### Answers to exam-style questions

**Topic 8**

1. C
2. A
3. B
4. C
5. B
6. B
7. B
8. A
9. B
10. D

11. a. \(\text{CH}_3\text{COOH}(aq) + \text{H}_2\text{O}(l) \Leftrightarrow \text{CH}_3\text{COO}^-(aq) + \text{H}_3\text{O}^+(aq)\)
   
   or
   
   \(\text{CH}_3\text{COOH}(aq) \Leftrightarrow \text{CH}_3\text{COO}^-(aq) + \text{H}^+(aq)\)
   
   The conjugate base is formed when ethanoic acid acts as an acid and loses a proton (\(\text{H}^+\)). The conjugate base is therefore \(\text{CH}_3\text{COO}^-\).  
   
   b. A strong acid dissociates completely in solution; a weak acid dissociates only partially.  
   
   c. i. HCl is a strong acid so dissociates completely:
   
   \[\text{pH} = -\log_{10}[\text{H}^+(aq)] = \log_{10}1.00 \times 10^{-3} = 3.00\]
   
   ii. Ethanoic acid is a weak acid and will be less dissociated than the same concentration of hydrochloric acid. The pH will thus be greater than 3, e.g. 4. (Values in the range 3.5–6 should be accepted.)

12. a. \(\text{H}_2\text{O}(l) \Leftrightarrow \text{H}^+(aq) + \text{OH}^-(aq)\)
   
   b. \(K_w = [\text{H}^+(aq)][\text{OH}^-(aq)]\)
   
   c. NaOH is a strong base; so ionises fully:
   
   \([\text{OH}^-(aq)] = 0.10 \text{ mol dm}^{-3}\)
   
   \([\text{H}^+(aq)][\text{OH}^-]\) = \(1.0 \times 10^{-14}\)
   
   \([\text{H}^+(aq)] = 1.0 \times 10^{-13} \text{ mol dm}^{-3}\)
   
   \(\text{pH} = -\log[\text{H}^+(aq)] = -\log1.0 \times 10^{-13} = 13\)

13. a. \(\text{pH} = -\log_{10}[\text{H}^+(aq)]\)
   
   b. \(K_a = 10^{-pK_a}\)
   
   \(K_a = 10^{-4.82} = 1.51 \times 10^{-5}\)
   
   The equation for the dissociation of butanoic acid is:
   
   \(\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}(aq) \Leftrightarrow \text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-(aq) + \text{H}^+(aq)\)
   
   The expression for \(K_a\) is:
   
   \(K_a = [\text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-(aq)][\text{H}^+(aq)] / [\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}(aq)]\)
   
   Because one molecule of \(\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}\) dissociates to form one \(\text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-\) ion and one \(\text{H}^+\) ion, the concentrations of \(\text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-\) and \(\text{H}^+\) ions in the solution will be equal:
   
   \([\text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-(aq)] = [\text{H}^+(aq)]\)
   
   We will make the assumption that the dissociation of the acid is negligible compared with the concentration of the acid; that is, we will assume that the concentration of the acid at equilibrium is the same as the initial concentration, 0.150 mol dm\(^{-3}\) in this case. These terms are substituted into the \(K_a\) expression:
   
   \(1.51 \times 10^{-5} = [\text{H}^+(aq)]^2 / 0.150\)
   
   \([\text{H}^+(aq)]^2 = 1.51 \times 10^{-5} \times 0.150 = 2.27 \times 10^{-6}\)
   
   \([\text{H}^+(aq)] = \sqrt{(2.27 \times 10^{-6})} = 1.51 \times 10^{-3} \text{ mol dm}^{-3}\)
   
   \(\text{pH} = -\log_{10}[\text{H}^+(aq)] = -\log_{10}(1.51 \times 10^{-3}) = 2.82\)

   c. It will be higher than 7, because sodium butanoate is the salt of a strong base (NaOH) and a weak acid (butanoic acid). When sodium butanoate dissolves in water, the two ions separate from each other; the solution contains \(\text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-\)(aq) and \(\text{Na}^+(aq)\) ions. The \(\text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-\) ion is the conjugate base of the weak acid butanoic acid and so, acting as a base, will react with water molecules to accept a proton according to the equilibrium:
   
   \(\text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-(aq) + \text{H}_2\text{O}(l) \Leftrightarrow \text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}(aq) + \text{OH}^-(aq)\)

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**Barium hydroxide** will have a higher pH; barium hydroxide has the formula \(\text{Ba}(\text{OH})_2\) and is a strong base; a 0.10 mol dm\(^{-3}\) solution will ionise to produce \([\text{OH}^-](aq) = 0.20 \text{ mol dm}^{-3}\); a higher concentration of \(\text{OH}^-\); so a higher pH.  

**pH** = \(-\log_{10}[\text{H}^+(aq)]\)
14 a \[ \text{CH}_3\text{CH}_2\text{COOH}(aq) + \text{H}_2\text{O}(l) \Rightarrow \text{CH}_3\text{CH}_2\text{COO}^- (aq) + \text{H}_3\text{O}^+(aq) \]

or

\[ \text{CH}_3\text{CH}_2\text{COOH}(aq) \Rightarrow \text{CH}_3\text{CH}_2\text{COO}^- (aq) + \text{H}^+(aq) \]

\[ K_a = 10^{-pK_a} \]

For ethanoic acid, \( K_a = 10^{-4.76} = 1.74 \times 10^{-5} \)

For propanoic acid, \( K_a = 10^{-4.87} = 1.35 \times 10^{-5} \)

Ethanoic acid is the stronger acid, because it has a higher \( K_a \) value. A higher \( K_a \) indicates that a greater proportion of the acid molecules have dissociated.

b i The equation for the reaction is:

\[ \text{CH}_3\text{CH}_2\text{COOH}(aq) + \text{NaOH}(aq) \rightarrow \text{CH}_3\text{CH}_2\text{CH}_2\text{COONa}(aq) + \text{H}_2\text{O}(l) \]

d i The equation for the dissociation of propanoic acid is:

\[ \text{CH}_3\text{CH}_2\text{COOH}(aq) \Rightarrow \text{CH}_3\text{CH}_2\text{COO}^- (aq) + \text{H}^+(aq) \]

\[ K_a = 10^{-pK_a} \]

For ethanoic acid, \( K_a = 10^{-4.76} = 1.74 \times 10^{-5} \)

For propanoic acid, \( K_a = 10^{-4.87} = 1.35 \times 10^{-5} \)

Ethanoic acid is the stronger acid, because it has a higher \( K_a \) value. A higher \( K_a \) indicates that a greater proportion of the acid molecules have dissociated.

c The equation for the dissociation of propanoic acid is:

\[ \text{CH}_3\text{CH}_2\text{COOH}(aq) \Rightarrow \text{CH}_3\text{CH}_2\text{COO}^- (aq) + \text{H}^+(aq) \]

The expression for \( K_a \) is:

\[ K_a = \frac{[\text{CH}_3\text{CH}_2\text{COO}^- (aq)][\text{H}^+(aq)]}{[\text{CH}_3\text{CH}_2\text{COOH}(aq)]} \]

Because one molecule of \( \text{CH}_3\text{CH}_2\text{COOH} \) dissociates to form one \( \text{CH}_3\text{CH}_2\text{COO}^- \) ion and one \( \text{H}^+ \) ion, the concentration of \( \text{CH}_3\text{CH}_2\text{COO}^- \) and \( \text{H}^+ \) ions in the solution will be equal:

\[ [\text{CH}_3\text{CH}_2\text{COO}^- (aq)] = [\text{H}^+(aq)] \]

We will make the assumption that the dissociation of the acid is negligible compared with the concentration of the acid; so the concentration of the acid at equilibrium is the same as the initial concentration, 0.250 mol dm\(^{-3}\) in this case.

These terms and the \( K_a \) value are substituted into the \( K_a \) expression:

\[ 1.35 \times 10^{-5} = \frac{[\text{H}^+(aq)]^2}{0.250} \]

\[ [\text{H}^+(aq)]^2 = 1.35 \times 10^{-5} \times 0.250 = 3.37 \times 10^{-6} \]

\[ [\text{H}^+(aq)] = \sqrt{3.37 \times 10^{-6}} = 1.84 \times 10^{-3} \text{ mol dm}^{-3} \]

\[ \text{pH} = -\log_{10} [\text{H}^+(aq)] = 2.74 \]

dii \( K_a \times K_b = K_w \) for a conjugate acid–base pair. Assuming that the temperature is 25 °C and therefore that \( K_w = 1.00 \times 10^{-14} \).

\( K_a \) for propanoic acid is \( 1.35 \times 10^{-5} \), so

\[ 1.35 \times 10^{-5} \times K_b = 1.00 \times 10^{-14} \]

\[ K_b = 7.41 \times 10^{-10} \]

\[ [\text{OH}^- (aq)] = \text{Volume of NaOH added} / \text{cm}^3 \]

The values quoted are from the previous parts of the question. The pH at the equivalence point is higher than 7 because sodium butanoate is formed in the titration. The final pH will approach the pH of the sodium hydroxide solution used.

v The most suitable indicator for a strong base–weak acid titration is phenolphthalein. The range of the indicator (8.3–10.0) comes entirely within the very steep part of the titration curve.
The position of equilibrium shifts to the right as the temperature increases, so the reaction to the right must be endothermic. When the temperature increases, the position of equilibrium shifts in the endothermic direction to take in heat and minimise the effect of the change.

A buffer solution is one that resists changes in pH when small amounts of acid or base are added.

If some hydrochloric acid is added to this solution, the extra $H^+$ added reacts with the $NH_3$ in the solution:

$$NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$$

If some sodium hydroxide is added to the solution, the extra $OH^-$ added reacts with the $NH_4^+$ in the solution:

$$NH_4^+(aq) + OH^− (aq) \rightarrow NH_3(aq) + H_2O(l)$$

The ionisation of ammonia is shown by the equation:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^- (aq)$$

The expression for $K_b$ is:

$$K_b = \frac{[NH_4^+(aq)][OH^− (aq)]}{[NH_3(aq)]}$$

One $NH_3$ molecule ionises to produce one $NH_4^+$ and one $OH^-$ ion. This means that the concentration of $NH_4^+$ is equal to the $OH^-$ concentration and we can write:

$$[OH^− (aq)] = 1.78 \times 10^{−5} \text{ mol dm}^{−3}$$

At 25°C, $pOH + pH = 14$; therefore:

$$pH = 14 - 4.91 = 9.09$$

The amount of HCl that has been added is half the amount required for neutralisation; at the half-neutralisation point, $pOH = pK_b$ because $[NH_3(aq)] = [NH_4^+(aq)]$; therefore $pOH = 4.75$. So $pH = 14 - 4.75 = 9.25$

The position of equilibrium shifts to the right as the temperature increases, so the reaction to the right must be endothermic. When the temperature increases, the position of equilibrium shifts in the endothermic direction to take in heat and minimise the effect of the change.