Acids, Bases and Salts

Question 1:

Select the proper choice from the given multipule choices :

Question 1.1:

What is formed by reaction of non-metal oxide with water ?

Solution :

A. Acid
An acid is formed when a metal reacts with the oxide of a non-metal.
Example:
SO_3 (g) + H_2O (l) → H_2SO_4 (aq)

Question 1.2:

Acid + Metal-oxide —> ?

Solution :

B. Salt + Water
When an acid reacts with a metal oxide, it forms salt and water.
Example:
2HCl (aq) + Na_2O(s) → 2NaCl + H_2O (l)

Question 1.3:

Which gas is produced by reaction of base with metal ?

Solution :

C. Dihydrogen
A strong base reacts with some amphoteric metals to form salt and releases dihydrogen gas.

Question 1.4:

500 ml aqueous solution is prepared by dissolving 2 moles of HC1 in water. What will be the molarity of this solution ?

Solution :

D. 4
1 mole of solute in 1 litre solution gives a solution with molarity 1 M.
In 500 ml solution, 0.5 moles of HCl gives molarity 1 M.
Thus, if 2 moles of HCl are added in 500 ml solution, the molarity is 4 M.

Question 1.5:

What is correct for acidic aqueous solution ?

Solution :

C. $[H_3O^+] > 10^{-7}M$
In acidic solutions, the concentration of H_3O+ ions is between 100 to 10-6.

Question 1.6:

Which of the following solutions is the most basic ?

Solution :

C. pH = 11.5
The pH of a basic solution ranges from 8 to 14. More the basicity, more the pH.

Question 1.7:

Which statement is incorrect ?

Solution :

C. pH scale is applicable to only non-aqueous solutions.
The pH scale is applicable to only aqueous solutions and ranges between 0 and 14.

Question 1.8:

How is the exact pH of an aqueous solution measured ?

Solution :

C. pH meter
A pH meter is an electronic instrument which is used in exact measurement of pH of an aqueous solution.

Question 1.9:

Which of the following substances is an antacid

Solution :

B. $Mg(OH)_2$
An antacidis a basic substance which is taken for the remedy of acidity. $Mg(OH)_2$ also called milk of magnesia is a commonly used antacid.

Question 1.10:

The aqueous solution having pH 11 is how many times less basic than aqueous solution having pH 8 ?

Solution :

D. 1000

The concentration of $[OH^-]$ ions in solution with pH 11 is 10^{3} and that of $[OH^-]$ ions in solution with pH 8 is 10^{-6}. Thus, the aqueous solution with pH 11 is 1000 times less basic than the aqueous solution with pH 8.

Question 1.11:

Which of the following is strong acid ?

Solution :

C. Nitric acid

Nitric acid dissociates completely in its aqueous solution. Thus, it is a strong acid.

Question 1.12:

What type of substance is NH_3 ?

Solution :

D. Weak base

NH_3 accepts a proton when dissolved in water. Thus, it is a base. However, its hydroxide dissociates partially in its aqueous solution. Thus, it is a weak base.

Question 1.13:

pH + pOH = ?

Solution :

C. 14

In distilled water at 298 K,

$[H_3O^+] = [OH^-] = 1 \times 10^{-7} M$

$\therefore [H_3O^+] \times [OH^-] = 10^{-7} \times 10^{-7} = 10^{-14} M$

$\therefore \log_{10}[H_3O^+] + \log_{10}[OH^-] = -14 \log_{10}10$

$\therefore -\log_{10}[H_3O^+] - \log_{10}[OH^-] = 14 \log_{10}10$

$\therefore$ pH + pOH = 14

Question 1.14:

Which formula is correct ?

Solution :

B. Mole = Weight/Molecular mass
One mole of a substance equals gram molecular mass of that substance.

Question 1.15:

What will be the pH of aqueous solution of NH_4 Cl ?

Solution :

C. pH < 7
NH_4Cl is the salt produced by the neutralisation reaction between a strong acid (HCl) and a weak base (NH_4OH). Thus, it is acidic in nature and its pH is less than 7.

Question 1.16:

Which of the following solutions will have pH = 2 ?

Solution :

A. 0.01 M HCl
HCl is a strong acid. Thus, 0.01 M HCl has 0.01 M H^+ ion concentration.
0.01 = 10^{-2} H^+ ions = pH 2

Question 1.17:

What will be OH“ concentration in aqueous solution having pH = 8 ?

Solution :

B. 1×10^{-6} M
The OH^- concentration in aqueous solution with pH = 8 is 1×10^{-6} M. The pOH of the solution is 6.

Question 1.18:

If the pH of aqueous solutions A, B, C and D are 1.9, 2.5, 2.1 and 3.0 then what will be the order of acidity ?

Solution :

C. D < B < C < A
The lesser the pH of a solution, the more the acidity of the solution. Thus, solution A with pH 1.9 is most acidic followed by solution C with pH 2.1, followed by solution B with pH 2.5 and solution D is the least acidic of the four solutions with pH 3.

Question 1.19:

Which solution will be basic ?

Solution :

B. $[H_3O^+] = 10^{-12}$ M
The $[H_3O^+]$ ion concentration of an acidic solution is between 100 M and 10^{-6} M. The solution with an $[H_3O+]$ ion concentration of 10^{-7} M is neutral. Thus, the basic solution will have an $[H_3O^+]$ ion concentration of 10^{-12} M.

Question 1.20:

Which substance is present in poison of honey bee ?

Solution :

C. Mellitin
Melittin is a polypeptide with 26 amino acids and is highly acidic. It is present in the poison of honey bee.

Question 2:

Answer the following questions in brief :

Question 2.1:

Write names of two Arrhenius acids and bases.

Solution :

Arrhenius acids:

- Nitric acid (HNO_3)
- Hydrochloric acid (HCl)

Arrhenius bases:

- Sodium hydroxide (NaOH)
- Potassium hydroxide (KOH)

Question 2.2:

Write products of reaction of acid with metal.

Solution :

When an acid reacts with a metal, the corresponding salt of the metal and dihydrogen gas are formed.
Acid + Metal → Salt of Metal + Dihydrogen gas

Question 2.3:

Mention names of four methods of expressing concentration of a solution.

Solution :

The methods of expressing the concentration of a solution are

- Molarity
- Normality
- Formality
- Percentage proportion

Question 2.4:

Which scientist presented pH scale ? Write the formula of pH.

Solution :

S. P. L. Sorensen presented the pH scale to conveniently express the concentration of hydronium ions in an aqueous solution.
The formula of pH is
$pH = -\log_{10}[H_3O^+]$

Question 2.5:

Write names of two methods to measure approximate pH of aqueous solution.

Solution :

Methods to measure the approximate pH of an aqueous solution are

- Use of pH paper
- Use of universal indicator

Question 2.6:

Mention names of two strong acids and two strong bases.

Solution :

Two strong acids:

- Sulphuric acid (H_2SO_4)
- Nitric acid (HNO_3)

Two strong bases:

- Sodium hydroxide (NaOH)
- Potassium hydroxide (KOH)

Question 2.7:

Write names of two weak acids and two weak bases.

Solution :

Two weak acids:

- Acetic acid (CH_3COOH)
- Carbonic acid (H_2CO_3)

Two weak bases:

- Ammonium hydroxide (NH_4OH)
- Ammonia (NH_3)

Question 2.8:

Give definitions :

Solution :

1. Arrhenius acid:
   An acid is a substance containing hydrogen which produces hydrogen ion (H^+) in its aqueous solution.
2. Arrhenius base:
   A base is a substance containing hydroxide which produces hydroxide ion (OH^-) in its aqueous solution.
3. Bronsted-Lowry acid:
   The substance which donates a proton (H^+) to the other substance is called a Bronsted-Lowry acid.
4. Bronsted-Lowry base:
   The substance which accepts a proton (H^+) from the other substance is called a Bronsted-Lowry base.
5. Concentration of a solution:
   The amount of solute in relation to the amount of solvent is called the concentration of a solution.
6. pH of a solution:
   The pH is defined as the negative logarithm to the base 10 of concentration of H_3O^+ in aqueous solution. pH = -log10[H_3O^+]
7. pOH of a solution:
   The pOH is defined as the negative logarithm to the base 10 of concentration of OH^- in aqueous solution. pOH = -log10[OH^-]
8. Strong acid:
   The acid which ionises or dissociates completely in an aqueous solution is called a strong acid.
9. Weak acid:

The acid which ionises or dissociates incompletely is called a weak acid.

10. Strong base:

    The base which ionises or dissociates completely in an aqueous solution is called a strong base.

11. Weak base:

    The base which ionises or dissociates partially or incompletely is called a weak base.

12. Neutralisation reaction:

    The reaction of an acid with a base or a base with an acid, resulting in the formation of salt and water is called a neutralisation reaction.

13. 1 M concentration:

    When 1 mole of a solute is dissolved in 1 litre of a solvent, the solution has 1 M concentration.

Question 2.9:

Mention the formula, name and physical state of the products of following reactions :

Solution :

1. CO_2 (g) + H_2O (l) → H_2CO_3 (aq)

   H_2CO_3– Carbonic acid
2. Li_2O (s) + H_2O (l) → 2LiOH (aq)

   LiOH – Lithium hydroxide
3. H_2SO_4 (aq) + 2KOH (aq) → K_2SO_4 (aq) + $2H_2O$(l)

   K_2SO_4 – Potassium sulphate

   $2H_2O$ – Water
4. 2HCl (aq) + Ca(s) → $CaCl_2$ (aq) + H_2 (g)

   $CaCl_2$ – Calcium chloride

   H_2 – Dihydrogen gas
5. $2HNO_3$ (aq) + CaO (s) → $Ca(NO_3)_2$ (aq) + H_2O (l)

   $Ca(NO_3)_2$ – Calcium nitrate

   H_2O – Water
6. 2HCl (aq) + Na_2CO_3 (aq) → 2NaCl (aq) + H_2O (l) + CO_2 (g)

   NaCl – Sodium chloride

   H_2O – Water

   CO_2 – Carbon dioxide
7. NaOH (aq) + HCl(aq) → NaCl(aq) + H_2O(l)

   NaCl – Sodium chloride

   H_2O – Water
8. $BaCl_2$ (aq) + H_2SO_4 (aq) → $BaSO_4$ (s) + HCl (l)

   $BaSO_4$ – Barium sulphate

   HCl – Hydrochloric acid
9. $Ca(OH)_2$ (aq) + CO_2 (g) → $CaCO_3$(s) + H_2O (l)

   $CaCO_3$ – Calcium carbonate

   H_2O – Water
10. HNO_3(aq) + NH_4OH (aq) → NH_4NO_3 (aq) + H_2O

    NH_4NO_3 – Ammonium nitrate

    H_2O – Water
11. Na_2SO_4 (aq) + $Ba(OH)_2$ (aq) → $BaSO_4$ (s) + 2NaOH

    $BaSO_4$ – Barium sulphate

    NaOH – Sodium hydroxide
12. $AgNO_3$ (aq) + NaCl (aq) → AgCl (s) + $NaNO_3$ (aq)

AgCl – Silver Chloride
$NaNO_3$ – Sodium Nitrate

Question 3 (A):

Answer the following questions :

Question 3.A.1:

Write two chemical properties of acid.

Solution :

Two chemical properties of an acid are

- An acid reacts with a base to form salt and water. This reaction is called a neutralisation reaction.

Acid + Base → Salt + Water

- An acid reacts with a metal carbonate or a metal hydrogen carbonate to produce salt water and carbon dioxide gas.

Acid + Metal carbonate/Metal hydrogen carbonate → Salt + Water + Carbon dioxide

Question 3.A.2:

Write two chemical properties of base.

Solution :

Two chemical properties of a base are

- A base reacts with a non-metal oxide to form salt and water.

Base + Non-metal oxide → Salt + Water

- A strong base reacts with a metal to form salt and hydrogen gas.

Base + Metal → Salt + Dihydrogen gas

Question 3.A.3:

Write two chemical properties of salt.

Solution :

Two chemical properties of salt are

- Salt reacts with an acid to form another salt and acid.

Salt + Acid → Other salt + Acid

- Salt reacts with a base to form another salt and base.

Salt + Base → Other salt + Base

Question 3.A.4:

Deduce pH + pOH = 14

Solution :

Distilled water is neither acidic nor basic. It is neutral.
In distilled water at 298 K,
$[H_3O^+] = [OH^-] = 1 \times 10^{-7} M$
$\therefore [H_3O^+] \times [OH^-] = 10^{-7} \times 10^{-7} = 10^{-14} M$
$\therefore \log_{10}[H_3O^+] + \log_{10}[OH^-] = -14 \log_{10}10$
$-\log_{10}[H_3O^+] - \log_{10}[OH^-] = 14 \log_{10}10$
$\therefore$ pH + pOH = 14

Question 3.A.5:

Explain the importance of pH in digestion of food.

Solution :

pH is important at different levels in the digestive system of the human body for the proper digestion of food.
In the stomach, hydrochloric acid is secreted as food enters the stomach. It turns the pH of the stomach between 1 and 3. This pH is important for the activation of the enzyme pepsin, which helps in the digestion of protein in food.

Question 3.A.6:

Write chemical equations of four neutralisation reactions.

Solution :

Chemical equations of four neutralisation reactions:
H_2CO_3 (aq) + $2NH_4OH$ (aq) → $(NH_4)2CO_3$ (aq)+ $2H_2O$ (l)
CH_3COOH (aq) + NH_4OH (aq) → CH_3COONH_4 (aq) + H_2O (l)
H_2CO_3 (aq) + 2NaOH(aq) → Na_2CO_3 (aq) + $2H_2O$ (l)
HNO_3 + KOH(aq) → KNO_3 (aq) + H_2O (l)

Question 3 (B):

Solve the following problems :

Question 3.B.1:

Calculate pH of an aqueous solution having H_30^+ concentration equal to 7.9×10^{-11} M. Which nature, acidic, basic or neutral will be possessed by this aqueous solution.

Solution :

In aqueous solution $[H_3O^+] = 7.9 \times 10^{11}$ M

By definition,

$pH = -\log_{10}[H_3O^+]$

$= -\log_{10}(7.9 \times 10^{-11})$

$= -\log_{10}(7.9) - \log_{10}(10^{-11})$

$= -\log_{10}(7.9) + 11 \log_{10}10$

$= -0.8976 + 11$

$= 10.1024$

10.10

Thus, the pH of the given solution with $[H_3O^+] = 7.9 \times 10^{-11}$ M is 10.1, so the solution is basic in nature.

Question 3.B.2:

Calculate pH of 0.00424 M aqueous solution of KOH.

Solution :

KOH is a strong base; thus, it gets completely ionised in the aqueous solution.

$KOH(aq) + H_2O \rightarrow K^+(aq) + OH^-(aq)$

The given solution is 0.00424 M.

Thus, $[OH^-] = 0.00424$ M

$= 4.24 \times 10^{-3}$ M

Now,

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

$= -\log_{10}(4.24 \times 10^{-3})$

$= -\log_{10}4.24 - \log_{10}10^{-3}$

$= -\log_{10}4.24 + 3\log_{10}10$

$= -0.6274 + 3$

$= 2.3726$

Now,

$pH + pOH = 14$

$pH = 14 - pOH$

$= 14 - 2.3726$

$= 11.6274$

11.63

Thus, the pH of 0.00424 M aqueous solution of KOH is 11.63.

Question 3.B.3:

How many times concentrated will be the aqueous solution having pH 11.9 as compared to aqueous solution having pH 8 ?

Solution :

For aqueous solution with pH 11.9, pOH is 2.1.

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

$2.1 = -\log_{10}[OH^-]$

$\log_{10}[OH^-] = -2.1$

$\log_{10}[H_3O^+] = 3 - 2.1 - 3$

$\log_{10}[H_3O^+] = -3 + 0.9$

$\log_{10}[H_3O^+] = + 0.9$

$[H_3O^+]$ = antilog (.9)

$[H_3O^+] = 7.943 \times 10^{3}$M

For aqueous solution with pH 8, pOH is 6.

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

$-\log_{10}[OH^-] = 6$

$\log_{10}[OH^-] = -6$

$[OH^-] = 1 \times 10^{6}$M ($\log_a b = m$ for $a^m = b$)

$$\frac{[H_3O^+]\text{in aqueous solution having pH 11.9}}{[H_3O^+]\text{in aqueous solution having pH 8}} = \frac{7.943\times10^{-3}}{1\times10^{-6}} = 7943$$

Thus, the aqueous solution with pH 8 is 7943 times more concentrated than the aqueous solution with pH 11.9.

Question 3.B.4:

How will you prepare 500 ml aqueous solution of 0.2 M H_2S0_4 ?

Solution :

Molar mass of H_2SO_4 = 2(Atomic mass of H) + (Atomic mass of S) + 4(Atomic mass of O)

= 2(1) + (32) + 4(16)

= 2 + 32 + 64

= 98

According to the definition of molarity, to prepare 1000 ml of 1 M H_2SO_4, 98 gms of H_2SO_4 will be required.

$\therefore$ To prepare 500 ml of 0.2 M solution, $\frac{500 \times 0.2 \times 98}{1000 \times 1}$ of H_2SO_4 will be required.

= 9.8 gms

Thus, 9.8 gms of H_2SO_4 should be dissolved in 500 ml of water to obtain an aqueous solution of H_2SO_4 of molarity 0.2 M.

Question 3.B.5:

How will you prepare 125 ml 0.03 M aqueous solution of KOH.

Solution :

Molar mass of KOH= (Atomic mass of K) + (Atomic mass of O) + (Atomic mass of H)

= (39) + (16) + (1)

= 56

According to the definition of molarity, to prepare 1000 ml of 1 M KOH, 56 gms of KOHwill be required.

$\therefore$ To prepare 125 ml of 0.03 M solution, $\frac{125 \times 0.03 \times 56}{1000 \times 1}$ of KOHwill be required.

= 0.21 gms

Thus, 0.21 gms of KOHshould be dissolved in 125 ml of water to obtain an aqueous solution of KOHof molarity 0.03 M.

Question 4 (A):

Answer the following questions in detail :

Question 4.A.1:

Explain giving example, Arrhenius acid-base theory. Mention the limitations of this theory.

Solution :

Arrhenius gave a definite concept about acids and bases. According to Arrhenius, an acid is a substance containing hydrogen which produces hydrogen ion (H^+) in its aqueous solution, and a base is a substance containing hydroxide which produces hydroxide ion (OH^-) in its aqueous solution.

For example, hydrogen nitrate (nitric acid) produces hydrogen ion in its aqueous solution, and sodium hydroxide produces hydroxide ion in its aqueous solution.
This hydrogen nitrate is an acid and sodium hydroxide is a base.

$$HNO_3 (l) \xrightarrow{H_2O} H^+ (aq) + NO_3^- (aq)$$

$$KOH (s) \xrightarrow{H_2O} K^+ (aq) + OH^- (aq)$$

Limitations of Arrhenius theory:

- It is applicable to aqueous solutions only.
- It could not explain the basic nature of ammonia in its aqueous solution, as ammonia (NH_3) does not have the hydroxide ion.
- An Arrhenius acid is based on the formation of the H^+ ion. However, it is highly unstable and does not exist independently. It combines immediately with the solvent and forms hydronium ions, i.e. H_3O^+.

Question 4.A.2:

Discuss Bronsted-Lowry Acid-base theory.

Solution :

Bronsted-Lowry presented a detailed concept of acid-base. The theory proposed by them is applicable to both aqueous and non-aqueous solutions. It overcomes all limitations of the Arrhenius acid-base theory.
A Bronsted-Lowry acid is a substance which donates a proton (H^+) to the other substance. A Bronsted-Lowry base is a substance which accepts a proton (H^+) from the other substance.
Thus, the Bronsted-Lowry theory is based on the concept of the transfer of protons from one substance to the other.

Example:
When hydrogen chloride gas (HCl) is dissolved in water, hydrogen chloride donates a proton to water.
$HCl (g) + H_2O (l) \rightarrow H_3O^+ (aq) + Cl^- (aq)$
Thus, hydrogen chloride which donates a proton is a Bronsted-Lowry acid and water which accepts a proton is a Bronsted-Lowry base.
When ammonia gas is dissolved in water, water donates a proton to ammonia.
$NH_3 (g) + H_2O (l) \rightarrow NH_4^+ (aq) + OH^- (aq)$
Thus, water which donates a proton is a Bronsted-Lowry acid and ammonia which accepts a proton is a Bronsted-Lowry base.

Question 4.A.3:

Discuss methods to measure pH of aqueous solution.

Solution :

The pH of an aqueous solution is measured by the following methods:

1. Measurement of approximate pH range:
Red and blue litmus papers are used to determine if the given solution is acidic, basic or neutral. A blue litmus paper turns red when dipped in a solution which is acidic; it remains blue if the solution is neutral.Similarly, a red litmus paper turns blue when dipped in a solution which is basic; it remains red if the solution is neutral.

2. Measurement of approximate pH values:
There are two methods which help to measure the approximate pH values of a given aqueous solution – pH paper and universal indicator.
Use of pH paper: A pH paper is a technically manufactured yellow-coloured paper which changes to different colours depending on the solution being tested. When pH paper is dipped in an aqueous solution, it changes its colour. The approximate pH of the solution can then be determined by matching the colour of the solution with the reference chart provided with the pH paper.
Use of universal indicator: Universal indicators are in the solution form and are added to the solution whose pH is to be determined. One to two drops of universal indicator are added to the solution, and the change in colour is noted. The colour change is compared with the chart on the bottle, and the approximate pH of the given solution is determined.

3. Measurement of exact pH values:
An instrument known as the pH meter is used for exact measurement of pH values. The pH meter has two electrodes which are dipped in the solution whose pH is to be measured. The electrodes are attached through a circuit to an indicating box where the pH of the solution is read.

Question 4(B):

Solve the following problems :

Question 4.B.1:

$[OH^-]$ in aqueous solution A is = 4.3×10^{-4} M and $[H_3O^+]$ in aqueous solution of B is 7.3×10^{-10} M, pH of which solution will be less ? Which solution will be more basic ?

Solution :

In aqueous solution A, $[OH^-] = 4.3 \times 10^{-4}$ M
By definition,
$pOH = -\log_{10}[OH^-]$
$= -\log_{10}(4.3 \times 10^{-4})$
$= -\log_{10}(4.3) - \log_{10}(10^{-4})$
$= -\log_{10}(4.3) + 4\log_{10}10$
$= -0.6334 + 4$
$= 3.3665$

Now,

pH + pOH = 14

pH = 14 – pOH

= 14 – 3.3665

= 10.6335

10.63

Thus, the pH of the aqueous solution A is 10.63.

In aqueous solution B, $[H_3O^+] = 7.3 \times 10^{10}$ M

By definition,

$pH = -\log_{10}[H_3O^+]$

$= -\log_{10}(7.3 \times 10^{-10})$

$= -\log_{10}(7.3) – \log_{10}(10^{-10})$

$= -\log_{10}(7.3) +10 \log_{10}10$

= -0.8633 + 10

= 9.1367

9.14

The pH of the aqueous solution B is 9.14.

Thus, we see that the pH of solution B is less and solution A is more basic.

Question 4.B.2:

Calculate the concentration of $OH^–$ in aqueous solution having pH value 9.3.

Solution :

The pH of the aqueous solution is 9.3

pH + pOH = 14

pOH = 14 – pH

pOH = 14 – 9.3

pOH = 4.7

By definition,

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

$4.7 = -\log_{10}[OH^-]$

$\therefore \log_{10}[OH^-] = -4.7$

$\therefore \log_{10}[OH^-] = 5 - 4.7 - 5$

$\therefore \log_{10}[OH^-] = -5 + 0.3$

$\therefore \log_{10}[OH^-] = \bar{5} + 0.3$

$\therefore [OH^-] = \text{antilog}(\bar{5}.3)$

$\therefore [OH^-] = 1.995 \times 10^{-5}$ M

Thus, the concentration of $OH^–$ in aqueous solutions with pH value 9.3 is 1.995×10^{-5} M.

Question 4.B.3:

8 gram NaOH is dissolved in water and the aqueous solution is made to 5 litres. Find pH of this solution.

Solution :

To calculate the pH of NaOH, we should know its H_3O^+ concentration.

Molar mass of NaOH is 40 gram mole^{-}

$$\text{Moles of KOH} = \frac{\text{Weight of NaOH}}{\text{Molecular mass of NaOH}}$$

$$= \frac{8 \text{ gram}}{40 \text{ gram/mole}}$$

$= 0.2$ mole

Concentration of NaOH in 5 litres of water

$$= \frac{\text{Moles of NaOH}}{\text{Volume of Solution (Litre)}}$$

$$= \frac{0.2}{5}$$

$= 0.04$ mole/litre

NaOH is a strong base and it ionises completely.

NaOH (aq) + H_2O (l) → Na^+ + OH^-

The given solution is 0.04 M.

Thus, $[OH^-] = 0.04M$

$= 4 \times 10^{-2}M$

Now,

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

$= -\log_{10}(4 \times 10^{-2})$

$= -\log_{10}4 - \log_{10}10^{-2}$

$= -\log_{10}4 + 2\log_{10}10$

$= -0.602 + 2$

$= 1.398$

Now,

pH + pOH = 14

pH = 14 – pOH

= 14 – 1.398

= 12.602

12.6

Thus, the pH of the aqueous solution of NaOH is 12.6.

Question 5(A):

Answer the following questions in detail :

Question 5.A.1:

With reference to pH scale,

Solution :

i. Formula of pH and pOH are

$pH = -\log_{10}[H_3O^+]$

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

ii. Mention pH and concentration of H_3O^+ or OH^- in acidic, basic and neutral aqueous solutions.

	pH	Concentration of H_3O^+	Concentration of OH^-
Acidic solution	0-6	10^0-10^{-6} M	–

Basic solution	8-14	–	10^{-6}-10^{0} M
Neutral solution	7	10^{-7} M	10^{-7} M

iii. Limitations of the pH scale are

- It is applicable to aqueous solutions only.
- It is applicable to aqueous solutions with concentration of hydronium ions less than 1 M. Hence, the pH scale is from 0 to 14.

Question 5.A.2:

Explain importance of pH in everyday life.

Solution :

pH is important in everyday life:

1. Existence of living beings
   Organisms require a specific pH for their ideal growth and development. In the human body, all the physiological reactions take place in the pH of 7-7.8. The existence of aquatic plants, animals and microbes is at risk when acid rains mix with natural water bodies.
2. Digestion of food
   The pH is important at different levels in the digestive system of the human body for the proper digestion of food. In the stomach, hydrochloric acid is secreted as food enters the stomach. It turns the pH of the stomach between 1 and 3. This pH is important for the activation of the enzyme pepsin, which helps in the digestion of protein in food.
3. Importance in soil
   The pH of soil is a critical factor in the growth of crops and other plants. pH 6.5 to 7.3 is ideal for the growth of shrubs and crops. If the pH of soil is less than 6.5, lime is added to it to neutralise its acidity. Similarly, if the pH of soil is more than 7.3, gypsum is added to neutralise its basicity.
4. Stopping tooth decay
   Teeth decay when the pH of the mouth falls below 5.5. Bacteria present in food and the mouth breakdown food particles left in the mouth and form acids. This results in lowering the pH of the mouth. Tooth powder and tooth paste which are used for cleaning of teeth are basic in nature. They neutralise the acid produced in the mouth and protect the teeth from decay.
5. Remedy for acidic effect of honeybee bite
   When a honeybee bites, it inserts an acidic substance Melittin in our body. To get relief from it, an aqueous solution of a basic substance such as baking soda is applied around the place of bite. It helps to neutralise the acidic poison.

Question 5.A.3:

‘The aqueous solution of the salt produced by neutralisation of weak acid and strong base possesses basic nature, while aqueous solution of salt produced by neutralisation of weak base and strong acid possesses acidic nature” – Explain.

Solution :

Neutralisation reaction is the reaction of an acid and base to produce salt and water.
The salt which is produced by the neutralisation of a weak acid and a strong base hydrolyses in water and produces OH^- ions; thus, it is basic in nature.
Example:
CH_3COOH (aq) + NaOH (aq) → CH_3COONa (aq) + H_2O (l)
(weak acid) (strong base) (salt) (water)

$$CH_3COONa\ (aq) \xrightarrow{\text{Ionisation}} CH_3COO^-\ (aq) + Na^+$$

$$CH_3COO^-\ (aq) + H_2O\ (l) \xrightarrow{\text{Hydrolysis}} CH_3COOH\ (aq) + OH^-\ (aq)$$

The salt which is produced by the neutralisation of a strong acid and a weak base hydrolyses in water and produces H_3O^+ ions; thus, it is basic in nature.
Example:
HNO_3 (aq) + NH_4OH (aq) → NH_4NO_3 (aq) + H_2O (l)
(strong acid) (weak base) (salt) (water)

$$NH_4NO_3\ (aq) \xrightarrow{\text{Ionisation}} NH_4^+\ (aq) + NO_3^-\ (aq)$$

$$NH_4^+\ (aq) + H_2O\ (l) \xrightarrow{\text{Hydrolysis}} NH_3\ (aq) + H_3O^+\ (aq)$$

Question 5(B):

Solve the following problems :

Question 5.B.1:

pH of aqueous solution of potassium hydroxide at 298 K temperature is 11.65. The initial volume of this solution is made 6 times by addition of water. What will be the pH of the diluted solution ?

Solution :

For aqueous solution with pH 11.65,
pH + pOH =14
pOH = 14 – pH
pOH = 14 – 11.65
pOH = 2.35
pOH = $-\log_{10}[OH^-]$
2.35 = $-\log_{10}[OH^-]$
$\therefore \log_{10}[OH^- = -2.35$
$\therefore \log_{10}[OH^-] = 3 - 2.35 - 3$
$\therefore \log_{10}[OH^-] = -3 + 0.65$
$\therefore \log_{10}[OH^-] = \bar{3} + 0.65$
$\therefore [OH^-] = \text{antilog}(\bar{3}.65)$
$\therefore [OH^-] = 4.467 \times 10^{-3}$ M

When water is added to an aqueous solution of potassium hydroxide, the initial volume is increased by 6 times, so the concentration of the solution is decreased by 6 times.

$$\therefore [OH^-] \text{ in diluted solution} = \frac{4.467 \times 10^{-3}}{6} = 7.445 \times 10^{-4}$$

Now,
pOH = $-\log_{10}[OH^-]$

$=-\log_{10}(7.445 \times 10^{-4})$

$=-\log_{10}7.445 - \log_{10}10^{-4}$

$=-\log_{10}7.445 + 4\log_{10}10$

$= -0.8719 + 4$

$= 3.1281$

Now, pH + pOH = 14

pH = 14 – 3.1281

pH = 10.87

Thus, the pH of diluted potassium hydroxide is 10.87.

Question 5.B.2:

What will be the change in the value of pH if the concentration of aqueous solution of $HN0_3$ is increased to 0.05 M from 0.03 M ?

Solution :

The pH of the HNO_3 solution with concentration 0.05 M (5×10^2M) is

$pH = -\log_{10}[H_3O^+]$

$=-\log_{10}(5 \times 10^{-2})$

$=-\log_{10}5 - \log_{10}10^{-2}$

$=-\log_{10}5 + 2\log_{10}10$

$= -0.699 + 2$

$= 1.301$

The pH of the HNO_3 solution with concentration 0.03 M (3×10^2M) is

$pH = -\log_{10}[H_3O^+]$

$=-\log_{10}(3 \times 10^{-2})$

$=-\log_{10}3 - \log_{10}10^{-2}$

$=-\log_{10}3 + 2\log_{10}10$

$= -0.477 + 2$

$= 1.52$

When the concentration of the aqueous solution of HNO_3 is increased to 0.05 M from 0.03 M, the pH increases from 1.3 to 1.52.

Hence, the change in the pH = 1.52 – 1.30 = 0.22.