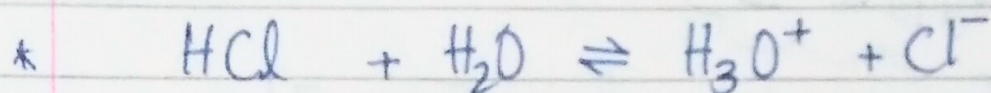
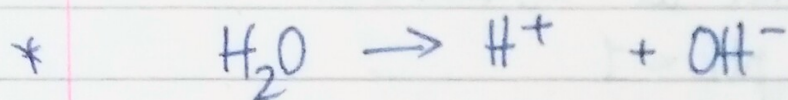


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Further aspects of equilibria



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{Cl}^-]}{[\text{HCl}]}$$



$$K = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$$

$$K_w = [\text{H}^+][\text{OH}^-] = [\text{H}^+]^2$$

↓
Ionic product of water.

$$K_w = 1 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-3} \text{ - pure water.}$$

$$\begin{aligned} K_w &= [\text{H}^+]^2 \\ [\text{H}^+] &= \sqrt{K_w} \\ &= \sqrt{1 \times 10^{-14}} \\ &= 1 \times 10^{-7} \text{ mol dm}^{-3} \text{ - pure water.} \end{aligned}$$

* pH: power of hydrogen.

$$* \text{ pH} = -\log_{10} [\text{H}^+]$$

$$* \text{ Pure water: } \text{pH} = -\log_{10} [1 \times 10^{-7}] \\ = 7$$

* Strong acids: HCl at 1.0 mol dm^{-3}
 $\text{pH} = 1$

$$\text{H}_2\text{SO}_4 \text{ at } 1.0 \text{ mol dm}^{-3} \\ \text{pH} = 1 \times 2 = 2$$

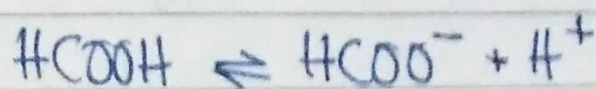
* Weak acids: $\text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{COO}^- + \text{H}^+$
 $\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+$

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]}$$

$$K_a = \frac{[\text{H}^+]^2}{[\text{CH}_3\text{COOH}]}$$

Calculate the pH of methanoic acid of acid dissociation constant 1.59×10^{-4} and concentration is 0.01 mol dm^{-3}

$$K_a = \frac{[H^+]^2}{[HCOOH]}$$



$$[H^+] = \sqrt{1.59 \times 10^{-4} \times 0.01}$$
$$= 1.26 \times 10^{-3}$$

$$pH = -\log_{10}(1.26 \times 10^{-3})$$
$$= \underline{2.9}$$

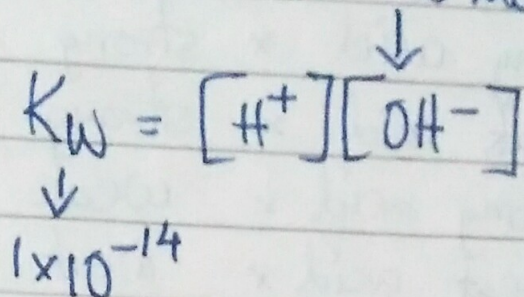
Calculate the pH of propanoic acid of conc. 0.05 mol dm^{-3} and acid dissociation constant 1.3×10^{-5}

$$K_a = \frac{[H^+]^2}{[CH_3CH_2COOH]}$$

$$[H^+] = \sqrt{1.3 \times 10^{-5} \times 0.05}$$
$$= 8.06 \times 10^{-4}$$

$$pH = -\log_{10}(8.06 \times 10^{-4})$$
$$= \underline{3.09}$$

* Strong base: NaOH at 1.0 mol dm^{-3}



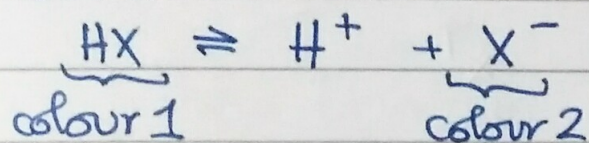
* $pH = -\log[H^+]$

$$pK_a = -\log[K_a]$$

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* Indicators

- changes colour due to pH of medium.
- are weak acids
- their dissociation is in equilibrium in solution.



→ When H^+ is high equilibrium shifts to HX side so colour 1 is shown.

→ When OH^- is added, it reacts with H^+ to give H_2O , equilibrium shifts to H^+ side colour 2 is visibly shown then.

→ Titration.

Strong acid x strong base

Weak acid x strong base

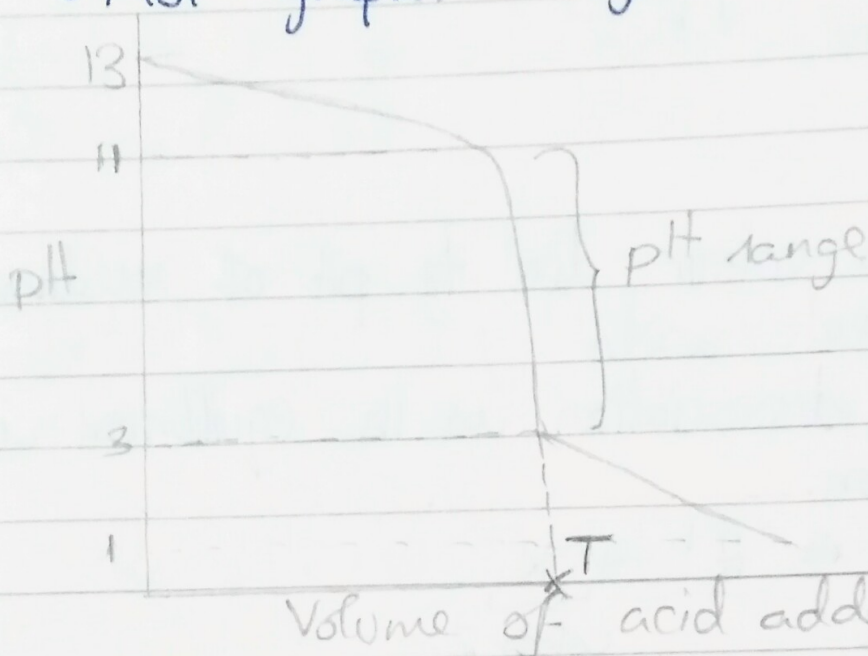
Strong acid x weak base

Weak acid x weak base.

→ pH curve. -

• Uses pH meter for every cm^3 of acid added in titration.

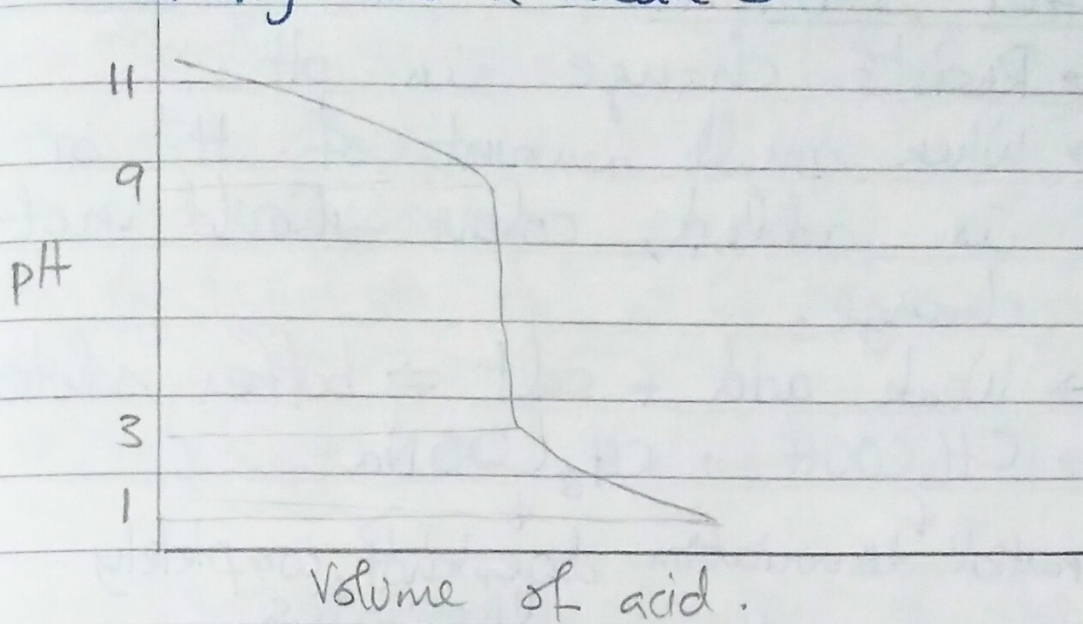
• Plot graph. - Strong acid x strong base.



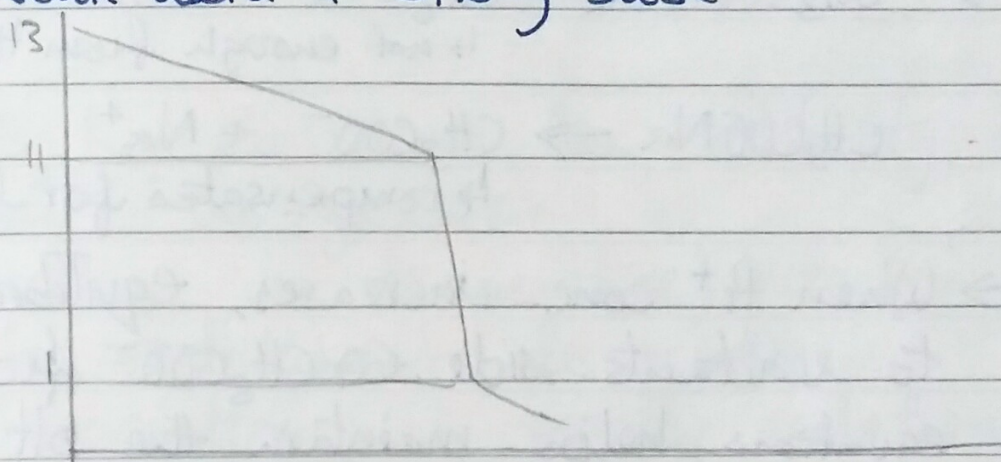
• Pick an indicator that has a similar pH range.

• $T \text{ cm}^3$ will be the titre value.

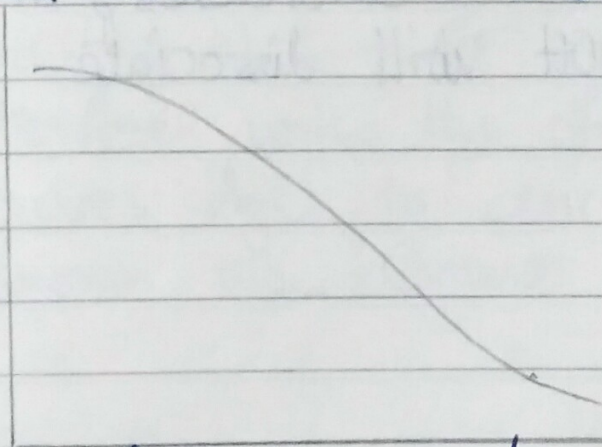
• Strong acid x Weak base.



• Weak acid x Strong base.



• Weak acid x weak base.



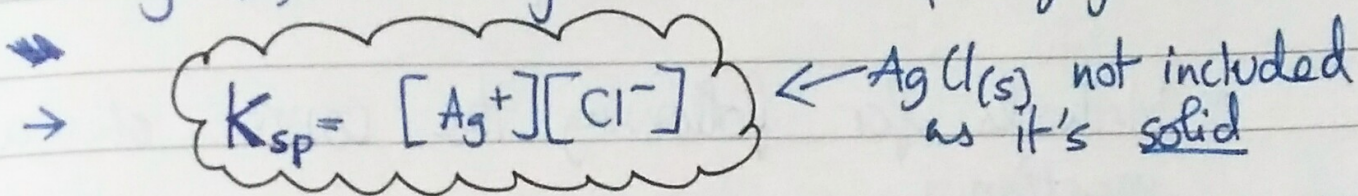
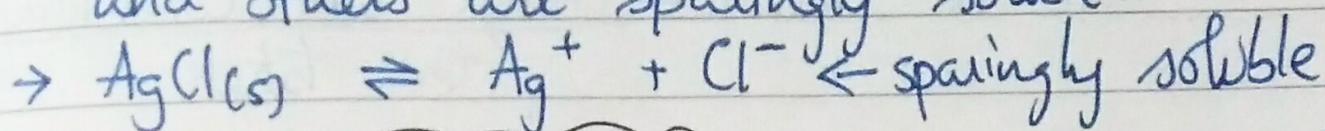
→ very difficult to get a pH range for this

* Buffer solution.

- Resists change in pH
- When small amount of H^+ or OH^- is added, colour should not change.
- Weak acid + salt \Rightarrow buffer solution.
- $CH_3COOH + CH_3COONa$
partial \downarrow dissociation. completely dissociates
- $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$
 \hookrightarrow not enough from this equation.
- $CH_3COONa \rightarrow CH_3COO^- + Na^+$
 \hookrightarrow compensates for loss
- When H^+ conc. increases, equilibrium shifts to reactants side. $\leftarrow CH_3COO^-$ from 2nd equation helps maintain the pH
- When OH^- conc. increases, H^+ will decrease so CH_3COOH will dissociate

* Solubility

- Dissolve substance in 100cm^3 of solvent
- Some substances are highly soluble and others are sparingly soluble.



K_{sp} : equilibrium constant of the sparingly soluble salt.

- As K_{sp} increases, the solubility of the salt is greater
- Increase in temperature, increases K_{sp} for endothermic dissociation.
- Dissociation is usually endothermic.
- Adding another salt.
 - $\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^- \rightarrow$ complete dissociation.
 - Interferes with the other equilibrium, and causes AgCl to increase
 - Known as "Common ion effect"