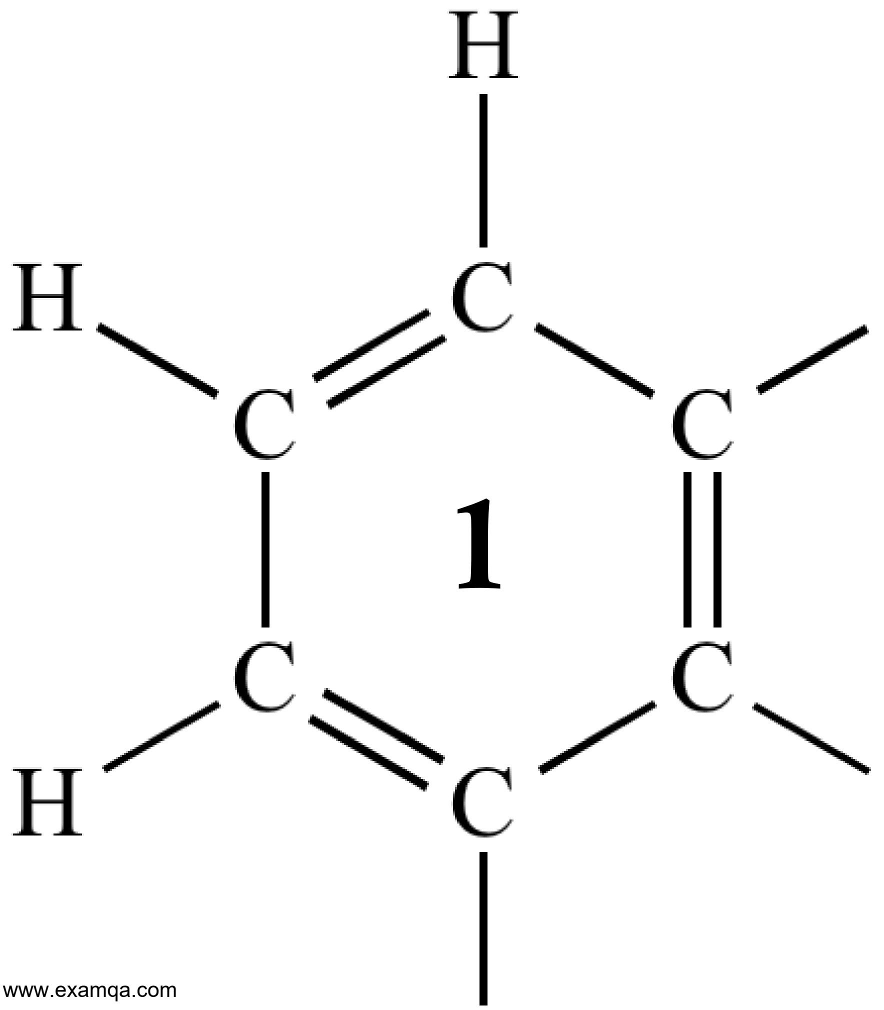
OCR A2 CHEMISTRY

# MODULE 5.3

ACIDS – BASES - pH



**1** Nitric acid ($HNO_{-3}$) is a strong acid. Ethanoic acid ($CH_3COOH$) is a weak acid.

(a) Write an equation to show how ethanoic acid behaves as a weak acid in its reaction with water.

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**(1)**

(b) When pure ethanoic acid reacts with pure nitric acid, ethanoic acid acts as a base.

Write an equation for this reaction.

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**(1)**

(c) Two beakers, **A** and **B**, each contain 100.0 $cm^{-3}$ of 0.0125 mol $dm^{-3}$ nitric acid.

(i) Calculate the pH of the solution formed after 50.0 $cm^{-3}$ of distilled water are added to beaker **A**.
Give your answer to 2 decimal places.

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**(2)**

(ii) Calculate the pH of the solution formed after 50.0 $cm^3$ of 0.0108 mol $dm^{-3}$ aqueous sodium hydroxide are added to beaker **B**.
Give your answer to 2 decimal places.

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**(4)**

(d) A third beaker, **C**, contains 100.0 $cm^3$ of 0.0125 mol $dm^{-3}$ ethanoic acid.
The acid dissociation constant $K_a$ for ethanoic acid has the value $1.74 \times 10^{-5}$ mol $dm^{-3}$ at 25 °C.

(i) Write an expression for $K_a$ for ethanoic acid and use it to calculate the pH of the ethanoic acid solution in beaker **C**.
Show your working. Give your answer to 2 decimal places.

$K_a$ ...............................................................................................................

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Calculation ...............................................................................................

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**(4)**

(ii) Aqueous sodium hydroxide is added to beaker **C** until the pH of the solution becomes 4.84.

Name the salt formed in the reaction of ethanoic acid with sodium hydroxide.

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**(1)**

(iii) Calculate the value of $\frac{[\text{salt}]}{[\text{ethanoic acid}]}$ in the solution with the pH of 4.84.

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**(3)**

(e) Explain why chloroethanoic acid is a stronger acid than ethanoic acid.

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(f) Explain why data books do not usually contain values of $K_a$ for strong acids.

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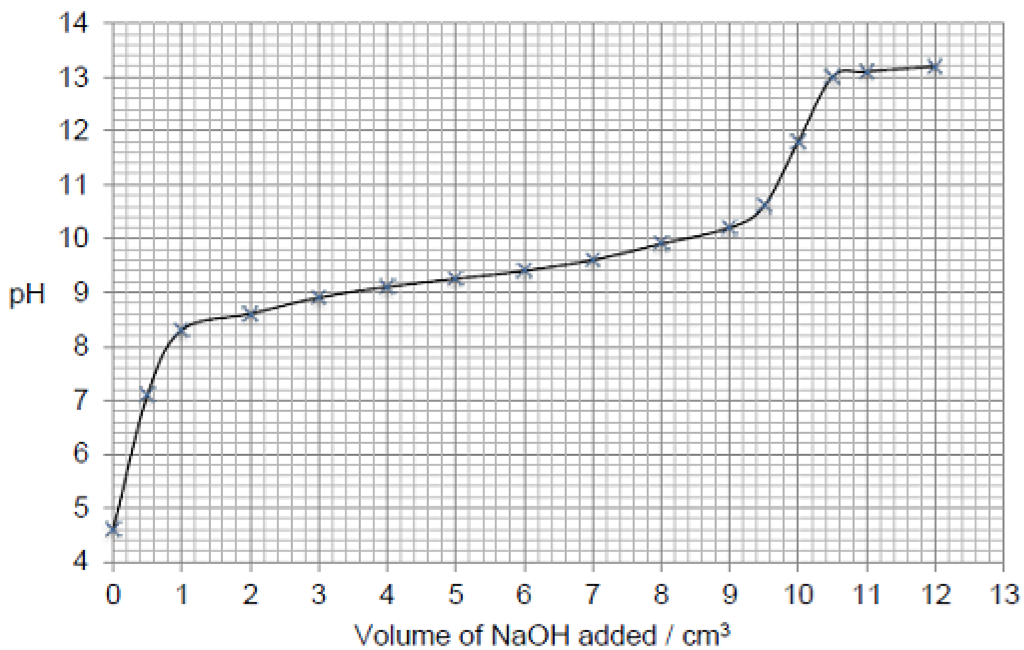
**(2)**
**(Total 20 marks)**

**2** Ammonium chloride, when dissolved in water, can act as a weak acid as shown by the following equation.

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

The following figure shows a graph of data obtained by a student when a solution of sodium hydroxide was added to a solution of ammonium chloride. The pH of the reaction mixture was measured initially and after each addition of the sodium hydroxide solution.




(a) Suggest a suitable piece of apparatus that could be used to measure out the sodium hydroxide solution.
Explain why this apparatus is more suitable than a pipette for this purpose.

Apparatus ..........................................................................................

Explanation ..........................................................................................

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**(2)**

(b) Use information from the curve in the figure above to explain why the end point of this reaction would be difficult to judge accurately using an indicator.

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**(2)**

(c) The pH at the end point of this reaction is 11.8.

Use this pH value and the ionic product of water, $K_w = 1.0 \times 10^{-14}$ mol$^2$ dm$^{-6}$, to calculate the concentration of hydroxide ions at the end point of the reaction.

Concentration = .............................. mol dm$^{-3}$

**(3)**

(d) The expression for the acid dissociation constant for aqueous ammonium ions is

$$K_a = \frac{[NH_3][H^+]}{[NH_4^+]}$$

The initial concentration of the ammonium chloride solution was 2.00 mol dm$^{-3}$.

Use the pH of this solution, before any sodium hydroxide had been added, to calculate a value for $K_a$

$K_a$ = .............................. mol dm$^{-3}$

**(3)**

(e) A solution contains equal concentrations of ammonia and ammonium ions.

Use your value of $K_a$ from part **(d)** to calculate the pH of this solution. Explain your working.

(If you were unable to calculate a value for $K_a$ you may assume that it has the value $4.75 \times 10^{-9}$ mol dm$^{-3}$. This is **not** the correct value.)

pH= ..............................

**(2)**
**(Total 12 marks)**

**3** Ethanedioic acid is a weak acid.
Ethanedioic acid acts, initially, as a monoprotic acid.

(a) Use the concept of electronegativity to justify why the acid strengths of ethanedioic acid and ethanoic acid are different.

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**(6)**

(b) A buffer solution is made by adding $6.00 \times 10^{-2}$ mol of sodium hydroxide to a solution containing $1.00 \times 10^{-1}$ mol of ethanedioic acid ($H_2C_2O_4$).
Assume that the sodium hydroxide reacts as shown in the following equation and that in this buffer solution, the ethanedioic acid behaves as a monoprotic acid.

$$H_2C_2O_4(aq) + OH^-(aq) \longrightarrow HC_2O_4^-(aq) + H_2O(l)$$

The dissociation constant $K_a$ for ethanedioic acid is $5.89 \times 10^{-2}$ mol dm$^{-3}$.

Calculate a value for the pH of the buffer solution.
Give your answer to the appropriate number of significant figures.

pH = ....................................

**(5)**

(c) In a titration, the end point was reached when 25.0 $cm^3$ of an acidified solution containing ethanedioic acid reacted with 20.20 $cm^3$ of $2.00 \times 10^{-2}$ mol $dm^{-3}$ potassium manganate(VII) solution.

Deduce an equation for the reaction that occurs and use it to calculate the original concentration of the ethanedioic acid solution.

Equation ..........................................................................................................

Calculation

Original concentration = .............................. mol $dm^{-3}$

**(4)**

**(Total 15 marks)**

**4** What is the pH of a 0.020 mol $dm^{-3}$ solution of a diprotic acid which is completely dissociated?

**A** 1.00 ☐

**B** 1.40 ☐

**C** 1.70 ☐

**D** 4.00 ☐

**(Total 1 mark)**

**5** The acid dissociation constant, $K_a$, of a weak acid HA has the value $2.56 \times 10^{-4}$ mol dm$^{-3}$.

What is the pH of a $4.25 \times 10^{-3}$mol dm$^{-3}$ solution of HA?

**A** 5.96 ⬭

**B** 3.59 ⬭

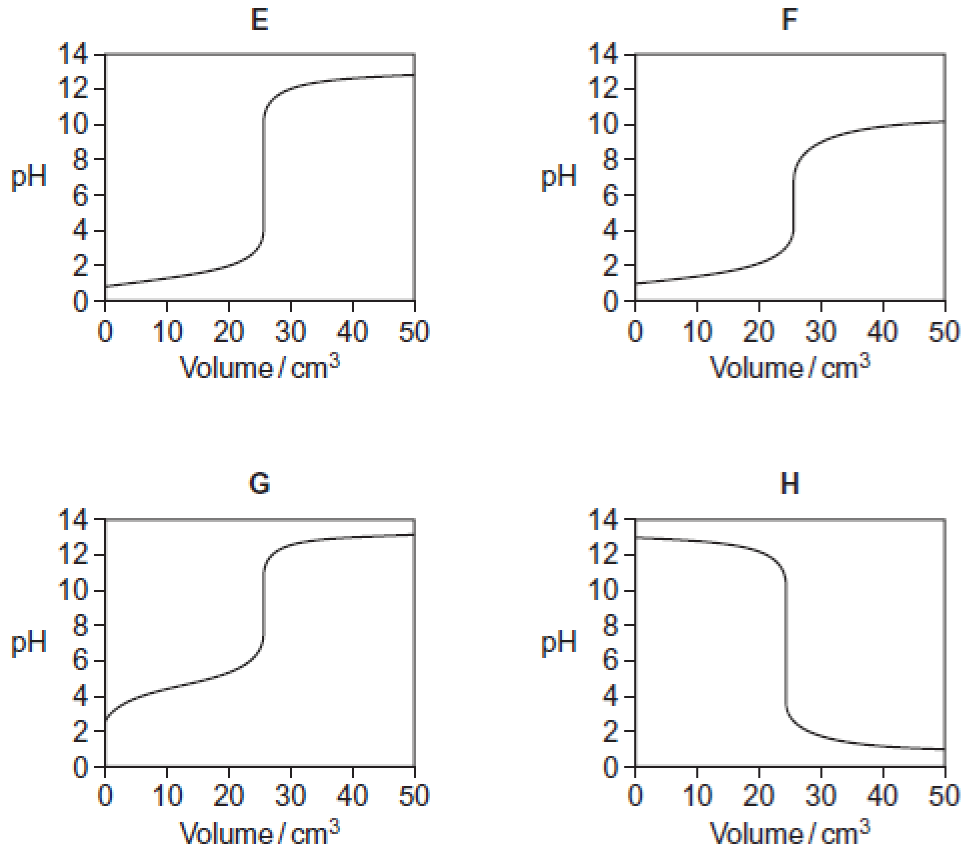
**C** 2.98 ⬭

**D** 2.37 ⬭

**(Total 1 mark)**

**6** Titration curves, labelled **E**, **F**, **G** and **H**, for combinations of different aqueous solutions of acids and bases are shown below.

All solutions have concentrations of 0.1 mol $dm^{-3}$.




(a) In this part of the question, write the appropriate letter in each box.

From the curves **E**, **F**, **G** and **H**, choose the curve produced by the addition of

(i) sodium hydroxide to 25 $cm^3$ of ethanoic acid ☐

**(1)**

(ii) ammonia to 25 $cm^3$ hydrobromic acid ☐

**(1)**

(iii) hydrochloric acid to 25 $cm^3$ of potassium hydroxide ☐

**(1)**

(b) The table shows information about some acid-base indicators.

| Indicator | pH range | Lower pH colour | Higher pH colour |
|---|---|---|---|
| pentamethoxy red | 1.2−3.2 | violet | colourless |
| naphthyl red | 3.7−5.0 | red | yellow |
| 4−nitrophenol | 5.6−7.0 | colourless | yellow |
| cresol purple | 7.6−9.2 | yellow | purple |

(i) Which indicator in the table could be used for the titration that produces curve **E** but **not** for the titration that produces curve **F**?

Tick (✓) **one** box.

pentamethoxy red ☐

naphthyl red ☐

4−nitrophenol ☐

cresol purple ☐

**(1)**

(ii) Give the colour change at the end point of the titration that produces curve **H** when naphthyl red is used as the indicator.

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**(1)**

(iii) A beaker contains 25 $cm^3$ of a buffer solution at pH = 6.0
Two drops of each of the four indicators in the table are added to this solution.

State the colour of the mixture of indicators in this buffer solution.
You should assume that the indicators do **not** react with each other.

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**(1)**
**(Total 6 marks)**

**7** Water dissociates slightly according to the equation:

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

The ionic product of water, $K_w$, is given by the expression

$$K_w = [H^+][OH^-]$$

$K_w$ varies with temperature as shown in the table.

| **Temperature / °C** | **$K_w$ / mol$^2$ dm$^{-6}$** |
|---|---|
| 25 | $1.00 \times 10^{-14}$ |
| 50 | $5.48 \times 10^{-14}$ |

(a) Explain why the expression for $K_w$ does **not** include the concentration of water.

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(b) Explain why the value of $K_w$ increases as the temperature increases.

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(c) Calculate the pH of pure water at 50 °C.
Give your answer to 2 decimal places.

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(d) Calculate the pH of 0.12 mol $dm^{-3}$ aqueous NaOH at 50 °C.
Give your answer to 2 decimal places.

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**(3)**
**(Total 10 marks)**

**8** The acid dissociation constant, $K_a$, for ethanoic acid is given by the expression

$$K_a = \frac{[H^+][CH_3COO^-]}{[CH_3COOH]}$$

The value of $K_a$ for ethanoic acid is $1.74 \times 10^{-5}$ mol dm$^{-3}$ at 25 °C.

(a) A buffer solution is prepared using ethanoic acid and sodium ethanoate. In the buffer solution, the concentration of ethanoic acid is 0.186 mol dm$^{-3}$ and the concentration of sodium ethanoate is 0.105 mol dm$^{-3}$.

Calculate the pH of this buffer solution.
Give your answer to 2 decimal places.

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(b) In a different buffer solution, the concentration of ethanoic acid is 0.251 mol $dm^{-3}$ and the concentration of sodium ethanoate is 0.140 mol $dm^{-3}$.

A sample of hydrochloric acid containing 0.015 mol of HCl is added to 1000 $cm^3$ of this buffer solution.

Calculate the pH of the buffer solution after the hydrochloric acid has been added.
You should ignore any change in total volume.
Give your answer to 2 decimal places.

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**(5)**
**(Total 8 marks)**

**9** A solution of chlorine in water is acidic. Swimming pool managers maintain pool water at a constant pH by using a buffer. They do so by adding sodium hydrogencarbonate and sodium carbonate.

(a) Hydrogen carbonate ions ($HCO_3^-$) act as a weak acid in aqueous solution. Write an equation for this equilibrium.

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**(1)**

(b) Use the equation in part (a) to explain how a solution containing sodium hydrogencarbonate and sodium carbonate can act as a buffer when small amounts of acid or small amounts of alkali are added.

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**(3)**
**(Total 4 marks)**