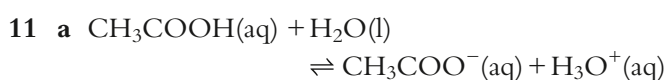


Answers to exam-style questions

Topic 8

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



or



The conjugate base is formed when ethanoic acid acts as an acid and loses a proton (H^+). The conjugate base is therefore CH_3COO^- [2]

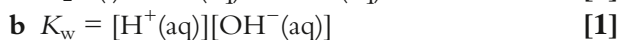
b A strong acid dissociates completely in solution; a weak acid dissociates only partially. [2]

c i HCl is a strong acid so dissociates completely:
 $[\text{H}^+(\text{aq})] = 1.00 \times 10^{-3} \text{ mol dm}^{-3}$
 $\text{pH} = -\log_{10}[\text{H}^+(\text{aq})] = \log_{10} 1.00 \times 10^{-3} = 3.00$

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.) [3]

d i $\text{Mg} + 2\text{CH}_3\text{COOH} \rightarrow (\text{CH}_3\text{COO})_2\text{Mg} + \text{H}_2$
magnesium ethanoate and hydrogen gas [2]

ii Equal concentrations of the two acids must be reacted with magnesium; the strong acid will react more vigorously with magnesium because the strong acid has a higher concentration of H^+ ions. [2]



c NaOH is a strong base; so ionises fully
 $[\text{OH}^-(\text{aq})] = 0.10 \text{ mol dm}^{-3}$
 $[\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})] = 1.0 \times 10^{-14}$
 $[\text{H}^+(\text{aq})] = 1.0 \times 10^{-13} \text{ mol dm}^{-3}$
 $\text{pH} = -\log[\text{H}^+(\text{aq})] = -\log 1.0 \times 10^{-13} = 13$ [3]

d 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 $[\text{OH}^-(\text{aq})] = 0.20 \text{ mol dm}^{-3}$; a higher concentration of $\text{OH}^-(\text{aq})$; so a higher pH. [2]

13 a $\text{pH} = -\log_{10}[\text{H}^+(\text{aq})]$ [1]

b $K_a = 10^{-\text{p}K_a}$
 $K_a = 10^{-4.82} = 1.51 \times 10^{-5}$
The equation for the dissociation of butanoic acid is:
 $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOH}(\text{aq})$



The expression for K_a is:

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

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 H^+ ion, the concentrations of $\text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-$ and H^+ ions in the solution will be equal:

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

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, $0.150 \text{ mol dm}^{-3}$ in this case.

These terms are substituted into the K_a expression:

$$1.51 \times 10^{-5} = \frac{[\text{H}^+(\text{aq})]^2}{0.150}$$

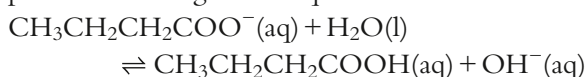
$$[\text{H}^+(\text{aq})]^2 = 1.51 \times 10^{-5} \times 0.150 = 2.27 \times 10^{-6}$$

$$[\text{H}^+(\text{aq})] = \sqrt{(2.27 \times 10^{-6})} = 1.51 \times 10^{-3} \text{ mol dm}^{-3}$$

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

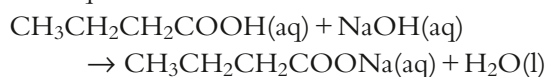
$$\text{pH} = -\log_{10}(1.51 \times 10^{-3}) = 2.82$$
 [4]

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 $\text{CH}_3\text{CH}_2\text{CH}_2\text{COO}^-(\text{aq})$ and $\text{Na}^+(\text{aq})$ 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:



The concentration of OH^- ions in the solution has been increased and the solution is alkaline. [3]

d i The equation for the reaction is:



$$\text{no. moles of butanoic acid} = \frac{25.00}{1000} \times 0.150$$

$$= 3.75 \times 10^{-3} \text{ mol}$$

From the chemical equation: one mole of butanoic acid react with one mole of sodium hydroxide. Therefore 3.75×10^{-3} mol butanoic acid reacts with 3.75×10^{-3} mol sodium hydroxide. So there are 3.75×10^{-3} mol present in 27.60 cm^3 of sodium hydroxide.

The concentration of sodium hydroxide is:

$$\frac{3.75 \times 10^{-3}}{(27.60/1000)} = 0.136 \text{ mol dm}^{-3} \quad [3]$$

ii NaOH is a strong base and therefore completely ionises in solution:

$$[\text{OH}^-(\text{aq})] = 0.136 \text{ mol dm}^{-3}$$

$$\text{pOH} = -\log_{10} [\text{OH}^-(\text{aq})]$$

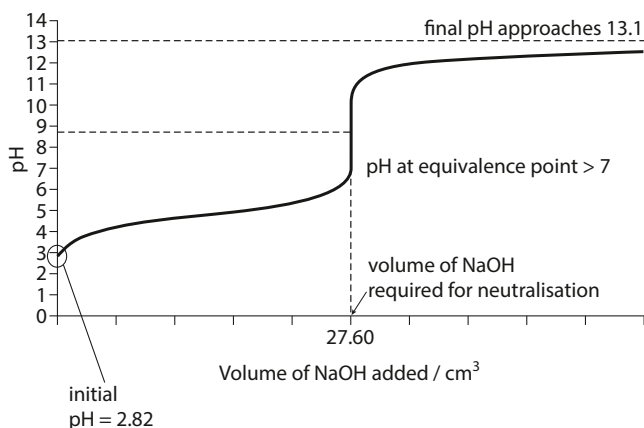
$$\text{pOH} = -\log_{10} 0.136 = 0.866$$

Assuming that the temperature is 25°C :

$$14 = \text{pH} + \text{pOH}$$

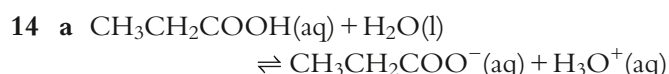
$$\text{pH} = 14 - 0.866 = 13.1 \quad [3]$$

iii

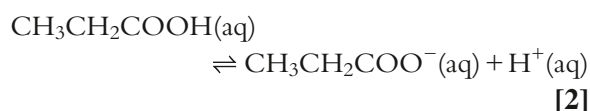


The values quoted are from the previous parts of the question. The pH at the equivalence point is higher than 7 because sodium butanoate is formed in the titration. The final pH will approach the pH of the sodium hydroxide solution used. [3]

iv 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. [2]



or



b $K_a = 10^{-\text{p}K_a}$

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. [4]

c The equation for the dissociation of propanoic acid is:



The expression for K_a is:

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

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 H^+ ion, the concentration of $\text{CH}_3\text{CH}_2\text{COO}^-$ and H^+ ions in the solution will be equal:

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

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, $0.250 \text{ mol dm}^{-3}$ in this case.

These terms and the K_a value are substituted into the K_a expression:

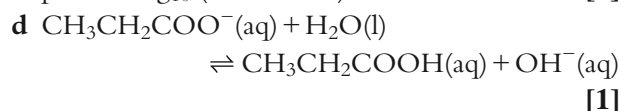
$$1.35 \times 10^{-5} = \frac{[\text{H}^+(\text{aq})]^2}{0.250}$$

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

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

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

$$\text{pH} = -\log_{10} (1.84 \times 10^{-3}) = 2.74 \quad [4]$$



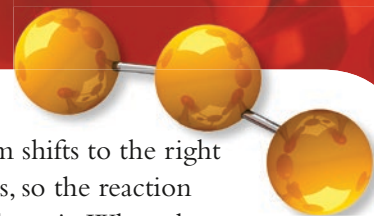
e $K_a \times K_b = K_w$ for a conjugate acid–base pair.

Assuming that the temperature is 25°C and therefore that $K_w = 1.00 \times 10^{-14}$.

K_a for propanoic acid is 1.35×10^{-5} , so

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

$$K_b = 7.41 \times 10^{-10} \quad [2]$$



f We can write an expression for K_b for $\text{CH}_3\text{CH}_2\text{COO}^-$

$$K_b = \frac{[\text{CH}_3\text{CH}_2\text{COOH}(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{CH}_3\text{CH}_2\text{COO}^-(\text{aq})]}$$

One $\text{CH}_3\text{CH}_2\text{COO}^-$ ion reacts with one H_2O molecule to form one $\text{CH}_3\text{CH}_2\text{COOH}$ molecule and one OH^- ion; therefore:

$$[\text{CH}_3\text{CH}_2\text{COOH}(\text{aq})] = [\text{OH}^-(\text{aq})]$$

We will use the approximation that the amount of $\text{CH}_3\text{CH}_2\text{COO}^-$ that reacts with water is negligible compared with its concentration and take $[\text{CH}_3\text{CH}_2\text{COO}^-(\text{aq})]$ to be $0.200 \text{ mol dm}^{-3}$. Substituting known values into the K_b expression:

$$7.41 \times 10^{-10} = \frac{[\text{OH}^-(\text{aq})]^2}{0.200}$$

$$[\text{OH}^-] = 1.22 \times 10^{-5} \text{ mol dm}^{-3}$$

$$\text{pOH} = -\log_{10} [\text{OH}^-(\text{aq})]$$

$$\text{pOH} = -\log_{10} 1.22 \times 10^{-5} = 4.91$$

$$\text{pOH} + \text{pH} = \text{p}K_w$$

At 25°C , $\text{pOH} + \text{pH} = 14$; therefore:

$$\text{pH} = 14 - 4.91 = 9.09 \quad [4]$$

15 a $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \quad [1]$

b $K_w = [\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})]$

Because one H_2O molecule dissociates to form one H^+ ion and one OH^- ion, in pure water the concentration of H^+ ions is equal to the concentration of OH^- ions. We can therefore write $K_w = [\text{H}^+(\text{aq})]^2$ for pure water. So, at 40°C :

$$5.48 \times 10^{-14} = [\text{H}^+(\text{aq})]^2$$

$$[\text{H}^+(\text{aq})] = \sqrt{5.48 \times 10^{-14}}$$

$$[\text{H}^+(\text{aq})] = 2.34 \times 10^{-7} \text{ mol dm}^{-3}$$

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

$$\text{pH} = -\log_{10} (2.34 \times 10^{-7}) = 6.63 \quad [3]$$

c pOH for a neutral solution at 323 K will be the same as the pH, i.e. 6.63. This solution has a pOH of 7.0. A higher pOH indicates a lower concentration of OH^- ions than in a neutral solution; therefore the solution is acidic.

or

K_w is 5.48×10^{-14} at 323 K . Therefore

$$\text{p}K_w = 13.26.$$

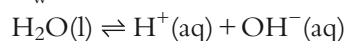
$$\text{p}K_w = \text{pH} + \text{pOH}$$

If $\text{pOH} = 7.0$, the pH of this solution is $13.26 - 7.0 = 6.26$.

This solution has a lower pH than the neutral pH at this temperature and is, therefore, acidic. [3]

d K_w for water at 288 K is $10^{-14.34}$, i.e. 4.57×10^{-15} .

K_w is 5.48×10^{-14} at 323 K .

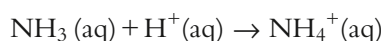


As the temperature increases, the value of K_w increases and therefore more water dissociates.

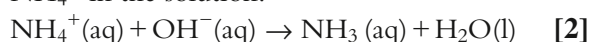
The position of equilibrium shifts to the right as the temperature increases, so the reaction to the right must be endothermic. When the temperature increases, the position of equilibrium shifts in the endothermic direction to take in heat and minimise the effect of the change. [3]

16 a A buffer solution is one that resists changes in pH when small amounts of acid or base are added. [2]

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



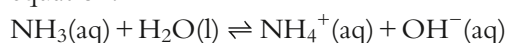
If some sodium hydroxide is added to the solution, the extra OH^- added reacts with the NH_4^+ in the solution:



c $K_b = 10^{-\text{p}K_b} \quad K_b = 10^{-4.75}$

Therefore, $K_b = 1.78 \times 10^{-5}$.

The ionisation of ammonia is shown by the equation:



The expression for K_b is:

$$K_b = \frac{[\text{NH}_4^+(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{NH}_3(\text{aq})]}$$

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:

$$K_b = \frac{[\text{OH}^-(\text{aq})]^2}{[\text{NH}_3(\text{aq})]}$$

We will make the approximation that the concentration of NH_3 at equilibrium is equal to the initial concentration, i.e. that the ionisation of the base is negligible compared with its concentration. Therefore we take $[\text{NH}_3(\text{aq})]$ to be $0.125 \text{ mol dm}^{-3}$. If we substitute this value and the value for K_b into the expression for K_b we get:

$$1.78 \times 10^{-5} = \frac{[\text{OH}^-(\text{aq})]^2}{0.125}$$

$$[\text{OH}^-(\text{aq})]^2 = 1.78 \times 10^{-5} \times 0.125 = 2.22 \times 10^{-6}$$

$$[\text{OH}^-(\text{aq})] = 1.49 \times 10^{-3} \text{ mol dm}^{-3}$$

$$\text{pOH} = -\log_{10} [\text{OH}^-(\text{aq})]$$

$$\text{pOH} = -\log_{10} (1.49 \times 10^{-3}) = 2.83 \quad [3]$$

d The amount of HCl that has been added is half the amount required for neutralisation; at the half-neutralisation point, $\text{pOH} = \text{p}K_b$ because $[\text{NH}_3(\text{aq})] = [\text{NH}_4^+(\text{aq})]$; therefore $\text{pOH} = 4.75$. So $\text{pH} = 14 - 4.75 = 9.25 \quad [3]$