$$\text{Buffer capacity} = \frac{\text{No. of moles of acids or base added per litre of buffer}}{\text{Change in pH}}$$

The range of pH over which the buffer solutions remain effective is called buffer range.

Buffer	**Buffer range in pH**
Acidic	$pK_a \pm 1$
Basic	$(pK_w - pK_b) \pm 1$

Example : pK_a of acetic acid is 4.74, then the range of buffer solution of acetic acid is

pH = 4.74 ± 1, *i.e.*, 3.74 to 5.74

Illustration 9

Calculate the amount of NH_3 and NH_4Cl required to prepare a buffer solution of pH = 9.0 when total concentration of buffering reagents is 0.6 mol litre^{-1}. (pK_b for NH_3 = 4.7, log 2 = 0.30).

Soln.: Applying, $pOH = pK_b + \log\frac{[\text{Salt}]}{[\text{Base}]}$

$pOH = 14 - pH = 14 - 9 = 5.0$

$5.0 = 4.7 + \log\frac{[\text{Salt}]}{[\text{Base}]}$, $\log\frac{[\text{Salt}]}{[\text{Base}]} = 5.0 - 4.7 = 0.3$

Taking antilog,

$\frac{[\text{Salt}]}{[\text{Base}]} = 2$, $[\text{Salt}] = 2 \times [\text{Base}]$

But it is given that [Salt] + [Base] = 0.6 M

On solving two above equations

[Base] = 0.2 mol litre^{-1}

[Salt] = 0.4 mol litre^{-1}

SOLUBILITY PRODUCT

- Solubility product of a sparingly soluble salt at a given temperature is defined as the product of molar concentration of its ions in a saturated solution, with each concentration term raised to the power equal to number of ions present in the chemical equation representing the equilibrium of dissociation of one molecule of the salt. It is represented as K_{sp}.

$$A_xB_y \rightleftharpoons xA^{y+} + yB^{x-}$$

Solubility product $(K_{sp}) = [A^{y+}]^x\,[B^{x-}]^y$

For example,

$$AgCl \rightleftharpoons Ag^+ + Cl^-\ ,\ K_{sp} = [Ag^+][Cl^-]$$

$$Al(OH)_3 \rightleftharpoons Al^{3+} + 3OH^-,\ K_{sp} = [Al^{3+}][OH^-]^3$$

Common Ion Effect

- Common ion effect is defined as the suppression of ionization of weak electrolytes by addition of strong electrolytes having an ion common to the weak electrolyte.

  For example weak base NH_4OH ionizes to a small extent

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

- When a strong electrolyte like NH_4Cl or NaOH is added to this solution, a common ion NH_4^+ and OH^- respectively is furnished so that the equilibrium is shifted to the right, according to Le Chatelier's principle and ionization of NH_4OH is suppressed.

$$NH_4OH \rightleftharpoons \boxed{NH_4^+} + OH^-$$

$$NH_4Cl \rightarrow \boxed{NH_4^+} + Cl^-$$

Common ion

Thus, degree of dissociation of an electrolyte decreases by common ion effect but dissociation constant of that electrolyte remains constant.

- **Applications of solubility product**
  - *Predicting precipitate formation:*
    if ionic product > K_{sp} precipitation takes place
    if ionic product < K_{sp} no precipitation takes place
  - *Predicting solubility of sparingly soluble salt:* Solubility of a sparingly soluble salt can be calculated from its solubility product at a given temperature.

$$XY \rightleftharpoons X^+ + Y^-$$

Let solubility is s s s

$K_{sp} = [X^+][Y^-] = (s)(s) = s^2$, $s = \sqrt{K_{sp}}$

Type of salt	**Solubility product**
AB type	$K_{sp} = s^2$
AB_2 or A_2B type	$K_{sp} = 4s^3$
AB_3 or A_3B	$K_{sp} = 27s^4$
AB_4 or A_4B	$K_{sp} = 256s^5$

where, s = solubility (mole/litre)

  - *Precipitation of common salt* : When HCl gas is passed through saturated solution of impure common salt, the concentration of Cl^- increases due to ionization of HCl.

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

    Thus ionic product $[Na^+]$ $[Cl^-]$ exceeds solubility product and pure NaCl precipitates out.
  - *Qualitative Analysis* : Various basic radicals are identified and separated by the principle of solubility product and common ion effect.

Illustration 10

Predict whether a precipitate will be formed or not on mixing equal volumes of 2×10^{-4} M $BaCl_2$ solution and 2×10^{-5} M Na_2SO_4 solution if solubility product of $BaSO_4$ is 1×10^{-10}.

Soln.: $BaSO_4 \rightleftharpoons Ba^{2+} + SO_4^{2-}$,

$BaCl_2 \rightleftharpoons Ba^{2+} + 2Cl^-$

$\therefore$ $Na_2SO_4 \rightleftharpoons 2Na^+ + SO_4^{2-}$

Since equal volumes of $BaCl_2$ and Na_2SO_4 are mixed, concentration of Ba^{2+} and SO_4^{2-} after mixing would be

$$[Ba^{2+}] = \frac{[BaCl_2]}{2} = \frac{2\times10^{-4}}{2} = 10^{-4}\,M$$

$$[SO_4^{2-}] = \frac{[Na_2SO_4]}{2} = \frac{2\times10^{-5}}{2} = 10^{-5}\,M$$

Ionic product of $BaSO_4$

$$= [Ba^{2+}][SO_4^{2-}] = [10^{-4}][10^{-5}] = 10^{-9}\,M$$

Ionic product (10^{-9} M) > K_{sp} (1×10^{-10})

Hence, precipitation will take place.