

HL Answers to pH curves questions

1. i. After 49.1 cm³ of NaOH(aq) has been added there will be 0.9 cm³ of 0.100 mol dm³ HCl(aq) remaining in 99.1 cm³ (say 100 cm³) of solution.

$$\text{Amount of HCl(aq)} = 0.9 \times 0.100 \times 1.0/1000 = 9 \times 10^{-5} \text{ mol}$$

$$\text{HCl(aq)} = [\text{H}^+(\text{aq})] = 9 \times 10^{-5} \times 1000/100 = 9 \times 10^{-4} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log_{10} 9 \times 10^{-4} = 3.05$$

ii. After 49.1 cm³ of NaOH(aq) has been added:

$$[\text{CH}_3\text{COOH(aq)}] = 0.9 \times 0.1 \times 100 = 9 \times 10^{-4} \text{ mol dm}^{-3}$$

$$[\text{CH}_3\text{COO}^-(\text{aq})] = 49.1 \times 0.100 \times 1/1000 \times 1000/100 = 4.91 \times 10^{-2} \text{ mol dm}^{-3}$$

$$K_a = [\text{H}^+(\text{aq})] \times [\text{CH}_3\text{COO}^-(\text{aq})] / [\text{CH}_3\text{COOH(aq)}] = 10^{-4.76} = [\text{H}^+(\text{aq})] \times 4.91 \times 10^{-2} / 9 \times 10^{-4}$$

$$[\text{H}^+(\text{aq})] = 9 \times 10^{-4} \times 10^{-4.76} / 4.91 \times 10^{-2} = 3.19 \times 10^{-7} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log_{10} 3.19 \times 10^{-7} = 6.50$$

2. i. After 50.1 cm³ of NaOH(aq) has been added there will be 0.1 cm³ of 0.100 mol dm³ NaOH(aq) present in 100.1 cm³ (say 100 cm³) of solution.

$$\text{Amount of NaOH(aq)} = 0.1 \times 0.100 \times 1.0/1000 = 1 \times 10^{-5} \text{ mol}$$

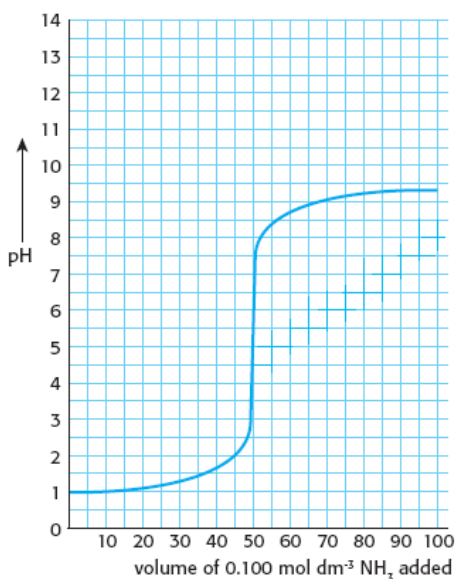
$$\text{NaOH(aq)} = [\text{OH}^-(\text{aq})] = 1 \times 10^{-5} \times 1000/100 = 1 \times 10^{-4} \text{ mol dm}^{-3}$$

$$\text{pOH} = -\log_{10} 1 \times 10^{-4} = 4.00 \text{ so pH} = 10.0$$

ii. The same answer as part i., i.e. pH = 10.0

3. Since this is the half-equivalence point (when [HCOOH(aq)] = [HCOO⁻(aq)]) then pH = pK_a = 3.75.

4.



5. i. There is no sharp change in pH as the equivalence point is approached, reached and passed.
- ii. The unknown solution is always titrated against a solution of known concentration. If a weak acid is being titrated then a strong base of known concentration can be used and *vice-versa*.
6. *Equivalence point* is when equimolar amounts of $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ have reacted.
End point is when the acid-base indicator just changes colour.
7. At the equivalence point the salt formed will be sodium ethanoate. This salt is completely dissociated into its ions:
 $\text{NaCH}_3\text{COO}(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$. Also present will be water which is very slightly dissociated into its ions:
 $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$. The ethanoate ions and the hydrogen ions will combine to form the weak, largely undissociated, ethanoic acid, $\text{CH}_3\text{COOH}(\text{aq})$. The hydroxide ions will remain in solution as sodium hydroxide is a strong base and therefore completely dissociated into its ions in solution. The solution will therefore have a $\text{pH} > 7$.
8. $\text{Na}_2\text{CO}_3(\text{s}) \rightarrow 2\text{Na}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$
 $\text{CO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq})$ (although this reaction is reversible the equilibrium lies very much on the products side as $\text{HCO}_3^-(\text{aq})$ is a weak acid)
9. Sodium cyanide is the salt of a weak acid, HCN, and a strong base, NaOH. When it is dissolved in water it dissociates completely to form sodium ions and cyanide ions. The cyanide ions react with water to form hydrogen cyanide.
 $\text{CN}^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCN}(\text{aq}) + \text{OH}^-(\text{aq})$
 Heating the solution helps to dissolve the salt more quickly but gases are less soluble in hot water than they are in cold water so the hydrogen cyanide is evolved as a gas and can be breathed in.

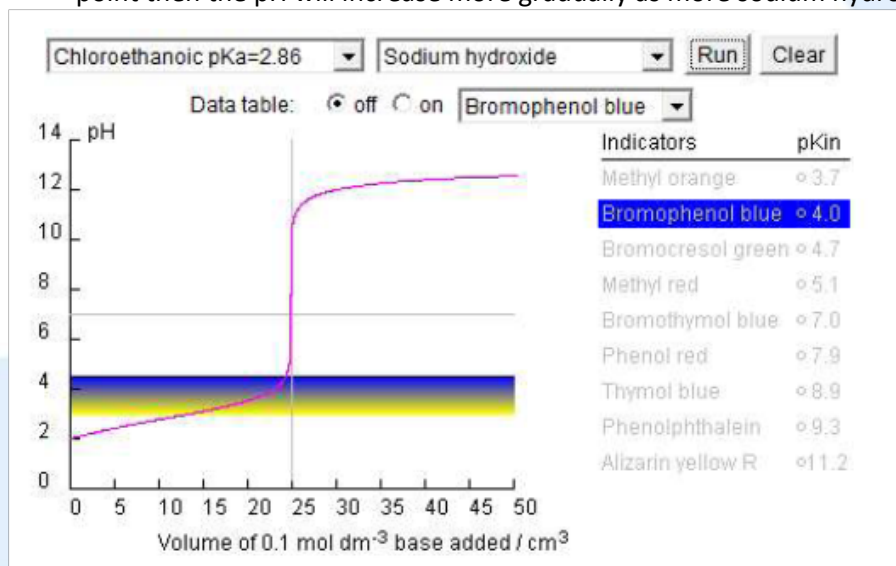
10. i. In acid solution when $[H^+(aq)]$ is high $[HIn(aq)] \gg [In^-(aq)]$ so $HIn(aq)$ is red and $In^-(aq)$ is yellow.

ii. $K_a = [H^+(aq)] \times [In^-(aq)] / [HIn(aq)]$. When $[In^-(aq)] = [HIn(aq)]$ $K_a = [H^+(aq)]$ and $pH = pK_a$
For there to be a colour change one of the species must dominate so the pH range will be around the pK_a value. Since the pK_a value is 5.1 the pH range of the indicator lies just either side of this value, i.e. 4.4 – 6.2.

iii. It will not be a suitable indicator. At the equivalence point the salt formed will be sodium ethanoate which is a basic salt with a pH greater than 7.0 and this does not lie in the 4.4 – 6.2 range of methyl red.

11. From Section 22 in the data booklet the pH range for phenolphthalein is 8.3 – 10. At the equivalence point the salt formed is ammonium chloride which is an acidic salt so the solution will have a pH value less than 7. The solution will be colourless as this is the colour of the indicator in acid solution. When one more drop of the weak alkali solution is added the pH of the solution will still not reach 8.3 so the solution will remain colourless.

12. i. This is a weak acid-strong base titration. It will start at about pH = 2 rise gently to about pH = 6 just below the equivalence point then rise very steeply to about pH = 9 or 10 around the equivalence point then the pH will increase more gradually as more sodium hydroxide is added.



ii. Any two from bromothymol blue, phenol red, thymol blue or phenolphthalein.