Atomic Energy Education Society, Mumbai

Class: XI **Topic: Chaper-7: EQUILIBRIUM** **Subject: Chemistry**

HANDOUTS OF MODULE 6 OF 6

pH Scale

- The pH scale is used to express concentration of H_3O^+ in an aqueous solution.
- For expressing the H_3O^+ ion concentration a logarithmic scale was devised by P.L.Sorensen.
- The pH of a solution is defined as the negative logarithm (log) of the hydronium ion (hydrogen ion) concentration.
- $pH = - \log [H_3O^+]$
- $[H_3O^+] = 1 \times 10^{-pH}$
- In similar way, pOH is the negative logarithm of the hydroxide ion concentration. $pOH = - \log [OH^-]$

- In pure water,
  - $K_w = [H_3O^+] [OH^-] = 1 \times 10^{-14} M^2$
  - $[H_3O^+] = [OH^-] = 1 \times 10^{-7} M$
  - $pH = pOH = - \log (1 \times 10^{-7})$
  - $pH = pOH = 7$
  - Therefore $pH + pOH = pK_w = 14$
- Concept is:
  - At 25 ^oC,
  - Acidic solution : pH < 7.0 ; pOH > 7.0
  - Basic solution : pH > 7.0 ; pOH < 7.0
  - Neutral solution : pH = 7.0 ; pOH = 7.0

Buffer Solution

- A solution that maintains its pH when a small amount of a strong acid or a strong base is added to it. It is also defined as a solution which contains weak acid or weak base with salt that has its conjugate pair.
- Two types of buffer solution:
- Acidic buffer solution
- Basic buffer solution
- **ACIDIC BUFFER SOLUTION**: Has pH < 7: An acidic buffer solution is made up of a weak acid and its salt (containing its conjugate base)
- Example: CH_3COOH / CH_3COONa, H_2CO_3 / $NaHCO_3$, H_3PO_4 / NaH_2PO_4, HCOOH / HCOONa
- equation, popularly known as the Henderson equation
- pH of acidic buffer = pKa + log ([salt]/[acid])
- **Basic Buffer Solution**: Has pH > 7: Basic buffer solution is made up of a weak base and its salt (containing its conjugate)
- The equation pOH of basic buffer = pKb + log ([salt]/[base]) is the Henderson-Hasselbalch
- Example: NH_4OH / NH_4Cl, NH_3 / NH_4Cl, NH_3 / $(NH_4)_2CO_3$
- **Buffer Capacity**: The number of millimoles of acid or base to be added to a litre of buffer solution to change the pH by one unit is the Buffer capacity of the buffer.
  Formula: Buffer Capacity = millimoles of acid or base / (ΔpH)

Solubility:

- The solubility of a substance is the amount of that substance that will dissolve in a given amount of solvent. "Solubility" may be considered to be equilibrium; the equilibrium is between solid and ions in solution. Any ionic solid is 100% ionized in aqueous solution; once it actually dissolves.

Factors Affecting Solubility:

- Temperature: Solubility generally increases with temperature.
- Common ion effect: Common ions reduce solubility
- pH of solution: pH affects the solubility of ionic compounds.
- Formation of complex ion: The formation of complex ion increases solubility
- Some salts are soluble but most are insoluble or slightly soluble in water.
- A saturated solution is a solution that contains the maximum amount of solute that can dissolve in a solvent.
- The solubility of a salt is the amount of solid that dissolved in a known value of saturated solution.
- The unit of solubility used may be g L^{-1} or mol L^{-1}
- **Molar solubility** is the maximum number of moles of solute that dissolves in a certain quantity of solvent at a specific temperature.

Different types of solution

- Unsaturated solution: More solute can be dissolved in it.
- Saturated solution: No more solute can be dissolved in it. Any more of solute you add will not dissolve. It will precipitate out.
- Super saturated solution: Has more solute than can be dissolved in it. The solute precipitates out.

Solubility product constant.

- K_{sp} is the product of the molar concentrations of the ions involved in the equilibrium, each raised to the power of its stoichiometric coefficient in the equilibrium equation.
- K_{sp} is called the solubility product constant.
- The degree of solubility of a salt is shown by the value of Ksp.
- Soluble salt such NaCl and KNO_3 has an extremely high value of K_{sp} .
- The smaller the value of K_{sp} the less soluble the compound in water.
- Temperature ↑ , solubility ↑, K_{sp} ↑
- **$Ksp = x^x y^y (S)^{x+y}$**
  - X= No of cation =1, y=No of Anion=2, S is solubility of salt
- If we mix a solution containing M^+ ions with one containing A^- ions, the ionic product, Q_{sp} is given by : $Q_{sp} = [M^+] [A^-]$
- Qsp < Ksp;Solution is not saturated. Solid will dissolve and no precipitate formed.
- Qsp = Ksp; Saturated solution formed. System is in equilibrium.
- Qsp > Ksp; Solution is supersaturated; Ions will form precipitate until the ionic concentration product of the system equals the Ksp (until the system reaches equilibrium).