and $50\ ml$ of $0.2\ N\ NaOH$ is **[KCET 1996]**

(a) $-\log 2$ (b) $-\log 0.2$

(c) 1.0 (d) 2.0

109. Which of the following is a buffer **[MP PMT 1991; BHU 1995]**

(a) $NaOH + CH_3COONa$

[CPMT 1987]

(b) $NaOH + Na_2SO_4$

(c) $K_2SO_4 + H_2SO_4$

(d) $NH_4OH + CH_3COONH_4$

110. If 4.0 gm $NaOH$ is present in 1 *litre* solution, then its pH will be **[CPMT 1989]**

(a) 6 (b) 13

(c) 18 (d) 24

111. The pOH of beer is 10.0. The hydrogen ion concentration will be **[MP PMT 1994]**

(a) 10^{-2} (b) 10^{-10}

(c) 10^{-8} (d) 10^{-4}

112. When a buffer solution of sodium acetate and acetic acid is diluted with water **[CPMT 1985]**

(a) Acetate ion concentration increases

(b) H^+ ion concentration increases

(c) OH^- ion concentration increases

(d) H^+ ion concentration remain unaltered

113. What is the pH of $Ba(OH)_2$ if normality is 10 **[CPMT 1996]**

[KCET 1996]

(a) 4 (b) 10

(c) 7 (d) 9

114. What will be the pH of a solution formed by mixing $40\ ml$ of $0.10\ M\ HCl$ with $10\ ml$ of $0.45\ M\ NaOH$ **[Manipal MEE 1995]**

(a) 12 (b) 10

(c) 8 (d) 6

115. The pH of a solution having $[H^+] = 10 \times 10^{-4}$ $moles/litre$ will be **[BHU 1981]**

(a) 1 (b) 2

(c) 3 (d) 4

116. If $0.4\,gm\ NaOH$ is present in 1 *litre* solution, then its pH will be **[CPMT 1985; BHU 1980]**

(a) 2 (b) 10

(c) 11 (d) 12

117. Which of the following is not a Bronsted acid **[BHU 1997]**

(a) $CH_3NH_4^+$ (b) CH_3COO^-

(c) H_2O (d) HSO_4^-

118. pH of 0.005 $M\ H_2SO_4$ solution will be **[NCERT 1980]**

(a) 0.005 (b) 2

(c) 1 (d) 0.01

119. A buffer solution is a mixture of **[MP PMT 1987]**

(a) Strong acid and strong base

(b) Weak acid and weak base

(c) Weak acid and conjugate acid

(d) Weak acid and conjugate base

120. When pH of a solution decreases, its hydrogen ion concentration **[MADT Bihar 1981]**

(a) Decreases (b) Increases

(c) Rapidly increases (d) Remains always constant

121. If the pH of a solution is 4.0 at 25^oC, its pOH would be $(K_w = 10^{-14})$ **[MP PMT 1989]**

(a) 4.0 (b) 6.0

(c) 8.0 (d) 10.0

122. An aqueous solution whose $pH = 0$ is **[CPMT 1976; DPMT 1982]**

(a) Alkaline (b) Acidic

(c) Neutral (d) Amphoteric

123. In a solution of acid H^+ concentration is $10^{-10}\,M$. The *pH* of this solution will be

(a) 8 (b) 6

(c) Between 6 and 7 (d) Between 3 and 6

124. The concentration of hydronium (H_3O^+) ion in water is **[CET Pune 1998]**

(a) Zero (b) $1\times10^{7}\ gm\ ion/litre$

(c) $1\times10^{-14}\ gm\ ion/litre$ (d) $1\times10^{-7}\ gm\ ion/litre$

125. A solution whose *pH* value is less than 7 will be

(a) Basic (b) Acidic

(c) Neutral (d) Buffer

126. When the pH of a solution is 2, the hydrogen ion concentration in moles per litre is **[NCERT 1973; MNR 1979]**

(a) 1×10^{-14} (b) 1×10^{-2}

(c) 1×10^{-7} (d) 1×10^{-12}

127. A base is dissolved in water yields a solution with a hydroxide ion concentration of 0.05 $mol\ litre^{-1}$. The solution is **[AFMC 1997]**

(a) Basic (b) Acid

(c) Neutral (d) Both (a) and (b)

128. The pH of a solution is increased from 3 to 6. Its H^+ ion concentration will be **[EAMCET 1998]**

(a) Reduced to half

(b) Doubled

(c) Reduced by 1000 times

(d) Increased by 1000 times

129. Pure water is kept in a vessel and it remains exposed to atmospheric CO_2 which is absorbed, then its pH will be **[MADT Bihar 1984; DPMT 2002]**

(a) Greater than 7

(b) Less than 7

(c) 7

(d) Depends on ionic product of water

130. The pH of a solution is 2. If its pH is to be raised to 4, then the $[H^+]$ of the original solution has to be **[MP PET 1994]**

(a) Doubled

(b) Halved

(c) Increased hundred times

(d) Decreased hundred times

131. Which of the following solutions cannot act as a buffer **[EAMCET 1998]**

(a) $NaH_2PO_4 + H_3PO_4$

(b) $CH_3COOH + CH_3COONa$

(c) $HCl + NH_4Cl$

(d) $H_3PO_4 + Na_2HPO_4$

132. Assuming complete ionisation, the pH of $0.1\,M\ HCl$ is 1. The molarity of H_2SO_4 with the same pH is **[KCET 1998]**

(a) 0.2 (b) 0.1

(c) 2.0 (d) 0.05

133. The pH of blood is

(a) 5.2 (b) 6.3

(c) 7.4 (d) 8.5

134. The *pH* of 10^{-8} molar aqueous solution of HCl is **[CPMT 1988; MNR 1983, 90; MP PMT 1987; IIT 1981; BHU 1995; AFMC 1998; MP PET 1989, 99; BCECE 2005]**

(a) -8

(b) 8

(c) $6 > 7$ (Between 6 and 7)

(d) $7 > 8$ (Between 7 and 8)

135. As the temperature increases, the pH of a KOH solution **[UPSEAT 2001]**

(a) Will decreases

(b) Will increases
(c) Remains constant
(d) Depends upon concentration of KOH solution

136. The hydrogen ion concentration in a given solution is 6×10^{-4}. Its pH will be **[EAMCET 1978]**

(a) 6
(b) 4
(c) 3.22
(d) 2

137. The pH of $\frac{N}{100} HCl$ would be approximately

[CPMT 1971; DPMT 1982, 83; MP PMT 1991; Bihar MEE 1996]

(a) 1
(b) 1.5
(c) 2
(d) 2.5

138. A solution which is resistant to change of pH upon the addition of an acid or a base is known as **[BHU 1979]**

(a) A colloid
(b) A crystalloid
(c) A buffer
(d) An indicator

139. $10^{-6} M \; HCl$ is diluted to 100 times. Its pH is

[CPMT 1984]

(a) 6.0
(b) 8.0
(c) 6.95
(d) 9.5

140. The pH of a 10^{-10} molar HCl solution is approximately

[NCERT 1977]

(a) 10
(b) 7
(c) 1
(d) 14

141. The H^+ ion concentration is 1.0×10^{-6} *mole/litre* in a solution. Its *pH* value will be

[MP PMT 1985; AFMC 1982]

(a) 12
(b) 6
(c) 18
(d) 24

142. The pH of a solution is the negative logarithm to the base 10 of its hydrogen ion concentration in **[Manipal MEE 1995]**

(a) Moles per litre
(b) Millimoles per litre
(c) Micromoles per litre
(d) Nanomoles per litre

143. When 10^{-8} mole of HCl is dissolved in one litre of water, the pH of the solution will be

[CPMT 1973, 94; DPMT 1982]

(a) 8
(b) 7
(c) Above 8
(d) Below 7

144. The pH of the solution containing 10 ml of 0.1 N $NaOH$ and 10 ml of 0.05 N H_2SO_4 would be

[CPMT 1987; Pb. PMT 2002,04]

(a) 0
(b) 1
(c) > 7
(d) 7

145. The pH of 0.001 molar solution of HCl will be

[MP PET 1986; MP PET/PMT 1988; CBSE PMT 1991]

(a) 0.001
(b) 3
(c) 2
(d) 6

146. Which salt can be classified as an acid salt **[CPMT 1989]**

(a) Na_2SO_4
(b) $BiOCl$
(c) $Pb(OH)Cl$
(d) Na_2HPO_4

147. Given a 0.1M solution of each of the following. Which solution has the lowest pH **[MNR 1987]**

(a) $NaHSO_4$
(b) NH_4Cl
(c) HCl
(d) NH_3

148. Out of the following, which pair of solutions is not a buffer solution

(a) $NH_4Cl + NH_4OH$
(b) $NaCl + NaOH$
(c) $Na_2HPO_4 + Na_3PO_4$
(d) $CH_3COOH + CH_3COONa$

149. If the dissociation constant of an acid HA is 1×10^{-5}, the pH of a 0.1 molar solution of the acid will be approximately **[NCERT 1979]**

(a) Three
(b) Five
(c) One
(d) Six

150. pH value of $N/10$ $NaOH$ solution is

[CBSE PMT 1996;Pb. CET 2001; Pb. PMT 2002]

(a) 10
(b) 11
(c) 12
(d) 13

151. A solution of sodium borate has a pH of approximately

[JIPMER 2001]

(a) < 7
(b) > 7
(c) = 7
(d) Between 4 to 5

152. If pH of A, B, C and D are 9.5, 2.5, 3.5 and 5.5 respectively, then strongest acid is **[AFMC 1995]**

(a) A
(b) C
(c) D
(d) B

153. At $25°C$ the pH value of a solution is 6. The solution is

[AFMC 2001]

(a) Basic
(b) Acidic
(c) Neutral
(d) Both (b) and (c)

154. A certain buffer solution contains equal concentration of X^- and HX. The K_a for HX is 10^{-8}. The pH of the buffer is

[UPSEAT 2001]

(a) 3
(b) 8
(c) 11
(d) 14

155. The dissociation constant of HCN is 5×10^{-10}. The pH of the solution prepared by mixing 1.5 *mole* of HCN and 0.15 *moles* of KCN in water and making up the total volume to $0.5\, dm^3$ is

(a) 7.302
(b) 9.302
(c) 8.302
(d) 10.302

156. Which buffer solution out of the following will have $pH > 7$

[MP PET 2001]

(a) $CH_3COOH + CH_3COONa$
(b) $HCOOH + HCOOK$

(c) CH_3COONH_4

(d) $NH_4OH + NH_4Cl$

157. The pK_a of equimolecular sodium acetate and acetic acid mixture is 4.74. If pH is **[DPMT 2001]**

(a) 7 (b) 9.2

(c) 4.74 (d) 14

158. pH of $NaCl$ solution is **[CET Pune 1998]**

(a) 7 (b) Zero

(c) >7 (d) <7

159. A solution of sodium chloride in contact with atmosphere has a pH of about **[NCERT 1972, 77]**

(a) 3.5 (b) 5

(c) 7 (d) 1.4

160. Which would decrease the pH of $25\,cm^3$ of a $0.01\,M$ solution of hydrochloric acid **[MH CET 2001]**

(a) The addition of $25\,cm^3$ $0.005\,M$ hydrochloric acid

(b) The addition of $25\,cm^3$ of $0.02\,M$ hydrochloric acid

(c) The addition of magnesium metal

(d) None of these

161. The condition for minimum change in pH for a buffer solution is **[RPMT 2000]**

(a) Isoelectronic species are added

(b) Conjugate acid or base is added

(c) $pH = pK_a$

(d) None of these

162. A buffer solution with pH 9 is to be prepared by mixing NH_4Cl and NH_4OH. Calculate the number of moles of NH_4Cl that should be added to one litre of $1.0\,M\ NH_4OH.$ $[K_b = 1.8 \times 10^{-5}]$ **[UPSEAT 2001]**

(a) 3.4 (b) 2.6

(c) 1.5 (d) 1.8

163. The ionization constant of a certain weak acid is 10^{-4}. What should be the [salt] to [acid] ratio if we have to prepare a buffer with pH = 5 using this acid and one of the salts **[CPMT 2000; KCET 2000; [illegible]CET 2003]**

(a) $1:10$ (b) $10:1$

(c) $5:4$ (d) $4:5$

164. Which solution has the highest pH value **[JIPMER 2000]**

(a) $1\,M\,KOH$

(b) $1\,M\,H_2SO_4$

(c) Chlorine water

(d) Water containing carbon dioxide

165. One weak acid (like CH_3COOH) and its strong base together with salt (like CH_3COONa) is a buffer solution. In which pair this type of characteristic is found **[AIIMS 1982; CPMT 1994; MP PET 1994]**

(a) HCl and $NaCl$ (b) $NaOH$ and $NaNO_3$

(c) KOH and KCl (d) NH_4OH and NH_4Cl

166. If the pH of a solution of an alkali metal hydroxide is 13.6, the concentration of hydroxide is **[JIPMER 2000]**

(a) Between 0.1 M and 1 M

(b) More than 1 M

(c) Less than 0.001 M

(d) Between 0.01 M and 1 M

167. The pK_a of acetylsalicylic acid (aspirin) is 3.5. The pH of gastric juice in human stomach is about 2-3 and the pH in the small intestine is about 8. Aspirin will be **[IIT 1988; KCET 2003]**

(a) Unionized in the small intestine and in the stomach

(b) Completely ionized in the small intestine and in the stomach

(c) Ionized in the stomach and almost unionized in the small intestine

(d) Ionized in the small intestine and almost unionized in the stomach

168. The concentration of hydrogen ion in water is **[MP PET 1990]**

(a) 8 (b) 1×10^{-7}

(c) 7 (d) $1/7$

169. pH of a 10 M solution of HCl is **[CBSE PMT 1995]**

(a) Less than 0 (b) 2

(c) 0 (d) 1

170. The pH of $1\ N\ H_2O$ is **[CPMT 1988]**

(a) 7 (b) >7

(c) <7 (d) 0

171. If H^+ ion concentration of a solution is increased by 10 times its pH will be **[DCE 2000]**

(a) Increase by one (b) Remains unchanged

(c) Decrease by one (d) Increase by 10

172. The gastric juice in our stomach contains enough HCl to make the hydrogen ion concentration about $0.01\ mole/litre$. The pH of gastric juice is **[NCERT 197**

(a) 0.01 (b) 1

(c) 2 (d) 14

173. Addition of which chemical will decrease the hydrogen ion concentration of an acetic acid solution **[MP PMT 1990]**

(a) NH_4Cl (b) $Al_2(SO_4)_3$

(c) $AgNO_3$ (d) HCN

174. The one which has the highest value of pH is

(a) Distilled water

(b) NH_3 solution in water

(c) NH_3

(d) Water saturated with Cl_2

175. The solution of Na_2CO_3 has pH **[AMU 1988]**

(a) Greater than 7 (b) Less than 7

(c) Equal to 7 (d) Equal to zero

176. Which is not a buffer solution **[CPMT 1990]**

(a) $NH_4Cl + NH_4OH$

(b) $CH_3COOH + CH_3COONa$

(c) CH_3COONH_4

(d) Borax + Boric acid

177. What will be hydrogen ion concentration in moles $litre^{-1}$ of a solution, whose pH is 4.58 **[UPSEAT 2001]**

(a) 2.63×10^{-5} (b) 3.0×10^{-5}

(c) 4.68 (d) None of these

178. Assuming complete dissociation, the pH of a 0.01 M $NaOH$ solution is equal to

[NCERT 1975; CPMT 1977; DPMT 1982; BHU 1997]

(a) 2.0 (b) 14.0

(c) 12.0 (d) 0.01

179. 50 ml of $2\,N$ acetic acid mixed with $10\,ml$ of $1\,N$ sodium acetate solution will have an approximate pH of

[MP PMT/PET 1988]

(a) 4 (b) 5

(c) 6 (d) 7

180. The hydrogen ion concentration of 0.001 M $NaOH$ solution is **[AFMC 1983]**

(a) $1 \times 10^{-2}\, mole / litre$ (b) $1 \times 10^{-11}\, mole / litre$

(c) $1 \times 10^{-14}\, mole / litre$ (d) $1 \times 10^{-12}\, mole / litre$

181. A weak monoprotic acid of 0.1 M, ionizes to 1% in solution. What will be the pH of solution **[MNR 1988]**

(a) 1 (b) 2

(c) 3 (d) 11

182. pH of a solution is 4. The hydroxide ion concentration of the solution would be

[NCERT 1981; CBSE PMT 1991; MP PMT 1994]

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

(c) 10^{-2} (d) 10^{-12}

183. The pH of an aqueous solution containing $[H^+] = 3 \times 10^{-3} M$ is **[MP PET 2001,04]**

(a) 2.471 (b) 2.523

(c) 3.0 (d) – 3

184. pH of blood is maintained constant by mechanism of

[MH CET 2002]

(a) Common ion effect (b) Buffer

(c) Solubility (d) All of these

185. The pH of normal KOH is **[MP PET 1990]**

(a) 1 (b) 0

(c) 14 (d) 7

186. The concentration of hydrogen ion $[H^+]$ in 0.01 M HCl is

[EAMCET 1979]

(a) 10^{12} (b) 10^{-2}

(c) 10^{-1} (d) 10^{-12}

187. A solution of weak acids is diluted by adding an equal volume of water. Which of the following will not change

[JIPMER 1997]

(a) Strength of the acid

(b) The value of $[H_3O^+]$

(c) pH of the solution

(d) The degree of dissociation of acid

188. Ka of H_2O_2 is of the order of **[DCE 2004]**

(a) 10^{-12} (b) 10^{-14}

(c) 10^{-16} (d) 10^{-10}

189. Equivalent weight of an acid **[UPSEAT 2004]**

(a) Depends on the reaction involved

(b) Depends upon the number of oxygen atoms present

(c) Is always same

(d) None of the above

190. pH scale was introduced by **[UPSEAT 2004]**

(a) Arrhenius (b) Sorensen

(c) Lewis (d) Lowry

191. Buffer solution is prepared by mixing **[MH CET 2003]**

(a) Strong acid + its salt of strong base

(b) Weak acid + its salt of weak base

(c) Strong acid + its salt of weak base

(d) Weak acid + its salt of strong base

192. The pH of millimolar HCl is **[MH CET 2004]**

(a) 1 (b) 3

(c) 2 (d) 4

193. Which of the following is a Lewis base **[CPMT 2004]**

(a) $NaOH$ (b) NH_3

(c) BCl_3 (d) All of these

194. What will be the pH value of 0.05 M $Ba(OH)_2$ solution

[CPMT 2004]

(a) 12 (b) 13

(c) 1 (d) 12.96

195. In a mixing of acetic acid and sodium acetate the ratio of concentration of the salts to the acid is increased ten times. Then the pH of the solution **[KCET 20**

(a) Increase by one (b) Decreases by one

(c) Decrease ten fold (d) Increases ten fold

196. The rapid change of pH near the stoichiometric point of an acid-base titration is the basis of indicator detection. pH of the solution is related to ratio of the concentrations of the conjugate acid (HIn) and base (In^-) forms of the indicator by the expression **[CBSE PMT 2004]**

(a) $\log\frac{[HIn]}{[In^-]} = pH - pK_{In}$ (b) $\log\frac{[In^-]}{[HIn]} = pH - pK_{In}$

(c) $\log\frac{[In^-]}{[HIn]} = pK_{In} - pH$ (d) $\log\frac{[HIn]}{[In^-]} = pK_{In} - pH$

197. Which of the following statement(s) is(are) correct

[IIT 1998]

(a) The pH of $1.0\times10^{-8}\,M$ solution of HCl is 8

(b) The conjugate base of $H_2PO_4^-$ is HPO_4^{2-}

(c) Autoprotolysis constant of water increases with temperature

(d) When a solution of a weak monoprotic acid is titrated against a strong base, at half neutralization point $pH = \frac{1}{2}pK_a$

198. An aqueous solution of sodium carbnate has a pH greater than 7 becaue **[DCE 2003]**

(a) It contains more carbonate ions than H_2O molecules

(b) Contains more hydroxide ions than carbonate ions

(c) Na^+ ions react with water

(d) Carbonate ions react with H_2O

199. A pH of 7 signifies **[CPMT 1974; DPMT 1982]**

(a) Pure water (b) Neutral solution

(c) Basic solution (d) Acidic solution

200. Assuming complete dissociation, which of the following aqueous solutions will have the same pH value

[Roorkee Qualifying 1998]

(a) $100\ ml$ of $0.01\ M\ HCl$

(b) $100\ ml$ of $0.01\ M\ H_2SO_4$

(c) $50\ ml$ of $0.01\ M\ HCl$

(d) Mixture of 50 ml of $0.02\ M\ H_2SO_4$ and $50\ ml$ of $0.02\ M\ NaOH$

201. A buffer solution can be prepared from a mixture of

[IIT 1999; KCET 1999; MP PMT 2002]

(a) Sodium acetate and acetic acid in water

(b) Sodium acetate and hydrochloric acid in water

(c) Ammonia and ammonium chloride in water

(d) Ammonia and sodium hydroxide in water

202. Which of the following will not function as a buffer solution

[Roorkee 2000]

(a) $NaCl$ and $NaOH$

(b) $NaOH$ and NH_4OH

(c) CH_3COONH_4 and HCl

(d) Borax and boric acid

203. Which one of the following statements is not true

[AIEEE 2003]

(a) The conjugate base of $H_2PO_4^-$ is HPO_4^{2-}

(b) $pH + pOH = 14$ for all aqueous solutions

(c) The pH of $1\times10^{-8}\,M\,HCl$ is 8

(d) 96,500 coulombs of electricity when passed through a $CuSO_4$ solution deposits 1 *gram* equivalent of copper at the cathode

204. The pH value of $0.1\ M\ HCl$ is approximately 1. What will be the approximate pH value of 0.05 $M\ H_2SO_4$

[MP PMT 1991]

(a) 0.05 (b) 0.5

(c) 1 (d) 2

205. The K_{sp} of $Mg(OH)_2$ is $1\times10^{-12}, 0.01 M\, Mg(OH)_2$ will precipitate at the limiting pH **[DPMT 2005]**

(a) 3 (b) 9

(c) 5 (d) 8

206. The pH of an aqueous solution having hydroxide ion concentration as 1×10^{-5} is **[MP PMT 1991]**

(a) 5 (b) 9

(c) 4.5 (d) 11

Critical Thinking

Objective Questions

1. The K_{SP} of AgI is 1.5×10^{-16}. On mixing equal volumes of the following solutions, precipitation will occur only with

[AMU 2000]

(a) $10^{-7}\,M\,Ag^+$ and $10^{-19}\,M\,I^-$

(b) $10^{-8}\,M\,Ag^+$ and $10^{-8}\,M\,I^-$

(c) $10^{-16}\,M\,Ag^+$ and $10^{-16}\,M\,I^-$

(d) $10^{-9}\,M\,Ag^+$ and $10^{-9}\,M\,I^-$

2. The strongest Bronsted base in the following anion is

[IIT 1981; MP PET 1992, 97; MP PMT 1994; RPMT 1999; KCET 2000; AIIMS 2001; UPSEAT 2002; AFMC 2002; Pb. CET 2004]

(a) ClO^- (b) ClO_2^-

(c) ClO_3^- (d) ClO_4^-

3. Which one of the following compound is not a protonic acid

[CBSE PMT 2003]

(a) $SO_2(OH)_2$ (b) $B(OH)_3$

(c) $PO(OH)_3$ (d) $SO(OH)_2$

4. Calculate the hydrolysis constant of the salt containing NO_2. Given the K_a for HNO_2 $= 4.5\times10^{-10}$

[UPSEAT 2001]

(a) 2.22×10^{-5} (b) 2.02×10^{5}

(c) 4.33×10^{4} (d) 3.03×10^{-5}

5. The molar solubility $(mol\ L^{-1})$ of a sparingly soluble salt MX_4 is $'s'$. The corresponding solubility product is K_{sp}. $'s'$ is given in terms of K_{sp} by the relation **[AIEEE 2004]**

(a) $s=(256K_{sp})^{1/5}$ (b) $s=(128K_{sp})^{1/4}$

(c) $s=(K_{sp}/128)^{1/4}$ (d) $s=(K_{sp}/256)^{1/5}$

6. Electrophiles are **[RPET 2000]**

(a) Lewis acids (b) Lewis base

(c) Bronsted acid (d) Bronsted base

7. Total number of moles for the reaction $2HI \rightleftharpoons H_2+I_2$. if α is degree of dissociation is **[CBSE PMT 1996]**

(a) 2 (b) $2-\alpha$

(c) 1 (d) $1-\alpha$

8. Which one is a Lewis acid **[RPMT 1997]**

(a) ClF_3 (b) H_2O

(c) NH_3 (d) None of these

9. Heat of neutralisation of weak acid and strong base is less than the heat of neutralisation of strong acid and strong base due to

(a) Energy has to be spent for the total dissociation of weak acid

(b) Salt of weak acid and strong base is not stable

(c) Incomplete dissociation of weak acid

(d) Incomplete neutralisation of weak acid

10. pK_a values of two acids A and B are 4 and 5. The strengths of these two acids are related as **[KCET 2001]**

(a) Acid A is 10 times stronger than acids B

(b) Strength of acid A : strengtha of acid $B=4:5$

(c) The strengths of the two acids can not be compared

(d) Acid B is 10 times stronger than acid A

11. 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 **[RPMT 2000]**

(a) 1 : 4 (b) 4 : 1

(c) 1 : 16 (d) 16 : 1

12. An aqueous solution of ammonium acetate is **[NCERT 1980, 81; RPMT 1999]**

(a) Faintly acidic (b) Faintly basic

(c) Fairly acidic (d) Almost neutral

13. The dissociation constant of a weak acid is 1.0×10^{-5}, the equilibrium constant for the reaction with strong base is

[MP PMT 1990]

(a) 1.0×10^{-5} (b) 1.0×10^{-9}

(c) 1.0×10^{9} (d) 1.0×10^{14}

14. The pH of $0.1\ M$ solution of the following salts increases in the order **[IIT 1999]**

(a) $NaCl<NH_4Cl<NaCN<HCl$

(b) $HCl<NH_4Cl<NaCl<NaCN$

(c) $NaCN<NH_4Cl<NaCl<HCl$

(d) $HCl<NaCl<NaCN<NH_4Cl$

15. Which of the following is the strongest acid **[AMU 1999; MH CET 1999, 2002]**

(a) $SO(OH)_2$ (b) $SO_2(OH)_2$

(c) $ClO_3(OH)$ (d) $PO(OH)_3$

16. The strongest of the four acids listed below is **[NCERT 1984]**

(a) $HCOOH$ (b) CH_3COOH

(c) $ClCH_2COOH$ (d) FCH_2COOH

17. Which equilibrium can be described as an acid-base reaction using the Lewis acid-base definition but not using the Bronsted-Lowry definition **[AIIMS 1980]**

(a) $2NH_3+H_2SO_4 \rightleftharpoons 2NH_4^+ + SO_4^{2-}$

(b) $NH_3+CH_3COOH \rightleftharpoons NH_4^+ + CH_3COO^-$

(c) $H_2O+CH_3COOH \rightleftharpoons H_3O^+ + CH_3COO^-$

[KCET 2002]

(d) $[Cu(H_2O)_4]^{2+}+4NH_3 \rightleftharpoons [Cu(NH_3)_4]^{2+}+4H_2O$

18. The hydride ion H^- is stronger base than its hydroxide ion OH^-. Which of the following reaction will occur if sodium hydride (NaH) is dissolved in water **[CBSE PMT 1997]**

(a) $H^-(aq)+H_2O\rightarrow H_2O$

(b) $H^-(aq)+H_2O(l)\rightarrow OH^-+H_2$

(c) $H^-+H_2O\rightarrow$ No reaction

(d) None of these

19. Dissociation constant of a weak acid is 1×10^{-4}. Equilibrium constant of its reaction with strong base is **[UPSEAT 2003]**

(a) 1×10^{-4} (b) 1×10^{10}

(c) 1×10^{-10} (d) 1×10^{14}

20. Arrange the acids (I) H_2SO_3 (II) H_3PO_3 and (III) $HClO_3$ in the decreasing order of acidity **[UPSEAT 2001]**

(a) I > III > II (b) I > II > III

(c) II > III > I (d) III > I > II

21. Self-ionisation of liquid ammonia occurs as, $2NH_3 \rightleftharpoons NH_4^+ + NH_2^-; K=10^{-10}$. In this solvent, an acid might be **[JIPMER 2001]**

(a) NH_4^+

(b) NH_3

(c) Any species that will form NH_4^+

(d) All of these

22. $\Delta H_f(H_2O) = X$; Heat of neutralisation of CH_3COOH and $NaOH$ will be **[BHU 2003]**

(a) Less than $2X$
(b) Less than X
(c) X
(d) Between X and $2X$

23. Which of the following oxides will not give OH^- in aqueous solution **[NCERT 1980]**

(a) Fe_2O_3
(b) MgO
(c) Li_2O
(d) K_2O

24. A precipitate of CaF_2 $(K_{sp} = 1.7 \times 10^{-10})$ will be obtained when equal volume of the following are mixed

[MP PMT 1990, 95; IIT 1982; MNR 1992]

(a) $10^{-4} M\, Ca^{2+}$ and $10^{-4} M\, F^-$
(b) $10^{-2} M\, Ca^{2+}$ and $10^{-3} M\, F^-$
(c) $10^{-5} M\, Ca^{2+}$ and $10^{-3} M\, F^-$
(d) $10^{-3} M\, Ca^{2+}$ and $10^{-5} M\, F^-$

25. The degree of hydrolysis of a salt of weak acid and weak base in its 0.1 *M* solution is found to be 50%. If the molarity of the solution is 0.2 *M*, the percentage hydrolysis of the salt should be**[AMU 1999]**

(a) 50%
(b) 35%
(c) 75%
(d) 100%

26. The pH of 0.1 *M* solution of a weak monoprotic acid 1% ionized is **[UPSEAT 2001; Pb. PMT 2001]**

(a) 1
(b) 2
(c) 3
(d) 4

27. Which one is the strongest acid

[MH CET 1999; AMU 1999,2000; Pb.CET 2001,03; MP PET 2001]

(a) $HClO$
(b) $HClO_2$
(c) H_2SO_4
(d) $HClO_4$

28. Which of the following is Lewis acid **[Pb. CET 2000]**

(a) S
(b) $:CH_2$
(c) $(CH_3)_3B$
(d) All of these

29. The solubility product of $Mg(OH)_2$ is 1.2×10^{-11}. The solubility of this compound in gram per $100\, cm^3$ of solution is

[Roorkee 2000]

(a) 1.4×10^{-4}
(b) 8.16×10^{-4}
(c) 0.816
(d) 1.4

30. What is $[H^+]$ of a solution that is $0.01 M$ in HCN and $0.02 M$ in $NaCN$

$(K_a$ for $HCN = 6.2 \times 10^{-10})$ **[MP PMT 2000]**

(a) 3.1×10^{10}
(b) 6.2×10^{5}
(c) 6.2×10^{-10}
(d) 3.1×10^{-10}

31. Which is neucleophile **[DPMT 2001; RPMT 2002]**

(a) BF_3
(b) NH_3
(c) $BeCl_2$
(d) H_2O

32. The solubility of $CuBr$ is $2 \times 10^{-4} mol/l$ at $25°C$. The K_{sp} value for $CuBr$ is **[AIIMS 2002]**

(a) $4 \times 10^{-8} mol^2 l^{-2}$
(b) $4 \times 10^{-11} mol^2 L^{-1}$
(c) $4 \times 10^{-4} mol^2 l^{-2}$
(d) $4 \times 10^{-15} mol^2 l^{-2}$

33. A 0.004 *M* solution of Na_2SO_4 is isotonic with a 0.010 *M* solution of glucose at same temperature. The apparent degree of disociation of Na_2SO_4 is

[IIT JEE Screening 2004]

(a) 25%
(b) 50%
(c) 75%
(d) 85%

34. K_{sp} for $Cr(OH)_3$ is 2.7×10^{-31}. What is its solubility in *moles / litre*. **[JEE Orissa 2004]**

(a) 1×10^{-8}
(b) 8×10^{-8}
(c) 1.1×10^{-8}
(d) 0.18×10^{-8}

35. pK_a of acetic acid is 4.74. The concentration of CH_3COONa is 0.01 *M*. The *p*H of CH_3OONa is

[Orissa JEE 2004]

(a) 3.37
(b) 4.37
(c) 4.74
(d) 0.474

36. If the solubility product of $AgBrO_3$ and Ag_2SO_4 are 5.5×10^{-5} and 2×10^{-5} respectively, the relationship between the solubilities of these can be correctly represented as **[EAMCET 1985]**

(a) $S_{AgBrO_3} > S_{Ag_2SO_4}$
(b) $S_{AgBrO_3} < S_{Ag_2SO_4}$
(c) $S_{AgBrO_3} = S_{Ag_2SO_4}$
(d) $S_{AgBrO_3} \approx S_{Ag_2SO_4}$

37. The ionisation constant of phenol is higher than that of ethanol because **[JIPMER 2002]**

(a) Phenoxide ion is bulkier than ethanoxide
(b) Phenoxide ion is stronger base than ethanoxide
(c) Phenoxide ion is stabilised through delocalisation
(d) Phenoxide ion is less stable than ethoxide

38. A weak acid *HX* has the dissociation constant $1 \times 10^{-5} M$. It forms a salt *NaX* on reaction with alkali. The degree of hydrolysis of 0.1 *M* solution of *NaX* is

[IIT JEE Screening 2004]

(a) 0.0001%
(b) 0.01%
(c) 0.1%
(d) 0.15%

39. In the equilibrium $A^- + H_2O \rightleftharpoons HA + OH^-$ $(K_a = 1.0 \times 10^{-5})$. The degree of hydrolysis of 0.001 *M* solution of the salt is **[AMU 1999]**

(a) 10^{-3}
(b) 10^{-4}
(c) 10^{-5}
(d) 10^{-6}

40. The sulphide ion concentration $[S^{2-}]$in saturated H_2S solution is 1×10^{-22}. Which of the following sulphides should be quantitatively precipitated by H_2S in the presence of dil. HCl

Sulphide	**Solubility Product**
(I)	1.4×10^{-16}
(II)	1.2×10^{-22}
(III)	8.2×10^{-46}
(IV)	5.0×10^{-34}

(a) I, II (b) III, IV

(c) II, III, IV (d) Only I

41. When equal volumes of the following solutions are mixed, precipitation of $AgCl\,(K_{sp}=1.8\times 10^{-10})$ will occur only with **[IIT 1988; CBSE PMT PMT 1992; DCE 2000]**

(a) $10^{-4}\,M\,Ag^+$ and $10^{-4}\,M\,Cl^-$

(b) $10^{-5}\,M\,Ag^+$ and $10^{-5}\,M\,Cl^-$

(c) $10^{-6}\,M\,Ag^+$ and $10^{-6}\,M\,Cl^-$

(d) $10^{-10}\,M\,Ag^+$ and $10^{-10}\,M\,Cl^-$

42. K_{sp} of an electrolyte AB is 1×10^{-10}. $[A^+]=10^{-5}\,M$, which concentration of B^- will not give precipitate of AB **[BHU 2003]**

(a) 5×10^{-6} (b) 1×10^{-5}

(c) 2×10^{-5} (d) 5×10^{-5}

43. One litre of water contains $10^{-7}\,mole$ hydrogen ions. The degree of ionization in water will be **[CPMT 1985, 88, 93]**

(a) $1.8\times 10^{-7}\%$ (b) $0.8\times 10^{-9}\%$

(c) $3.6\times 10^{-7}\%$ (d) $3.6\times 10^{-9}\%$

44. If the solubility products of $AgCl$ and $AgBr$ are 1.2×10^{-10} and 3.5×10^{-13} respectively, then the relation between the solubilities (denoted by the symbol$'S'$) of these salts can correctly be represented as**[MP PET 1994]**

(a) S of $AgBr$ is less than that of $AgCl$

(b) S of $AgBr$ is greater than that of $AgCl$

(c) S of $AgBr$ is equal to that of $AgCl$

(d) S of $AgBr$ is 10^6 times greater than that of $AgCl$

45. If the solubility product of lead iodide (PbI_2) is 3.2×10^{-8}, then its solubility in moles/litre will be **[MP PMT 1990]**

(a) 2×10^{-3} (b) 4×10^{-4}

(c) 1.6×10^{-5} (d) 1.8×10^{-5}

46. Calculate the solubility of $AgCl$ (s) in $0.1\,M\,NaCl$ at $25°C$. $K_{sp}(AgCl)=2.8\times 10^{-10}$ **[UPSEAT 2001]**

(a) $3.0\times 10^{-8}\,ML^{-1}$ (b) $2.5\times 10^{-7}\,ML^{-1}$

(c) $2.8\times 10^{-9}\,ML^{-1}$ (d) $2.5\times 10^{7}\,ML^{-1}$

47. The solubility product of a binary weak electrolyte is 4×10^{-10} at $298\,K$. Its solubility in mol dm^{-3} at the same temperature is **[KCET 2001]**

(a) 4×10^{-5} (b) 2×10^{-5}

(c) 8×10^{-10} (d) 16×10^{-20}

48. Solubility of $AgCl$ at 20^oC is $1.435\times 10^{-3}\,gm\ per\ litre$. The solubility product of $AgCl$ is

[CPMT 1989; BHU 1997; AFMC 2000; CBSE PMT 2002]

(a) 1×10^{-5} (b) 1×10^{-10}

(c) 1.435×10^{-5} (d) 108×10^{-3}

49. pK_a value for acetic acid at the experimental temperature is 5. The percentage hydrolysis of $0.1\,M$ sodium acetate solution will be

(a) 1×10^{-4} (b) 1×10^{-2}

(c) 1×10^{-3} (d) 1×10^{-5}

50. At 30^oC, the solubility of $Ag_2CO_3\,(K_{sp}=8\times 10^{-12})$ would be greatest in one litre of **[MP PMT 1990]**

(a) $0.05\,M\,Na_2CO_3$ (b) $0.05\,M\,AgNO_3$

(c) Pure water (d) $0.05\,M\,NH_3$

51. The values of K_{sp} for CuS, Ag_2S and HgS are 10^{-31},10^{42} and 10^{-54} respectively. The correct order of their solubility in water is **[MP PMT 2003]**

(a) $Ag_2S > HgS > CuS$ (b) $HgS > CuS > Ag_2S$

(c) $HgS > Ag_2S > CuS$ (d) $Ag_2S > CuS > HgS$

52. The pH of a soft drink is 3.82. Its hydrogen ion concentration will be **[MP PET 1990]**

(a) $1.96\times 10^{-2}\ mol/l$ (b) $1.96\times 10^{-3}\ mol/l$

(c) $1.5\times 10^{-4}\ mol/l$ (d) $1.96\times 10^{-1}\ mol/l$

53. The pH of a solution at $25°C$ containing $0.10\,m$ sodium acetate and $0.03\,m$ acetic acid is (pK_a for $CH_3COOH=4.57$) **[AIIMS 2002; BHU 2002]**

(a) 4.09 (b) 5.09

(c) 6.10 (d) 7.09

54. A weak acid is 0.1% ionised in 0.1 M solution. Its pH is **[BVP 2004]**

(a) 2 (b) 3

(c) 4 (d) 1

55. The solubility product of As_2S_3 is 2.8×10^{-72}. What is the solubility of As_2S_3

(a) $1.09\times 10^{-15}\ mole/litre$

(b) $1.72\times 10^{-15}\ mole/litre$

(c) $2.3\times 10^{-16}\ mole/litre$

(d) $1.65\times 10^{-36}\ mole/litre$

56. For a weak acid HA with dissociation constant 10^{-9}, pOH of its 0.1 M solution is **[CBSE PMT 1989]**

(a) 9 (b) 3

(c) 11 (d) 10

57. The dissociation of water at $25^\circ C$ is $1.9\times10^{-7}\%$ and the density of water is $1.0\,g/cm^3$. The ionisation constant of water is **[IIT 1995]**

(a) 3.42×10^{-6} (b) 3.42×10^{-8}

(c) 1.00×10^{-14} (d) 2.00×10^{-16}

58. What is the pH of 0.01 M glycine solution? For glycine, $Ka_1 = 4.5\times10^{-3}$ and $Ka_2 = 1.7\times10^{-10}$ at 298 K **[AIIMS 2004]**

(a) 3.0 (b) 10.0

(c) 6.1 (d) 7.2

59. The concentration of $[H^+]$ and concentration of $[OH^-]$ of a 0.1 aqueous solution of 2% ionised weak acid is

[Ionic product of water $= 1\times10^{-14}$]

[DPMT 2004; CBSE PMT 1999]

(a) $2\times10^{-3}\,M$ and $5\times10^{-12}\,M$

(b) $1\times10^{3}\,M$ and $3\times10^{-11}\,M$

(c) $0.02\times10^{-3}\,M$ and $5\times10^{-11}\,M$

(d) $3\times10^{-2}\,M$ and $4\times10^{-13}\,M$

60. If solubility product of $HgSO_4$ is 6.4×10^{-5}, then its solubility is **[BHU 2004]**

(a) $8\times10^{-3}\,mole/litre$ (b) $6.4\times10^{-5}\,mole/litre$

(c) $6.4\times10^{-3}\,mole/litre$ (d) $2.8\times10^{-6}\,mole/litre$

61. At 298K a 0.1 M CH_3COOH solution is 1.34% ionized. The ionization constant K_a for acetic acid will be **[AMU 2002; AFMC 2005]**

(a) 1.82×10^{-5} (b) 18.2×10^{-5}

(c) 0.182×10^{-5} (d) None of these

62. Hydrogen ion concentration of an aqueous solution is $1\times10^{-4}\,M$. The solution is diluted with equal volume of water. Hydroxyl ion concentration of the resultant solution in terms of *mol* dm^{-3} is **[KCET 2001]**

(a) 1×10^{-8} (b) 1×10^{-6}

(c) 2×10^{-10} (d) 0.5×10^{-10}

63. Which one of the following is not a buffer solution **[AIIMS 2003]**

(a) $0.8\,M\,H_2S + 0.8\,M\,KHS$

(b) $2M\,C_6H_5NH_2 + 2M\,C_6H_5\overset{+}{N}H_3\,Br$

(c) $3\,M\,H_2CO_3 + 3\,M\,KHCO_3$

(d) $0.05\,M\,KClO_4 + 0.05\,M\,HClO_4$

64. The hydrogen ion concentration of a $0.006\,M$ benzoic acid solution is $(K_a = 6\times10^{-5})$ **[MP PET 1994]**

(a) 0.6×10^{-4} (b) 6×10^{-4}

(c) 6×10^{-5} (d) 3.6×10^{-4}

65. Calculate the amount of $(NH_4)_2SO_4$ in grams which must be added to 500 *ml* of $0.200\,M\,NH_3$ to yield a solution with $pH = 9.35$ (K_b for $NH_3 = 1.78\times10^{-5}$) **[UPSEAT 2001]**

(a) 10.56 *gm* (b) 15 *gm*

(c) 12.74 *gm* (d) 16.25 *gm*

66. *pH* of a solution produced when an aqueous solution of pH 6 is mixed with an equal volume of an aqueous solution of pH 3 is about **[KCET 2001]**

(a) 3.3 (b) 4.3

(c) 4.0 (d) 4.5

67. Calculate the H^+ ion concentration in a 1.00 (M) HCN *litre* solution $(K_a = 4\times10^{-10})$ **[Bihar CEE 1995]**

(a) $4\times10^{-14}\,mole/litre$

(b) $2\times10^{-5}\,mole/litre$

(c) $2.5\times10^{-5}\,mole/litre$

(d) None of these

68. The number of moles of hydroxide (OH^-) ion in 0.3 *litre* of 0.005 M solution of $Ba(OH)_2$ is **[JIPMER 2001]**

(a) 0.0050 (b) 0.0030

(c) 0.0015 (d) 0.0075

69. The *pH* of pure water or neutral solution at $50°C$ is .... $(pK_w = 13.26 = 13.26$ at $50°C)$

[Pb. PMT 2002; DPMT 2002]

(a) 7.0 (b) 7.13

(c) 6.0 (d) 6.63

70. *pH* of 0.1 M solution of a weak acid (HA) is 4.50. It is neutralised with $NaOH$ solution to decrease the acid content to half *pH* of the resulting solution **[JIPMER 2002]**

(a) 4.50 (b) 8.00

(c) 7.00 (d) 10.00

71. If $50\,ml$ of $0.2\,M\,KOH$ is added to $40\,ml$ of $0.5\,M\,HCOOH$, the pH of the resulting solution is $(K_a = 1.8\times10^{-4})$ **[MH CET 2000]**

(a) 3.4 (b) 7.5

(c) 5.6 (d) 3.75

72. The Bronsted acids in the reversible reaction are $HCO_3^-(aq.) + OH^-(aq.) \rightleftarrows CO_3^{2-}(aq.) + H_2O$ **[DPMT 2002]**

(a) OH^- and CO_3^{2-}

(b) OH^- and H_2O

(c) HCO_3^- and H_2O

(d) HCO_3^- and CO_3^{2-}

73. A 0.1N solution of an acid at room temperature has a degree of ionisation 0.1. The concentration of OH^- would be **[MH CET 1999]**

(a) $10^{-12}\,M$ (b) $10^{-11}\,M$

(c) $10^{-9}\,M$ (d) $10^{-2}\,M$

74. Increasing order of acidic character would be **[RPMT 1999]**

(a) $CH_3COOH < H_2SO_4 < H_2CO_3$

(b) $CH_3COOH < H_2CO_3 < H_2SO_4$

(c) $H_2CO_3 < CH_3COOH < H_2SO_4$

(d) $H_2SO_4 < H_2CO_3 < CH_3COOH$

75. The correct order of increasing $[H_3O^+]$ in the following aqueous solutions is **[UPSEAT 2000]**

(a) 0.01 M H_2S <0.01 M H_2SO_4 < 0.01 M $NaCl$ < 0.01 M $NaNO_2$

(b) 0.01 M $NaCl$ <0.01 M $NaNO_2$ < 0.01 M H_2S < 0.01 M H_2SO_4

(c) 0.01 M $NaNO_2$ <0.01 M $NaCl$ < 0.01 M H_2S < 0.01 M H_2SO_4

(d) 0.01 M H_2S < 0.01 M $NaNO_2$ < 0.01 M $NaCl$ < 0.01 M H_2SO_4

76. A base dissolved, in water, yields a solution with a hydroxyl ion concentration of $0.05\, mol\, litre^{-1}$. The solution is

[CBSE PMT 2000]

(a) Basic (b) Acid

(c) Neutral (d) Either (b) or (c)

77. In the given reaction, the oxide of sodium is $....\begin{bmatrix} 4Na + O_2 \rightarrow 2Na_2O \\ Na_2O + H_2O \rightarrow 2NaOH \end{bmatrix}$ **[Orissa JEE 2002]**

(a) Acidic (b) Basic

(c) Amphoteric (d) Neutral

78. What is the pH of a 1M CH_3COOH a solution K_a of acetic acid $= 1.8 \times 10^{-5}$.

$K = 10^{-14}\, mol^2 litre^{-2}$ **[DPMT 2002]**

(a) 9.4 (b) 4.8

(c) 3.6 (d) 2.4

AR Assertion & Reason

For AIIMS Aspirants

Read the assertion and reason carefully to mark the correct option out of the options given below :

(a) *If both assertion and reason are true and the reason is the correct explanation of the assertion.*
(b) *If both assertion and reason are true but reason is not the correct explanation of the assertion.*
(c) *If assertion is true but reason is false.*
(d) *If the assertion and reason both are false.*
(e) *If assertion is false but reason is true.*

1. Assertion : *p*H of hydrochloric acid solution is less than that of acetic acid solution of the same concentration.

Reason : In equimolar solutions, the number of titrable protons present in hydrochloric acid is less than that persent in acetic acid.

[NDA 1999]

2. Assertion : A ionic product is used fro any types of electrolytes whereas solubility product is applicable only to sparingly soluble salts.

Reason : Ionic product is defined at any stage of the raction whereas solubility product is only appicable to the saturation stage.

[AIIMS 2001]

3. Assertion : A solution of $FeCl_3$ in water produce brown precipitate on standing.

Reason : Hydrolysis of $FeCl_3$ takes place in water

4. Assertion : $BaCO_3$ is more soluble in HNO_3 than in plain water.

Reason : Carbonate is a weak base and reacts with the H^+ from the strong acid, casuing the barium salt to dissociate.

5. Assertion : $CHCl_3$ is more acidic than CHF_3.

Reason : The conjugate base of $CHCl_3$ is more stable than CHF_3.

6. Assertion : Addition of silver ions to a mixture of aqueous sodium chloride and sodium bromide solution will first precipitate $AgBr$ rather than $AgCl$.

Reason : K_{sp} of $AgCl < K_{sp}$ of $AgBr$. **[AIIMS 2004]**

7. Assertion : The pK_a of acetic acid is lower than that of phenol.

Reason : Phenoxide ion is more resonance stabilized. **[AIIMS 2004]**

8. Assertion : Sb (III) is not precipitated as sulphide when in its alkaline solution H_2S is passed.

Reason : The concentration of S^{2-} ion in alkaline medium is inadequate for precipitation.

[AIIMS 2004]

9. Assertion : Ionic reactions are not instantaneous.

Reason : Oppositely charged ions exert strong forces.

Answers

Electrical conductors, Arrenius theory and Ostwald's dilution law

1	c	2	b	3	b	4	b	5	b
6	a	7	b	8	d	9	a	10	a
11	c	12	d	13	b	14	b	15	d
16	d	17	b	18	b	19	c	20	c
21	a	22	c	23	b	24	d	25	c
26	d	27	a	28	d	29	d	30	c
31	b	32	c	33	c	34	c	35	a
36	c								

Acids and Bases

1	a	2	d	3	a	4	c	5	b
6	d	7	b	8	c	9	b	10	a
11	a	12	c	13	a	14	d	15	d
16	b	17	b	18	d	19	a	20	b
21	d	22	b	23	b	24	c	25	d
26	c	27	c	28	d	29	b	30	b
31	b	32	c	33	d	34	a	35	a
36	a	37	c	38	c	39	c	40	a
41	c	42	d	43	c	44	a	45	a
46	c	47	c	48	d	49	a	50	b
51	b	52	a	53	a	54	a	55	d
56	b	57	ab	58	a	59	c	60	d
61	a	62	b	63	b	64	b	65	a
66	a	67	d	68	a	69	a	70	b
71	b	72	d	73	b	74	d	75	b
76	b	77	a	78	a	79	b	80	a
81	d	82	b	83	b	84	d	85	bd
86	a	87	d	88	a	89	d	90	c
91	a	92	a	93	a	94	d	95	b
96	c	97	b	98	c	99	d	100	c
101	a	102	a	103	c	104	c	105	a
106	a	107	d	108	b	109	c	110	d

111	b	112	c	113	c	114	a	115	c
116	d	117	d	118	c	119	c	120	c
121	a	122	d	123	b	124	a	125	b
126	c	127	c	128	d	129	c	130	b
131	b	132	c	133	a	134	c	135	a
136	c	137	c	138	b	139	b	140	b
141	b	142	a	143	b	144	a	145	a
146	a	147	b	148	b	149	c	150	a
151	c	152	c	153	d	154	c	155	a
156	b	157	a	158	d	159	b	160	d
161	b	162	d	163	c	164	d	165	a

Common ion effect, Isohydric solutions, Solubility product, Ionic product of water and Salt hydrolysis

1	b	2	c	3	b	4	b	5	b
6	d	7	c	8	a	9	c	10	a
11	b	12	b	13	a	14	d	15	d
16	b	17	c	18	c	19	d	20	c
21	d	22	d	23	d	24	b	25	a
26	d	27	a	28	d	29	b	30	b
31	a	32	d	33	a	34	c	35	d
36	b	37	b	38	c	39	b	40	a
41	a	42	d	43	b	44	a	45	b
46	d	47	b	48	c	49	c	50	c
51	b	52	d	53	a	54	d	55	c
56	d	57	b	58	c	59	d	60	c
61	b	62	b	63	a	64	bc	65	d
66	a	67	d	68	a	69	d	70	c
71	c	72	d	73	d	74	d	75	a
76	c	77	b	78	d	79	b	80	d
81	c	82	c	83	d	84	c	85	d
86	d	87	b	88	c	89	b	90	a
91	a	92	a	93	c	94	c	95	a
96	d	97	d	98	b	99	b	100	ab
101	d	102	c	103	c	104	a	105	c
106	c	107	c	108	a	109	b	110	b
111	b	112	a	113	a	114	a	115	b
116	a								

Hydrogen ion concentration - *pH* scale and Buffer solution

1	a	2	c	3	d	4	b	5	b
6	d	7	a	8	c	9	b	10	d
11	a	12	d	13	c	14	a	15	b
16	d	17	b	18	c	19	c	20	d
21	b	22	c	23	a	24	d	25	b
26	a	27	a	28	d	29	b	30	b
31	d	32	b	33	a	34	a	35	c
36	c	37	c	38	a	39	a	40	d
41	a	42	a	43	b	44	d	45	b
46	a	47	b	48	a	49	c	50	a
51	a	52	c	53	c	54	c	55	d
56	a	57	c	58	b	59	b	60	c
61	b	62	c	63	b	64	b	65	c
66	a	67	b	68	d	69	a	70	c
71	d	72	a	73	b	74	a	75	b
76	a	77	a	78	c	79	b	80	a
81	b	82	c	83	b	84	b	85	c
86	b	87	a	88	a	89	a	90	d
91	b	92	c	93	a	94	a	95	c
96	b	97	b	98	b	99	b	100	b
101	c	102	c	103	a	104	a	105	c
106	a	107	c	108	c	109	d	110	b
111	d	112	d	113	b	114	a	115	c
116	d	117	b	118	b	119	d	120	c
121	d	122	b	123	c	124	d	125	b
126	b	127	a	128	c	129	b	130	d
131	c	132	d	133	c	134	c	135	a
136	c	137	c	138	c	139	c	140	b
141	b	142	a	143	d	144	b	145	b
146	d	147	c	148	b	149	a	150	d
151	b	152	d	153	b	154	b	155	c
156	d	157	c	158	a	159	c	160	d
161	c	162	d	163	b	164	a	165	d
166	a	167	d	168	b	169	c	170	a
171	c	172	c	173	d	174	b	175	a
176	c	177	a	178	c	179	a	180	b
181	c	182	b	183	b	184	b	185	c
186	b	187	b	188	a	189	a	190	b
191	d	192	b	193	b	194	b	195	a
196	b	197	b	198	b	199	a	200	ad
201	ac	202	a bc	203	c	204	c	205	b
206	b								

Critical Thinking Questions

1	b	2	a	3	b	4	a	5	d
6	a	7	c	8	a	9	a	10	a
11	b	12	d	13	c	14	d	15	c
16	d	17	d	18	b	19	b	20	d
21	a	22	b	23	a	24	b	25	a
26	c	27	d	28	d	29	b	30	d
31	b	32	a	33	c	34	a	35	a
36	b	37	c	38	b	39	a	40	b
41	a	42	a	43	a	44	a	45	a
46	c	47	b	48	b	49	b	50	c
51	d	52	c	53	b	54	c	55	a
56	d	57	d	58	c	59	a	60	a

61	a	62	c	63	d	64	b	65	a
66	a	67	b	68	b	69	d	70	b
71	a	72	c	73	a	74	c	75	c
76	a	77	b	78	a				

Assertion & Reason

1	c	2	b	3	a	4	a	5	a
6	c	7	c	8	a	9	e		

Answers and Solutions

Electrical conductors, Arrhenius theory and Ostwald's dilution law

1. (c) $C_{12}H_{22}O_{11}$ is a sugar and non-electrolyte.

2. (b) It is a weak electrolyte since it is slightly ionized.

3. (b) It is a weak electrolyte since slightly ionized.

4. (b) Because the degree of dissociation is inversely proportional to the concentration of the electrolyte.

5. (b) Electrolytes are those substances which on dissolving in water give ions.

6. (a) $K=\dfrac{\alpha^2C}{1-\alpha}$; $\alpha=\dfrac{0.01}{100}\approx 1$ $\therefore K=\alpha^2C=\left[\dfrac{0.01}{100}\right]^2\times 1$

$=1\times 10^{-8}$.

7. (b) As $NaCl$ ionises completely to yield free ions.

8. (d) $CH_3COONa \rightleftharpoons CH_3COO^- + Na^+$

$H_2O \rightleftharpoons H^+ + OH^-$

$CH_3COOH + NaOH$

9. (a) $NaCl$, being a salt, is a strong electrolyte.

10. (a) We can determine by measurement of very dilute HF solutions.

11. (c) According to the Ostwald's dilution formula $\alpha^2=\dfrac{K(1-\alpha)}{C}$. But for weak electrolytes α is very small. So that $(1-\alpha)$ can be neglected. So that $\alpha=\sqrt{\dfrac{K_a}{C}}$.

12. (d) Arrhenius proposed the theory of ionisation.

13. (b) higher the dielectric constant of a solvent more of its ionising power.

14. (b) $\alpha \propto$ dilution of solution.

15. (d) Generally ionic compound are conduct electricity in fused state.

16. (d) According to Ostwald's dilution law because degree of ionization is directly proportional to the dilution.

17. (b) The degree of ionisation of a solute depends upon its nature, concentration, and temperature.

18. (b) Mathematical form of Ostwald's dilution law.

19. (c) It is a weak electrolyte because it's ionization is very less.

20. (c) When we add NH_4OH in NH_4Cl solution ionization of NH_4OH is decreased due to common ion effect.

21. (a) $BaCl_2 \rightleftharpoons Ba^{2+} + 2Cl^-$

Initially 1 0 0

After dissociation $a-\alpha$ α 2α

Total = $1-\alpha+\alpha+2\alpha=1+2\alpha$

$\alpha=\dfrac{1.98-1}{\alpha}=\dfrac{0.98}{\alpha}=0.49$

for a mole $\alpha=0.49$

For 0.01 mole $\alpha=\dfrac{0.49}{0.01}=49$

22. (c) It is an ionic salt.

23. (b) Due to common ion effect of H^+.

24. (d) Current does not affect the degree of ionization.

25. (c) Polar solvent facilitate ionisation of strong electrolytes due to dipole-ion attraction.

26. (d) $HA \rightleftharpoons H^+ + A^-$

$K=Ca^2=0.1\times(10^{-4})^2=10^{-9}$

28. (d) Degree of dissociation α = ?

Normality of solution = 0.1 $N=\dfrac{1}{10}N$

Volume = 10 *litre*

Dissociation constant $K=1\times 10^{-5}$

$K=\dfrac{\alpha^2}{V}$; $\alpha=\sqrt{KV}=\sqrt{1\times 10^{-5}\times 10}$; $\alpha=1\times 10^{-2}$

29. (d) $MgCl_2 \rightleftharpoons Mg^{++}+2Cl^-$ in aqueous solution it is dissociated into ions.

30. (c) Ostwald's dilution formula is $\alpha^2=K(1-\alpha)/C$ but for weak electrolyte α is very small. So that $(1-\alpha)$ is neglected for weak electrolytes. So for weak electrolyte the dilution formula is $\alpha=\sqrt{\dfrac{K}{C}}$.

31. (b) $K_a=C\alpha^2=0.2\times\left(\dfrac{32}{100}\right)^2=2.048\times 10^{-4}$.

32. (c) $H_2SO_4 \rightleftharpoons H^+ + H^-SO_4 \rightleftharpoons H^+ + SO_4^{--}$

Because of it is completely ionised.

33. (c) Colour of electrolyte depends on the nature of both ions.

e.g. $CuSO_4$ is blue because Cu^{2+} ions are blue.

34. (c) Ionisation depends upon dilution, when dilution increases then ionisation is also increased.

35. (a) More is K_a, lesser is $pK_a\,(pK_a = -\log K_a)$ more is acidic strength.

Acids and Bases

1. (a) CO doesn't have a vacant *d*-orbital.

2. (d) $\underbrace{HClO_4 + H_2O \rightleftharpoons H_3O^+ + ClO_4^-}_{\text{Conjugate acid and base pair}}$

3. (a) $FeCl_3 + 3H_2O \rightleftharpoons Fe(OH)_3 + 3HCl$. Strong acid and weak base.

4. (c) $Na_2CO_3 + 2H_2O \rightleftharpoons 2NaOH + H_2CO_3$

5. (b) Those substance accept the proton are called Bronsted base and which is donate the proton are called Bronsted acid.

$HCO_3^- + H^+ \rightleftharpoons H_2CO_3$ Bronsted base.

$HCO_3^- \rightleftharpoons H^+ + CO_3^{--}$ Bronsted acid.

7. (b) The value of pK_a for strong acid is less.

8. (c) Because it is a salt of strong base and weak acid.

9. (b) Because it is conjugate base of weak acid.

$CH_3COOH \rightleftharpoons CH_3COO^- + H^+$.

11. (a) Those compound which accept H^+ is called bronstad base NO_3^- accept H^+ and form HNO_3. So it is a base.

13. (a) Larger the size of halogen atom less is the back donation of electrons into empty 2*p* orbital of *B*.

14. (d) $H_2O + NH_3 \rightleftharpoons NH_4^+ + OH^-$

16. (b) $NH_2^- \rightleftharpoons NH^{-2} + H^+$

Conjugate acid, base pair.

17. (b) Those substances which lose proton are called acid.

e.g. strong acid have a strong tendency to donate a proton.

18. (d) Electron donating species called nucleophile. NH_3 have a lone pair of electron.

19. (a) H_2O acts as acid as it provides H^+ to NH_3.

20. (b) $CH_3COOH + HF \rightleftharpoons CH_3COOH_2^+ + F^-$. HF gives H^+ to the CH_3COOH. So it is a conjugate base of *HF*.

22. (b) $Ba(NO_3)_2$ does not undergo hydrolysis.

23. (b) H_2SO_4 is a mineral acid.

24. (c) Because it is a electron pair acceptor it's central atom have a vacant *d*-orbital.

25. (d) $HClO_4$ is a acid and their conjugate base is ClO_4^-.

26. (c) It is completely ionised because their ionization is very high.

27. (c) $NaHCO_3$ in water is alkaline in nature due to hydrolysis of HCO_3^- ion. $NaCO_3 \rightleftharpoons Na^+ + HCO_3^-$

28. (d) In this reaction H_2O acts as a acid.

29. (b) The basic character of hydride decreases down the group.

30. (b) Its ionization is very less.

32. (c) Hydrolysis of Cu^{2+} produces H^+ ions in solution.

$Cu^{2+} + 2H_2O \rightarrow Cu(OH)_2 + 2H^+$

33. (d) Both possess the tendancy to accept proton.

34. (a) HF does not give proton easily.

35. (a) In weak electrolyte the degree of dissociation is very small. So it increases with increasing dilution.

36. (a) Because it is a electron pair acceptor.

37. (c) $H_3PO_4 \xrightleftharpoons{\text{I step}} H^+ + H_2PO_4^-$

$H_2PO_4^- \xrightleftharpoons{\text{II step}} H^+ + HPO_4^{--}$

$HPO_4^{--} \xrightleftharpoons{\text{III step}} H^+ + PO_4^{---}$

38. (c) Na_2HPO_4 on hydrolysis of HPO_4^{2-} ion produces free OH^- ion in solution.

39. (c) $H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$.

40. (a) $Fe^{3+} + 3H_2O \rightarrow Fe(OH)_3 + 3H^+$

41. (c) $\underset{\text{Conjugate acid}}{H_2PO_4^-} \rightleftharpoons H^+ + HPO_4^{2-}$ action.

42. (d) $H^+ + OH^- \rightleftharpoons H_2O,$ it is a neutralization reaction.

43. (c) It is a conjugate base of the strong acid.

44. (a) Because those acid have higher pK_a value are weak acid.

45. (a) Conjugate acid is obtained from the base by gain of H^+.

46. (c) CH_3COOH is weak acid shows dissociation equilibrium as $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

47. (c) $pK_a = \log_{10}\dfrac{1}{K_a}$

48. (d) A weak acid and strong base.

49. (a) In NaH_2PO_2 both hydrogen are bonded with '*P*,' so it is not replaceable.

50. (b) Aqueous solution of Na_2CO_3 is alkaline due to hydrolysis of CO_3^{--}.

51. (b) Because they give CO_3^{--} in solution.

52. (a) M.eq. of 0.2*M* $H_2SO_4 = \frac{2\times 0.2M}{1000}\times 100 = 0.04$ *m*/*l*

M.eq. of .2*M* $NaOH = \frac{0.2}{1000}\times 100 = 0.02$ *m*/*l*

left $[H^+] = .04 - .02 = .02$.

Total volume $= 200 = \frac{.02}{200} = .0001 = 10^{-4}\,M$

$pH = 4$.

53. (a) H_3BO_3 is a weak monobasic acid it does not act as a H^+ donor but behaves as a Lewis acid.

54. (a) Because $SnCl_2$ is a electron acceptor according to Lewis concept.

55. (d) *ROH* is a Lewis base because it has an lone pair of electron.

56. (b) $H_2SO_4 + 2H_2O \rightleftharpoons 2H_3O^+ + SO_4^{--}$

$NaOH \rightleftharpoons Na^+ + OH^-$

1 mole of H_2SO_4 acid gives 2 moles of H_3O^+ ions. So 2 moles of OH^- are required for complete neutralization.

57. (ab) Diprotic solvents give $2H^+$ ions or OH^- ions.

58. (a) $N_{NaOH} = 1\times 1 = 1N$

$N_{H_2SO_4} = 2\times 10 = 20\,N$

M.eq. of $NaOH = 1\times 100 = 100$

M.eq. of $H_2SO4 = 20\times 10 = 200$

Thus M.eq. of acid are left and therefore *pH*< 7, so the resulting mixture will be acidic.

60. (d) PH_3 is a Lewis base.

61. (a) Because of CH_3COONa is a salt of weak acid and strong base.

62. (b) Acid $\xrightarrow{-H^+}$ conjugate base.

Base $\xrightarrow{+H^+}$ conjugate acid.

63. (b) $\underset{\text{Acid}}{HCl} \rightarrow \underset{\text{Base}}{Cl^-}$

65. (a) $\underset{\text{Aciedic character decreases}}{HClO_4 > H_2SO_4 > HCl > HNO_3}$.

66. (a) Those substances which accept the H^+ are called conjugate base.

67. (d) NH_3 is a Lewis base, which donate a lone pair of electron.

69. (a) $FeCl_3 + 3H_2O \rightleftharpoons Fe(OH)_3 + 3HCl$. Strong acid have less than 7 *pH*.

70. (b) It donates their e^- pair.

71. (b) The strength of the acid will depend upon the proton donation.

73. (b) Because it is a salt of strong acid and strong base.

$H^+{}_{(aq)} + Cl^-{}_{(aq)} + Na^+{}_{(aq)} + OH^-{}_{(aq)} \rightleftharpoons$

$H_2O_{(l)} + Na^+{}_{(aq)} + Cl^-$

74. (d) CCl_4 is not a Lewis or bronsted acid. It does not contain H^+.

75. (b) $NH_4Cl + H_2O \rightleftharpoons \underset{\text{Weak base}}{NH_4OH} + \underset{\text{Strong acid}}{HCl}$. So it is acidic in nature.

76. (b) *Cu*(II) complexes are blue. The four water molecules are attached with secondary valencies of the metal atom e.g. $[Cu(H_2O)_4]SO_4 . H_2O$.

77. (a) The species which can accept as well as donate H^+ can act both as an acid and a base.

$\underset{\text{base}}{HSO_4^-} + H^+ \rightleftharpoons H_2SO_4$

$\underset{\text{acid}}{HSO_4^-} \rightleftharpoons SO_4^{2-} + H^+$

78. (a) NH_4^+ is the weakest acid. So its conjugate base is strongest.

79. (b) Ag^+ is an electron deficient compound and hence is a Lewis acid.

80. (a) $\underset{\text{Conjugate acid}}{H_3PO_4} \rightleftharpoons H^+ + H_2PO_4^-$

81. (d) $\underset{\text{Conjugate acid}}{HS_2O_8^-} \rightleftharpoons H^+ + \underset{\text{Conjugate base}}{S_2O_8^{--}}$

82. (b) PH_3 donates electron pair to BCl_3.

83. (b) The conjugate base of weak acid is a strong base.

84. (d) $OH^- \rightarrow \underset{\text{Conjugate base of } OH^-}{O^{2-} + H^+}$

85. (bd) Presence of lone pair of electron and they donate two electron pairs.

86. (a) *On increasing oxidation number Acidic strength increases.*

87. (d) H_3PO_4 is shows +5 maximum oxidation state.

88. (a) $C_2H_5^-$ is a strongest base.

89. (d) *NaOCl* is a mixture of strong base and weak acid.

90. (c) NH_4OH gives minimum OH^- ion. So it is a weak base.

91. (a) Solution become acidic and methyl orange act on acidic *pH*.

92. (a) Larger is bond length, more is acidic nature (for halogen acids). *HF* bond length is small.

93. (a) *HCl* is accepting proton in *HF* medium and acts as weak base.

95. (b) For oxoacids of the same element, the acidic strength increases with increase in the oxidation number of that element.

$\overset{+1}{HClO} < \overset{+3}{HClO_2} < \overset{+5}{HClO_3} < \overset{+7}{HClO_4}$

96. (c) Because their conjugate base and conjugate acids are strong.

97. (b) *HCl* is a strong acid and their conjugate base is a very weak base.

100. (c) $NH_3 \rightleftharpoons NH_2^- + H^+$

101. (a) $H_2SO_4 \rightleftharpoons H^+ + HSO_4^- \rightleftharpoons H^+ + SO_4^{--}$

HSO_4^- is conjugate base. But it is also an acid because it lose H^+.

102. (a) BF_3 is a Lewis acid because '*B*' has incomplete octet.

104. (c) H_3O^+, however it exists as $H_9O_4^+$.

105. (a) $Al_2(SO_4)_3$ is a salt of weak base $Al_2(OH)_3$ and strong acid H_2SO_4.

106. (a) Al^{+++} of $AlCl_3$ undergoes hydrolysis.

107. (d) $H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$

109. (c) Conjugate base of H_2SO_4 is HSO_4^-.

110. (d) Presence of lone pair.

111. (b) According to Bronsted principle HNO_3 is acid they give H^+ in aqueous solution and form NO_3^-.

112. (c) $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$.

113. (c) NH_4^+ is a conjugate acid;

$NH_4^+ \rightleftharpoons NH_3 + H^+$

114. (a) $AlCl_3 + 3H_2O \rightleftharpoons Al(OH)_3 + \underset{\text{Stron acid}}{3HCl}$

116. (d) *HCl* is a strong acid its conjugate base means Cl^- is a weak base.

117. (d) Hydrazoic acid (HN_3) is a Lewis acid.

118. (c) Smaller the pK_a value than. Stronger the acid.

120. (c) BF_3 is acidic because due to Lewis concept it accept a lone pair of electron.

122. (d) Because it is a weak electrolyte.

126. (c) $AlCl_3$ and SO_2 both are example of Lewis theory.

127. (c) $Na_2CO_3 + 2H_2O \rightleftharpoons 2NaOH + H_2CO_3$. It is a strong base and weak acid so it is a basic.

129. (c) $\underset{\text{Acid}}{H_2O} + \underset{\text{base}}{NH_3} \rightleftharpoons NH_4^+ + OH^-$.

In this reaction H_2O acts as acid because it donate a proton.

130. (b) $\underbrace{H_2SO_4 + H_2O \rightleftharpoons H_3O^+ + HSO_4^-}_{\text{Conjugate acid and base}}$

131. (b) $Al_2(SO_4)_3 \rightleftharpoons 2Al^{3+} + 3SO_4{}^{2-}$

$\underset{\text{Weak base}}{Al(OH)_3} + \underset{\text{Strong acid}}{H_2SO_4}$

132. (c) $NaOH + HCl \xrightleftharpoons[\text{Reaction}]{\text{Neutralization}} \underset{\text{Salt}}{NaCl + H_2O}$

133. (a) Conjugate acid is obtained from the base by gain of H^+.

134. (c) *KCl* is a ionic compound .

135. (a) $H_3BO_3 + H_2O \rightleftharpoons [B(OH)_4]^- + H^+$

H_3BO_3 is a weak monobasic acid if does not act as a H^+ donor but behaves as a Lewis acid.

136. (c) Because it is not accept the proton.

137. (c) NH_4Cl is a salt of weak base (NH_4OH)and strong acid (*HCl*).

138. (b) Because it accept electron pair from the PH_3

139. (b) $NH_4^+ \rightleftharpoons NH_3 + H^+$

141. (b) Hydrolysis of Fe^{3+}

$Fe^{3+} + 3H_2O \rightarrow Fe(OH)_3 + 3H^+$

142. (a) According to the Lewis concept.

143. (b) Conjugate base is obtained from the acid by loss of H^+.

144. (a) BF_3 accept electron pair from NH_3 so it is Lewis acid.

145. (a) CH_4 has almost no acidic nature and thus CH_3^- is strongest base.

146. (a) $CuSO_4$ is a salt of weak base, $(Cu(OH)_2)$and strong acid (H_2SO_4).

147. (b) Weak acid consists of highest pK_a value and strongest acid consist of less pK_a value.

149. (c) Because it gain and also lose the proton

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

150. (a) $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$

151. (c) $HBr + H_2O \rightleftharpoons H_3O^+ + Br^-$

152. (c) Because both are strong acid and strong base.

153. (d) $LiAlH_4$ is a nucleophilic and capable of donating electron pair, thus acts as a Lewis base.

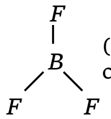
154. (c) The solvent which neither accept proton nor donates.

155. (a) Because of F^- is a highly electronegative. So it is easily lose the electron and reaction occur rapidly.

156. (b) Strong acid can be used titrate both strong and weak base.

159. (b) For a weak acid value of pKa will be vary high but in case of strong acid value of pKa will be vary low.

160. (d) Boron halides behave as Lewis acid because of their electron deficient nature *eg.*, as



(Deficiency of two electron for inert configuration)

161. (b) Gaseous HCl does not give H^+ but liquid HCl gives H^+ in aquous solution there for gaseous HCl is not a Arrhenius acid due to covalent bonding in gaseous condition.

162. (d) $\underset{\text{Acid}}{H_2O} \rightarrow H^+ + OH^-$ [Acid due to donation of proton]

$\underset{\text{Base}}{H_2O} + H^+ \rightarrow H_3O^+$ [Basic due to gaining of proton]

163. (c) F^- strongest conjugate base due to it smallest size in a group and gain proton due to most electronegative capacity.

$H^+ + F^- \rightarrow HF$

164. (d) $\underset{\text{Conjugate acid}}{H_2PO_4^-} \rightarrow H^+ + \underset{\text{Conjugate base}}{HPO_4^{2-}}$

165. (a) $\underset{\text{Conjugate acid}}{HSO_4^-} \rightarrow H^+ + \underset{\text{Conjugate base}}{SO_4^{2-}}$

Common ion effect, Isohydric solutions, Solubility product, Ionic product of water and Salt hydrolysis

1. (b) Solubility of $Al_2(SO_4)_3$

$Al_2(SO_4)_3 \rightleftharpoons 2Al^{+++} + 3SO_4^{--}$

$K_{sp} = [Al^{3+}]^2\,[SO_4^{2-}]^3$

2. (c) Due to common ion effect.

3. (b) $MX_2 \rightleftharpoons \underset{S}{M^{2+}} + \underset{2S}{2X^-}$

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

$\Rightarrow S = 2\sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{4\times10^{-12}}{4}} = 1.0\times10^{-4}\,M$.

5. (b) $MX_2 \rightleftharpoons \underset{(S)}{M^+} + \underset{(2S)^2}{2X^-}$; $K_{sp} = 4S^3$

$S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{1\times10^{-11}}{4}} = 1.35\times10^{-4}$

8. (a) $Mg(OH)_2 \rightleftharpoons \underset{(X)}{Mg^{++}} + \underset{(2X)^2}{2OH^-}$

$K_{sp} = 4X^3$

9. (c) $BaSO_4 \rightleftharpoons \underset{\substack{(S)\\0.01}}{Ba^{++}} + \underset{\substack{(S)\\5}}{SO_4^{--}}$

$K_{sp} = S^2 = S\times S = 0.01\times S$

$S_{(SO_4^{2-})} = \frac{K_{sp}}{S_{(Ba^{++})}} = \frac{1\times10^{-9}}{0.01} = 10^{-7}$ mole/litre

10. (a) $AB_2 \rightleftharpoons \underset{1\times10^{-5}}{A^+} + \underset{2\times10^{-5}}{2B^-}$

$K_{sp} = [1\times10^{-5}][2\times10^{-5}]^2 = 4\times10^{-15}$

11. (b) $CaF_2 \rightleftharpoons \underset{(S)}{Ca^{++}} + \underset{(2S)^2}{2F^-}$; $K_{sp} = 4S^3$

12. (b) Due to common ion effect.

13. (a) $Ag_2CrO_4 \rightleftharpoons [2Ag^+]+[CrO_4^{--}]$

Hence $K_{sp} = [Ag^+]^2\,[CrO_4^{--}]$

14. (d) K_{sp} for $CaF_2 = 4s^3 = 4\times[2\times10^{-4}]^3 = 3.2\times10^{-11}$.

15. (d) The concentration of S^{2-} ions in group II is lowered by maintaining acidic medium in the presence of NH_4Cl. The ionization of H_2S is supressed due to common ion effect. So the ionic product is less than solubility product.

16. (b) NH_4Cl is hydrolysed and give $[H^+]$

$NH_4Cl + H.OH \rightleftharpoons NH_4OH + HCl$

$NH_4{}^+ + H.OH \rightleftharpoons NH_4OH + H^+$

17. (c) $FeCl_3$ is a salt of weak base $(Fe(OH)_3)$ and strong acid (HCl).

18. (c) For the precipitation of an electrolyte, it is necessary that the ionic product must exceed its solubility product.

19. (d) $K_{sp} = [Ag^+]^2[Cro_4^{--}] = [2S]^2[0.01]$

$= 4S^2[0.01] = 4[2\times10^{-8}]^2\times0.01 = 16\times10^{-18}$.

20. (c) Complex salts contain two different metallic elements but give test for only one of them. *e.g.* $K_4Fe(CN)_6$ does not give test for Fe^{3+} ions.

21. (d) $10^{-3}N$ KOH will give $[OH^-] = 10^{-2}\,M$

$pOH = 2$

$\because pH + pOH = 14$, $pH = 14 - 2 = 12$

22. (d) It is $FeSO_4(NH_4)_2SO_4 . 10H_2O$.

23. (d) Salt of a strong base with a weak acid.

24. (b) NH_4CN is a salt of weak acid and weak base and thus for it

25. (a) Because it is a salt of strong base with a weak acid.

26. (d) Because CCl_4 is a organic solvent and $AgNO_3$ is insoluble in organic solvent.

27. (a) $SnS_2 \rightleftharpoons Sn^{4+} + 2S^{2-}$

$\therefore K_{sp} = [Sn^{4+}]\,[S^{2-}]^2$

28. (d) It does not dissociate much or its ionization is very less.

29. (b) $NaHCO_3$ has one replaceable hydrogen.

30. (b) $CaOCl_2$ has two anions Cl^- and OCl^- along with Ca^{2+} ions.

31. (a) $K_{sp} = 4S^3$, $S^3 = \frac{4\times10^{-9}}{4} = 10^{-9}$

$\therefore S = 10^{-3}\,M$.

32. (d) $Be(OH)_2$ has lowest solubility and hence lowest solubility product.

33. (a) Because it is a salt of strong acid and strong base.

34. (c) $NH_4OH \rightleftharpoons \boxed{NH_4^+} + OH^-$

$NH_4Cl \rightleftharpoons NH_4^+ + Cl^-$ Common ion

35. (d) It is a less ionic, so that least soluble in water.

36. (b) pH of 9 means the salt solution should be fairly basic.

37. (b) $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$

On adding CH_3COONa, $[H^+]$ decreases.

38. (c) 0.01 M $CaCl_2$ gives maximum Cl^- ions to keep K_{sp} of $AgCl$ constant, decrease in $[Ag^+]$ will be maximum.

39. (b) Due to the common ion effect.

40. (a) $K_{sp} = 4s^3$

$$S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{1.0\times 10^{-6}}{4}} = 6.3\times 10^{-3}.$$

42. (d) K_w increases with increase in temperature.

43. (b) It contains two cations and one anion.

44. (a) $HgSO_4$ of $K_{sp} = S^2$

$S = \sqrt{K_{sp}}$; $S = \sqrt{6.4\times 10^{-5}}$; $S = 8\times 10^{-3}$ *m/l.*

45. (b) The solubility of $BaSO_4$ in *g/litre* is given 2.33×10^{-3}

$\because$ in *mole/litre.* $n = \dfrac{W}{m.wt} = 1\times 10^{-5} = \dfrac{2.33\times 10^{-3}}{233}$

Because $BaSO_4$ is a compound

$K_{sp} = S^2 = [1\times 10^{-5}]^2 = 1\times 10^{-10}$

46. (d) $\underset{a}{AgCl} \rightleftharpoons \underset{a}{Ag^+} + \underset{a}{Cl^-}$

$\underset{0.02}{NaCl} \rightleftharpoons \underset{0.02}{Na^+} + \underset{0.02}{Cl^-}$

$K_{sp}\ AgCl = 1.20\times 10^{-10}$

$K_{sp}\ AgCl = [Ag^+][Cl^-] = a\times[a+0.2] = a^2 + 0.2a$

a^2 is a very small so it is a neglected.

$K_{sp}\ AgCl = 0.2a$

$1.20\times 10^{-10} = 0.2a$

$a = \dfrac{1.20\times 10^{-10}}{0.20} = 6\times 10^{-10}$ *mole*

47. (b) Solubility is decreased due to common ion effect.

$AgI \rightleftharpoons Ag^+ + \boxed{I^-}$

$NaI \rightleftharpoons Na^+ + I^-$ Common ion

48. (c) K_{sp} of $BaSO_4 = 1.5\times 10^{-9}$; $Ba^{++} = 0.01M$

$$SO_4^{--} = \frac{1.5\times 10^{-9}}{0.01} = 1.5\times 10^{-7}$$

49. (c) $AgCrO_4 \rightleftharpoons \underset{(2S)^2}{2Ag^+} + \underset{S}{CrO_4^-}$

$K_{sp} = 4S^3$ given $2S = 1.5\times 10^{-4}$

$\therefore K_{sp} = (2S)^2\times S$

$$= (1.5\times 10^{-4})^2\times\left(\frac{1.5\times 10^{-4}}{2}\right) = 1.6875\times 10^{-12}$$

50. (c) $PbCl_2 \rightleftharpoons \underset{S}{Pb^{2+}} + \underset{S}{2Cl^-}$

K_{sp} of $PbCl_2 = [Pb^{2+}]\times[Cl^-]^2$; $K_{sp} = S\times(2S)^2$

$K_{sp} = S\times 4S^2 = 4S^3$; $S^3 = \dfrac{K_{sp}}{4}$; $S = \sqrt[3]{\dfrac{K_{sp}}{4}}$

51. (b) $AgCl \rightleftharpoons [Ag^+][Cl^-]$; $K_{sp} = S\times S$; $K_{sp} = S^2$

$S = \sqrt{K_{sp}} = \sqrt{1.44\times 10^{-4}} = 1.20\times 10^{-2}\ M.$

52. (d) By formula $BA_2 \rightarrow B^+ + 2A^-$

$K_{sp} = 4x^3$

53. (a) $\underset{S}{AgCrO_4} \rightarrow \underset{2S}{2Ag^+} + \underset{S}{CrO_4^{--}}$

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

$$S = \left(\frac{K_{sp}}{4}\right)^{\frac{1}{3}} = \left(\frac{32\times 10^{-12}}{4}\right)^{\frac{1}{3}} = 2\times 10^{-4}\ M.$$

54. (d) Common ion effect is noticed only for weak electrolyte dissociation. H_2SO_4 is strong electrolyte.

55. (c) When we added barium ion in chromate ion solution we obtained yellow ppt of $BaCrO_4$.

$$BaCl_2 + K_2CrO_4 \rightarrow \underset{\text{Yellow ppt.}}{BaCrO_4\downarrow} + 2KCl$$

56. (d) AB *is a* binary electrolyte.

$S = \sqrt{K_{sp}} = \sqrt{1.21\times 10^{-6}} = 1.1\times 10^{-3}\,M$

57. (b) Precipitation occurs when ionic product > solubility.

58. (c) For a binary electrolyte, so that

$K_{sp} = S\times S = S^2$

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

61. (b) CH_3COONa is a salt of weak acid and strong base. Hence its aqueous solution is alkaline.

62. (b) $BaSO_4 \rightleftharpoons Ba^{2+} + SO_4^{--}$

Solubility constant $= S \times S$

$1.5 \times 10^{-19} = S^2$; $S = \sqrt{1.5 \times 10^{-19}}$; $S = 3.9 \times 10^{-5}$

65. (d) $Ca(OH)_2 \rightleftharpoons \underset{(S)}{Ca^{++}} + \underset{(2S)^2}{2OH^-}$

$K_{sp} = 4S^3 = 4 \times \sqrt{3} \times \sqrt{3} \times \sqrt{3} = 12\sqrt{3}$

66. (a) Due to common ion effect.

67. (d) $PbCl_2 \rightleftharpoons \underset{(S)}{Pb^{2+}} + \underset{(2S)^2}{2Cl^-}$

$K_{sp} = 4S^3 = 4 \times (2 \times 10^{-2})^3 = 3.2 \times 10^{-5}$

68. (a) $Ag_2S \rightleftharpoons 2Ag^+ + S^{--}$

$K_{sp} = 4S^3$

$$S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{3.2 \times 10^{-11}}{4}} = 2 \times 10^{-6}$$

69. (d) $\underset{S}{CaCO_3} \rightleftharpoons \underset{S}{Ca^{++}} + \underset{S}{CO_3^{--}}$

Solubility product of $CaCO_3$

$K_{sp} = S^2$; $S = \sqrt{K_{sp}}$

It is a binary electrolyte.

$S^2 = K_{sp}$; $(3.05 \times 10^{-4})^2 = K_{sp}$; $K_{sp} = 9.3 \times 10^{-8}$

71. (c) $PbCl_2 \rightarrow \underset{S}{Pb^{++}} + \underset{2S}{2Cl^-}$

$K_{sp} = S \times (2S)^2 = [6.3 \times 10^{-3}] \times [12.6 \times 10^{-3}]^2$.

72. (d) A salt of strong acid and strong base cannot be hydrolysed. In this case the equilibrium cannot shifted towards the backward.

73. (d) If we mixed any substance into the solution. Then the value of pH is increased these substance is a salt of weak acid and strong base.

74. (d) It is a salt of strong base and weak acid.

75. (a) $K_{sp} = 4s^3 = 4 \times [2.5 \times 10^{-2}]^3 = 62.5 \times 10^{-6}$.

76. (c) $NaCl \rightleftharpoons \underset{S}{Na^+} + \underset{S}{Cl^-}$

$K_{sp} = S^2$, $S = \sqrt{K_{sp}} = \sqrt{36} = 6$.

78. (d) $\underset{K_{sp}}{PbI_2} \rightleftharpoons \underset{S}{Pb^{++}} + \underset{2S}{2I^-}$

$K_{sp} = 4S^3 = 4 \times [2 \times 10^{-3}]^3 = 32 \times 10^{-9}$.

79. (b) When ionic product is greater than K_{sp} then precipitation occur

$K_{sp} < 10^{-2} M\ Ca^{2+} + 10^{-3} M\ F^-$

80. (d) In IVth group the S^{2-} concentration increase when added the NH_4OH because

$NH_4OH \rightleftharpoons NH_4^+ + OH^-$

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

$OH^- + H^+ \rightleftharpoons H_2O$. So that S^{2-} is increased.

81. (c) $BaSO_4 \rightleftharpoons Ba^{2+} + SO_4^{2-}$

$K_{sp} = S^2 \Rightarrow S = \sqrt{K_{sp}}$; $K_{sp} = [Ba^{2+}] \times [SO_4^{2-}]$

$4 \times 10^{-10} = [1 \times 10^{-4}] \times [SO_4^{2-}]$

$$[SO_4^{2-}] = \frac{4 \times 10^{-10}}{1 \times 10^{-4}} = 4 \times 10^{-6} .$$

82. (c) $AB_2 \rightleftharpoons \underset{(S)}{A^{2+}} + \underset{(2S)^2}{2B^-}$

$K_{sp} = 4S^3$

$$S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{4 \times 10^{-12}}{4}} = 1 \times 10^{-4}\ gm.mol/litre$$

83. (d) $[B] = \dfrac{K_{sp}AB}{[A]} = \dfrac{1 \times 10^{-8}}{10^{-3}} = 1 \times 10^{-5} M$

Where ionic product > K_{sp}, ppt formed

∴ 8 should be more then $10^{-5} M$.

84. (c) $NaCl_{(s)} \rightleftharpoons Na^+_{(aq)} + Cl^-_{(aq)}$

$HCl \rightleftharpoons H^+ + Cl^-$. The increase in $[Cl^-]$ brings in an increase in $[Na^+]\,[Cl^-]$ which will lead for backward reaction because

$K_{sp}(NaCl) = [Na^+]\,[Cl^-]$

means Ionic product $\geq K_{sp}$

85. (d) $BaSO_4 \rightleftharpoons \underset{(S)}{Ba^{++}} + \underset{(S)}{SO_4^{--}}$

$K_{sp} = S^2$; $S = \sqrt{K_{sp}} = \sqrt{1.3 \times 10^{-9}}$

$= 3.6 \times 10^{-5}\ mol/litre$

87. (b) Alkaline,

$CH_3COONa + H_2O \rightleftharpoons \underset{\text{Weak acid}}{CH_3COOH} + \underset{\text{Strong base}}{NaOH}$

88. (c) Because it is a strong base.

89. (b) For pure water $[H^+] = [OH^-]$, $\therefore K_w = 10^{-12}$ s

90. (a) $MX_2 \rightleftharpoons \underset{(S)}{M^{2+}} + \underset{(2S)^2}{2X^-}$; $4S^3 = 4 \times (0.5 \times 10^{-4})^3$

$= 5 \times 10^{-13}$

91. (a) Solubility coefficient $= [Pb^{2+}]\,[Cl^-]^2$

92. (a) Solubility of $Al(OH)_3$ is lesser than $Zn(OH)_2$.

93. (c) $NaCl_{(s)} \rightleftharpoons Na^{+}{}_{(aq)} + Cl^{-}{}_{(aq)}$

$HCl \rightleftharpoons H^+ + Cl^-$

The increase in $[Cl^-]$ brings in an increase in $[Na^+]\ [Cl^-]$ which will lead for backward reaction because $K_{sp}\ NaCl = [Na^+]\,[Cl^-]$.

94. (c) Common ion effect.

95. (a) $\underset{S}{CaF_2} \rightleftharpoons \underset{S}{Ca^{++}} + \underset{(2S)^2}{2F^-}$

$K_{sp}=4S^3$

$S=\sqrt[3]{\frac{K_{sp}}{4}}=\sqrt[3]{\frac{3.2\times10^{-11}}{4}}=2\times10^{-4}$ *m/l.*

96. (d) In aqueous solution following euilibrium is exist. $H_2S\rightleftharpoons H^++HS^-$

While adding the dilute *HCl* solution

$(HCl\rightleftharpoons H^++Cl^-)$ equilibrium is shift to the left side in $H_2S\rightleftharpoons H^++HS^-$

97. (d) $\underset{K_{sp}}{M_2X_3}\rightleftharpoons\underset{(2y)^2}{2M^{+++}}+\underset{(3y)^3}{3X^{--}}$

Solubility product $K_{sp}=108y^5\ mol\ \frac{d}{m^3}$

98. (b) Solubility is directly proportional to the K_{sp}.

99. (b) $\underset{S}{PbCl_2}\rightleftharpoons\underset{S}{Pb^{++}}+\underset{(2S)^2}{2Cl^-}$

$K_{sp}=S\times(2S)^2=4S^3$

$S=\sqrt[3]{\frac{K_{sp}}{4}}=\sqrt[3]{\frac{1.5\times10^{-4}}{4}}=3.34\times10^{-2}$.

101. (d) $AlCl_3$ on hydrolysis gives weak base and strong acid among all.

102. (c) Fe^{3+} ions are hydrolysed to develop acidic nature.

103. (c) $K_h=\frac{K_w}{K_a\times K_b}$

104. (a) *KCN* is salt of strong base and weak acid.

105. (c) Sulphides of Group-II radicals have low solubility product.

106. (c) Because NH_3 acts as Lewis acid and they give electron pair to H_3O^+ ion. H_3O^+ is a Lewis base. Which accept the electron pair from NH_3.

107. (c) Due to common ion effect.

108. (a) For $Ag_2SO_4\rightarrow\underset{2x}{2Ag^+}+\underset{x}{SO_4^{--}}$

$K_{sp}=(2x)^2.x$; $K_{sp}=4x^3$; $K_{sp}=4\times(2.5\times10^{-2})^3$

$K_{sp}=62.5\times10^{-6}$

109. (b) For $AgCl\rightarrow\underset{x}{Ag^+}+\underset{x}{Cl^-}$

$K_{sp}=x^2$; $x=\sqrt{K_{sp}}$,

$\sqrt{1\times10^{-6}}=1\times10^{-3}\ mole/litre$.

110. (b) $AgCl\rightarrow\underset{x}{Ag^=}+\underset{x}{Cl^-}$

After *NaCl* is added x $x+1\times10^{-4}$

That is why Ag^+ will be less.

111. (b) Because of ionic product of $AgI>>$ solubility product of its.

112. (a) $AX_2\rightarrow\underset{x}{A}+\underset{2x}{2X}$

$K_{sp}=4x^3$; $x=\sqrt[3]{\frac{3.2\times10^{-11}}{4}}$;

$x=2\times10^{-4}$ *mole/litre.*

113. (c) $C_6H_5COONH_4\rightarrow C_6H_5COO^-+NH_4^+$

$0.5-\frac{.25}{100}\quad\frac{.25}{100}\quad\frac{.25}{100}$

According to Ostwald dilution law.

$K=\frac{\alpha^2C}{1-\alpha}$ $\left(\therefore\ \alpha=\frac{.25}{100}\right)$

$K=\alpha^2C$ $(\because\ 1-\alpha=\text{Very small})$

$K=\frac{.25}{100}\times\frac{.25}{100}\times.5$; $K=3.125\times10^{-6}$

114. (a) $Sb_2S_3\rightarrow\underset{2x}{2Sb^{+2}}+\underset{3x}{3S^{--}}$; $K_{sp}=(2x)^2.(3x)^3$

$K_{sp}=108x^5$; $K_{sp}=108\times(1\times10^{-5})^5=108\times10^{-25}$.

115. (b) When increasing the temperature the value of ionic product also increases.

116. (a) Hydrolysis constant $h=\frac{K_w}{K_a}$

Hydrogen ion concentration - *p*H scale and Buffer solution

1. (a) *pH* of blood does not change because it is a buffer solution.

2. (c) 0.001 *M* of *NaOH* means $[OH^-]=.001$

$=10^{-3}M\Rightarrow pOH=3$

$pH+pOH=14\Rightarrow pH=14-3=11$

3. (d) $[H_3O^+]$ means $[H^+]=6.2\times10^{-9}\ mol/l$

$pH=-\log(6.2\times10^{-9})=8.21$

4. (b) $CH_3NH_2+HCl\longrightarrow CH_3NH_3^+Cl^-$

0.1 0.08 0

0.02 0 0.08

(Basic buffer solution)

$pOH=pK_b+\log\frac{0.08}{0.02}$

$=pK_b+0.602$

$=3.30+0.602=3.902$

$\therefore pH=10.09$

$[H^+]=7.99\times10^{-11}\approx8\times10^{-11}\ M$

5. (b) $pH+pOH=pK_w$

6. (d) $pH=-\log[H^+]$

$5.4=-\log[H^+]$; $[H^+]=3.98\times10^{-6}$.

7. (a) $KCN+H_2O\rightleftharpoons KOH+HCN$. *KOH* is a strong base and *HCN* is a weak acid.

8. (c) $[H^+] = 10^{-3} M,\ pH = -\log[10^{-3}],\ pH = 3$

9. (b) $[H^+] = [OH^-]$

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

$\therefore [H^+] = 10^{-7}\ ,\ pH = -\log[H^+] = 7$.

10. (d) $pH = 5$ means $[H^+] = 10^{-5}$

$pOH = 14 - pH = 14 - 5 = 9$

$[OH^-] = 10^{-pOH} = 10^{-9}$

11. (a) $pH = -\log\ [H^+]\,;[H^+] = 0.01\ N$

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

12. (d)

$$\begin{array}{lcccc} & BOH & \rightleftharpoons & B^+ & + & OH^- \\ \text{Initial} & C & & 0 & & 0 \\ \text{At eq.} & C - C\alpha & & C\alpha & & C\alpha \end{array}$$

$K_b = \dfrac{C^2\alpha^2}{C(1-\alpha)} = C\alpha^2$ assuming $\alpha << 1\,;1-\alpha \simeq 1$

$10^{-12} = 10^{-2} \times \alpha^2\ ;\ \alpha^2 = 10^{-10}\ ;\ \alpha = 10^{-5}$

$[OH^-] = C\alpha = .01 \times 10^{-5} = 10^{-7}$

13. (c) $pH = 4$ means; $[H^+] = 10^{-4}\ mol$

14. (a) Buffer solution is a mixture of weak acid and its conjugate base.

15. (b) Adding Na_2CO_3 to water makes the solution basic and hence a *pH* increases from 7.

16. (d) $NaClO_4$ is a salt of strong acid $HClO_4$. So it is a strong acid salt.

17. (b) *NaOH* is a base, so that its $pH > 7$

18. (c) It is a strong base.

1 *M NaOH* has maximum $[OH^-]$ and minimum $[H^+]$ and maximum *pH*.

19. (c) When *pH* = 7 means neutral, *pH* < 7 means acidic, *pH* > 7 means basic.

20. (d) As the solution is acidic, $pH < 7$. This is because $[H^+]$ from $H_2O[10^{-7}M]$ cannot be neglected in comparison to $10^{-10}\ M$

22. (c) It is a strong acid and they lose proton in a solution.

23. (a) $[OH^-] = 10^{-2}\,M\ ;\ pOH = 2$

$pH + pOH = 14\ ;\ pH = 14 - pOH$

$pH = 14 - 2 = 12$

24. (d) Order of acidic strength is $H_2Te > H_2Se > H_2S > H_2O$

Na_2O is a salt of NaOH $+H_2O$ and H_2O is least acidic among given acids hence pH in this case will be $\max^m$.

25. (b) *pH* of the solution *A* = 3

$[H^+]_A = 10^{-3}\ M$.

pH of the solution *B* = 2

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

$[H^+] = 10^{-3} + 10^{-2} = 10^{-3} + 10 \times 10^{-3} = 11 \times 10^{-3}$.

$pH = -\log(11 \times 10^{-3}) = 3 - \log 11$

$= 3 - 1.04 = 1.95$

26. (a) $CN^- + H_2O \to HCN + OH^-$

Because OH^- concentration is increased.

27. (a) On dilution the *pH* of acid *A* increases while *pH* of base *B* decreases.

28. (d) CH_3COONa is a salt of weak acid, (CH_3COOH) and strong base $(NaOH)$.

29. (b) The equilibrium will shift in the backward direction.

30. (b) $K_a = 10^{-5}\ ;\ pH = 6$

$pH = -\log K_a + \log\dfrac{[\text{Salt}]}{[\text{Acid}]}\ ;\quad 6 = -\log 10^{-5} + \log\dfrac{[\text{Salt}]}{[\text{Acid}]}$

$6 = 5\log 10 + \log\dfrac{[\text{Salt}]}{[\text{Acid}]}\ ;\ 6 = 5 + \log\dfrac{[\text{Salt}]}{[\text{Acid}]}$

$\log\dfrac{[\text{Salt}]}{[\text{Acid}]} = 6 - 5 = 1\ ;\ \dfrac{[\text{Salt}]}{[\text{Acid}]} = \dfrac{10}{1}$

31. (d) All are true – (a) is true for acid buffer, (b) for basic buffer, (c) is called buffer solution.

32. (b) $\dfrac{N}{100} = 0.01\ N\ HCl\ ;[H^+] = 10^{-2}\,M\ ;\ pH = 2$

$[OH] = 10^{-2}\,M$ for *NaOH*

$pH + pOH = 14\ ;\ pH = 14 - 2\ ;\ pH = 12$

34. (a) It is a buffer solution of strong acid and its weak conjugate base.

35. (c) $HA \rightleftharpoons H^+ + A^-$

$[H^+] = 0.1\ M\ ;\ [H^+]^2 = K_a \times C$

$[H^+] = \sqrt{K_a \times C} = \sqrt{1 \times 10^{-5} \times 0.1} = \sqrt{10^{-6}}$

$[H^+] = 10^{-3}M;\ pH = 3$

36. (c) As the solution is acidic, $pH < 7$. This is because $[H^+]$ from H_2O cannot be neglected in comparison to 10^{-8} .

37. (c) Human body contain buffer solution. Its *pH*= 6.8

38. (a) It is a neutral solution and its *pH* = 7

39. (a) $pH = 5,$ means $[H^+] = 10^{-5}M.$

After dilution $[H^+] = 10^{-5}/100 = 10^{-7}M$

$[H^+]$ from H_2O cannot be neglected.

Total $[H^+] = 10^{-7} + 10^{-7} = 2 \times 10^{-7}$

$pH = 7 - 0.3010 = 6.6990 = 7$ (neutral).

40. (d) $[H^+] = \alpha.C = \dfrac{2}{100} \times .02\ ;\ [H^+] = 4 \times 10^{-4}\ M$

$pH = -\log\ [H^+] = 4 - \log\ 4\ ;\ pH = 3.3979$

41. (a) $pH = pK_a + \log\left[\dfrac{\text{salt}}{\text{acid}}\right]$

$= 9.30 + \log\left[\frac{0.2}{0.1}\right] = 9.30 + 0.3010 = 9.6.$

42. (a) $pH = pK_a + \log\frac{[\text{Salt}]}{[\text{Acid}]}$

$pH = -\log(1.8\times10^{-5}) + \log\frac{[10]}{[100]}$

$= -\log 1.8 + 5 + \log 10^{-1}$

$= -0.2553 + 5 - 1 = 3.7447 \text{ or } = 4$

43. (b) 20 *ml*. of 0.1 $NHCl = \frac{0.1}{1000}\times 20g$ eq. $= 2\times10^{-3} g$ *eq*.

20*ml*. of 0.001 $KOH = \frac{0.001}{1000}\times 20\, gm$ eq.

$= 2 \times 10^{-5} g\ eq.$

$\therefore$ HCl left unneutralised $= 2(10^{-3} - 10^{-5})$

$= 2\times10^{-3}(1-0.01) = 2\times0.99\times10^{-3} = 1.98\times10^{-3} g\, eq.$

Volume of solution = 40 *ml*.

$\therefore [HCl] = \frac{1.98\times10^{-3}}{40}\times1000\, M = 4.95\times10^{-2}$

$\therefore pH = 2 - \log 4.95 = 2 - 0.7 = 1.3.$

45. (b) $10^{-7} M$ $NaOH$ means $[OH^-] = 10^{-7}$; $pOH = 7$

$pH = 14 - 7 = 7$

46. (a) $[H^+] = c\times\alpha = 0.1\times\frac{30}{100} = 0.03\ M$

48. (a) The pH of buffer solution never changed.

49. (c) $[H^+] = \frac{10^{-14}}{10^{-1}} = 10^{-13}\ mol/litre\quad pH = 13.$

50. (a) $pH = -\log[H^+]$; $7.4 = -\log[H^+]$; $[H^+] = 4\times10^{-8} M$

51. (a) The pH of 0.1M $HCl = 1$ Ionization of H_2SO_4 takes place in two steps.

$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$; $HSO_4^- \rightleftharpoons H^+ + SO_4^{--}$

52. (c) 1N $NaOH$ solution have highest pH

$[OH^-] = 1$; $pOH = 0$; $pH + pOH = 14$

$pH = 14 - 0 = 14$

53. (c) $H_2O \rightleftharpoons [H^+][OH^-]$

$HCl \rightleftharpoons [H^+][Cl^-]$

Total $[H^+] = [H^+]_{H_2O} + [H^+]_{HCl} = 10^{-7} + 10^{-8}$

$= 10^{-7}[1 + 10^{-1}]$

$[H^+] = 10^{-7}\times\frac{11}{10}$

$pH = -\log[H^+] = -\log\left(10^{-7} + \frac{11}{10}\right)$; $pH = 6.958$

54. (c) pK_a = − log K_u, pK_b = − log K_b

pH = $-\frac{1}{2}[\log K_a + \log K_w - \log K_b]$

$= -\frac{1}{2}[-5 + \log(1\times10^{-14}) - (-5)]$

$= -\frac{1}{2}[-5 - 14 + 5] = -\frac{1}{2}(-14) = 7$

55. (d) BaO, CaO and Na_2O are shows more than 7 pH because of their basic nature.

56. (a) $MgCl_2 + 2H_2O \rightleftharpoons Mg(OH)_2 + 2HCl$

57. (c) H_2SO_4 ionized in two step.

58. (b) $pH = pK_a + \log\frac{[\text{Salt}]}{[\text{Acid}]}$

$5.8 = 4.8 + \log\frac{[\text{Salt}]}{[\text{Acid}]}$ or $\log\frac{[\text{Salt}]}{[\text{Acid}]} = 1.0$

$\frac{[\text{Salt}]}{[\text{Acid}]} = \text{antilog } 1.0 = 10$

$\therefore \quad \frac{[\text{Acid}]}{[\text{Salt}]} = \frac{1}{10} = 0.1$

59. (b) It contains replacable H atom.

60. (c) (i) 20 *ml* of 0.5 N HCl

$0.5N \Rightarrow 1000\ ml\ 0.5\ mole\ HCl$ is present in $20\ ml$

$= \frac{20\times0.5}{1000} = 1.0\times10^{-2}$

(ii) 35 *ml* of 0.1 N $NaOH$

$0.1N \Rightarrow 1000\, ml$ of 0.1 mole $NaOH$ is 35 *ml*

$= \frac{35\times0.1}{1000} = 0.35\times10^{-2}$

Total = 20 + 35 = 55 *ml*.

$\Rightarrow (1.0-0.35)10^{-2} = 0.65\times10^{-2}$mole HCl

$HCl = H^+ + Cl^-$

$\Rightarrow [HCl] = [H^+] + [Cl^-]$

55 *ml* contains 0.65×10^{-2} mole of H^+ ions

$1000\ ml - \frac{0.65\times10^{-2}\times10^3}{55} = \frac{6.5}{55}$

$pH = -\log[H^+] = -\log(6.5/55)$

$= \log 55 - \log 6.5 = 0.92$

Due to acidic nature of solutions the colour of phenolphthalein becomes pink.

61. (b) $[H^+] = 2\times10^{-2} M$

$\because pH = -\log[2\times10^{-2}]$;

$pH = 1.7$ *i.e.* in between 1 and 2.

63. (b) $pH = 4, (H^+) = 10^{-pH} = 10^{-4} M$

65. (c) $NaOH \rightleftharpoons Na^+ + OH^-$

$[OH^-] = 10^{-5} M$; $[H^+][OH^-] = 10^{-14}$

$[H^+] = \frac{10^{-14}}{10^{-5}}$; $[H^+] = 10^{-9} M$; $pH = 9$.

67. (b) $pH = pK_a + \log\frac{[\text{Salt}]}{[\text{Acid}]}$; $pH = 4.75 + \log\frac{0.1}{0.1}$

$pH = 4.75 + \log 1$; $pH = 4.75$

68. (d) A weak acid and its salt with a strong base maintain pH 4 - 5

69. (a) $NaOH \rightleftharpoons Na^+ + OH^- = [OH^-] = 10^{-8}\ M.$

70. (c) $[OH^-] = 0.0001\ N$, $pOH = 4$, $pH + pOH = 14$

$pH = 14 - pOH \ = 14 - 4 = 10$

71. (d) 0.001 M KOH solution

$[OH^-] = 0.001\ M = 1\times10^{-3}\ M$

$[H^+]\times[OH^-] = 1\times10^{-14}$

$[H^+] = \frac{1\times10^{-14}}{[OH^-]}$

$[H^+] = \frac{1\times10^{-14}}{1\times10^{-3}} = 1\times10^{-14}\times10^{+3}$

$[H^+] = 10^{-11}\ M$

$pH = 11$

72. (a) An acid buffer solution consists of solution of weak acid with strong base of its salt.

73. (b) An acid buffer solution consists of a weak acid and its salt with strong base. *i.e.* $CH_3COOH + CH_3COONa$

74. (a) $pOH = pK_b + \log\frac{[\text{salt}]}{[\text{base}]}$

$= 5 + \log\frac{0.02}{0.2} = 5 + \log\frac{1}{10} = 5 + (-1) = 4$

$pH = 14 - pOH = 14 - 4 = 10$

75. (b) [Salt]= 0.1 M, [Acid]= 0.1 M

$K_a = 1.8\times10^{-5}$; $pH = -\log K_a + \log\frac{[\text{Salt}]}{[\text{Acid}]}$

$= -\log 1.8\times10^{-5} + \log\frac{0.1}{0.1} = -\log\ 1.8\times10^{-5}$

$pH = 4.7$.

76. (a) NH_4Cl and NH_4OH is a buffer solution (weak base and salt of strong acid).

77. (a) $pH + pOH = 14$; $pH = 14 - pOH$

$\because [OH^-] = 10^{-7}$

$pOH = 7$

$\therefore pH = 14 - 7 = 7$.

78. (c) 0.01 M $Ba(OH)_2 = 0.02\,N\ Ba(OH)_2$

$N_1V_1 = N_2V_2$

$[0.02\,N]\times[50\ ml] = N_2 \times 100\ ml$

$N_2 = \frac{0.02\times50}{100} = 10^{-2}\,N$; $[OH^-] = 10^{-2}\,N$

$pOH = 2$ or $pH = 12$

79. (b) $pH = -\log\,[H^+]$.

80. (a) Na_2CO_3 is a mixture of weak acid and strong base, so it is a base.

81. (b) $10^{-7}\,N\ HCl$ means $(H^+) = 10^{-7}\ M$

$pH = -\log(H^+)$, $pH = 7$

82. (c) $pH = 2$; $pH = -\log\,[H^+]$; $2 = -\log\ [H^+]$

$[H^+] = 10^{-2} = 0.01\ N$

83. (b) pH does not change on addition of some concentration of HCl.

84. (b) Solution of CH_3COONa on addition to acid shows a decrease in dissociation of acid due to common ion effect. To decrease in $[H^+]$ or increase pH.

85. (c) $pH + pOH = 14$; $pH = 14 - pOH$; $pH = 14 - 6 = 8$.

86. (b) $[H^+]_{\text{I}} = 10^{-5}$ $[H^+]_{\text{II}} = 10^{-2}$

Thus increase in $[H^+] = \frac{10^{-2}}{10^{-5}} = 1000$ times

87. (a) The HCl is a strong acid and they lose easily H^+ in solution.

88. (a) $X^- + H_2O \rightleftharpoons OH^- + HX$

$K_b = \frac{[OH^-]\,[HX]}{[X^-]}$

$HX \rightleftharpoons H^+ + X^-$

$K_a = \frac{[H^+]\,[X^-]}{[HX]}$

$\therefore K_a \times K_b = [H^+]\,[OH^-] = K_w = 10^{-14}$

Hence $K_a = 10^{-4}$

Now as $[X^-] = [HX]$, $pH = pK_a = 4$.

90. (d) Buffer solution is formed. So the pH will not change.

91. (b) Na_2CO_3 when react with water form strong base and weak acid. So its aqueous solution is basic.

92. (c) $K_w = [H_3O^+]\,[OH^-]$

Concentration of H_3O^+ in distilled water $= 1\times10^{-6}$ *mol/l*.

Now $[H_3O^+] = [OH^-]$

$K_w = [1\times10^{-6}]\times[1\times10^{-6}] = 1\times10^{-12}$.

93. (a) $[OH^-] = 10^{-1}M$; $pOH = 1$

$pH + pOH = 14$; $pH = 14 - 1 = 13$.

94. (a) Maximum pH $HClO$ is a weak acid all of these. So that the salt of weak acid is also weak.

95. (c) As the solution is acidic $pH < 7$. This is because $[H^+]$ from $H_2O\,[10^{-7}]$ cannot be neglected in comparison to 10^{-12} M.

96. (b) [Normal salt + acidic salt] is a buffer solution.

97. (b) 100 *ml* of $\frac{M}{10} NaOH$ = 50*ml* of $\frac{M}{5} NaOH$. They exactly neutralise 50 *ml* $\frac{N}{5} HCl$. Hence *pH* of resulting solution = 7.

98. (b) $M_1 = 6.0\, M$ of HCl ; $V_1 = ?$

M_2 =0.30 *M* is H^+ concentration in solution.

$V_2 =$ 150 *ml* of solution.

$M_1V_1 = M_2V_2$; $6.0 \times V_1 = .30 \times 150$

$V_1 = \frac{.30 \times 150}{6} = 7.5$ *ml.*

99. (b) $pH = 3$, $[H^+] = 10^{-3} M$

$\because [H^+] = \sqrt{K \times c}$

$[10^{-3}]^2 = K \times c$; $\frac{[10^{-6}]}{0.1} = K = 10^{-5}$

100. (b) When ratio of concentration of acid to salt is increased *pH* decrease.

101. (c) For NH_4OH .

$[OH^-] = C.\alpha$; $C = \frac{1}{10} M$, $\alpha = 0.2$

$[OH^-] = \frac{1}{10} \times 0.2 = 2 \times 10^{-2}\, M$

$pOH = -\log\, [OH^-] = \log\, [2 \times 10^{-2}]$; $pOH = 1.7$

$pH = 14 - pOH \quad = 14 - 1.7 = 12.30$.

102. (c) $pH = pK_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$. For small concentration of buffering agent and for maximum buffer capacity $\frac{[\text{Salt}]}{[\text{Acid}]} \approx 1$.

103. (a) $[H^+]$ = increased ten fold means *pH* of solution decreased by one.

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

104. (a) Because the *pH* of buffer are not changed.

105. (c) $pH = pK_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$; $5.5 = 4.5 + \log \frac{[\text{Salt}]}{[0.1]}$

$\log \frac{[\text{Salt}]}{0.1} = 5.5 - 4.5 = 1$

$\frac{[\text{Salt}]}{0.1}$ = antilog 1 = 10 ; $[\text{Salt}] = 1$

106. (a) Moles of $H_2SO_4 = \frac{0.49}{98} = 5 \times 10^{-3}$ moles of H_2SO_4 present per litre of solution (molarity) $= \frac{.005}{1} = .005$ *M.*

$H_2SO_4 + 2H_2O \rightleftharpoons 2H_3O^+ + SO_4^{--}$

one mole of H_2SO_4 give 2 moles of H_3O^+ ions.

$H_3O^+ = 2 \times (H_2SO_4) = 2 \times 0.005 = 0.01\, M$

$[H^+] = 10^{-2}\, M$; $pH = 2$

107. (c) CH_3COONH_4 is a simple buffer and called salt of weak acid.

108. (c) N.eq. for $HCl = \frac{0.4}{1000} \times 50 = 0.02$

N.eq. for $NaOH = \frac{0.2}{1000} \times 50 = 0.1$

Now $[OH^-]$ left $= 0.1 - 0.02$

$[OH^-] = .08 = 8 \times 10^{-2}\, M$

$pOH = -\log 8 \times 10^{-2}\, M$; $pOH = 1.0$

109. (d) Buffer is mixture of weak base and its acid salt.

110. (b) $[NaOH] = 0.4/40$ $mole/l. = 0.1\, M$

$[OH^-] = 10^{-1} M$, $[H^+] = 10^{-13} M$, $pH = 13$

111. (d) $pH + pOH = 14$, *pH* = 4 , $H^+ = 10^{-4}$ *mole/litre.*

112. (d) Buffer solution have constant *pH*. When we add the water into this buffer solution. So no effect on it.

113. (b) $Ba(OH)_2 \rightleftharpoons Ba^{2+} + 2OH^-$

One molecule on dissociation furnishes $2OH^-$ ions.

So, $[OH^-] = 2 \times 10^{-4}\, N$

$N = M \times 2$; $M = \frac{N}{2} = \frac{2 \times 10^{-4}}{2} = 10^{-4}$

$pOH = -\log[OH^-] = -\log(1 \times 10^{-4}) = -4$

$pH + pOH = 14$; $pH = 14 - 4 = 10$.

114. (a) M.eq. of 0.10 $M\ HCl = \frac{0.10}{1000} \times 40 = 0.004\, M$

M.eq. of 0.45 $M\ NaOH = \frac{0.45 \times 10}{1000} = 0.0045\, M$

Now left $[OH^-] = 0.0045 - 0.004 = 5 \times 10^{-4}\, M$

Total volume = 50 *ml.*

$[OH^-] = \frac{5 \times 10^{-4}}{50} \times 1000$; $[OH^-] = 1 \times 10^{-2}$

pOH = 2 ; *pH* = 14 – *pOH* = 12.

115. (c) 0.001 $M\, HCl = 10^{-3}\, M [H^+]$, $pH = 3$.

116. (d) $[NaOH] = \frac{0.4}{40} = 0.01M$; $[OH^-] = 10^{-2} M$

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

117. (b) Those substance which give a proton is called Bronsted acid while CH_3COO^- doesn't have proton so it is not a Bronsted acid.

120. (c) $pH=-\log\,[H^+]$

121. (d) $pH+pOH=14, pH=4.0$

$pOH=14-pH$; $pOH=14-4.0=10.0$

122. (b) $pH=0$ means $[H^+]=10^o=1M$. Hence solution is strongly acidic.

123. (c) As the solution is acidic, pH < 7. This is because $[H^+]$ from $H_2O\,(10^{-7}M)$ cannot be neglected in comparison to 10^{-10} *M HCl*.

124. (d) $H_3O^+ \rightleftharpoons OH^- + H_2$

$pOH+pH=14$; $7+7=14$; $[H^+]+[OH^-]=10^{-14}$

$10^{-7}+10^{-7}=10^{-14}$; $[OH^-]=10^{-7}$ *gm ion/l.*

126. (b) When $pH=2$, $[H^+]=10^{-2}M$

127. (a) $[OH^-]$ ion conc. $=0.05\,\frac{mol}{l}\;=5\times10^{-2}\,\frac{mol}{l}$

$pOH=-\log\,[OH^-]\;=-\log\,[5\times10^{-2}]$

$pOH=1.30$ **;** $pH+pOH=14$

$pH=14-pOH\;=14-1.30=12.7$

128. (c) When $pH=3$, then $[H^+]=10^{-3}M$

after that we increased the *pH* from 3 to 6 then $[H^+]=10^{-6}M$ means reduced 1000 times.

129. (b) CO_2 is acidic oxide which on dissolution in water develops acidic nature.

130. (d) If *pH* of any solution is 2.

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

If *pH* of any solution is just double then $pH=4$ and $[H^+]$ will be 10^{-4}.

131. (c) A strong acid is not used to make a buffer.

132. (d) $pH=1$ means $[H^+]=10^{-1}M$

Hence $[H_2SO_4]=\frac{10^{-1}}{2}=\frac{1}{20}=0.05\;M$

133. (c) The *pH* of blood is 7.4 due to presence of bicarbonates ions

134. (c) As the solution is acidic, *pH* < 7. This is because $[H^+]$ from $H_2O\,[10^{-7}]$ cannot be neglected in comparison to 10^{-8}

135. (a) *pH* will decrease because $[OH^-]$ increased due to this *pOH* is decreased.

136. (c) $[H^+]=6\times10^{-4}M$

$pH=-\log\,[H^+]\;=-\log\,[6\times10^{-4}]=3.22$.

137. (c) $0.01\,M\,HCl=10^{-2}\,M\,[H^+],\;pH=2$.

138. (c) Because buffer solution have a constant *pH*.

139. (c) $10^{-6}\,M\,HCl=10^{-8}\,M[H^+]$. Also from H_2O

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

Total $[H^+]=10^{-7}+10^{-8}=10^{-7}[1+0.1]=10^{-7}[1.1]$

Hence $pH=7-0.0414=6.96$.

140. (b) $10^{-10}\,M\;HCl=10^{-10}\;M[H^+]$. But $pH\neq10$ because solution is acidic. This is because H^+ from $H_2O(10^{-7}M)$ cannot be neglected.

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

$=10^{-7}+(1+10^{-3})=10^{-7}(1.001)$

That is why $pH=7$ (slightly less than 7)

141. (b) $[H^+]=1.00\times10^{-6}$ *mole/litre*

$pH=-\log\,[1.00\times10^{-6}]$; pH = 6.

142. (a) $[H^+]$ is in moles per litre.

143. (d) As the solution is acidic, $pH<7$. This is because $[H^+]$ from H_2O $(10^{-7}M)$ cannot be neglected in comparison to 10^{-8}.

145. (b) *pH* of 0.001 *M HCl* $=10^{-3}M[H^+]$, *pH* = 3.

146. (d) Because it can furnish H^+ ions in solutions.

147. (c) Because it is a strong acid.

$H^+=10^{-1}$

$pH=-\log\,[H^+]\;=-\log\,[10^{-1}]$; $pH=1$.

148. (b) Buffer solution is a combination of weak acid and conjugate base. *NaCl* is a salt and *NaOH* is the base.

149. (a) $[H^+]=\sqrt{Kc}=\sqrt{10^{-5}\times0.1}=10^{-3},\;pH=3$.

150. (d) In $\frac{N}{10}NaOH$ have $[OH^-]=10^{-1}M$ means

$pOH=1$ and then $pH+pOH=14$

$pH=14-pOH=13$.

151. (b) Borate ions are hydrolyzed to develop alkaline nature in solution.

152. (d) Less the *pH*, more acidic is the solution.

154. (b) The equal conc. of salt and acid.

155. (c) $pH=-\log\,K_a+\log\frac{[KCN]}{[HCN]}$

$pH=-\log[5\times10^{-10}]+\log\left(\frac{0.15}{1.5}\right)=8.302$

157. (c) $pH=pK_a+\log\frac{[\text{Salt}]}{[\text{Acid}]}$ equimolar means

$\frac{[\text{Salt}]}{[\text{Acid}]}=1$; $\quad pH=4.74+0=4.74$

158. (a) Because of *NaCl* is a salt of strong acid and strong base. So that it is neutral.

159. (c) When strong acid and strong base are react neutral salt are formed. So that *NaCl* is a neutral salt.

162. (d) $pH = -\log K_b + \log \frac{[\text{Salt}]}{[\text{Acid}]}$

$pH = -\log[1.8 \times 10^{-5}] + \log \frac{[\text{Salt}]}{1.0}$

$9 = 4.7 + \log \frac{[\text{Salt}]}{1.0}$; $\log \frac{[\text{Salt}]}{1.0} = 4.7 - 9 = -4.3$

$\frac{[\text{Salt}]}{1.0} = \text{Antilog} \frac{1}{4.3}; [S\text{alt}] = 1.8$

163. (b) $pH = -\log K_b + \log \frac{[\text{salt}]}{[\text{acid}]}$

$5 = -\log 10^{-4} + \log \frac{[\text{salt}]}{[\text{acid}]}$

$\log \frac{[\text{salt}]}{[\text{acid}]} = 1$

$\frac{[\text{salt}]}{[\text{acid}]} = \text{antilog } 1 = 10 : 1$

164. (a) 1 *M KOH* show highest *pH* value because it is a strong base.

165. (d) NH_4OH is a weak acid and NH_4Cl is a strong base salt.

166. (a) $pH = 13.6$

$pOH = 14 - 13.6 = 0.4$

$[OH^-] = \text{Antilog}(-0.4) = 0.3979$. So the value of $[OH^-]$ between $0.1\ M$ and $1\ M$

167. (d) Aspirin is a weak acid. Due to common ion effect it is unionised in acid medium but completely ionised in alkaline medium.

168. (b) $[H^+][OH^-] = 10^{-14}$; $(10^{-7})(10^{-7}) = 10^{-14}$

169. (c) $HCl = 10^o M$ has $pH = 0$. The value of *pH* decreases as concentration further increases.

170. (a) Because pure water has a 7 *pH*.

171. (c) When concentration of $[H^+]$ increased then the value of *pH* is decreases.

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

172. (c) The concentration of $[H^+] = 10^{-2}$ mole/litre

$pH = -\log[H^+] = -\log[10^{-2}]$; $pH = 2$

173. (d) Due to common ion effect.

174. (b) In water solution.

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

concentration of OH^- is increased so that solution become more basic and the *pH* is increased.

175. (a) Na_2CO_3 is basic in nature. So its *pH* is greater than 7.

176. (c) It is not a mixture of weak acid or base and their strong salt.

177. (a) $[H^+] = \text{Antilog}(-4.58)$;

$[H^+] = 2.63 \times 10^{-5}$ *moles / litre*

178. (c) $10^{-2}\ M\ NaOH$ will give $[OH^-] = 10^{-2}$

$\therefore pOH = 2$, Also $pH + pOH = 14$

$\therefore pH = 12$.

179. (a) $pH = pK_a + \log \frac{[\text{Salt}]}{[\text{Acid}]} = -\log 2 \times 10^{-5} + \log \frac{10 \times 1}{50 \times 2} = 4$.

180. (b) 0.001 *M NaOH* means $[OH^-] = 10^{-3}$; $pOH = 3$

$pH + pOH = 14; pH = 14 - 3$

$pH = 11$; $[H^+] = 10^{-11}$ *mole-litre*$^{-1}$

181. (c) $[H^+] = C.\alpha$

$[H^+] = 0.1 \times \frac{1}{100} = 10^{-3}$

$pH = -\log[H^+] = -\log 10^{-3} = 3$

182. (b) $pH = 4$

$pH + pOH = 14$; $pOH = 14 - pH$

$pOH = 14 - 4 = 10$; $[OH^-] = 10^{-10}\ M$

183. (b) $pH = \log \frac{1}{[H^+]} = \log \frac{1}{[3 \times 10^{-3}]} = 2.523$

185. (c) It is a strong base.

186. (b) $[H^+]$Concentration in 0.01*M HCl* is $10^{-2} M$ because 0.01 *M HCl* have only H^+

$HCl \rightleftharpoons H^+ + Cl^-$.

187. (b) The value of H_3O^+ ions will not changed.

$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$.

188. (a) H_2O_2 (Hydrogen peroxide) is a corrosive volatile liquid. It is slightly acidic in nature. Its pKa value is approximately 10^{-12} .

192. (b) $pH = \log \frac{1}{H^+}$; $pH = \log \frac{1}{10^{-3}}$; $pH = 3$.

193. (b) NH_3 is Lewis base because of one lone pair of electron.

194. (b) $Ba(OH)_2 \rightarrow \underset{.05M}{Ba^{+2}} + \underset{2 \times 0.5M}{2OH^-}$

$pOH = \log \frac{1}{[OH]^-} = \log \frac{1}{.1} = 1$

$pH + pOH = 14$; $pH + 1 = 14$; $pH = 14 - 1 = 13$

195. (a) If concentration of acid is increases ten times in a buffer then *pH* of the solution is increase by one.

198. (b) $pH > 7 =$ Basic

It means contain more hydroxide ions than carbonate ions.

199. (a) At 7*pH* the concentration of OH^- and H^+ are equal.

200. (a,d) M.eq. of 0.01 $M\ HCl = \frac{.01\times 100}{1000} = 1\times 10^{-3}$

$pH = 3$

M.eq. of .02 $M\ H_2SO_4 = \frac{.04\times 50}{1000} = 2\times 10^{-3}$

M.eq. of .02 $M\ NaOH = \frac{0.02\times 50}{1000} = 1\times 10^{-3}$

Left $[H^+] = 2\times 10^{-3} - 1\times 10^{-3} = 1\times 10^{-3}$; $pH = 3$

202. (a,b,c) Because buffer solution are mixture of weak acid or weak base and their salt.

203. (c) Because $pH = 8$ is basic nature but HCl is a strong acid.

204. (c) $H_2SO_4 = 0.05\times 2$

$\therefore [H^+] = 0.1$ and $pH = 1$

205. (b) $Mg(OH)_2 \rightleftharpoons Mg^{2+} + 2OH^-$

$K_{sp} = [Mg^{2+}][OH^-]^2$

$1\times 10^{-12} = 0.01[OH^-]^2$

$[OH^-]^2 = 1\times 10^{-10} \Rightarrow [OH^-] = 10^{-5}$

$[H^+] = 10^{-14}/10^{-5} = 10^9$

$pH = -\log[H^+] = -\log[10^{-9}] = 9$

206. (b) $[OH^-] = 1\times 10^{-5}$

$pOH = -\log[OH^-] = 5$

$pH + pOH = 14 \Rightarrow pH = 14 - 5 = 9$.

Critical Thinking Questions

1. (b) K_{sp} of $AgI = 1.5\times 10^{-16}$

$10^{-8}\ M\ Ag^+$ and $10^{-8}\ M\ I^-$

Ionic product $= 10^{-16}$

$K_{sp} =$ Ionic product

2. (a) $HClO$ is the weakest acid. Its conjugate base ClO^- is the strongest base.

3. (b) $B(OH)_3$ not have H^+

4. (a) $h = \frac{k_w}{k_a} = \frac{1\times 10^{-14}}{4.5\times 10^{-10}} = 2.22\times 10^{-5}$

5. (d) $MX_4 \rightarrow \underset{s}{M} + \underset{4s}{4X}$; $K_{sp} = (4s)^4 s$; $K_{sp} = 256 s^5$

$$s = \left(\frac{K_{sp}}{256}\right)^{1/5} .$$

6. (a) Lewis acid are electrophiles because they accept election pairs.

7. (c) $\begin{array}{lccc} & 2HI & \rightleftharpoons & H_2 + I_2 \\ \text{Initial} & 2 & & 0 \quad 0 \end{array}$

At equilibrium $2-\alpha \quad \frac{\alpha}{2} \quad \frac{\alpha}{2} \quad = 2-\alpha+\frac{\alpha}{2}+\frac{\alpha}{2} = 2$.

8. (a) Because they have vacant *d*-orbital in central atom.

9. (a) Energy has to be spent for the total dissociation of weak acid.

10. (a) pK_a of acid $A = 4$; pK_a of acid $B = 5$

We know that $pK_a = -\log K_a$

$\therefore$ Acid A $K_a = 10^{-4}$

Acid B $K_a = 10^{-5}$

Hence *A* is ten times stronger than that of *B*.

11. (b) $\frac{\alpha_1}{\alpha_2} = \sqrt{\frac{K_{a_1}}{K_{a_2}}} = \sqrt{\frac{3.14\times 10^{-4}}{1.6\times 10^{-5}}} = 4:1$

12. (d) It is a salt of weak acid and weak base.

13. (c) $HA \rightleftharpoons H^+ + A^-$; $K_a = \frac{[H^+][A^-]}{[HA]}$ ......(i)

neutralization of the weak acid with strong base is

$HA + OH^- \rightleftharpoons A^- + H_2O$

$K = \frac{[A^-]}{[HA][OH^-]}$ ......(ii)

dividing (i) by (ii) $\frac{K_a}{K} = [H^+][OH^-] = K_w = 10^{-14}$

$K = \frac{K_a}{K_w} = \frac{10^{-5}}{10^{-14}} = 10^9$.

14. (d) NH_4Cl undergoes cationic hydrolysis hence *pH* is >7 because the solution due to cationic hydrolysis in acids.

NaCN undergoes anionic hydrolysis hence *pH* is >7.

HCl is strong acid and *NaCl* is neutral solution.

Hence the *pH* of given solutions will increases.

$HCl < NaCl < NaCN < NH_4Cl$

15. (c) It is a $HClO_4$

16. (d) Smaller the $_pK_a$ value, strong the acid.

17. (d) It involves gain and loss of electron pair (Lewis concept).

18. (b) $H^-_{\text{(aq.)}} + H_2O_{(l)} \rightarrow OH^- + H_2$.

19. (b) $K_H = \frac{K_a}{K_w} = \frac{10^{-4}}{10^{-14}} = \frac{1}{10^{-10}} = 10^{10}$.

20. (d) Acidity is directly propotional to oxidation number. As the O.No. of S, P and Cl in H_2SO_3, H_3PO_3 & $HClO_3$ is +4, +3 & +5 respectively so decreasing order of acidity will be III > I > II.

21. (a) A substance which can donate a proton is known as acid so NH_4^+ will be a acid.

22. (b) Acetic acid is a weak acid.

23. (a) Because it is a acidic oxide.

24. (b) $[Ca^{2+}][F^-]^2 = \left[\frac{10^{-2}\times V}{2V}\right]\times\left[\frac{10^{-3}\times V}{2V}\right]^2 = 1.25\times 10^{-9}$.

25. (a) The degree of hydrolysis of a salt of weak acid and weak base is independent of concentration of salt.

26. (c) $C = 0.1M;\ \alpha = 1\%;\ (H^+) = C\times\alpha$

$= 0.1\times\frac{1}{100} = 10^{-3}$; $(H^+) = 10^{-3}$; $pH = 3$.

27. (d) $HClO_4$ is a strong acid, because its oxidation no is + 7.

28. (d) Acceptor of electron pair is known as lewis acid. $S,\ :CH_2, (CH_3)_3B$ all can accept an electron pair so answer is (d).

29. (b) $\underset{}{Mg(OH)^2} \rightleftharpoons \underset{(s)}{Mg^{2+}} + \underset{(2s)^2}{2OH^-}$

$K_{sp} = 4S^3 \Rightarrow S = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{1.2\times10^{-11}}{4}}$

$S = 8.16\times10^{-4}$

30. (d) $K_a = \frac{[H^+][CN^-]}{[HCN^-]}$

$6.2\times10^{-10} = \frac{[H^+][0.02]}{[0.01]}$

$[H^+] = \frac{6.2\times10^{-10}\times0.01}{0.02} = 3.1\times10^{-10}$

31. (b) $\overset{..}{N}H_3$ presence of lone pair of electrons.

32. (a) $\underset{K_{sp}}{CuBr} \rightleftharpoons \underset{(S)}{Cu^+} + \underset{(S)}{Br^-}$

$K_{sp} = S^2 = (2\times10^{-4})^2 = 4\times10^{-8}\ \frac{mol^2}{l^2}$

33. (c) $\underset{(0.004-x)}{Na_2SO_4} \rightleftharpoons \underset{2x}{2Na^+} + \underset{x}{SO_4^{2-}}$

Since both the solution are isotonic $0.004 + 2x = 0.01$

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

$\therefore$ Percent dissociation $= \frac{3\times10^{-3}}{0.004}\times100 = 75\%$.

34. (a) $Cr(OH)_3 \rightarrow \underset{x}{Cr^{+3}} + \underset{3x}{3OH^-}$

$K_{sp} = x.(3x)^3 = 27x^4$

$x = \sqrt[4]{\frac{K_{sp}}{27}}$; $x = \sqrt[4]{\frac{2.7\times10^{-31}}{27}}$

$x = 1\times10^{-8}$ *mole/litre.*

35. (a) $H^+ = c.\alpha = \sqrt{K_a.c}$

$pH = -\log\left(\sqrt{K_a.c}\right)^{1/2} = \frac{1}{2}[-\log K_a - \log c]$

$= \frac{1}{2}[4.74 - \log10^{-2}] = \frac{1}{2}[4.74 + 2] = 3.37$.

36. (b) $Ag_2SO_4 \rightleftharpoons \underset{4S^2}{2Ag^+} + \underset{S}{SO_4^{--}}$

$K_{sp} = 4S^3;\ K_{sp} = 2\times10^{-5}$

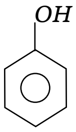
$S = \sqrt[3]{\frac{2\times10^{-5}}{4}} = 0.017\ m/l = 1.7\times10^{-2}$

$Ag\,BrO_3 \rightleftharpoons \underset{S}{Ag^+} + \underset{S}{BrO_3^-}$

$K_{sp} = S^2;\ K_{sp} = 5.5\times10^{-5}$

$S = \sqrt{5.5\times10^{-5}} = 7.4\times10^{-3}\ m/l.$

37. (c)



Phenol are more acidic than ethanol because of stabilization of phenoxide as compare to ethoxide ion.

38. (b) $X^- + H_2O \rightleftharpoons HX + OH^-$

$K_h = \frac{10^{-14}}{10^{-5}}$ so $h = \sqrt{\frac{10^{-9}}{10^{-1}}} = 10^{-4}$

$100\times10^{-4} = 10^{-2}$

So, 0.01%.

39. (a) $K_a = 1.0\times10^{-5}$

K_h = hydrolysis constant

$K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{10^{-5}} = 10^{-9}$

degree of hydrolysis $(h) = \sqrt{\frac{K_h}{C}}$

$= \sqrt{\frac{10^{-9}}{0.001}} = \sqrt{10^{-6}} = 10^{-3}$; $h = 10^{-3}$

40. (b) Basic radicals of group II & IV are precipitated by H_2S in the form of their sulphides. IInd group in acidic medium & IV group in alkaline medium. They precipitate when ionic product increases than solubility product.

41. (a) After mixing $[Ag^+][Cl^-] > K_{sp}$

42. (a) K_{sp} > ionic product $1\times10^{-10} > 5\times10^{-11}$

43. (a) 1 litre of water contains 1000/18 mole.

So degree of ionization $= \frac{10^{-7}\times18}{1000} = 1.8\times10^{-7}$ %.

44. (a) $AgCl\ K_{sp} = 1.2\times10^{-10}$

$S = \sqrt{1.2\times10^{-10}}$; $S = 1.09\times10^{-5}$

$AgBr\ K_{sp} = 3.5\times10^{-13}$

$S = \sqrt{3.5\times10^{-13}} = 5.91\times10^{-6}$

So that S of $AgBr$ is less than that of $AgCl$.

45. (a) $K_{sp} = 4S^3$

$4S^3 = 3.2\times10^{-8}$; $S = 2\times10^{-3}M$.

46. (c) $\frac{2.8\times10^{-10}}{0.1} = 2.8\times10^{-9}\,ML^{-1}$.

47. (b) $AB \rightleftharpoons A^+ + B^-$; $K_{sp} = S^2$

$S = \sqrt{K_{sp}} = \sqrt{4\times10^{-10}} = 2\times10^{-5}$

48. (b) $S = 1.435\times10^{-3}\ g/l$, $= \frac{1.435\times10^{-3}}{143.5} = 10^{-5}\ M$

$K_{sp} = S\times S = 10^{-10}$

49. (b) $pKa = 5$, so $Ka = 1\times10^{-5}$

$\alpha = \sqrt{\frac{Ka}{C}} = \sqrt{\frac{1\times10^{-5}}{0.1}} = 1\times10^{-2}$

50. (c) Presence of common ion decreases the solubility of salt.

51. (d) $Ag_2S > CuS > HgS$.

Solubility of CuS $= \sqrt{10^{-31}} = 3.16\times10^{-16}\ mol/lit.$

Solubility of Ag_2S

$= \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{10^{-42}}{4}} = 6.3\times10^{-5}\,mole/litre$

Solubility of HgS $= \sqrt{K_{sp}} = \sqrt{10^{-54}} = 10^{-27}\ mol/litre.$

52. (c) $pH = 3.82 = -\log[H^+]$

$\therefore [H^+] = 1.5\times10^{-4}$ *mole/litre.*

53. (b) $pH = pK_a + \log\frac{[\text{salt}]}{\text{acid}} = 4.57 + \log\frac{0.10}{0.03} = 5.09$

54. (c) For a monobasic acid

$[H^+] = C\alpha$

$= \frac{1}{10}\times0.001 = 10^{-4} \Rightarrow pH = 4$

55. (a) $K_{sp} = [As^{3+}][S^{2-}]$, $S = \sqrt[5]{\frac{K_{sp}}{108}}$

$= \sqrt[5]{\frac{2.8\times10^{-72}}{108}} = 1.09\times10^{-15}$

56. (d) Dissociation constant of $HA = 10^{-9}$

$HA \rightleftharpoons H^+ + A^-$

$[H^+] = \sqrt{\frac{K_a}{C}} = \sqrt{\frac{10^{-9}}{0.1}}$; $[H^+] = 10^{-4}$

$\therefore pH = 4$

$\because pH + pOH = 14$

$pOH = 14 - pH = 14 - 4$; $pOH = 10$

57. (d) $\alpha = 1.9\times10^{-9}$; $C = \frac{1000}{18}$

$K = \frac{[H^+][OH^-]}{(H_2O)} = C\alpha^2$

$= 1.9\times10^{-9}\times1.9\times10^{-9}\times\frac{1000}{18} = 2.0\times10^{-16}$.

58. (c) $K = Ka_1\times Ka_2 = 4.5\times10^{-3}\times1.7\times10^{10}$

$H^+ = \sqrt{Kc} = \sqrt{4.5\times10^{-3}\times1.7\times10^{-10}\times.01}$

$= .87\times10^{-7}$

$pH = -\log 0.87\times10^{-7} = 7 - 0.93 = 6.07$.

59. (a) Given that

Concentration of solution =.1

Degree of ionisation $= 2\% = \frac{2}{100} = .02$

Ionic product of water $= 1\times10^{-14}$

Concentration of $[H^+]$= Concentration of solution X degree of ionisation $= .1\times.02 = 2\times10^{-3}M$

Concentration of $[OH^-] = \frac{\text{Ionic product of water}}{[H^+]}$

$= \frac{1\times10^{-14}}{2\times10^{-3}} = 0.5\times10^{-11} = 5\times10^{-12}M$.

60. (a) $HgSO_4 \rightarrow \underset{x}{Hg^{++}} + \underset{x}{SO_4^{--}}$

$K_{sp} = x^2$; $x = \sqrt{K_{sp}}$; $x = \sqrt{6.4\times10^{-5}}$

$x = 8\times10^{-3}$ *mole/litre.*

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

62. (c) $[H^+] = 1\times10^{-4}M \Rightarrow [H^+][OH^-] = 1\times10^{-14}$

$\therefore [OH^-] = \frac{1\times10^{-14}}{0.5\times10^{-4}} = 2\times10^{-10}\ M$

63. (d) Because $HClO_4$ is a strong acid. While buffer is a mixture of weak acid and their salt.

64. (b) Initial concentration = 0.006 M $\Rightarrow$ $K_a = 6 \times 10^{-5}$

equilibrium reaction.

$$\begin{array}{lccc} & C_6H_5COOH & \rightleftharpoons C_6H_5COO^- + & H^+ \\ & C & O & O \ \text{initial} \\ & C(1-a) & Ca & Ca \end{array}$$

$K_a = C\alpha^2$ $(\because 1-\alpha \approx 1)$ ***for weak electrolyte***

$\alpha = \sqrt{\dfrac{K_a}{C}}$ **;** $\alpha = \sqrt{\dfrac{6 \times 10^{-5}}{0.006}}$ **;** $\alpha = 10^{-1}$

$\because [H^+] = C\alpha = 0.006 \times 10^{-1} = 6 \times 10^{-4}$ M.

65. (a) $pOH = pKb + \log \dfrac{[Salt]}{[Base]}$

$14 - 9.35 = -\log(1.78 \times 10^{-5}) + \log \dfrac{[Salt]}{100}$

$[Salt] = 79.9 \Rightarrow \dfrac{w}{132} \times 1000 = 79.9 \Rightarrow w = 10.56$

66. (a) $pH = 6$ means $[H^+] = 10^{-6} M$

$pH = 3$ means $[H^=] = 10^{-3} M$

After mixing,

Total $[H^+] = \dfrac{(10^{-6} + 10^{-3})}{2} = \dfrac{1.001 \times 10^{-3}}{2}$

$= 5.005 \times 10^{-4}$

$pH = 4 - \log 5.005$; $pH = 3.301$.

67. (b) $[H^+] = \sqrt{K \times C}$

$[H^+] = \sqrt{4 \times 10^{-10} \times 1}$; $[H^+] = 2 \times 10^{-5}$ *mole/litre*

68. (b) Moles of $[OH^-] = M \times V$

Number of $NaOH = 0.3 \times 0.005 \times 2 = 0.0030$

69. (d) $[H^+][OH^-] = 10^{-13.26}$

$\therefore$ $[H^+] = [OH^-]$, $\therefore$ $[H^+]^2 = 10^{-13.26}$

$[H^+] = 10^{\frac{-13.26}{2}}$

pH = 6.63.

70. (b) $pH = pKa + \log \dfrac{[Salt]}{[Acid]}$

$pH = pKa$

$Ka = 0.1 \times (10^{-3.5})^2 = 0.1 \times 10^{-7} = 10^{-8} \Rightarrow pH = 8$

71. (a) $pH = -\log K_a + \log \dfrac{[\text{Salt}]}{[\text{Acid}]}$

$[\text{Salt}] = \dfrac{0.2 \times 50}{1000} = 0.01$; $[\text{Acid}] = \dfrac{0.5 \times 40}{1000} = 0.02$

$pH = -\log (1.8 \times 10^{-4}) + \log \dfrac{0.01}{0.02}$

$pH = 4 - \log (1.8) + \log 0.5$

$pH = 4 - \log (1.8) - 0.301$

pH = 3.4

72. (c) HCO_3^- and H_2O

73. (a) $[H^+] = C \times \alpha = 0.1 \times 0.1 = 10^{-2} M$

pH = 2; pOH = 12 ; $[OH^-] = 10^{-12} M$.

74. (c) Decreasing order of acidic character is $H_2SO_4 > CH_3COOH > H_2CO_3$

75. (c) The acidic nature increases in the order.

$NaNO_2 < NaCl < H_2S < H_2SO_4$

76. (a) $[OH^-] = 0.05 = 5 \times 10^{-2} M$

$pOH = 2 - \log 5 = 1.3$

$pH + pOH = 14$

$pH = 14 - 1.3 = 12.7$

77. (b) Na_2O form $NaOH$. So that it is basic oxide.

78. (a) $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

$\therefore [OH^-] = c \times h;$ $h = \sqrt{\dfrac{K_w}{K_a} \times c} = \sqrt{\dfrac{10^{-14}}{1.8 \times 10^{-5}} \times 1}$

$= 2.35 \times 10^{-5}$

$\therefore pOH = 4.62$; $pH = 9.38 \approx 9.4$

Assertion & Reason

1. (c) HCl is a strong electrolyte since it will produce more H^+, comparison than that of CH_3COOH . Hence assertion is true but reason false.

2. (b) For sparingly soluble salts, reason is not a correct explanation. Hence both assertion and reason are true but reason is not a correct explanation of assertion.

3. (a) Aq. solution of $FeCl_3$ on standing produce brown ppt. Due to hydrolysis it produce ppt. of $Fe(OH)_3$ which is of brown colour. Hence both are correct and reason is a correct explanation.

4. (a) Barium carbonate is more soluble in HNO_3 than in water become carbonate is a weak base and reacts with the H^+ ion of HNO_3 causing the barium salt to dissociate.

$BaCO_3 + HNO_3 \rightarrow Ba(NO_3)_2 + CO_2 + H_2O$

5. (a) The conjugate base of $CHCl_3$ is more stable than conjugate base of $CHF_3 (CF_3)$. CCl_3 stabilized by $-I$ effect of chlorine atoms as well as by the electrons. But conjugate base of $CH_3 (CH_3)$ is stabilized only by $-I$ effect of flourine atoms. Here both assertion and reason are true and reason is correct explanation of assertion.

6. (c) Ionic product of $AgBr$ is greater than that of $AgCl$ in comparison with there solubility product $AgBr$ will precipitate. First rather than that of $AgCl$.

9. (e) It is fact that ionic reactions are instantaneous due to the fact that oppositely charged ions exert strong forces and combine immediately.

Ionic Equilibrium

SET Self Evaluation Test - 9

1. The most important buffer in the blood consists of **[BHU 1981]**

(a) HCl and $Cl^{\oplus}$ (b) H_2CO_3 and $HCO_3^{\ominus}$

(c) H_2CO_3 and $Cl^{\ominus}$ (d) HCl and $HCO_3^{\ominus}$

2. The solubility product of AgI at $25°C$ is $1.0 \times 10^{-16} mol^2 L^{-2}$. The solubility if AgI in $10^{-4} N$ solution of KI at $25°C$ is approximately (in mol l^{-1}) **[CBSE PMT 2003]**

(a) 1.0×10^{-8} (b) 1.0×10^{-16}

(c) 1.0×10^{-12} (d) 1.0×10^{-10}

3. The pH of the solution: $5\ mL$ of $\frac{M}{5}$, $HCl + 10\ mL$ of $\frac{M}{10}\ NaOH$ is **[MH CET 2004]**

(a) 5 (b) 3

(c) 7 (d) 8

4. Given that the dissociation constant for H_2O is $K_w = 1 \times 10^{-14} mole^2/litre^2$. What is the pH of a 0.001 *molar KOH* solution **[UPSEAT 2000; MP PET 2001]**

(a) 10^{-11} (b) 10^{-3}

(c) 3 (d) 11

5. The pH of 0.1 M solution of the following salts increases in the order **[Pb. CET 2004]**

(a) $NaCl < NH_4Cl < NaCN < HCl$

(b) $HCl < NH_4Cl < NaCl < NaCN$

(c) $NaCN < NH_4Cl < NaCl < HCl$

(d) $HCl < NaCl < NaCN < NH_4Cl$

6. The degree of hydrolysis in hydrolytic equilibrum $A^- + H_2O \rightleftharpoons HA + OH^-$ at salt concentration of 0.001 M is $\left(K_a = 1 \times 10^{-5}\right)$ **[UPSEAT 2004]**

(a) 1×10^{-3} (b) 1×10^{-4}

(c) 5×10^{-4} (d) 1×10^{-6}

7. If pK_b for fluoride ion at 25^oC is 10.83, the ionisation constant of hydrofluoric acid in water at this temperature is **[IIT 1997]**

(a) 1.74×10^{-3} (b) 3.52×10^{-3}

(c) 6.75×10^{-4} (d) 5.38×10^{-2}

8. If the hydrogen ion concentration of a given solution is $5.5 \times 10^{-3} mol\ litre^{-1}$, the pH of the solution will be **[AMU 1985]**

(a) 2.26 (b) 3.40

(c) 3.75 (d) 2.76

9. Henderson's equation is $pH = pK_a + \log\frac{[\text{salt}]}{[\text{acid}]}$. If the acid gets half neutralized the value of pH will be : $[pK_a = 4.30]$ **[RPMT 2000]**

(a) 4.3 (b) 2.15

(c) 8.60 (d) 7

10. The pH of a $0.01 M$ solution of acetic acid having degree of dissociation 12.5% is **[JIPMER 2000]**

(a) 5.623 (b) 2.903

(c) 3.723 (d) 4.509

11. Which of the following solutions will have pH close to 1.0 **[IIT 1992; MP PET 1993; AMU 1999]**

(a) $100\ ml$ of $\frac{M}{10} HCl + 100\ ml$ of $\frac{M}{10} NaOH$

(b) $55\ ml$ of $\frac{M}{10} HCl + 45\ ml$ of $\frac{M}{10} NaOH$

(c) $10\ ml$ of $\frac{M}{10} HCl + 90\ ml$ of $\frac{M}{10} NaOH$

(d) $75\ ml$ of $\frac{M}{5} HCl + 25\ ml$ of $\frac{M}{5} NaOH$

12. In which of the following solvents will $AgBr$ have the highest solubility **[CBSE PMT 1992]**

(a) $10^{-3}\ M\ NaBr$ (b) $10^{-3}\ M\ NH_4OH$

(c) Pure water (d) $10^{-3}\ M\ HBr$

13. How many grams of CaC_2O_4 will dissolve in distilled water to make one litre of saturated solution ? (Solubility product of CaC_2O_4 is $2.5 \times 10^{-9} mole^2 litre^{-2}$ and its molecular weight is 128) **[MP PET 1993; MP PMT 2000]**

(a) 0.0064 *gm* (b) 0.0128 *gm*

(c) 0.0032 *gm* (d) 0.0640 *gm*

14. The solubility product of CuS, Ag_2S, HgS are 10^{-31}, 10^{-44}, 10^{-54} respectively. The solubilities of these sulphides are in the order

(a) $Ag_2S > CuS > HgS$ (b) $Ag_2S > HgS > Cus$

(c) $HgS > Ag_2S > Cus$ (d) $CuS > Ag_2S > HgS$

15. The solubility product constant K_{sp} of $Mg(OH)_2$ is 9.0×10^{-12}. If a solution is $0.010\ M$ with respect to Mg^{2+} ion, what is the maximum hydroxide ion concentration which could be present without causing the precipitation of $Mg(OH)_2$

(a) $1.5 \times 10^{-7}\ M$ (b) $3.0 \times 10^{-7}\ M$

(c) $1.5 \times 10^{-5}\ M$ (d) $3.0 \times 10^{-5}\ M$

16. If the K_b value in the hydrolysis reaction $B^+ + H_2O \rightleftharpoons BOH + H^+$ is 1.0×10^{-6}, then the hydrolysis constant of the salt would be **[Roorkee Qualifying 1998]**

(a) 1.0×10^{-6} (b) 1.0×10^{-7}

(c) 1.0×10^{-8} (d) 1.0×10^{-9}

17. For a sparingly soluble salt A_pB_q, the relationship of its solubility product (L_S) with its solubility (S) is

[IIT Screening 2001]

(a) $L_s = S^{p+q}.p^p.q^q$ (b) $L_s = S^{p+q}.p^q.q^p$

(c) $L_s = S^{pq}.p^p.q^q$ (d) $L_s = S^{pq}.(p.q)^{p+q}$

18. Arrange NH_4^+, H_2O, H_3O^+, HF and OH^- in increasing order of acidic nature **[BVP 2003]**

(a) $H_3O^+ < NH_4^+ < HF < OH^- < H_2O$

(b) $NH_4^+ < HF < H_3O^+ < H_2O < OH^-$

(c) $OH^- < H_2O < NH_4^+ < HF < H_3O^+$

(d) $H_3O^+ > HF > H_2O > NH_4^+ > OH^-$

19. How many grams of CaC_2O_4 (molecular weight = 128) on dissolving in distilled water will give a saturated solution $[K_{sp}(CaC_2O_4) = 2.5 \times 10^{-9}\ mol^2 l^{-2}]$ **[KCET 2003]**

(a) 0.0064 *g* (b) 0.1280 *g*

(c) 0.0128 *g* (d) 1.2800 *g*

20. If the concentration of CrO_4^- ions in a saturated solution of silver chromate is 2×10^{-4}. Solubility product of silver chromate will be

(a) 4×10^{-8} (b) 8×10^{-12}

(c) 12×10^{-12} (d) 32×10^{-12}

21. According to Bronsted-Lowry concept, the correct order of relative strength of bases follows the order

[Pb. PMT 2001]

[CBSE PMT 1997]

(a) $CH_3COO^- > Cl^- > OH^-$

(b) $CH_3COO^- > OH^- > Cl^-$

(c) $OH^- > CH_3COO^- > Cl^-$

(d) $OH^- > Cl^- > CH_3COO^-$

22. $H_2SO_4^- + OH^- \rightarrow SO_4^{2-} + H_2O$ Which is correct about conjugate acid base pair **[JEE Orissa 2004]**

[Pb. PMT 1998]

(a) HSO_4^{2-} is conjugate acid of base SO_4^{2-}

(b) HSO_4^- is conjugate base of acid SO_4^{2-}

(c) SO_4^- is conjugate acid of base HSO_4^-

(d) None of these

23. Which may be added to one litre of water to act as a buffer

[JIPMER 2000]

(a) One mole of $HC_2H_3O_2$ and 0.5 mole of $NaOH$

(b) One mole of NH_4Cl and one mole of HCl

(c) One mole of NH_4OH and one mole of $NaOH$

(d) One mole of $HC_2H_3O_2$ and one mole of HCl

24. Which of the following base is weakest **[DCE 2003]**

(a) $NH_4OH : K_b = 1.6 \times 10^{-6}$

(b) $C_6H_5NH_2 : K_b = 3.8 \times 10^{-10}$

(c) $C_2H_5NH_2 : K_b = 5.6 \times 10^{-4}$

(d) $C_6H_7N : K_b = 6.3 \times 10^{-10}$

25. $HClO$ is a weak acid. The concentration of H^+ ions in $0.1M$ solution of $HClO\ (K_a = 5 \times 10^{-8})$ will be equal to

[CPMT 1993]

(a) $7.07 \times 10^{-5}\ m$ (b) $5 \times 10^{-9}\ m$

(c) $5 \times 10^{-7}\ m$ (d) $7 \times 10^{-4}\ m$

26. Upto what *pH* must a solution containing a precipitate of $Cr(OH)_3$ be adjusted so that all of precipitate dissolves

(When $Cr^{3+} = 0.1 mol/l, K_{sp} = 6 \times 10^{-31}$) **[MP PET 2003]**

[MP PET 1992; CPMT 1993]

(a) Upto 4.4 (b) Upto 4.1

(c) Upto 4.2 (d) Upto 4.0

27. NH_4Cl is acidic, because **[JEE Orissa 2004]**

(a) On hydrolysis NH_4Cl gives weak base NH_4OH and strong acid HCl

(b) Nitrogen donates a pair of electron

(c) It is a salt of weak acid and strong base

(d) On hydrolysis NH_4Cl gives strong base and weak acid

28. A solution of weak acid HA containing 0.01 *moles* of acid *per litre* of solutions has $pH = 4$. The percentage degree of ionisation of the acid and the ionisation constant of acid are respectively **[UPSEAT 2001]**

(a) $1\%, 10^{-6}$ (b) $0.01\%, 10^{-4}$

(c) $1\%, 10^{-4}$ (d) $0.01\%, 10^{-6}$

29. The pH of a buffer solution containg 0.2 *mole per litre* CH_3COONa and 1.5 *mole per litre* CH_3COOH is (Ka for acetic acid is 1.8×10^{-5}) **[CPMT 2001]**

(a) 4.87 (b) 5.8

(c) 2.4 (d) 9.2

30. 100 *mL* of $0.04\ N\ HCl$ aqueous solution is mixsed with 100 *mL* of 0.02 *N NaOH* solution. The *p*H of the resulting solution is **[UPSEAT 2004]**

(a) 1.0 (b) 1.7

(c) 2.0 (d) 2.3

31. An alcoholic drink substance $pH = 4.7$ then OH ion concentration of this solution is $(K_w = 10^{-14} mol^2/l^2)$

[RPMT 2002]

(a) 3×10^{-10} (b) 5×10^{-10}

(c) 1×10^{-10} (d) 5×10^{-8}

32. In its 0.2 *M* solution, an acid ionises to an extent of 60%. Its hydrogen ion concentration is

(a) 0.6 *M* (b) 0.2 *M*

(c) 0.12 *M* (d) None of these

33. *p*H of $0.1\ M\ NH_3$ aqueous solution is

$(K_b = 1.8 \times 10^{-5})$ **[UPSEAT 2004]**

(a) 11.13 (b) 12.5

(c) 13.42 (d) 11.55

34. 40 *mg* of pure sodium hydroxide is dissolved in 10 litres of distilled water. The *p*H of the solution is

[Kerala PMT 2004]

(a) 9.0 (b) 10

(c) 11 (d) 12

(e) 8

35. Solubility of PbI_2 is 0.005 *M*. Then, the solubility product of PbI_2 is **[BVP 2004]**

(a) 6.8×10^{-6}

(b) 6.8×10^{6}

(c) 2.2×10^{-9}

(d) None of these

36. A monoprotic acid in a $0.1\ M$ solution ionizes to 0.001%. Its ionisation constant is

[MP PET 1985,88,99; MP PMT1988; CPMT 2003]

(a) 1.0×10^{-3} (b) 1.0×10^{-6}

(c) 1.0×10^{-8} (d) 1.0×10^{-11}

37. Select the pK_a value of the strongest acid from the following

[KCET 2004]

(a) 1.0 (b) 3.0

(c) 2.0 (d) 4.5

38. At 90°*C*, pure water has H_3O^+ ion concentration of $10^{-6} mol/L^{-1}$. The K_w at 90°*C* is **[DCE 2004]**

(a) 10^{-6} (b) 10^{-14}

(c) 10^{-12} (d) 10^{-8}

39. By adding 20 $ml\ 0.1\ N\ HCl$ to $20\ ml\ 0.1\ N\ KOH$, the pH of the obtained solution will be **[CPMT 1975, 86, 93]**

(a) 0 (b) 7

(c) 2 (d) 9

Answers and Solutions

(SET -9)

1. (b) Blood consists of $H_2CO_3 + HCO_3^-$ buffer solution.

2. (c) $AgI \rightleftharpoons \underset{(s)}{Ag^+} + \underset{(s)}{I^-}$; $K_{sp} = S^2 = 10^{-4} \times S$

$$S = \frac{1.0 \times 10^{-16}}{10^{-4}} = 1 \times 10^{-12} \frac{mol^2}{l^2}$$

3. (c) Milliequivalents of $HCl = 5 \times \frac{1}{5} = 1$

Milliequivalents of *NaOH* $= 10 \times \frac{1}{10} = 1$

$$\therefore\ 5ml \frac{M}{5} HCl = 10\ ml \frac{M}{5} HCl$$

Hence the solution will be neutral *i.e.*, $pH = 7$.

4. (d) $pH = 14 - pOH = 14 - 3 = 11$

5. (b) *HCl* is strong acid. In its .1*M* solution, $[H^+] = 0.1M$ and hence, $pH = 1$

$NH_4Cl_{(aq)}$ hydrolyses in solution and give acidic solution which is less acidic than .1*M HCl*. *NaCl* is not hydrolysed in aqueous solutions. Its $pH = 7$ *NaCN* undergoes hydrolysis in solution to give alkaline solution. So that *pH* increases in the order, $HCl < NH_4Cl < NaCl < NaCN$

6. (a) $K_h = \frac{K_w}{K_a} = \frac{10^{-14}}{1 \times 10^{-5}} = 10^{-9}$

$$K_h = \alpha^2 C\ ;\ \alpha = \sqrt{\frac{K_h}{C}} = \sqrt{\frac{1 \times 10^{-9}}{.001}} = 1 \times 10^{-3}$$

7. (c) $K_a \times K_b = K_w$

$$\therefore K_a = \frac{K_w}{K_b} = \frac{10^{-14}}{1.48 \times 10^{-11}} = 6.75 \times 10^{-4}$$

8. (a) $[H^+] = 5.5 \times 10^{-3}$ *mole/litre*

$$pH = -\log[H^+]\ ;\ pH = -\log[5.5 \times 10^{-3}]\ ;\ pH = 2.26$$

9. (a) $pH = pK_a + \log\frac{[\text{Salt}]}{[\text{Acid}]}$

$$pH = 4.3 + \log\frac{\frac{1}{2}}{\frac{1}{2}} = 4.3 + \log 1\ ;\ pH = 4.3 + 0 = 4.3$$

10. (b) $[H^+] = C\alpha = 0.01 \times \frac{1.25}{100}$

$$H^+ = 1.25 \times 10^{-3}\ ;\ pH = \text{between}\ \ 2 \text{ or } 3 = 2.90$$

11. (d) M.eq. of $HCl = \frac{1}{5} \times 75 = 15$

M.eq. of $NaOH = 25 \times \frac{1}{5} = 5$

Total No. of eq. $= 15 - 5 = 10$

Total volume = 100

Normality $= \frac{10}{100} = \frac{1}{10}, [H^+] = 10^{-1}\ M$

12. (b) *AgBr* are not dissolved in *NaBr* and *HBr* due to common ion effect. And pure water is a neutral solvent. They do not have ions.

13. (a) CaC_2O_4 is a binary electrolyte. Then solubility is

$$S = \sqrt{K_{sp}} = \sqrt{2.5 \times 10^{-9}}$$

$= 5 \times 10^{-5}$ *mole/l.* $= 0.0064$ *gm/l.*

14. (a)

15. (d) $\underset{K_{sp}}{Mg(OH)_2} \underset{=}{\rightleftharpoons} \underset{S}{Mg^{++}} + \underset{(2S)^2}{2OH^-}$

$$K_{sp} = S \times 4S^2$$

$$\frac{K_{sp}}{S \times 4} = S^2 = \frac{9 \times 10^{-12}}{.010 \times 4} = 2.25 \times 10^{-10}$$

$S = \sqrt{2.25 \times 10^{-10}} = 1.5 \times 10^{-5}\ m/l$

16. (c) For hydrolysis of B^+; $K_H = \frac{K_w}{K_b} = \frac{10^{-14}}{10^{-6}} = 10^{-8}$.

17. (a) $A_pB_q \rightleftharpoons pA^{1+} + qB^{p-}$

$L_s = [A^{q+}]^p[B^{p-}]^q = (p \times S)^p(q \times S)^q = S^{p+q}.p^p.q^q$.

18. (c) $H_3O^+ > HF > NH_4^+ > H_2O > OH^-$.

Acidic nature is decreasing order.

19. (a) Solubility of $CaC_2O_4 = \sqrt{K_{sp}} = \sqrt{2.5 \times 10^{-9}}$

$= 5 \times 10^{-5} molL^{-1}$

$= 5 \times 10^{-5} \times 128 = 640 \times 10^{-5} = 0.0064\ g$

20. (d) K_{sp} of $Ag_2CrO_4 = [Ag^+]^2\,[Cro_4^{--}]$

$CrO_4^{--} = 2 \times 10^{-4}$ then $Ag^+ = 2 \times 2 \times 10^{-4}$

$K_{sp} = (4 \times 10^{-4})^2\,(2 \times 10^{-4}) = 32 \times 10^{-12}$

21. (c) Relative strength of bases can be shown by their conjugated acids.

Conjugate acid of OH^- is H_2O which is a weak acid conjugate acid of CH_3COO^- is CH_3COOH which is stronger than H_2O. while conjugate acid of Cl^- is HCl which is strongest out of there. so the order of relative strength of bases is $OH^- > CH_3COO^- > Cl^-$.

22. (a) $\underset{\text{Conjugate acid}}{HSO_4^-} + OH^- \rightarrow \underset{\text{Conjugate base}}{SO_4^{2-}} + H_2O$

23. (a) One mole oxalic acid & 0.5 mole of $NaOH$ will make.

24. (b) Smallest value of K_b indicates that aniline $(C_2H_5NH_2)$ is the weakest base.

25. (a) $[H^+]^2 = C.\alpha = 0.1 \times 5 \times 10^{-8}$

$H^+ = \sqrt{5 \times 10^{-9}} = 7.07 \times 10^{-5}\ M.$

26. (d) $K_{sp} = [Cr^{3+}][OH^-]^3$

$[OH]^{-3} = K_{sp/Cr^{3+}} = \frac{6 \times 10^{31}}{1 \times 10^{-1}} = 6 \times 10^{-30}$

$[OH]^- = 1.8 \times 10^{-10}$

$pOH = (\log 1.8 + \log 10^{10}) = 10 + 0.25 + 1 = 11.25$

$pH = 14 - 11.25 = 2.27$

27. (a) $NH_3Cl + H_2O \rightleftharpoons NH_4OH + HCl$

NH_4Cl is a salt of weak base & strong acid so solution will be acidic.

28. (a) $H^+ = C\alpha$

$\alpha = \frac{H^+}{C} = \frac{10^{-4}}{10^{-2}} = 10^{-6}$

29. (a) $pH = -\log K_a + \log \frac{[\text{Salt}]}{[\text{Acid}]}$

$= -\log\ [1.8 \times 10^{-5}] + \log \frac{0.2}{0.1} = 4.87$

30. (c) $N_1V_1 = .04 \times 100 = 4$

$N_2V_2 = .02 \times 100 = 2$

$N_1V_1 - N_2V_2 = N_3V_3$

$4 - 2 = N_3 \times 200$, $N_3 = 10^{-2}\,M$

$pH = \log 10 \frac{1}{H^+} = \log 10 \frac{1}{10^{-2}} = 2$.

31. (b) $pH = 4.7$

$pH + pOH = 14$; $pH = 14 - 4.7$; $pOH = 9.3$

$[OH^-] = \text{Antilog}\ \ [-pOH] = \text{Antilog}\ \ [-9.3]$

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

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32. (c) $[H^+] = C.\alpha, \ = 0.2 \times 0.60 = 0.12\,M$

33. (a) $NH_4OH \rightleftharpoons NH_4^+ + OH^-$

$K_b = C\alpha^2$; $\frac{1.8 \times 10^{-5}}{.1} = \alpha^2$; $\alpha = 1.34 \times 10^{-3}$

$[OH^-] = \alpha\ .\ C = 1.34 \times 10^{-3} \times .1$

$pOH = \log 10 \frac{1}{1.34 \times 10^{-4}}$; $pOH = 2.87$

$pH + pOH = 14$; $pH + 2.87 = 14$

$pH = 14 - 2.87$; $pH = 11.13$

34. (b) $M = \frac{\text{Solute in 1 litre solution}}{\text{Molecular weight of solute}}$

$= \frac{40 \times 10^{-3}}{40} \times \frac{1}{10} = 10^{-4}\,M$

$pOH = \log 10 \frac{1}{[OH^-]} = \log 10 \frac{1}{10^{-4}} = 4$

$pH + pOH = 14$; $pH + 4 = 14 \Rightarrow pH = 10$.

35. (d) $PbI_2 \rightarrow \underset{x}{Pb} + \underset{2x}{I_2}$

$K_{sp} = 4x^3 = 4(.005)^3 = 4 \times .005 \times .005 = .4 \times 10^{-6}$.

36. (d) $\therefore$ Monoprotic acid HA

$HA \rightleftharpoons H^+ + A^-$

Ionisation constant = ?

$$\alpha = 0.001\,\% = \frac{0.001}{100} = 10^{-5}$$

$$K = \frac{\alpha^2}{V} = \frac{[10^{-5}]^2}{10} = 10^{-11}\,.$$

37. (a) $pKa <<$ then strongest acid

$pKa >>$ then weak acid

$$pKa \propto \frac{1}{\text{Acidic strength}}$$

38. (c) $\underset{10^{-6}}{H_3O^+} \rightarrow \underset{10^{-6}}{H_2O} + \underset{10^{-6}}{H^+}$

$$K_w = [H_2O][H^+] = [10^{-6}][10^{-6}] = 10^{-12}$$

39. (b) Neutralization reaction will takes place and form salt of strong acid and strong base. Which does not hydrolysed and thus $pH = 7$.