VIVEKANANDA COLLEGE
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NAAC ACCREDITED ‘A’ GRADE

Topic: Ioinic Equilibrium

Course Title: General Physical Chemistry

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Name of the Department: Biochemistry

Ionic Equilibrium:

⇒ Standard solution:

A solution containing a known weight of a solute in a known/definite volume of a solution.

In a standard solution, the concentration of the solute is precisely known.

It is usually made by dissolving a primary standard solute in a solvent such that the net solution has a definite volume & a definite mass of the standard solute.

→ Primary standard:

A solute which can be weighed accurately. It has the following properties:

a) Highly Stable (no reactivity or decomposition)

b) Highly pure (no. of moles can be obtained from direct weighing)

c) High Equivalent Weight (minimal weighing error)

d) Non-hygroscopic (Does not absorb moisture)

e) Cheap & readily available

eg. Sodium oxalate ($Na_2C_2O_4$), $K_2Cr_2O_7$, KIO_3, $KBrO_3$, Oxalic acid, etc.

⇒ Molar solution: (1 g mole)

A solution containing 1 g molecular weight of a solute in 1 litre of solution is called a 1 M (molar) solution.

e.g. NaCl has a molar mass of $23 + 35.5 = 58.5$

∴ g-molecular weight of NaCl $= 58.5\,g$ = weight of 1 mole in grams

1(M) NaCl = 58.5 g NaCl in 1 L solution

⇒ Normal solution:

A solution containing 1 g-equivalent of a solute in 1 L of a solution is called a 1 N (normal) solution.

e.g. NaCl has a g-eq weight of

$$\frac{23}{1} + \frac{35.5}{1} = 58.5\,g$$

∴ 1 (N) NaCl = 58.5 g NaCl in 1 L solution.

[For NaCl, (M) & (N) are identical since its valency of its ions is 1. However, for H_2SO_4, they are different. Try to calculate the amount of H_2SO_4 in a 1 (M) and 1 (N) solution and compare]

⇒ Molal solution:

1 (m) (molal) solution refers to a solution containing 1 mole of a solute in 1 kg of solvent or 1000 g of solvent.

i.e. mol/g or mol/kg

$$\text{molality} = \frac{\text{no. of moles}}{\text{amount of solvent (g or kg)}}$$

Thus it is only dependent on mass and is independent of T and p. Thus it is a better scale than normality (N) or molarity (M).

⇒ Formal solution:

The number of formula mass of a solute dissolved in a litre of solution gives a formal solution.

e.g. for ionic compounds, which donot have definite molecular entities, their simplest formula is used to represent them.

Like $Na_{12}Cl_{12}$ / Na_6Cl_6 is represented as NaCl.

∴ formula mass of NaCl = mass of 1 simple formula

= 23 + 35.5 = 58.5

∴ g-formula mass = 58.5 g.

∴ if a solution has 58.5 g NaCl in 1 L, it may be called a 1 formal solution.

⇒ Percent solution:

It is the amount or volume of a solute per **100 ml** of solution.

⇒ It can be (mass/volume)% or

(m/v)%

e.g. 1 g solute in 100 ml solution = 1% (m/v)

Unit: g/ml or $g\,cc^{-1}$.

⇒ It can also be (volume/volume)% or (v/v)%

e.g. 1 ml ethanol in 100 ml solution = 1% (v/v)

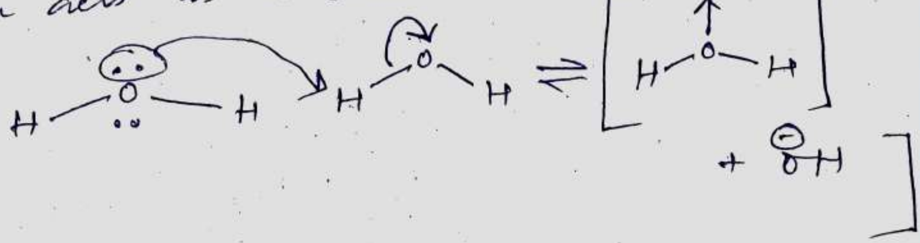
Unit: Unitless, dimensionless

⇒ Ionic product of Water :

Considering self-ionization of water molecules, we can write an equilibrium as

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

[In this reaction, a proton (H^+) is donated from 1 water molecule to another. Thus one molecule acts as an acid and the other acts as a base.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The dissociation/eq. constant

$$K = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

Since water practically is unionized, its concentration remains constant.

$\therefore$ $[H_2O]$ = constant

$[H_2O]^2$ = (constant)^2 = constant

$\therefore$ $K \times [H_2O]^2$ = constant = K_w

$$\therefore \boxed{K_w = [H_3O^+][OH^-]}$$

K_w is the Ionic product of water.

$K_w = 1.008 \times 10^{-14}$ mol^2 L^{-2} at 25 °C.

For pure water, $[H_3O^+] = [OH^-]$

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

$\therefore$ $[H_3O^+] = [OH^-] = 10^{-7}$ mol L^{-1} at 25 °C.

⇒ Unit of K_w : $mol^2 L^{-2}$
or M^2 / $(molar)^2$

⇒ The concentration of protons $[H^+]$ or hydronium ions $[H_3O^+]$ determine the acidity / basicity of a solution.

→ For a neutral solution:

$$[H_3O^+] = [OH^-] = \sqrt{K_w}$$

→ For an acidic solution:

$$[H_3O^+] > [OH^-]$$

$$\Rightarrow [H_3O^+] > \sqrt{K_w}$$

→ For a basic solution:

$$[H_3O^+] < [OH^-]$$

$$\Rightarrow [H_3O^+] < \sqrt{K_w}$$

Thus at 25°C / 298 K :

Neutral solution : $[H^+] = 10^{-7}$ M
Acidic " : $[H^+] > 10^{-7}$ M
Basic " : $[H^+] < 10^{-7}$ M

⇒ pH scale:

$$pH = -\log[H^+] \text{ or } -\log[H_3O^+]$$

Generally, unit of $[H^+]$ is M (molar). pH is a number which is dimensionless and just expresses the acidity / basicity of a solution. So it is more appropriate to write

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

i.e. dividing the concentration term with the corresponding concentration to make it dimensionless.

⇒ At 25 °C, $pH = -\log[10^{-7}] = 7$.

⇒ $pOH = -\log[OH^-]$

⇒ At 25 °C, $pOH = -\log[10^{-7}] = 7$.

Now, $K_w = [H^+][OH^-]$

∴ $\log K_w = \log[H^+] + \log[OH^-]$

⇒ $-\log K_w = -\log[H^+] - \log[OH^-]$

⇒ $\boxed{pK_w = pH + pOH}$

$$pK_w = -\log\left\{\frac{K_w}{M^2}\right\}$$

At 25 °C, $pK_w = -\log[10^{-14}] = 14$