Back titration :

Back titration is a titration done in reverse; instead of titrating the original sample, a known excess of standard reagent is added to the solution, and the excess is titrated. for example

$$Cl^- + Ag^+ \rightarrow AgCl\,(s)$$

$$Ag^+ + SCN^- \rightarrow AgSCN\,(s)$$

Acid And Base titration :

The acid reacts with the base in an aqueous medium , Upon reaching the end point, the number of moles of the acid have reacted with the necessary moles of the base The result is usually water and salt according to the following equation .

$$HCl + NaOH \rightarrow NaCl + H_2O$$

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

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

Dissociation constant in 25c:

$$K_w = 1.0 \times 10^{-14}$$

In Pure water $[H]^+ = [OH]^- = 1.0 \times 10^{-7}$

Acid-base equilibria in water

When adding the acid or base to the water, both of them will dissolve in the water. If the acid is strong, then it will be completely ionized ... as in the equation.

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

While a base like ammonia is weak, it is partially ionized

$$NH_3 + H_2O \Leftrightarrow NH_4^+ + OH^-$$

Stability of the base equilibrium

$$Kb = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

PH-Scale

$$pH = -\log [H^+]$$

$$poH = -\log [oH^-]$$

ex. Calculate pH ,poH for NaoH [5×10^{-2}] M
sol.

$[oH^-] = [5 \times 10^{-2}]$

$poH = -\log [oH^-]$

$poH = -\log [5 \times 10^{-2}]$

$poH = 1.3$

$PH + poH = 14$

$PH + 1.3 = 14$

$PH = 12.7$

Acid and Base strength

Strong acid and strong base

Strong acid		Strong base	
HCl	**Hydrochloric acid**	NaOH	Sodium hydroxide
H_2SO_4	Sulfuric Acid	$Ca(OH)_2$	Calcium hydroxide
HNO_3	Nitric acid	KOH	Potassium hydroxide

Weak acid and weak base

$$HA \Leftrightarrow H^+ + A^-$$

The above equation for weak acid thus a stable equilibrium of acid is :

$$Ka = \frac{[H^+]^2}{Ca}$$

Weak base equation

$$B + H_2O \Leftrightarrow HB + OH^-$$

stable equilibrium of base

$$K_b = \frac{[OH^-]^2}{Cb}$$

Ex. If the dissociation constant for acetic acid 1.8×10^{-5} calculate PH when acid Concentration 0.1 M ?

Sol.

$$Ka = \frac{[H^+]^2}{Ca}$$

$1.8\times10^{-5} = [H^+]^2 / 0.1$

$[H^+]^2 = 1.8 \times10^{-6}$

$[H^+] = 1.3 \times10^{-3}$

$pH = - \log [H^+]$

$pH = - \log 1.3 \times10^{-3}$

$pH = 2.88$

Weak acid		Weak base	
Hf	hydrofluoric acid	NH_3	Ammonia
CH_3COOH	Acetic acid	C_5H_5N	Pyridine