Salts in water and acid-base titrations

$NH_4Cl(s) + H_2O(l) \rightleftharpoons NH_4^+(aq) + Cl^-(aq) \Rightarrow$ solution is acidic

weak acid (NH_3 is weak base)

very weak base (HCl is strong acid)

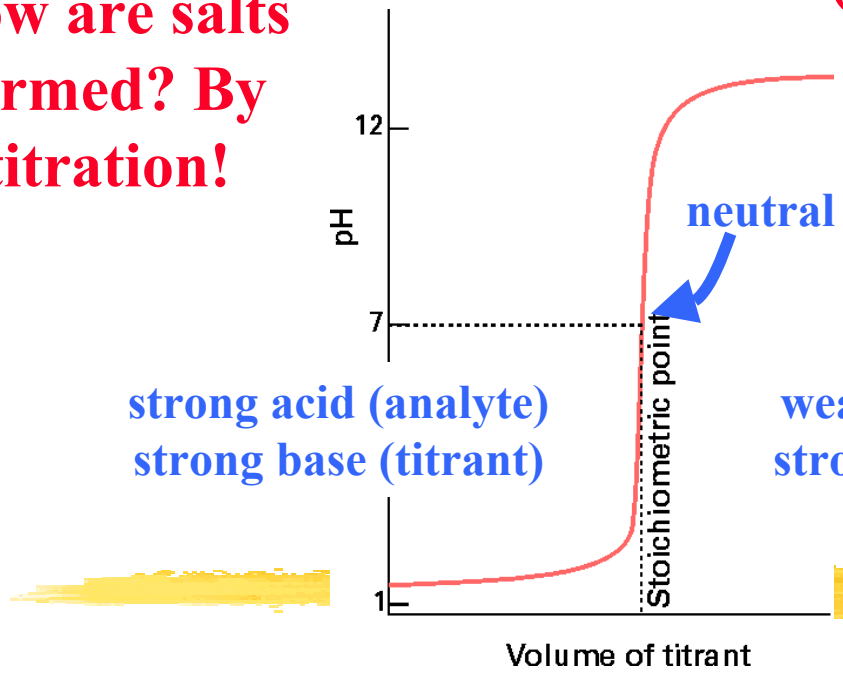
$pK_a + pK_b = pK_w = 14$

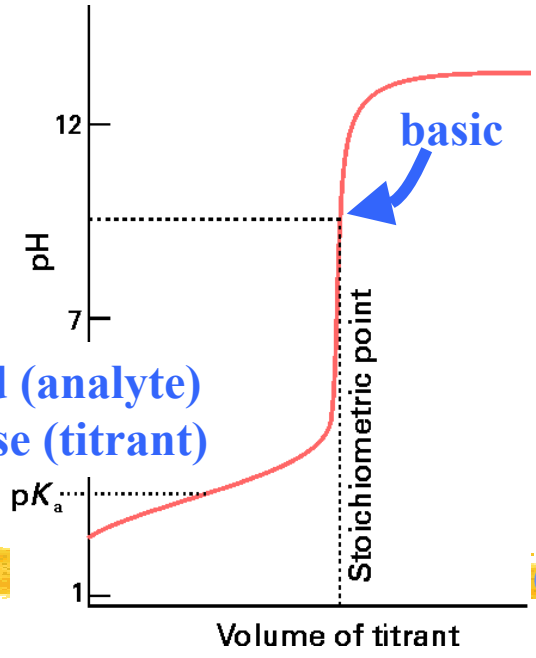
$CH_3COONa(s) + H_2O(l) \rightleftharpoons Na^+(aq) + CH_3COO^-(aq) \Rightarrow$ solution is basic

neutral ion

weak base (CH_3COOH is weak acid)

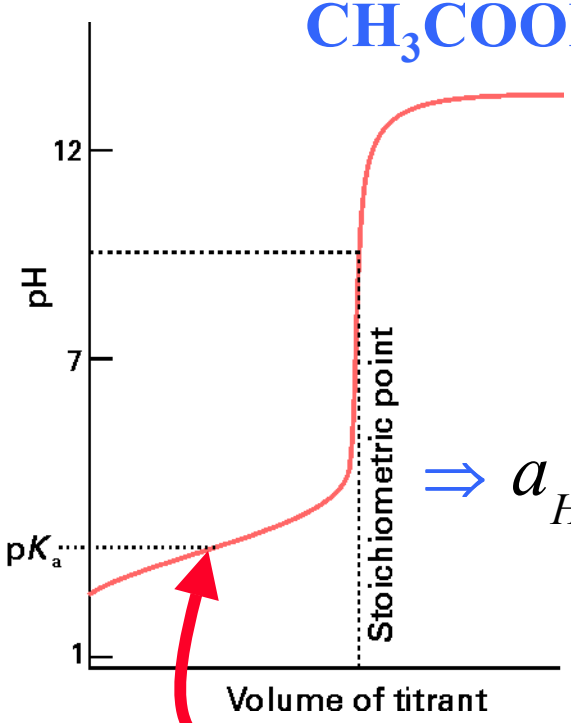
How are salts formed? By titration!





Acid-base titrations quantitatively

$$CH_3COOH(aq) + OH^-(aq) \rightleftharpoons CH_3COO^-(aq) + H_2O(aq)$$



$$K_a = \frac{a_{CH_3COO^-} \times a_{H_3O^+}}{a_{CH_3COOH}} \approx \frac{[base] \times a_{H_3O^+}}{[acid]}$$

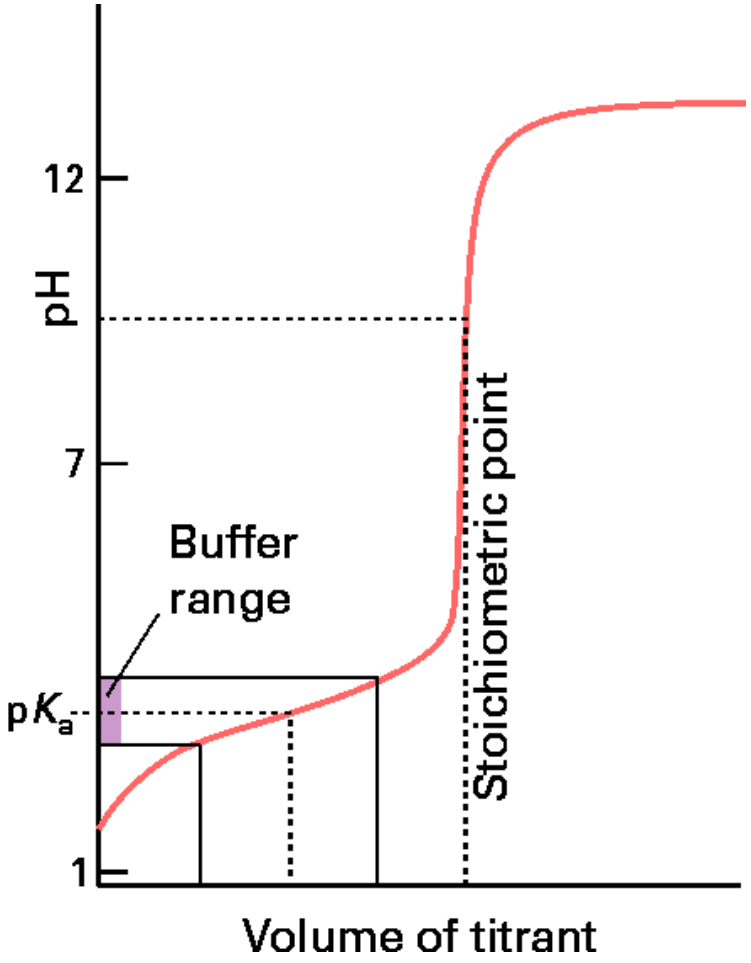
$$\Rightarrow a_{H_3O^+} \approx \frac{K_a \times [acid]}{[base]} \Rightarrow pH \approx pK_a - \log\frac{[acid]}{[base]}$$

Henderson-Hasselbalch equation

$$[acid] = [base] \Rightarrow pH \approx pK_a$$

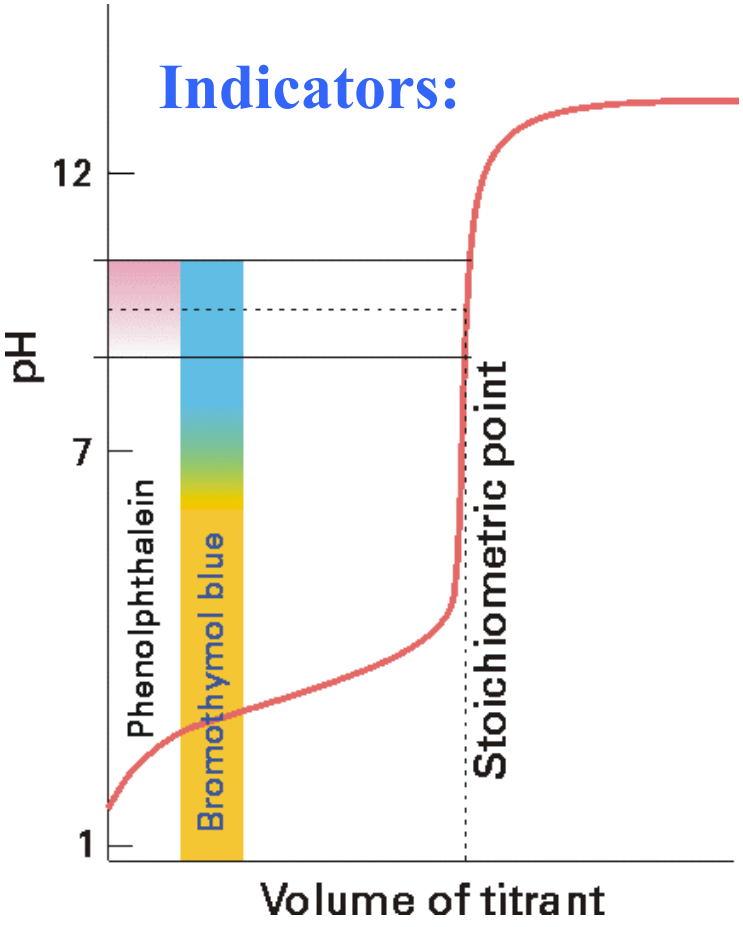
Determination of the pK_a of an acid

Buffers and indicators



acid buffer: stabilizes pH @ below pH 7 (around pK_a)

Indicators:



$$\log \frac{[acid]}{[base]} = pH - pK_{In}$$

Solubility equilibria

$Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2\ OH^-(aq)$

solubility constant $K_s = a_{Ca^{2+}} \times a_{OH^-}{}^2$

molar solubility $S = [Ca^{2+}] = \frac{1}{2}[OH^-]$

$\Rightarrow K_s = S \times (2S)^2 = 4S^3 \quad \Rightarrow S \approx (\frac{1}{4} K_s)^{1/3}$ (ignoring ion-ion interactions; a ≈ [ ])

$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$

$K_s = a_{Ag^+} \times a_{Cl^-} \approx [Ag^+][Cl^-] = 1.6 \times 10^{-10}\ M^2$

$S = [Ag^+] \approx \frac{K_s}{[Cl^-]}$

AgCl can be precipitated by addition of Cl^- as a common ion from, e.g., NaCl (Le Chatelier!)