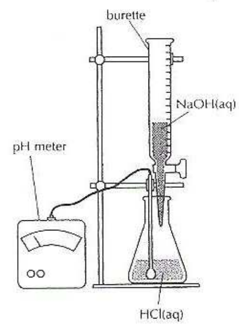
Titration curves and indicators

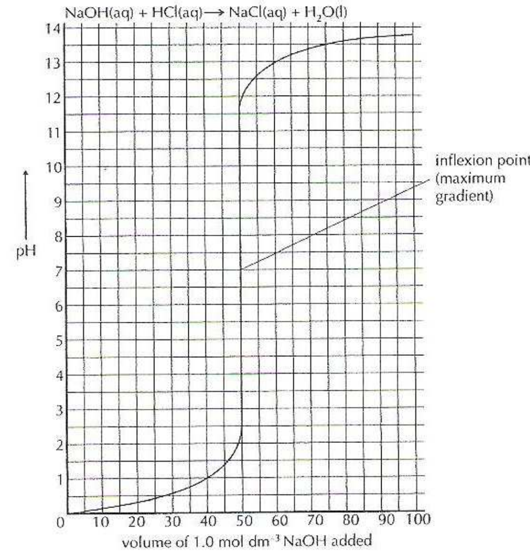
STRONG ACID – STRONG BASE TITRATION

The change in pH during an acid–base titration can be followed using a pH meter. Consider starting with 50 cm^3 of 1.0 mol dm^{-3} hydrochloric acid. Since [H^+(aq)] = 1.0 mol dm^{-3} the initial pH will be 0. After 49 cm^3 of 1.0 mol dm^{-3} NaOH have been added there will be 1.0 cm^3 of the original 1.0 mol dm^{-3} hydrochloric acid left in 99 cm^3 of solution. At this point [H^+(aq)] $\approx 1.0 \times 10^{-2}$ mol dm^{-3} so the pH = 2.

When 50 cm^3 of the NaOH solution has been added the solution will be neutral and the pH will be 7. This is indicated by the point of inflexion, which is known as the equivalence point. It can be seen that there is a very large change in pH around the equivalence point. Almost all of the common acid–base indicators change colour (reach their end point) within this pH region. This means that it does not matter which indicator is used.



This curve shows what happens when 1.0 mol dm^{-3} sodium hydroxide is added to 50 cm^3 of 1.0 mol dm^{-3} hydrochloric acid



WEAK ACID – STRONG BASE TITRATION

Consider titrating 50.0 cm^3 of 1.0 mol dm^{-3} CH_3COOH with 1.0 mol dm^{-3} NaOH.

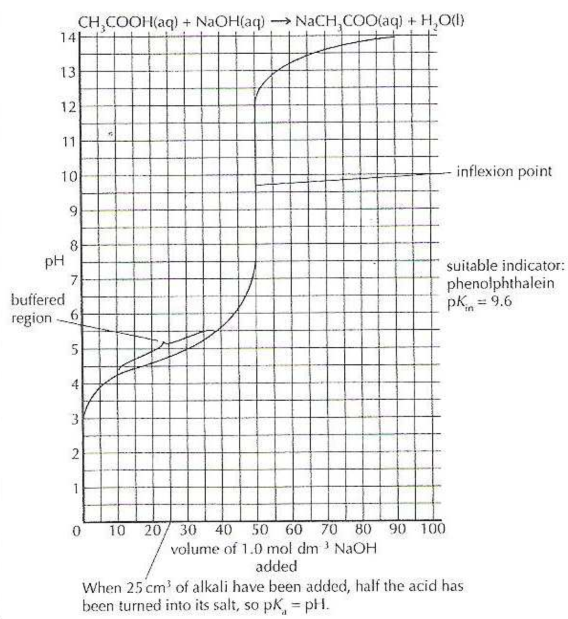
$K_a = 1.8 \times 10^{-5}$. Making the usual assumptions the initial $[H^+] = \sqrt{K_a \times [CH_3COOH]}$ and pH = 2.37.

When 49.0 cm^3 of the 1.0 mol dm^{-3} NaOH has been added $[CH_3COO^-] \approx 0.50$ mol dm^{-3} and $[CH_3COOH] \approx 1.0 \times 10^{-2}$ mol dm^{-3}.

$$[H^+] = \frac{K_a \times [CH_3COOH]}{[CH_3COO^-]} \approx \frac{1.8 \times 10^{-5} \times 1 \times 10^{-2}}{0.50}$$

$$= 3.6 \times 10^{-7} \text{ mol dm}^{-3} \text{ and pH} = 6.44$$

After the equivalence point the graph will follow the same pattern as the strong acid – strong base curve as more sodium hydroxide is simply being added to the solution.



INDICATORS

An indicator is a weak acid (or base) in which the dissociated form is a different colour to the undissociated form.

HIn(aq)	⇌	H^+(aq) + In^-(aq)
colour A		colour B
(colour in acid solution)		(colour in alkali solution)

$K_{in} = [H^+] \times \dfrac{[In^-]}{[HIn]}$ Assuming the colour changes when $[In^-] \approx [H_{in}]$ then the end point of the indicator will be when $[H^+] \approx K_{in}$, i.e. when pH $\approx pK_{in}$. Different indicators have different K_{in} values and so change colour within different pH ranges.

Indicator	**pK_{in}**	**pH range**	**Use**
methyl orange	3.7	3.1–4.4	titrations with strong acids
phenolphthalein	9.6	8.3–10.0	titrations with strong bases

Similar arguments can be used to explain the shapes of pH curves for strong acid – weak base, and weak acid – weak base titrations. Since there is no sharp inflexion point titrations involving weak acids with weak bases should not be used in analytical chemistry.