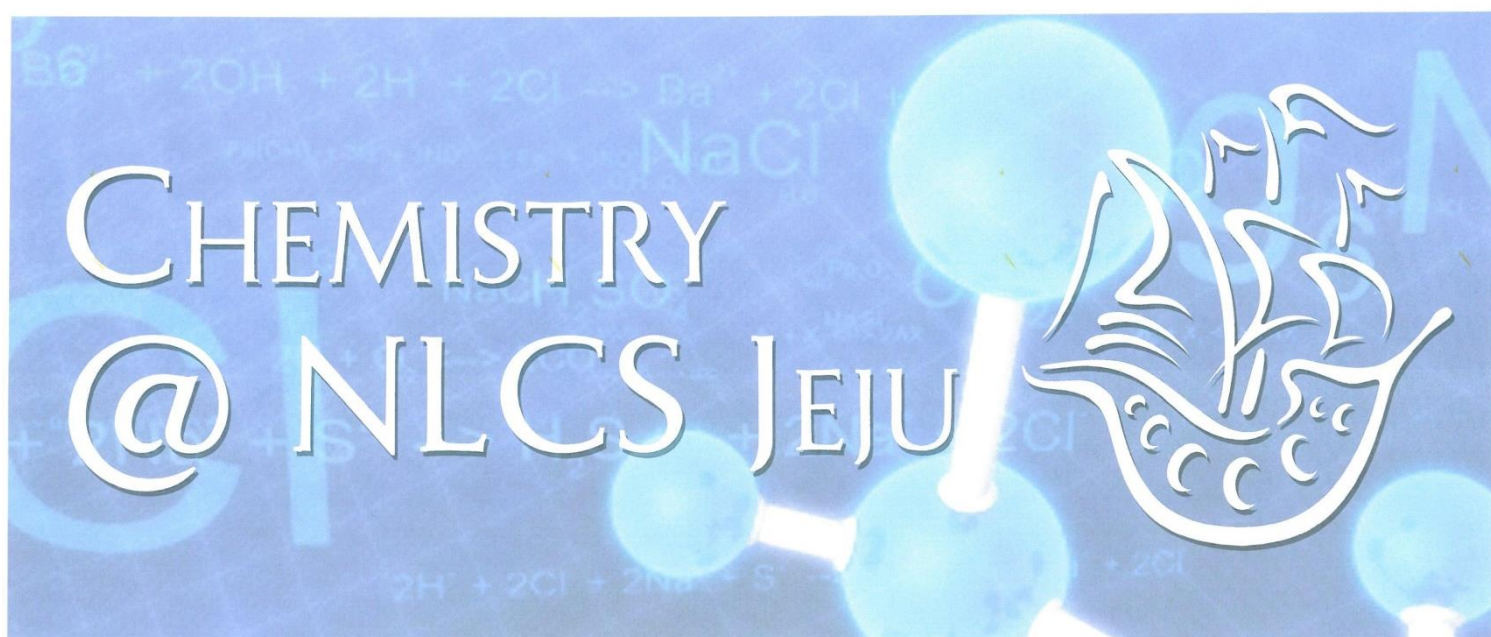# Topic 8

HIGHER Level

Summer & Winter Papers
Summer 1999 to Summer 2013

Name:_____

Topic Exam Statistics (Paper 2):

Section	Marks	% of All Marks	Last four exams marks	Last four exams %
A	105/1080	**10%**	/160	%
B	280/2700	**10%**	/400	%
TOTAL	385/3780	**10%**	/560	%

Total number of papers represented here is 27, each with 40 marks of Section A and 100 marks of section B (4 questions from which you chose to answer only 2)

3. Buffer solutions are widely used in both chemical and biochemical systems.

(a) Describe the composition of an acidic buffer solution. [1]

(b) Determine the pH of a buffer solution, correct to **two** decimal places, showing your working, consisting of 10.0 g of CH_3COOH and 10.0 g of CH_3COONa in 0.250 dm³ of solution. K_a for $CH_3COOH = 1.8 \times 10^{-5}$ at 298 K. [5]

4. Hypochlorous acid, HOCl(aq), is an example of a weak acid.

    (a) State the expression for the ionic product constant of water, K_w. [1]

    (b) A household bleach contains sodium hypochlorite, NaOCl(aq), at a concentration of 0.705 mol dm^{-3}. The hypochlorite ion, OCl$^-$(aq) is a weak base.

$$OCl^-(aq) + H_2O(l) \rightleftharpoons HOCl(aq) + OH^-(aq)$$
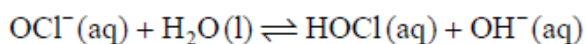
    (i) The pK_a value of HOCl(aq) is 7.52. Determine the K_b value of OCl$^-$(aq) assuming a temperature of 298 K. [1]

    (ii) Determine the concentration of OH$^-$(aq), in mol dm^{-3}, at equilibrium and state **one** assumption made in arriving at your answer other than a temperature of 298 K. [3]

(iii)   Calculate the pH of the bleach.   [2]

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HL SECTION A 10s

3

(d) (i) State an equation for the reaction of ethanoic acid with water. [1]

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(ii) Calculate the pH of 0.200 mol dm^{-3} ethanoic acid (pK_a = 4.76). [3]

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(e) Determine the pH of a solution formed from adding 50.0 cm^3 of 1.00 mol dm^{-3} ethanoic acid, CH_3COOH (aq), to 50.0 cm^3 of 0.600 mol dm^{-3} sodium hydroxide, NaOH (aq). [4]

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(f) Explain how the solution formed in part (e) can act as a buffer. Use equations to support your answer. [2]

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HL SECTION A 09s

TABLE 14 and 15

2. (a) The pK_a value for propanoic acid is given in Table 15 of the Data Booklet.

(i) State the equation for the reaction of propanoic acid with water. [1]

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(ii) Calculate the hydrogen ion concentration (in mol dm^{-3}) of an aqueous solution of 0.100 mol dm^{-3} propanoic acid. [2]

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(b) The graph below shows a computer simulation of a titration of 25.0 cm^3 of 0.100 mol dm^{-3} hydrochloric acid with 0.100 mol dm^{-3} sodium hydroxide and the pH range of phenol red indicator.
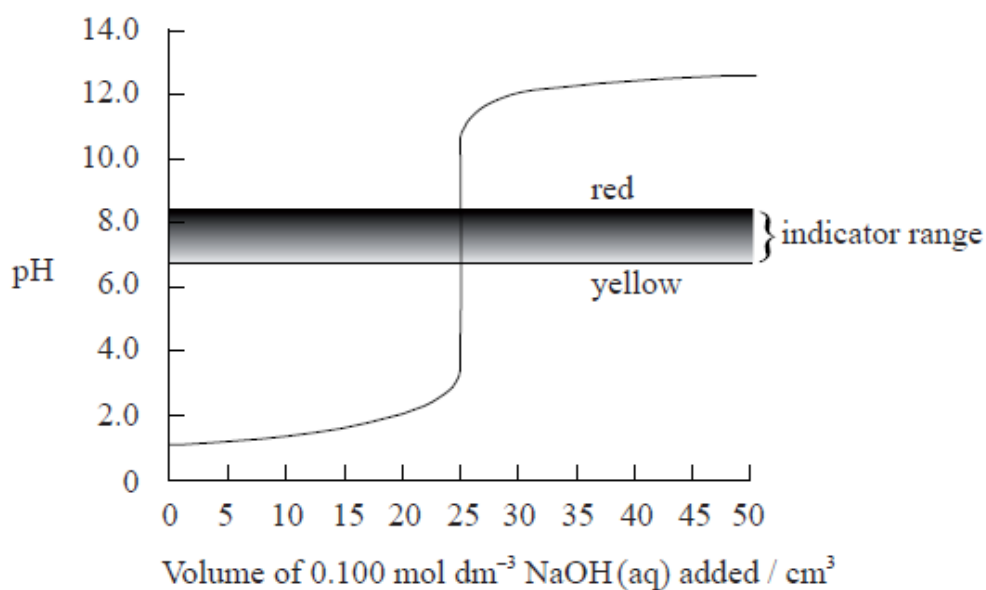
Sketch the graph that would be obtained for the titration of 25.0 cm^3 of 0.100 mol dm^{-3} propanoic acid with 0.100 mol dm^{-3} potassium hydroxide using bromophenol blue as an indicator. (The pH range of bromophenol blue can be found in Table 16 of the Data Booklet). [3]
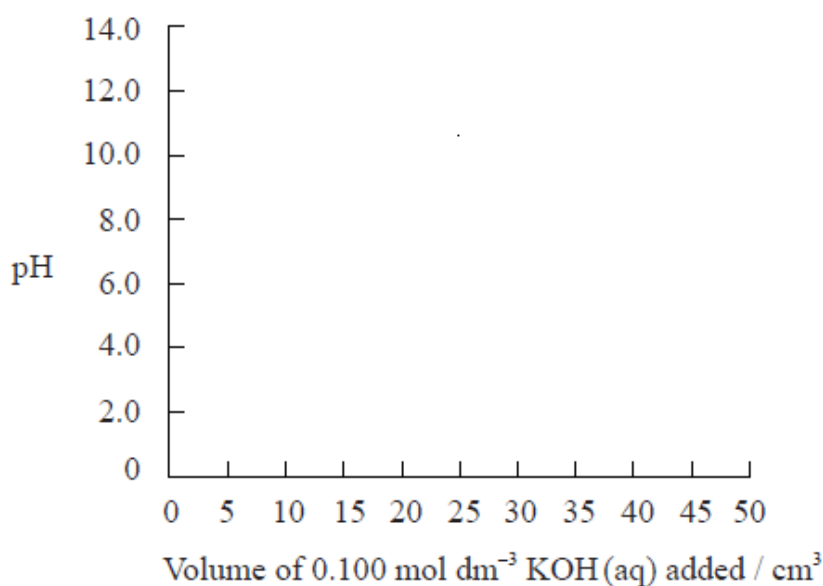
HL SECTION A 09s

5. Sodium oxide, Na$_2$O, is a white solid with a high melting point.

(c) (i) State the acid-base nature of sodium oxide. *[1]*

. . . .

(ii) State the equation for the reaction of sodium oxide with water. *[1]*

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HL SECTION A 08w

3. (a) The equation for the reaction that occurs when ammonia gas dissolves in water is shown below.

$$NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$
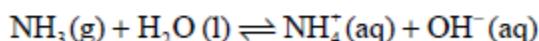
(i) Define the term *Brønsted-Lowry acid*. [1]

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(ii) Identify a conjugate pair present in the above equation. [1]

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(iii) Identify **one** species in the equation above that acts as a Lewis base and name the type of bond it forms in the reaction. [2]

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(b) The ionization constant expression for methylamine is shown below.

$$K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]} = 4.37 \times 10^{-4} \ mol \ dm^{-3}$$

(i) Write an equation for the reaction between methylamine and water. [1]

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(ii) Calculate the pOH of a $0.0500 \ mol \ dm^{-3}$ aqueous solution of methylamine. State any assumptions made in your calculation. [4]

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(iii) Calculate the pH of a buffer solution made by mixing together 0.025 mol of CH_3NH_2 and 0.010 mol of HCl in 1.0 dm^3 of solution. [5]

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HL SECTION A 08s

3. Lactic acid, $CH_3CH(OH)COOH$, is a weak monoprotic acid ($K_a = 1.40 \times 10^{-4}$ mol dm^{-3}).

(a) Write an equation for the reaction of lactic acid with water. [2]

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(b) State the ionization constant expression, K_a, for lactic acid. [1]

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(c) Calculate the $[H^+]$ of a 0.250 mol dm^{-3} solution of lactic acid. [2]

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(d) Calculate the $[H^+]$ of a buffer solution of 1.00 dm^3 volume, containing 0.250 mol dm^{-3} of lactic acid and 0.125 mol dm^{-3} of sodium lactate. [2]

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(e) Explain why the two values of $[H^+]$ in (c) and (d) are different using Le Chatelier's principle. [2]

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HL SECTION A 07s

TABLE 16

5. Propanoic acid is classified as a weak acid.

(a) State the meaning of the term *weak acid*. [1]

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(b) State an equation for the reaction of propanoic acid with water. Identify **one** conjugate *Brønsted-Lowry* pair. [2]

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(c) State, giving a reason in each case, **two** methods other than measuring pH, that could be used to distinguish between $0.100 \text{ mol dm}^{-3}$ propanoic acid and $0.100 \text{ mol dm}^{-3}$ nitric acid. [2]

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(d) With reference to Table 16 in the Data Booklet, determine the pH of a $0.100 \text{ mol dm}^{-3}$ solution of propanoic acid. [3]

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HL SECTION A 04w

4. The indicator bromophenol blue, $HIn\ (aq)$, has a form that is yellow and an $In^-\ (aq)$ form that is blue.

    (a) Write an equation to show how bromophenol blue acts as an indicator. [1]

    . . . .

    (b) State and explain the colour of bromophenol blue [3]

        (i)   on the addition of a strong acid.

        . . . .

        . . . .

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        (ii) at the equivalence point of a titration.

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HL SECTION A 03w

4. (a) (i) Calculate the K_a value of methanoic acid, HCOOH, using table 16 in the Data Booklet. [1]

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(ii) Based on its K_a value, state and explain whether methanoic acid is a strong or weak acid. [2]

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(iii) Calculate the hydrogen ion concentration and the pH of a $0.010 \text{ mol dm}^{-3}$ methanoic acid solution. State **one** assumption made in arriving at your answer. [4]

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(b) Explain how you would prepare a buffer solution of pH3.75 starting with methanoic acid. [3]

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HL SECTION A 02s

3.   In aqueous solution, hydrochloric acid is a strong acid and ethanoic acid is a weak acid.

(a)   Use the Brønsted–Lowry theory to state why **both** substances are classified as acids.   *[1]*

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(b)   Solutions of $0.1 \ mol \ dm^{-3}$ hydrochloric acid and $0.1 \ mol \ dm^{-3}$ ethanoic acid have different electrical conductivities.

(i)   State and explain which solution has the greater conductivity.   *[1]*

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(ii)   Calculate the pH value of $0.1 \ mol \ dm^{-3}$ hydrochloric acid, and suggest a value for the pH of $0.1 \ mol \ dm^{-3}$ ethanoic acid.   *[2]*

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(c)   Write an equation to show the reaction of ethanoic acid with water and classify **each** product as a Brønsted–Lowry acid or base.   *[2]*

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(d)   Use information from Table 16 of the Data Booklet to calculate the value of the ionisation constant, K_a, of ethanoic acid.   *[1]*

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(e)   Write the expression for the ionisation constant, K_a, of ethanoic acid.   *[1]*

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(f) Use your answers to (d) and (e) to calculate the pH value of a 0.050 $mol\,dm^{-3}$ solution of ethanoic acid. [2]

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HL SECTION A 00s

3. (a) In the reaction

$$2H_2O(l) \rightarrow H_3O^+(aq) + OH^-(aq)$$
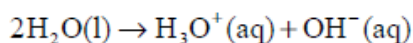
use the Brønsted-Lowry Theory to discuss the acidic and/or basic nature of water. [2]

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(b) What is the conjugate **base** of the hydroxide ion, OH^-? [1]

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(c) State **one** method which could be used to decide whether a solution of 0.10 $mol\,dm^{-3}$ acid is strong or weak. Give the results expected in each case. [3]

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(d) In a titration experiment, 40.0 cm^3 of 0.150 $mol\,dm^{-3}$ NaOH was added to 60.0 cm^3 of 0.200 $mol\,dm^{-3}$ $CH_3(CH_2)_3COOH$ $(K_a = 1.38 \times 10^{-5}\ mol\,dm^{-3})$. Calculate the pH of this mixture. [4]

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HL SECTION A .99w

Propanoic acid has the formula: CH_3CH_2COOH

2. (a) Write an equation to show the ionisation of propanoic acid in water. [1]

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(b) Give the equilibrium expression for this reaction. [1]

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(c) Using information from Table 16 in the Data Booklet, determine the pH of a 0.200 $mol\,dm^{-3}$ aqueous solution of propanoic acid. State the approximation that you have made in arriving at your answer. [3]

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(d) What mass of sodium propanoate, $Na^+CH_3CH_2COO^-$, is required in 500 cm^3 of a solution of 0.200 $mol\,dm^{-3}$ propanoic acid to give a pH of 4.87? [2]

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(e) Explain, with equations, why the pH of the above solution in (d) will remain relatively constant even if small amounts of acid or alkali are added. [2]

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HL B 13s

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(d) (i) Define an *acid* according to the Lewis theory. *[1]*

(ii) State and explain the acid–base character of PCl_3 according to the Lewis theory. *[2]*

HL B 12s

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TABLE 16 of the DATA BOOKLET needed

(d) (i) Hydrogen bromide forms a strong acid when dissolved in water whereas hydrogen fluoride forms a weak acid. Distinguish between the terms *strong acid* and *weak acid*. State equations to describe the dissociation of each acid in aqueous solution. *[3]*

(ii) A student titrated 25.00 cm³ of a 0.100 mol dm⁻³ solution of hydrofluoric acid, HF (aq), with 0.100 mol dm⁻³ NaOH (aq). Some of his data are presented below.

	A	B	C	D	E	F	G	H
1								
2	Volume of NaOH (0.100 mol dm⁻³) / cm³	pH						
3	0.00	2.88						
4	0.50	3.08						
5	1.00	3.39						
6	1.50	3.57						
7	2.00	3.71						
8	2.50	3.82						
9	3.00	3.90						
10	3.50	3.98						
11	4.00	4.05						
12	4.50	4.11						
13	5.00	4.17						
14	5.50	4.22						
15	6.00	4.27						
16	6.50	4.32						
17	7.00	4.36						
18	7.50	4.40						
19	8.00	4.44						
20	8.50	4.48						
21	9.00	4.52						
22	9.50	4.56						
23	10.00	4.59						
24	10.50	4.63						
25	11.00	4.66						
26	11.50	4.70						
27	12.00	4.73						
28	12.50	4.77						
29	13.00	4.80						
30	13.50	4.84						
31	14.00	4.87						

The graph (in cells C2–H16) plots pH against Volume of NaOH (0.100 mol dm⁻³) / cm³.

Two different data points can be used to determine a value for the pK_a of HF (aq). Identify the data points and determine the pK_a using two different calculations. [6]

(iii) Identify an indicator which could be used to find the equivalence point of the titration using Table 16 of the Data Booklet and explain your choice. [2]

HL B 11s

TABLE 16 of DATA BOOKLET is needed

8. (a) (i) Define the terms *acid* and *base* according to the Brønsted-Lowry theory. Distinguish between a weak base and a strong base. State **one** example of a weak base. [3]

(ii) Weak acids in the environment may cause damage. Identify a weak acid in the environment **and** outline **one** of its effects. [2]

(iii) The graph below indicates the pH change during the titration of 20.0 cm³ of 0.100 mol dm⁻³ of CH_3COOH (aq) with 0.100 mol dm⁻³ KOH (aq). From the graph, identify the volume of KOH (aq) and the pH at the equivalence point. [2]
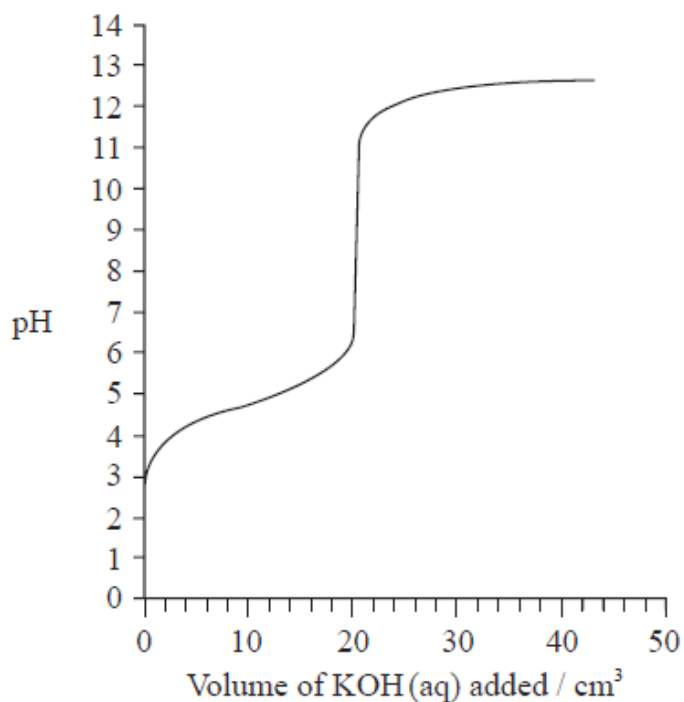
(iv) Explain how the graph could be used to determine the pK_a of ethanoic acid **and** determine the pK_a value for these data. [2]
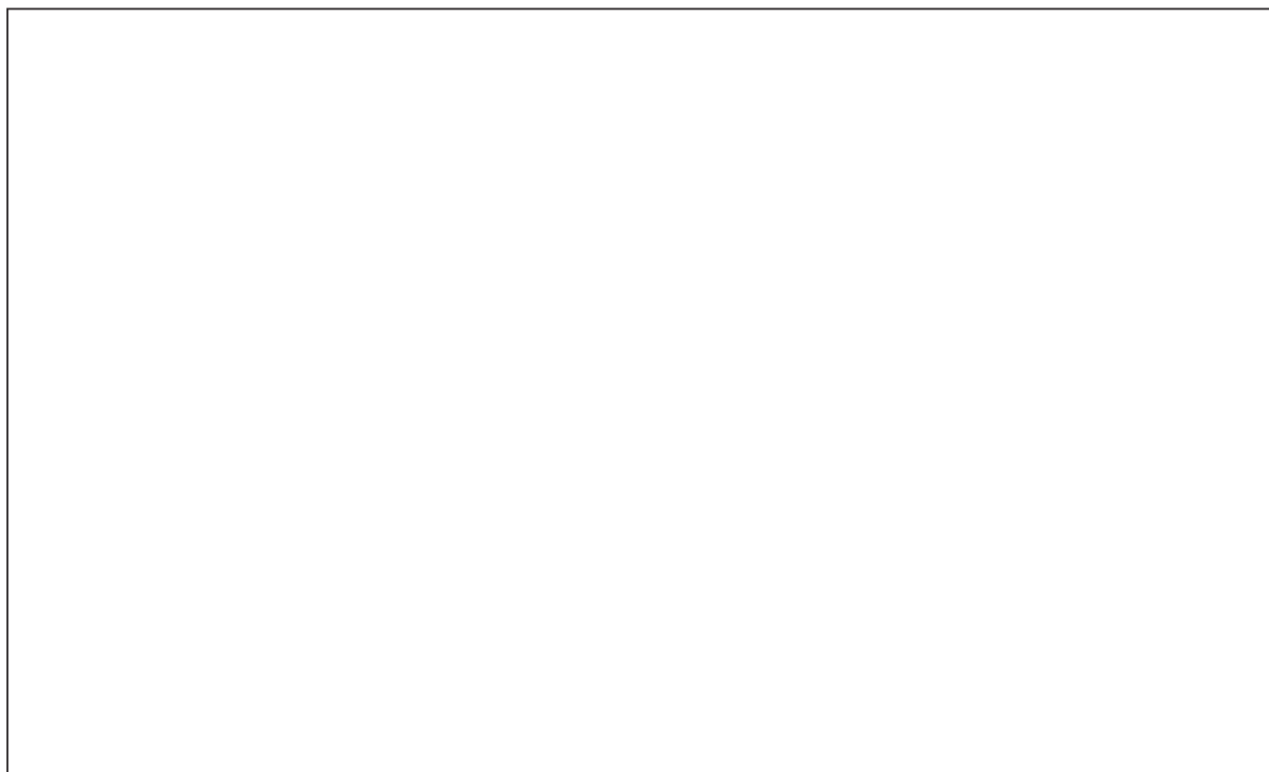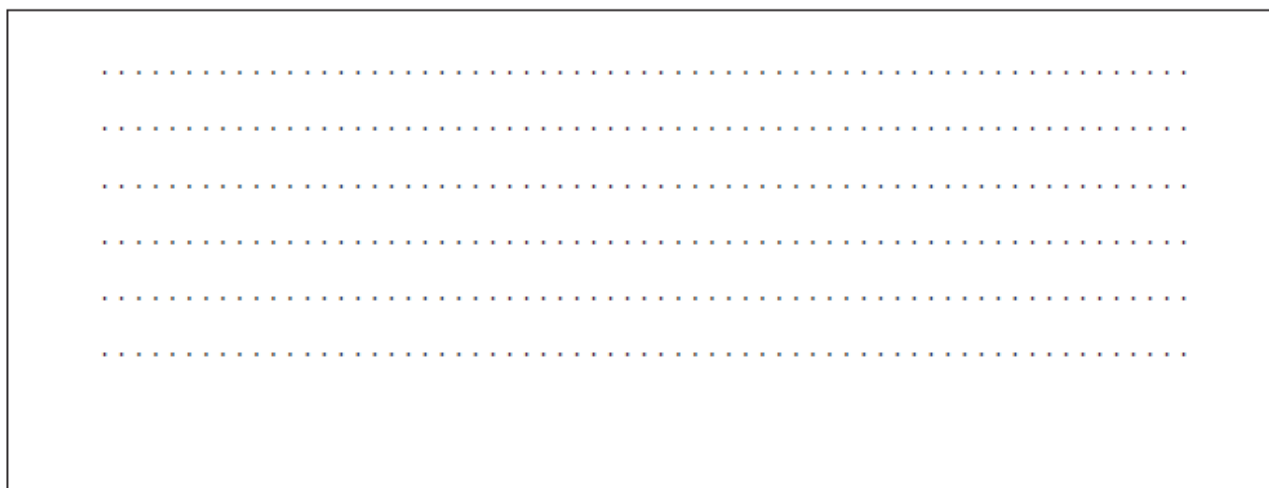
(v) Sketch a graph, similar to the graph on the previous page, to indicate the change in pH during a titration of 25.0 cm^3 of 0.100 mol dm^{-3} HNO$_3$(aq) with 0.100 mol dm^{-3} KOH(aq). On your graph, clearly indicate the starting pH value, the equivalence point, the pH at the equivalence point and the final pH reached. [4]

(b) (i) Describe how an indicator works. [3]

(ii) Using Table 16 of the Data Booklet, identify the most appropriate indicator for the titration of ethanoic acid with potassium hydroxide. Explain your choice. [2]

(c) Explain, using an equation, whether a solution of 0.10 mol dm^{-3} FeCl$_3$(aq) would be acidic, alkaline or neutral. [2]

(d) Determine the pH of the solution resulting when 100 cm^3 of 0.50 mol dm^{-3} HCl(aq) is mixed with 200 cm^3 of 0.10 mol dm^{-3} NaOH(aq). [5]

(d) (i) State whether aqueous solutions of magnesium oxide and magnesium chloride are acidic, alkaline or neutral. *[1]*

    (ii) State an equation for the reaction between magnesium oxide and water. *[1]*

HL B 10w

6

(b) Ammonia can be converted into nitric acid, $HNO_3(aq)$, and hydrocyanic acid, $HCN(aq)$. The pK_a of hydrocyanic acid is 9.21.

    (i) Distinguish between the terms *strong* and *weak acid* and state the equations used to show the dissociation of each acid in aqueous solution. *[3]*

    (ii) Deduce the expression for the ionization constant, K_a, of hydrocyanic acid and calculate its value from the pK_a value given. *[2]*

    (iii) Use your answer from part (b) (ii) to calculate the $[H^+]$ and the pH of an aqueous solution of hydrocyanic acid of concentration $0.108\ mol\,dm^{-3}$. State **one** assumption made in arriving at your answer. *[4]*

(c) A small piece of magnesium ribbon is added to solutions of nitric and hydrocyanic acid of the same concentration at the same temperature. Describe **two** observations that would allow you to distinguish between the two acids. [2]

(d) A student decided to investigate the reactions of the two acids with separate samples of 0.20 mol dm^{-3} sodium hydroxide solution.

   (i) Calculate the volume of the sodium hydroxide solution required to react exactly with a 15.0 cm^3 solution of 0.10 mol dm^{-3} nitric acid. [1]

   (ii) The following hypothesis was suggested by the student: "Since hydrocyanic acid is a weak acid it will react with a smaller volume of the 0.20 mol dm^{-3} sodium hydroxide solution." Comment on whether or not this is a valid hypothesis. [1]

   (iii) Use Table 16 of the Data Booklet to identify a suitable indicator for the titration of sodium hydroxide and hydrocyanic acid. [1]

(e) The graph below shows how the conductivity of the two acids changes with concentration.
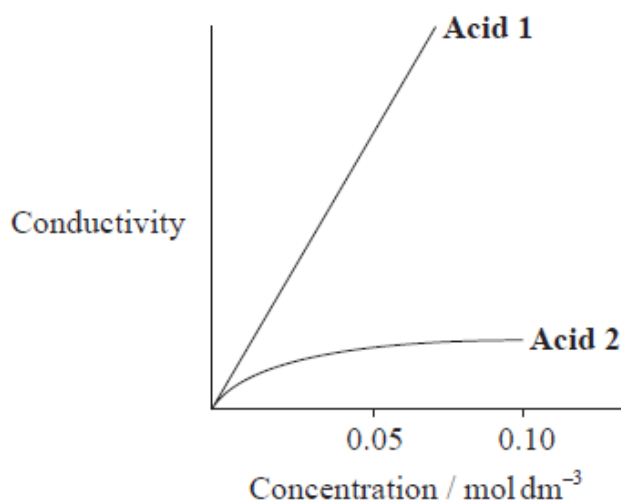
Identify **Acid 1** and explain your choice. [2]

HL B 09w

8. (a) A 25.0 cm³ solution of a weak monoprotic acid, HA (aq), is titrated with 0.155 mol dm⁻³ sodium hydroxide, NaOH (aq), and the following graph is obtained.

Volume of NaOH / cm³
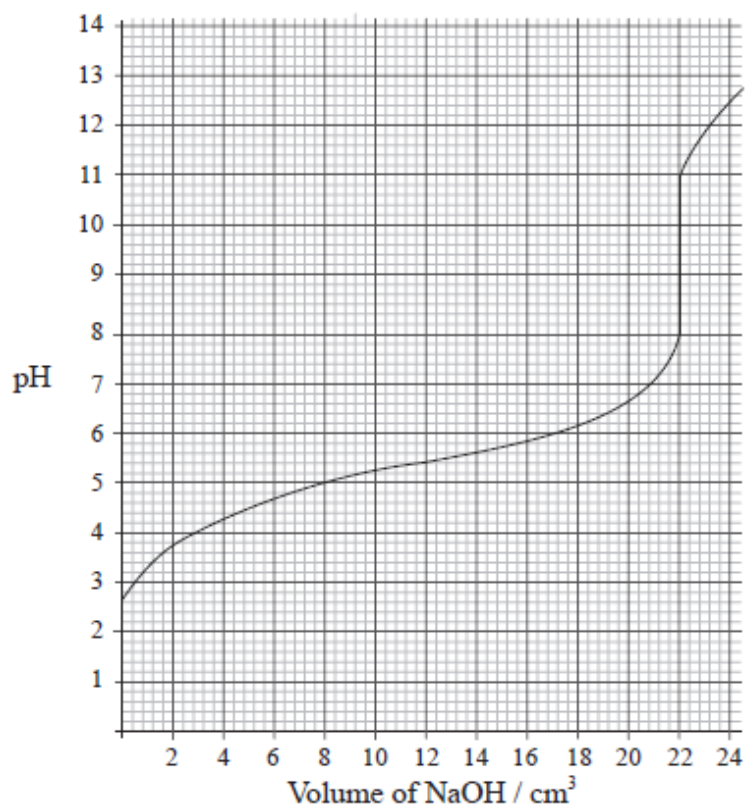
   (i)   Determine the pH at the equivalence point.   [1]

   (ii) Explain, using an equation, why the equivalence point is not at pH = 7.   [3]

   (iii) Calculate the concentration of the weak acid before the addition of any NaOH (aq).   [2]

   (iv) Estimate, using data from the graph, the dissociation constant, K_a, of the weak acid, HA, showing your working.   [3]

   (v)   Suggest an appropriate indicator for this titration.   [1]

   (b) Describe qualitatively the action of an acid-base indicator.   [3]

   (c) (i)   Explain what is meant by the term *buffer solution*.   [2]

        (ii) Calculate the pH of a solution prepared by mixing 50.0 cm³ of 0.200 mol dm⁻³ CH_3COOH (aq) and 50.0 cm³ of 0.100 mol dm⁻³ NaOH (aq), showing your working.   [3]

(d) State whether $AlCl_3$ is acidic, basic or neutral in an aqueous solution. Write an equation to support your answer. [2]

(e) 0.100 mol of ammonia, NH_3, was dissolved in water to make 1.00 dm^3 of solution. This solution has a hydroxide ion concentration of 1.28×10^{-3} $mol\,dm^{-3}$.

    (i) Determine the pH of the solution. [2]

    (ii) Calculate the base dissociation constant, K_b, for ammonia. [3]

HL B 08w

8

For the following questions use information from Tables 16 and 17 in the Data Booklet where relevant.

(g) A 25.0 cm^3 sample of an aqueous solution of benzoic acid needed 17.0 cm^3 of 0.0300 $mol\,dm^{-3}$ aqueous sodium hydroxide for complete neutralization.

    (i) Calculate the acid dissociation constant, K_a, for benzoic acid. [1]

    (ii) Write an equation for the neutralization reaction. [1]

    (iii) Identify a suitable indicator for the titration and explain your choice. [2]

    (iv) Calculate the concentration, in $mol\,dm^{-3}$, of the benzoic acid solution. [3]

HL B 07w

9. (a) A solution containing ammonia requires 25.0 cm³ of 0.100 mol dm⁻³ hydrochloric acid to reach the equivalence point of a titration.

(i)   Write an equation for the reaction of ammonia with hydrochloric acid. [1]

(ii) Calculate the amount (in mol) of hydrochloric acid and ammonia that react. [2]

(iii) Calculate the mass of ammonia in the solution. [2]

(b) 0.100 mol dm⁻³ hydrochloric acid solution is added to 25.0 cm³ 0.100 mol dm⁻³ ammonia solution and the pH is recorded until a total of 35.0 cm³ hydrochloric acid has been added.

(i)   Sketch a graph to show how the pH changes as hydrochloric acid is added to the ammonia solution. Use a pH scale of 0 – 14, and an acid volume scale of 0 – 35 cm³. Explain the shape of the curve. [6]

(ii) Use table 17 of the Data Booklet to suggest an indicator that could be used in the titration, explaining your choice. [2]

(c) (i)   State the composition of an acidic buffer solution. [1]

(ii) Suggest the identity of an acid and its amount that could be added to a solution containing 0.10 mol ammonia in order to prepare a buffer. [2]

(iii) Explain how the solution you prepare in (c) (ii) can act as a buffer solution when a strong acid and a strong base are added to separate portions of it. Write an equation to illustrate the buffer action in **each** case. [4]

(iv) Write an equation for the reaction of ammonia with water, and write its K_b expression. [2]

(v)   Given pK_b (ammonia) = 4.75, determine the pH when half the ammonia in a sample of cleaning solution has been neutralized. [3]

HL B 06w

8

(e) A titration was carried out to determine the concentration of 25.0 cm³ of an aqueous solution of ethanoic acid. The pH value of the liquid in the flask was measured during the titration. The results are shown on the graph below.

Volume of NaOH (aq)/cm³
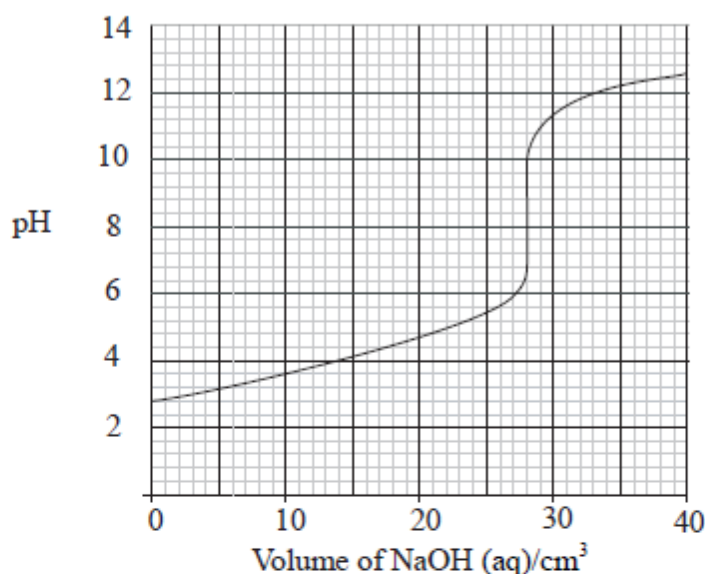
(i) Use the graph to determine the values of pH and $[H^+]$ of the ethanoic acid solution. [2]

(ii) Use the graph to determine the volume of 0.100 mol dm⁻³ sodium hydroxide solution needed to exactly neutralize the ethanoic acid. [1]

(iii) Calculate the concentration, in mol dm⁻³, of the ethanoic acid. [2]

(iv) Identify an indicator that could be used to detect the equivalence point of the titration. Using the formula HIn to represent the indicator, explain why the indicator changes colour during the titration. [3]

(f) The solution formed when the ethanoic acid is exactly half neutralized can act as a buffer solution.

(i) Using information from Table 16 of the Data Booklet, calculate the value of $[H^+]$ in this buffer solution, showing your working. [2]

(ii) Write an equation to show the buffer action of this solution when a small amount of acid is added. [1]

HL B 06s

8. (a) (i) Define the term pH. [1]

(ii) A 25.0 cm³ sample of 0.100 mol dm⁻³ hydrochloric acid was placed in a conical flask, and 0.100 mol dm⁻³ sodium hydroxide is added until a total of 50.0 cm³ had been added. Sketch a graph of pH against volume of NaOH(aq) added, clearly showing the volume of NaOH(aq) needed for complete reaction and the pH values at the start, the equivalence point and finish. [4]

(iii) The experiment in (a) (ii) was repeated, but with a 25.0 cm³ sample of 0.100 mol dm⁻³ ethanoic acid in the conical flask instead of the hydrochloric acid. Use information from Table 16 of the Data Booklet to calculate the pH at the start of the experiment. State the approximate pH value at the equivalence point. [5]

(b) (i) Describe how an indicator, HIn, works. [3]

(ii) Name a suitable indicator for the reaction between ethanoic acid and sodium hydroxide. Use information from Table 17 in the data booklet to explain your choice. [2]

(c) (i) Identify **two** substances that can be added to water to form a basic buffer solution. [1]

(ii) Describe what happens when a small amount of acid solution is added to the buffer solution prepared in (i). Use an equation to support your explanation. [2]

(d) Define the terms *Brønsted-Lowry acid* and *Lewis acid*. For each type of acid, identify one example other than water and write an equation to illustrate the definition. [5]

(e) Predict and explain whether an aqueous solution of 0.10 mol dm⁻³ AlCl₃ will be acidic, alkaline or neutral. [2]

HL B 05w

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(d) The equilibrium reached when ethanoic acid is added to water can be represented by the following equation:

$$CH_3COOH(l) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$
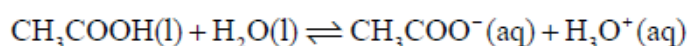
Define the terms Brønsted-Lowry acid and Lewis base, and identify **two** examples of each of these species in the equation. [4]

(e) The pH of a solution is 4.8. Using information from Table 17 of the Data Booklet, deduce and explain the colours of the indicators bromophenol blue and phenol red in this solution. [3]

(f) Calculate the pH of a buffer solution containing 0.0500 mol dm⁻³ of ethanoic acid $(K_a = 1.74 \times 10^{-5})$ and 0.100 mol dm⁻³ of sodium ethanoate. [3]

HL B 05s

6. An experiment was carried out to determine the concentration of an aqueous solution of ammonia by titrating it with a solution of sulfuric acid of concentration $0.150 \, \text{mol dm}^{-3}$. It was found that $25.0 \, \text{cm}^3$ of the ammonia solution required $20.1 \, \text{cm}^3$ of the sulfuric acid solution for neutralization.

(a) Write the equation for the reaction and calculate the concentration, in mol dm^{-3}, of the ammonia solution. [4]

(b) Several acid-base indicators are listed in Table 17 of the Data Booklet. State and explain which one of the following indicators should be used for this experiment: bromocresol green, phenol red or phenolphthalein. [3]

(c) Determine the pOH of a solution with an ammonia concentration of $0.121 \, \text{mol dm}^{-3}$. ($pK_b$ of ammonia is 4.75.) [4]

(d) (i) State what is meant by the term *buffer solution*, and describe the composition of an acid buffer solution in general terms. [3]

(ii) Calculate the pH of a mixture of $50 \, \text{cm}^3$ of ammonia solution of concentration $0.10 \, \text{mol dm}^{-3}$ and $50 \, \text{cm}^3$ of hydrochloric acid solution of concentration $0.050 \, \text{mol dm}^{-3}$. [4]

(e) Choosing suitable examples from the following:

$$NH_3, \, O^{2-}, \, Cu^{2+}, \, OH^-, \, NH_2^-, \, H_2O$$

explain, using a different equation in each case, the meaning of the terms below.

(i) Brønsted-Lowry acid [2]

(ii) Lewis acid [2]

(iii) conjugate acid-base pair (Identify each member of both acid-base pairs.) [3]

HL B 03s

8. (a) Define the term *pH*.   [1]

    (b) Predict whether each of the following solutions would be acidic, alkaline or neutral. In each
         case explain your reasoning.

         (i) $0.1 \, \text{mol dm}^{-3} \, FeCl_3 (aq)$

         (ii) $0.1 \, \text{mol dm}^{-3} \, NaNO_3 (aq)$

         (iii) $0.1 \, \text{mol dm}^{-3} \, Na_2CO_3 (aq)$ [6]

    (c) The following graph shows how the pH changes during the titration of $10 \, \text{cm}^3$ of a solution of a
         weak acid (HA) with $0.10 \, \text{mol dm}^{-3} \, NaOH$.

Volume of $0.10 \, \text{mol dm}^{-3} \, NaOH / cm^3$
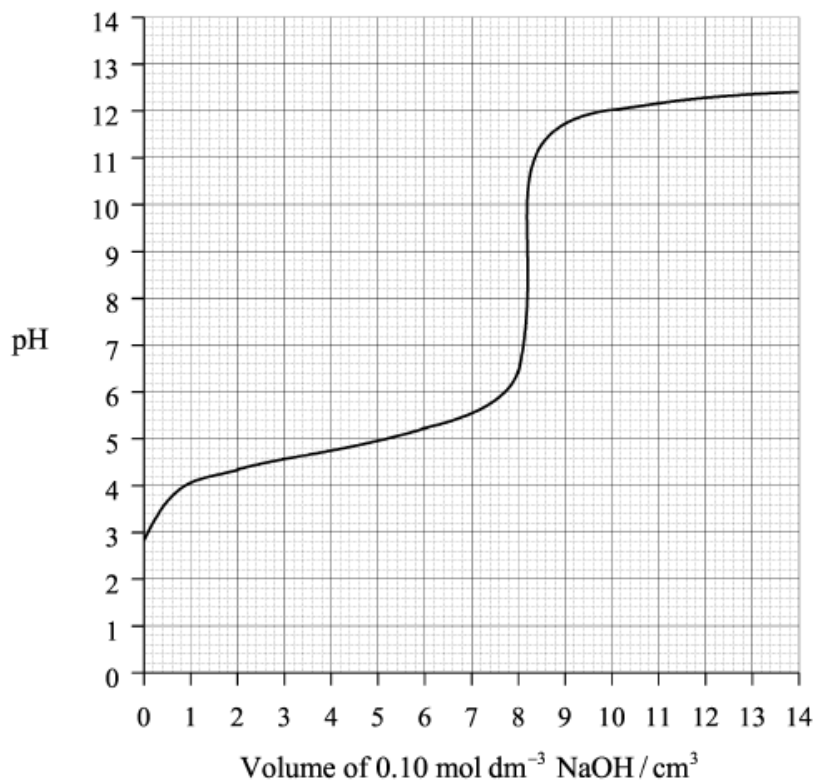
         (i)   State the pH at the equivalence point and explain why the pH changes rapidly in this
                region.   [2]

         (ii)   Calculate the initial concentration of the acid (HA).   [3]

         (iii) Calculate the $[H^+]$ of the acid before any sodium hydroxide is added. Use this value to
                determine the K_a value and the pK_a value of the acid.   [5]

    (d) A buffer solution can be made by dissolving 0.25 g of sodium ethanoate in $200 \, \text{cm}^3$ of
         $0.10 \, \text{mol dm}^{-3}$ ethanoic acid. Assume that the change in volume is negligible.

         (i)   Define the term *buffer solution*.   [2]

         (ii)   Calculate the concentration of the sodium ethanoate in mol dm^{-3}.   [3]

         (iii) Calculate the pH of the resulting buffer solution by using information from Table 16 of
                the Data Booklet.   [3]

8. Benzoic acid is a weak *monoprotic* acid.

    (a) Explain the term *monoprotic* acid. [1]

    (b) The experimentally determined graph below shows the change in pH when $0.10 \ mol \ dm^{-3}$ aqueous sodium hydroxide is added to $25 \ cm^3$ of $0.10 \ mol \ dm^{-3}$ aqueous benzoic acid.

Volume of NaOH added / cm^3
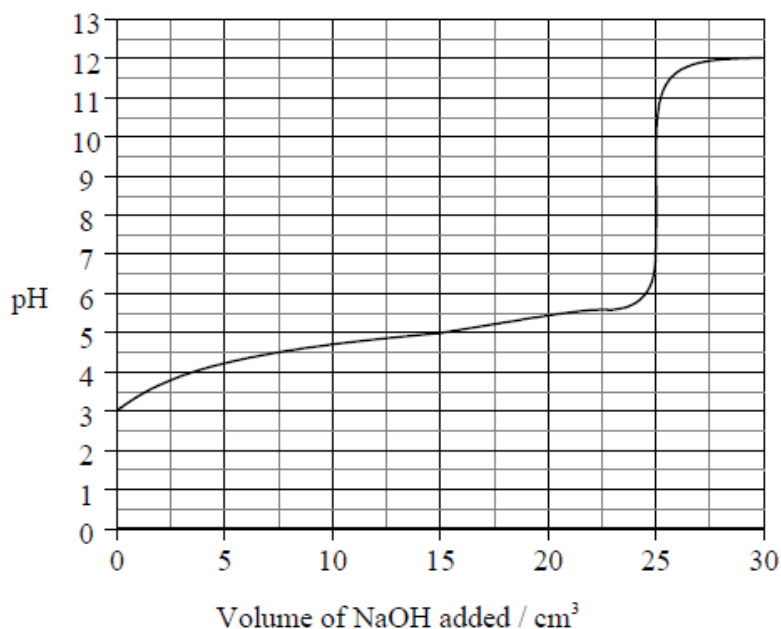
    (i) Calculate the pH when the benzoic acid is half-neutralised and explain how you arrived at your answer. [2]

    (ii) Explain by reference to Table 17 of the Data Booklet why phenolphthalein is a suitable indicator for this titration. [2]

    (c) The experiment was repeated using $25 \ cm^3$ of $0.10 \ mol \ dm^{-3}$ hydrochloric acid instead of benzoic acid.

    (i) Sketch the graph you would expect from the results of this second experiment. [3]

    (ii) State and explain any similarities and differences between the two graphs. [4]

    (d) Explain how an indicator works using

$$HIn \rightleftharpoons H^+ + In^-$$

    where HIn represents the formula of the indicator. [2]

(e)   A mixture of benzoic acid and sodium benzoate can act as a *buffer solution*.

   (i)   Define the term *buffer solution* and describe what happens when acid is added to a buffer solution.   [5]

   (ii)   Calculate the pH of a solution containing 7.2 g of sodium benzoate in $1.0 \, dm^3$ of $2.0 \times 10^{-2} \, mol \, dm^{-3}$ benzoic acid, $(K_a = 6.3 \times 10^{-5} \, mol \, dm^{-3})$ stating any assumptions that you have made.   [6]

HL B 01s

6. (a) State the definitions of an acid and a base according to the Brønsted–Lowry **and** Lewis theories. Give a different equation to illustrate an acid–base reaction for **each** theory, identifying clearly the acid and the base. State the type of bond formed in a Lewis acid-base reaction. [7]

(b) State the difference between a strong acid and a weak acid and give **one** example of **each**. [2]

(c) Explain qualitatively how an acid-base indicator works. [4]

(d) Sodium hydroxide solution is added to aqueous hydrochloric acid. The graph of pH against volume of sodium hydroxide solution added is shown below:
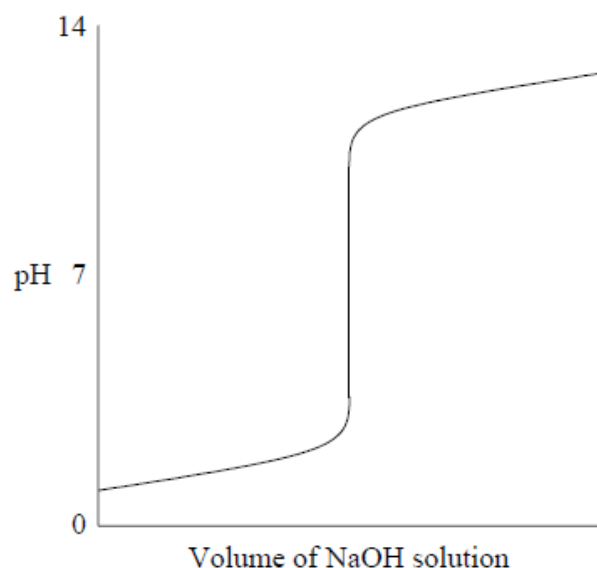
Sketch clearly labelled corresponding graphs for each of the following and suggest a suitable indicator in each case:

(i) The addition of sodium hydroxide solution to aqueous ethanoic acid. [3]

(ii) The addition of ammonia solution to aqueous hydrochloric acid. [3]

(e) $30 \, cm^3$ of $0.100 \, mol \, dm^{-3}$ CH_3COOH is placed in a beaker and mixed with $10 \, cm^3$ of $0.100 \, mol \, dm^{-3}$ $NaOH$.

(i) Explain, with the help of an equation, how the solution formed acts as a buffer solution when a small quantity of acid is added to it. [2]

(ii) Calculate the pH of the buffer solution. (K_a of $CH_3COOH = 1.74 \times 10^{-5} \, mol \, dm^{-3}$) [4]

HL B 00w

5

(c)   Classify the acid–base character of **one** oxide of **each** of the elements in the period from Na to S. Illustrate your answer by writing balanced chemical equations for the reaction of magnesium oxide and of a phosphorus oxide with water. Explain why 'pure' rain water is slightly acidic (pH 5.7).   [6]

(d)   (i)   Write balanced equations to show how aluminium oxide reacts with hydrochloric acid and with sodium hydroxide.   [2]

      (ii)   Write a balanced equation to show what happens when $FeCl_3$ is added to water.   [1]

HL B 00w

7. (a) The titration of 25.0 cm^3 of 0.100 mol dm^{-3} monoprotic acid HA with a base MOH gives the following graph:

volume of base added / cm^3
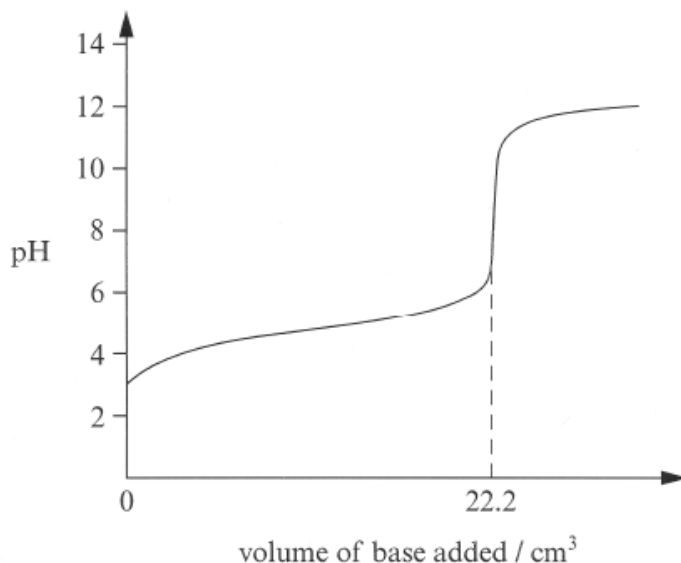
(i) State whether the acid and base used are weak or strong. Explain your answer. [4]

(ii) Use the above data to determine the concentration of the base and give its units. [2]

(iii) Using HIn as an example, explain qualitatively how an acid–base indicator works. [3]

(iv) Write the equilibrium expression for HIn and show how the pK_a value of the indicator relates to the pH value at which it changes colour. [3]

(v) State the pK_a value of an indicator that will be most suitable for use in the above titration. [1]

(b) Explain why an aqueous sodium ethanoate solution is basic whereas an aqueous ammonium ethanoate solution is approximately neutral. [4]

(c) If the pH of water in a swimming pool goes above 8, aluminium sulfate, $Al_2(SO_4)_3$, is added to it to adjust its pH. With the help of formulas and acid–base properties of the ions present, explain how this is achieved. [3]

(d) A household cleaner contains aqueous ammonia. A 2.447 g sample of the cleaner is diluted with water to 20.00 cm^3. This solution requires 28.51 cm^3 of 0.4040 mol dm^{-3} sulfuric acid to reach the equivalence point.

(i) Write a balanced chemical equation for the reaction of sulfuric acid with ammonia to form ammonium sulfate. [1]

(ii) Calculate the amount (moles) of sulfuric acid required for this reaction, and the amount (moles), mass and percentage by mass of ammonia present in the household cleaner. [4]

HL B .99s

7. This question is concerned with acids and bases.

(a) Some of the active ingredients in commercial antacids are $NaHCO_3$, $Al(OH)_3$, and $CaCO_3$.

    (i) Explain why $NaHCO_3$ could be described as both a Brønsted-Lowry base and a Lewis base. Which of these two descriptions applies to $CaCO_3$? [3]

    (ii) Write balanced equations showing the reactions that occur when excess stomach acidity, represented by the formula HCl, is decreased by each antacid. [4]

    (iii) Calculate the number of moles of antacid present in 1 gram of each sample and hence compare their neutralising power. [6]

    (iv) Give **two** reasons why these antacids are used in preference to sodium hydroxide. [2]

(b) Zinc hydroxide is a white solid that reacts with both strong acids and strong bases.

    (i) What name is given to this type of behaviour? [1]

    (ii) Write ionic equations for the reaction of zinc hydroxide with strong acid and strong base. Give another example of a compound which exhibits this behaviour. [5]

    (iii) Define a Lewis acid. Explain why transition metal ions can function as Lewis acids and give an equation to support your answer. [4]

IB HL 8 EQ Paper 2 s99 to s13 incl W Mark Scheme
HL SECTION A 12w

3. (a) (solution containing significant/equal amounts of a) <u>weak</u> acid **and** its salt / (solution containing) strong base to which excess of <u>weak</u> acid has been added / *OWTTE*; **[1]**
*Accept (solution containing) weak acid **and** conjugate base.*
Do not accept descriptions with specific compounds alone (e.g. CH₃COOH and CH₃COONa) unless compounds are stated as <u>weak</u> acid and its salt.
Accept answer such as (solution containing) x mol of <u>weak</u> acid and $\frac{1}{2}x$ mol of strong base.

(b) $M_r(CH_3COOH) = 60.06$ **and** $M_r\, CH_3COONa = 82.04$;
$[CH_3COOH] = 6.66 \times 10^{-1}/0.666\ mol\,dm^{-3}$;

$[CH_3COO^-] = 4.88 \times 10^{-1}/0.488\ mol\,dm^{-3}$;
$[H_3O^+]/[H^+] = (1.8 \times 10^{-5} \times 6.66 \times 10^{-1})/4.88 \times 10^{-1} = 2.46 \times 10^{-5}/0.0000246\ mol\,dm^{-3}$;

$pH = \left(-\log[H_3O^+] = -\log(2.46 \times 10^{-5}) = \right)4.61\,(2dp)$;
Award [5] for correct final answer of pH = 4.61 with some working shown.
Award [2 max] for pH = 4.61 without any working at all shown.
<u>*Two decimal places*</u> *are required for M5.*

OR

$M_r(CH_3COOH) = 60.06$ **and** $M_r\, CH_3COONa = 82.04$;
$[CH_3COOH] = 6.66 \times 10^{-1}/0.666\ mol\,dm^{-3}$;

$[CH_3COO^-] = 4.88 \times 10^{-1}/0.488\ mol\,dm^{-3}$;

$pH = -\log(1.8 \times 10^{-5}) + \log\dfrac{[salt]}{[acid]}$;

$= \left(4.74 + \log\dfrac{0.488}{0.666} = 4.74 - 0.135 = \right)4.61\,(2dp)$; **[5]**
M4 can be scored even if not explicitly stated if M5 is correct based on previous values.
Award [5] for correct final answer of pH = 4.61 with some working shown.
Award [2 max] for pH = 4.61 without any working at all shown.
<u>*Two decimal places*</u> *are required for M5.*

HL SECTION A 11w

4. (a) $(K_w =)[H^+(aq)][OH^-(aq)]$;   [1]

   Do not penalize if (aq) not stated.
   H_3O^+ may be given instead of H^+.
   Do not mark awarded if square brackets are omitted or are incorrect.

   (b) (i) $(pK_b = (14.00 - 7.52 =) 6.48$ and$)$ $K_b = (10^{-6.48}) = 3.3 \times 10^{-7}$;   [1]
             Do not award mark if answer just left as $10^{-6.48}$.

        (ii) $K_b = \dfrac{[HOCl][OH^-]}{[OCl^-]} = \dfrac{x^2}{0.705} = 3.3 \times 10^{-7}$;

             $[OH^-] = 4.8 \times 10^{-4} (mol\, dm^{-3})$;
             Award [2] for correct value of $[OH^-]$.

             OCl^- only partially hydrolysed / x negligible (compared to $[OCl^-]$) / *OWTTE*;   [3]
             Accept $[HOCl] = [OH^-]$.

        (iii) $[H_3O^+]/[H^+] = \dfrac{K_w}{[OH^-]} = \dfrac{1.00 \times 10^{-14}}{4.8 \times 10^{-4}} = 2.1 \times 10^{-11}$;

             $pH = (-\log_{10}[H_3O^+]/-\log_{10}[H^+] = -\log_{10}(2.1 \times 10^{-11}) =)10.68$;   [2]
             Award [2] for correct final answer.

HL SECTION A 10s

3

   (d) (i) $CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$;

             OR

             $CH_3COOH(l) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$;

             OR
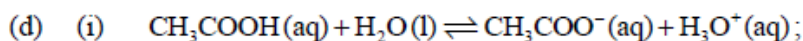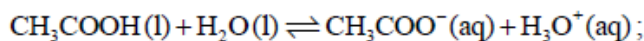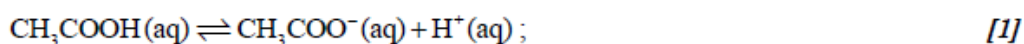
             $CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$;   [1]
             Must include $\rightleftharpoons$.
             Ignore state symbols.

        (ii) $K_a = 10^{-4.76} / 1.74 \times 10^{-5}$

             $1.74 \times 10^{-5} = \dfrac{[H^+]^2}{0.200}$ / $[H^+] = 0.00187$;

             $pH = 2.73$;   [3]
             Award [3] for correct final answer, allow mark for correct conversion of $[H^+]$
             to pH even if $[H^+]$ incorrect.

(e) (initial) $[CH_3COOH] = 0.500$ mol dm^{-3} and) eqm $[CH_3COOH] = 0.200$ mol dm^{-3};
(initial) $[CH_3COO^-] = 0.300$ mol dm^{-3} and) eqm $[CH_3COO^-] = 0.300$ mol dm^{-3};
Allow 0.02 moles and 0.03 moles instead of 0.200 and 0.300 mol dm^{-3}.

$$[H^+] = K_a \frac{[CH_3COOH]}{[CH_3COO^-]} = 1.16 \times 10^{-5} \text{ mol dm}^{-3} / \text{ pH} = pK_a + \log\frac{[SALT]}{[ACID]};$$

pH = 4.94;   [4]
Award [3 max] for correct final answer if no working shown.

(f) (if acid added) $CH_3COO^- + H^+ \rightarrow CH_3COOH$;

(if alkali added) $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$;   [2]

Explanation marks cannot be awarded without equations.
Accept $H^+ + OH^- \rightarrow H_2O$ as OH^- reacts with H^+ in the buffer to form water.

HL SECTION A 09s

2. (a) (i) $CH_3CH_2COOH + H_2O \rightleftharpoons CH_3CH_2COO^- + H_3O^+$
/ $CH_3CH_2COOH \rightleftharpoons CH_3CH_2COO^- + H^+$;   [1]
$\rightleftharpoons$ *required for mark.*

(ii) (pK_a for propanoic acid = 4.87)
$[H^+]^2 = 0.100 \times K_a$;
$[H^+] = 1.16 \times 10^{-3}$ (mol dm^{-3});   [2]

(b) sketch to show:

Volume of 0.100 mol dm^{-3}KOH added / cm^3
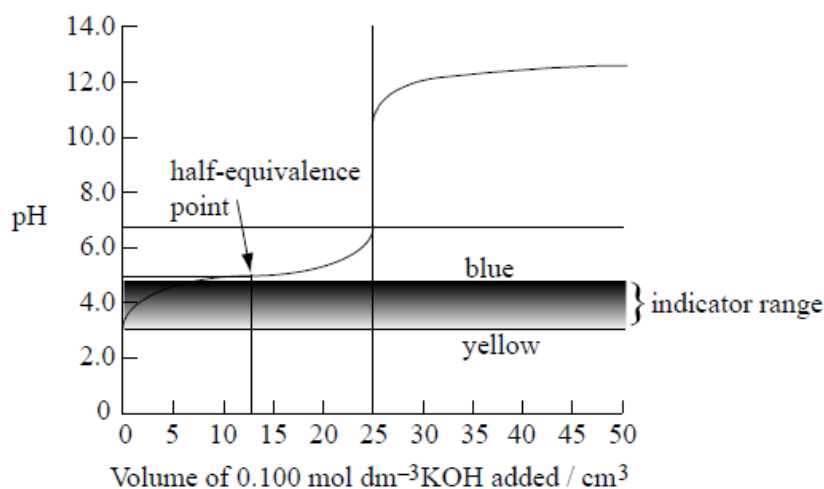
indicator range between pH 3.0 and pH 4.6 (with "yellow" at pH 3.0 and "blue" at pH 4.6);
initial pH of acid at 2.9 ± 1.0 (when no KOH has been added);
half-equivalence point (does not need to be named) at pH 4.9 when 12.5 cm^3 of KOH have been added;
equivalence point at approx pH 8.5 – 9.0 when 25.0 cm^3 of KOH(aq) added;
upper part of curve from 25.0 – 50.0 cm^3 added identical to original curve;   [3 max]
Award [1] each for any three points.

HL SECTION A 09s

5. (a) in the solid state ions are in fixed positions / there are no moveable ions / *OWTTE*;   [1]
        Do not accept answer that refers to atoms or molecules.

    (b) $2O^{2-} \rightarrow O_2 + 4e^- / O^{2-} \rightarrow \frac{1}{2}O_2 + 2e^-$;   [1]
        Accept e instead of e^-.

    (c) (i) basic;   [1]
            Allow alkaline

        (ii) $Na_2O + H_2O \rightarrow 2NaOH / Na_2O + H_2O \rightarrow 2Na^+ + 2OH^-$;   [1]
            Do not accept $\rightleftharpoons$

HL SECTION A 08w

3. (a) (i) proton / H^+ / hydrogen ion donor;   [1]

        (ii) NH_3 and NH_4^+ / H_2O and OH^- ;   [1]

        (iii) NH_3 / OH^- ;
            covalent / dative / co-ordinate;   [2]

    (b) (i) $CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$;   [1]
            Do not accept $\rightarrow$ in place of $\rightleftharpoons$
            Ignore state symbols.

        (ii) $[OH^-] = \sqrt{K_b \, [CH_3NH_2]} / \sqrt{(4.37 \times 10^{-4}) \times 0.0500}$;
            $[OH^-] = 4.67 \times 10^{-3} (mol\,dm^{-3})$;
            $pOH = 2.33$;
            Correct final answer scores [3] marks

            $[CH_3NH_2]$ does not change / *OWTTE* / appropriate mathematical abbreviation;   [4]
            If quadratic equation used award final mark.

        (iii) n $CH_3NH_3^+$ at eq (= n HCl added) =0.010;
            n CH_3NH_2 at eq (=0.025 − 0.010) =0.015;
            $[OH^-] = \dfrac{K_b \times [CH_3NH_2]}{[CH_3NH_3^+]} / \dfrac{4.37 \times 10^{-4} \times 0.015}{0.010} / = 6.6 \times 10^{-4}$ (mol dm^{-3});
            $pOH = 3.2 / [H^+] = 1.5 \times 10^{-11}$;
            $pH = 10.8$;   [5]
            Correct final answer scores full marks.

HL SECTION A 08s

3. (a) $CH_3CH(OH)COOH + H_2O \rightleftharpoons CH_3CH(OH)COO^- + H_3O^+$ **[2]**

Award [1] for correct reactants and products and [1] for $\rightleftharpoons$ sign.
Accept $CH_3CH(OH)COOH \rightleftharpoons CH_3CH(OH)COO^- + H^+$.

(b) $K_a = \dfrac{[CH_3CH(OH)COO^-][H_3O^+]}{[CH_3CH(OH)COOH]}$; **[1]**

Accept $[H^+]$ instead of $[H_3O^+]$.

(c) $1.40 \times 10^{-4} = \dfrac{[H^+]^2}{0.250}$;

$[H^+] = 5.92 \times 10^{-3} (mol\ dm^{-3})$; **[2]**

Award [2] for the correct final answer.

(d) $1.40 \times 10^{-4} \times \dfrac{0.250}{0.125}$;

$[H^+] = 2.80 \times 10^{-4} (mol\ dm^{-3})$; **[2]**

Award [2] for the correct final answer.

(e) equilibrium shifts to the left/side of the reactants;
increase in the concentration of lactate ions reduces the $[H^+]$ in (b) / *OWTTE*; **[2]**

HL SECTION A 07s

5. (a) an acid that partially dissociates/ionizes / doesn't fully dissociate/ionize; **[1]**

(b) $CH_3CH_2COOH + H_2O \rightleftharpoons CH_3CH_2COO^- + H_3O^+ / CH_3CH_2COOH \rightleftharpoons CH_3CH_2COO^- + H^+$;
$\rightleftharpoons$ *required for mark.*
CH_3CH_2COOH and $CH_3CH_2COO^- / H_3O^+$ and H_2O ; **[2]**

(c) conductivity - propanoic acid will be lower because lower ion concentration / less
dissociated;
reaction with metal/metal carbonate/metal hydrogencarbonate - propanoic acid will
react slower/less vigorously because lower $[H^+]$ / less dissociated;
reaction with alkali - temperature change will be less for propanoic acid because lower
$[H^+]$ / less dissociated; **[2 max]**
Award [1] mark each for two.

(d) $(pK_a \text{ (propanoic)} = 4.87)$
$K_a = \dfrac{[CH_3CH_2COO^-][H_3O^+]}{[CH_3CH_2COOH]}$;
$[H_3O^+] = 1.16 \times 10^{-3} (mol\ dm^{-3})$;
$pH = 2.94$; **[3]**
Award [3] for correct answer.

HL SECTION A 04w

4. (a) $HIn(aq) \rightleftharpoons H^+(aq) + In^-(aq)$; [1]
$\rightleftharpoons$ *needed for mark.*
State symbols not essential.

(b) (i) yellow as equilibrium shifts to left to remove (added) $H^+(aq)$; [1]
Colour and explanation needed for the mark.

(ii) green / blue-yellow;
both $HIn(aq)$ and $In^-(aq)$ are present; [2]

HL SECTION A 03w

4. (a) (i) $pK_a = 3.75$, therefore $K_a = 1.78 \times 10^{-4}$ *(accept* 1.8×10^{-4} *)* [1]
No units required.

(ii) weak acid;
less $[H^+]$ / partial dissociation / more reactants / less products / [2]
$K_a \ll 1$ / small K_a;

(iii) $(HCOOH(aq) \rightleftharpoons H^+(aq) + HCOO^-(aq))$

$K_a = \dfrac{[H^+][HCOO^-]}{[HCOOH]} = \dfrac{x^2}{0.010}$;

$(x^2 = 1.78 \times 10^{-6})$

$x = 1.33 \times 10^{-3}$ mol dm^{-3} $= [H^+]$ *(no mark without units)*;
ECF from (a) (i). [4]
No penalty for incorrect significant figures.
$pH = 2.88 / 2.9$ *(ECF)*;
assume $x \ll 0.010 / 25°C$ / negligible dissociation;

(b) add strong base / sodium hydroxide or other named alkali / salt of methanoic acid / HCOONa to methanoic acid;

in equimolar amounts / quantities / so that $[HCOOH] = [HCOO^-]$;
(from K_a expression) $pH = pK_a$ $(= 3.75)$; [3]

HL SECTION A 02s

3. (a) Proton / H^+ donors.   [1]

    (b) (i) hydrochloric acid, greater concentration of ions / *OWTTE* / strong electrolyte /
              dissociates completely.   [1]

         (ii) 1 *[1]*;
               > value for HCl < 7 *[1]*.   [2]

    (c) $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ *[1]* ($\rightleftharpoons$ *is essential*)
                         base   acid   *[1]*   [2]

    (d) $10^{-pK_a} = 1.74 \times 10^{-5}$ (*ignore units*)
         (*minimum 2 sig figs*)   [1]

    (e) $\dfrac{[CH_3COO^-][H^+]}{[CH_3COOH]}$ (*accept [H₃O⁺] in the equation*)   [1]

    (f) $[H^+] = \sqrt{1.74 \times 10^{-5} \times 0.050} = 9.33 \times 10^{-4}$ $(mol\,dm^{-3})$ *[1]*
         $pH = -\log[H^+] = -\log 9.33 \times 10^{-4} = 3.03$ *[1]*   [2]

HL SECTION A 00s

3. (a) Acidic because H^+ donor and basic because H^+ acceptor.   [1]
         Suitable equation OWTTE involving water   [1]

    (b) O^{2-}/oxide ion (*allow* O^{-2})   [1]

    (c) Electrical conductivity   OR   pH meter (or indicator paper)   [1]
         Strong: good conductor   OR   Strong: low pH   [1]
         Weak: poor conductor   OR   Weak: high pH   [1]
         (*Allow full range indicator, do not allow litmus, use judgement on other methods.*)

    (d) Ratio of moles $= 60 \times 0.2 : 40 \times 0.15$ [1]
         (acid : alkali) $=$ 2 : 1

         Acid is in excess and reacts to form salt.
         Moles acid : moles salt $= 1:1$ [1]

         $K_a = \dfrac{[H^+][A^-]}{[HA]}$ / $[H^+] = K_a \dfrac{[HA]}{[A^-]}$ [1]

         $pH = 4.86$ [1]

         (*Weak acid calculation based on excess acid giving a pH of 3.04 [3 max];*
         Weak acid calculation based on total acid giving a pH of 2.89 [2 max];
         Solution based on ½ neutralisation worth [4];
         Working must be shown.)

HL SECTION A .99w

HL B 13s

(d) (i) electron pair acceptor; *[1]*

  (ii) Lewis base;
      has non-bonding/lone pair of electrons; *[2]*
      No ECF from (i).

HL B 12s

6

(d) (i) *Strong acid:* acid/electrolyte (assumed to be almost) 100%/completely
      dissociated/ionized (in solution/water) / *OWTTE* **and** *Weak acid:*
      acid/electrolyte only partially/slightly dissociated/ionized (in solution/water)
      / *OWTTE*;
      $HBr\,(aq) \rightarrow H^+\,(aq) + Br^-\,(aq)$;
      $HF\,(aq) \rightleftharpoons H^+\,(aq) + F^-\,(aq)$; *[3]*

  (ii) *Data points:*
      (0.00, 2.88) **and** (12.50, 4.77);

      For first point also accept volume = 0.00 with pH in range 2.8–2.9.
      For second point also accept either:
      volume in range between 12.5–12.6 and pH 4.8–4.9
      OR
      volume in range between 25.0–25.2 and pH 8.0–10.0

      For (0.00, 2.88):
      $[H^+] = [F^-]$;

      $pK_a = 2pH - 1$ / $K_a = \dfrac{[H^+]^2}{0.100}$;

      For (12.50, 4.77):
      $[HF] = [F^-]$;
      $pK_a = pH$;

      $pK_a = 4.77$; *[6]*
      Accept any value in range 4.7–4.9 if consistent with data points.
      Accept alternative calculation method if other data points from the table or
      graph are used and the pK_a in correct range.

  (iii) bromothymol blue / phenol red / phenolphthalein;
      pK_a/end point of indicator in range 7–10 as pH at equivalence in range 7–10; *[2]*

HL B 11s

8. (a) (i) *Acid*: proton/H^+ donor **and** *Base*: proton/H^+ acceptor;
             Do not accept OH⁻ for base.

             Weak base: (base/electrolyte) partially dissociated/ionized (in solution/water) **and** *Strong base:* (base/electrolyte assumed to be almost) completely/100% dissociated/ionized (in solution/water) / *OWTTE*;
             NH_3 / $CH_3CH_2NH_2$;   *[3]*
             Allow either name or formula or other suitable example.

        (ii) sulfurous acid/H_2SO_3;
              corrodes marble/limestone buildings/statues / leaching in soils / harms/kills plants;

              OR

              nitrous acid/HNO_2;
              corrodes marble/limestone buildings/statues / leaching in soils / harms/kills plants;

              OR

              carbonic acid/H_2CO_3;
              corrodes marble/limestone buildings/statues / acidification of lakes;   *[2]*

              Do not allow oxides (e.g. CO₂ etc.).
              Do not accept just corrodes or damages.

        (iii) *Volume of KOH:* 20 (cm^3);
               Allow any value between 20 and 21 (cm^3).

               pH at the equivalence point: 8.0–10.0;   *[2]*

        (iv) At half-equivalence point $[CH_3COOH] = [CH_3COO^-]$ so $pH = pK_a$;
              $pK_a = 4.7$;   *[2]*
              Accept in range 4.2 to 5.2.
              M2 can only be scored if M1 correct (i.e. no marks for just Data Booklet value of 4.76).

(v)
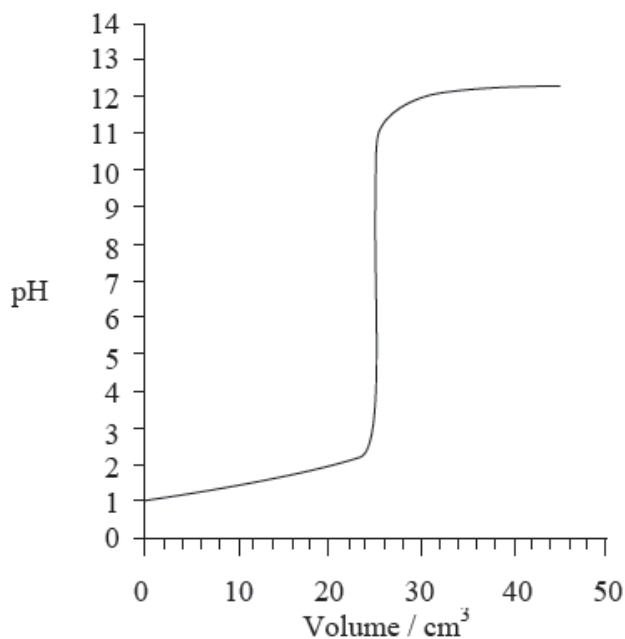
Starting pH: 1;
Equivalence point: pH = 7 **and** 25 cm^3;
Final pH reached: 12–13;
correct curve shape;   [4]
Do not award M4 if turn in curve is seen at low volumes (suggesting weak acid–strong base titration).
Award [4] if values corresponding to M1, M2 and M3 are labelled on graph (e.g using X) and correct shape of curve shown.

(b) (i) HIn is a weak acid / weak base;

$$HIn \rightleftharpoons H^+ + In^-$$
colour 1   colour 2   ;
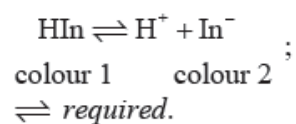
$\rightleftharpoons$ *required.*
Award [2] for M2 alone.

in base equilibrium moves to right / in acid equilibrium moves to left;   [3]

(ii) phenolphthalein;
indicator colour change occurs in range of pH at the equivalence point / *OWTTE*;   [2]
M2 can be scored independently even if indicator is incorrect.

(c)   acidic;
      $[Fe(H_2O)_6]^{3+} \rightleftharpoons [Fe(H_2O)_5(OH)]^{2+} + H^+$ /
      $[Fe(H_2O)_6]^{3+} + H_2O \rightleftharpoons [Fe(H_2O)_5(OH)]^{2+} + H_3O^+$;   [2]
      Accept equations indicating the formation of
      $[Fe(H_2O)_4(OH)_2]^+$
      $[Fe(H_2O)_3(OH)_3]$
      $[Fe(H_2O)_2(OH)_4]^-$
      Do not penalize $\rightarrow$.

(d) $n(HCl) = (0.100 \times 0.50) = 0.050$ (mol);
      $n(NaOH) = (0.200 \times 0.10) = 0.020$ (mol);
      $n(HCl)_{remaining} = (0.050 - 0.020) = 0.030$ (mol);
      $[HCl] = \left(\dfrac{0.030}{0.30}\right) = 0.10$ (mol dm^{-3});
      pH = 1.0;   [5]
      Award [2 max] for just pH = 1.0 without working.

HL B 10w

4

(d)   (i) $MgCl_2$ (weakly) acidic **and** MgO alkaline;   [1]

      (ii) $MgO + H_2O \rightarrow Mg(OH)_2$;   [1]
            Ignore state symbols.

HL B 10w

6

(b)   (i)   strong acid completely dissociated/ionized **and** weak acid partially dissociated/
            ionized;
            $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$;
            $HCN(aq) \rightleftharpoons H^+(aq) + CN^-(aq)$;   [3]
            Insist on both arrows as shown.
            State symbols not needed.
            Accept H_2O and H_3O^+.

      (ii) $K_a = \dfrac{[H^+][CN^-]}{[HCN]}$;
            Allow H_3O^+ instead of H^+.

            $K_a = 10^{-9.21} = 6.17 \times 10^{-10}$;   [2]

(iii) $[H^+] = \sqrt{K_a[HCN]} / \sqrt{(6.17 \times 10^{-10} \times 0.108)}$;

$= 8.16 \times 10^{-6}$;

Allow in the range 8.13×10^{-6} to 8.16×10^{-6}.

pH = 5.09 ;

OR

$pH = \frac{1}{2}(pK_a - \log[HCN]) / \frac{1}{2}(9.21 - \log 0.108)$;

$= 5.09$;

$[H^+] = 10^{-5.09} = 8.16 \times 10^{-6}$;

Allow in the range 8.13×10^{-6} to 8.16×10^{-6}.

If expression for $[H^+]$ missing but both answers correct, award [3], if one answer correct, award [2].

assume $[H^+] \ll 0.108$ / negligible dissociation; **[4]**

(c) *With HNO_3 :*
faster rate of bubble/hydrogen/gas production;
faster rate of magnesium dissolving;
higher temperature change; **[2 max]**
Accept opposite argument for HCN.
Reference to specific observations needed.
Award [1] if 2 observations given but acid is not identified.

(d) (i) (nitric acid) 7.5 cm^3 ; **[1]**

(ii) not valid as hydrocyanic acid reacts with same volume/ 7.5 cm^3 ; **[1]**

(iii) bromothymol blue / phenol red / phenolphthalein; **[1]**

(e) HNO_3 ;
(higher conductivity for solutions with same concentration as) there are more <u>ions</u> in solution; **[2]**

HL B 09w

8. (a) (i) 9.5;   *[1]*
             Accept any value in the range 9.4–9.6.

     (ii) titration involves a weak acid **and** a strong base;
            salt formed at equivalence point is basic due to hydrolysis;
            $A^- (aq) + H_2O (l) \rightleftharpoons HA (aq) + OH^- (aq)$;   *[3]*
            Ignore state symbols.

     (iii) $\dfrac{0.155 \times 22.0 \times 1}{25.0 \times 1}$;
            $= 0.136 \, (mol \, dm^{-3})$;   *[2]*

     (iv) at half neutralization point, $pH = pK_a$;
            $pK_a = 5.3$;
            Accept any value in the range 5.2–5.4.

            $K_a = 5.0 \times 10^{-6} \, (mol \, dm^{-3})$;   *[3]*
            Accept calculations based on initial pH or on pH of salt.

     (v) phenolphthalein;   *[1]*
            Accept thymolphthalein.
            Allow ECF from (a)(i).

  (b) $HIn (aq) \rightleftharpoons H^+ (aq) + In^- (aq)$;
      colour A   colour B
      in presence of acid/H^+, equilibrium shifts to left, colour A;
      in presence of base/OH^-, equilibrium shifts to right, colour B;   *[3]*

  (c) (i) buffer solution resists change in pH;
            on addition of <u>small</u> amount of acid or base;   *[2]*

     (ii) after mixing $[CH_3COO^-] = [CH_3COOH] = 0.050 \, mol \, dm^{-3}$;
            $K_a = [H^+] / pK_a = pH$;
            $pH = 4.76$;   *[3]*
            Working must be shown to score [3].
            Award [1] if 4.76 stated with no working.

  (d) acidic;
      $[Al(H_2O)_6]^{3+} (aq) \rightleftharpoons [Al(H_2O)_5(OH)]^{2+} (aq) + H^+ (aq)$;   *[2]*
      Accept $AlCl_3 + 3H_2O \rightarrow Al(OH)_3 + 3HCl$.

  (e) (i) $[H^+] = \dfrac{1.00 \times 10^{-14}}{1.28 \times 10^{-3}} = 7.81 \times 10^{-12} \, mol \, dm^{-3}$ / $pOH = -\log 1.28 \times 10^{-3} = 2.90$;
            $pH = (14.0 - 2.90) = 11.1$;   *[2]*
            Award [2] for the correct final answer

     (ii) $K_b = \dfrac{[NH_4^+][OH^-]}{[NH_3]}$;
            $= \dfrac{(1.28 \times 10^{-3})^2}{0.100 - 0.00128} / \dfrac{(1.28 \times 10^{-3})^2}{0.100}$;
            $= 1.66 \times 10^{-5} / 1.64 \times 10^{-5}$;   *[3]*

HL B 08w

8

(g) (i) $K_a (=10^{-4.20}) = 6.31 \times 10^{-5}$ (mol dm^{-3}); [1]

(ii) $C_6H_5COOH + NaOH \rightarrow C_6H_5COONa + H_2O$; [1]
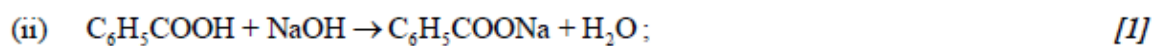
(iii) phenolphthalein;
Accept phenol red.

weak acid-strong base titration / *OWTTE*; [2]

(iv) n NaOH $= 0.0300 \times 0.0170 = 5.10 \times 10^{-4}$ (mol) = n benzoic acid;

$$[HA] = \frac{5.10 \times 10^{-4}}{0.0250} ;$$

$= 0.0204$ (mol dm^{-3}) ; [3]
Correct final answer scores [3] marks.

HL B 07w

9.　(a)　(i)　$NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$;　　　　　　　　　　[1]
　　　　　　States not required for mark

　　　(ii)　$n(HCl) = cV = 0.100\,mol\,dm^{-3} \times 0.0250\,dm^3 = 0.00250\,mol$;
　　　　　　$n(NH_3) = n(HCl) = 0.00250\,mol$;　　　　　　　　　　[2]
　　　　　　ECF

　　　(iii)　$(M(NH_3) = 14.01 + 3(1.01) =)\ 17.04\,/\,17.0\,(g\,mol^{-1})$;
　　　　　　$m(NH_3) = 0.00250\,mol \times 17.04\,g\,mol^{-1} = 0.0426\,g\,/\,0.0425\,g$;　　[2]
　　　　　　ECF

　　(b)　(i)
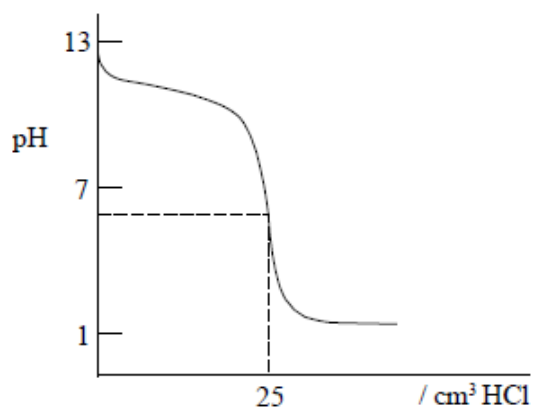
　　　　　　graph starting at $pH < 13$;
　　　　　　Award [0] for pH = 13.

　　　　　　equivalence point $pH < 7$;
　　　　　　Accept anything between 4 − 6

　　　　　　bottom end of graph: pH between 3 and 1;
　　　　　　NH_3 is a weak base / partially dissociated / $[OH^-] < / \ll 0.10\,mol\,dm^{-3}$
　　　　　　(therefore, pH < 13);
　　　　　　NH_4^+ formed is a weak acid / $NH_4^+ \rightleftharpoons NH_3 + H^+$ / NH_4^+ dissociates into a
　　　　　　weak base and a strong acid (thus acidic at equivalence point);
　　　　　　HCl is a strong acid, thus graph finishes close to $pH = 1$;　　　　[6]

　　　(ii)　methyl orange / bromocresol green / bromophenol blue / methyl red;
　　　　　　pK_a of indicator centred around pH at equivalence / end point / indicator pH
　　　　　　range falls where there is a sharp pH change / *OWTTE*;　　　　[2]

(c) (i) weak acid **and** salt of the weak acid/its conjugate base; *[1]*

(ii) HCl / HNO_3 / H_2SO_4;
Amount $< 0.10\,mol$ for HCl / HNO_3 / < 0.05 mol for H_2SO_4; *[2]*

(iii) (added) OH^- reacts with NH_4^+ present/acid of buffer;
(added) H^+ reacts with NH_3 present/base of buffer;
$OH^- + NH_4^+ \rightarrow NH_3 + H_2O$ (strong base replaced by weak base);
$H^+ + NH_3 \rightarrow NH_4^+$ (strong acid replaced by weak acid); *[4]*

(iv) $NH_3\,(aq) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$;
States not required for mark

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]};$$ *[2]*

(v) at half neutralization: $[NH_3] = [NH_4^+]$;
therefore, $K_b = [OH^-]$ / $pK_b = pOH$;
$pH = 14.00 - 4.75 = 9.25$;
Award only [2] for pH = 9.25 if working is not shown *[3]*

HL B 06w

8

(e) (i) $pH = 2.8$;
$[H^+] = 1.58 \times 10^{-3}$ (mol dm^{-3}); *[2]*
Apply ECF from pH to [H$^+$].

(ii) 28 (cm^3); *[1]*

(iii) $nNaOH$ / CH_3COOH $(= 0.100 \times 0.0280) = 2.80 \times 10^{-3}$ (mol);
ECF from value in (ii).
$[CH_3COOH]$ $(= 2.80 \times 10^{-3} \div 0.025) = 0.112$ (mol dm^{-3}); *[2]*
ECF from n above.
Correct final answer scores [2].

(iv) phenolphthalein / phenol red;
$HIn \rightleftharpoons H^+ + In^-$ and HIn and In^- have different colours;
in acid equilibrium shifts to left/to HIn / in alkali equilibrium shifts to right/to In^-; *[3]*

(f) (i) $pK_a = pH$ at half-neutralization;
$[H^+] = 10^{-pH} / 10^{-4.76} = 1.74 \times 10^{-5}$ (mol dm^{-3}); *[2]*

(ii) $CH_3COO^- + H^+ \rightarrow CH_3COOH$ / $CH_3COO^- + H_3O^+ \rightarrow CH_3COOH + H_2O$; *[1]*

HL B 06s

8. (a) (i) $pH = -\log[H^+]$; [1]

    (ii) curve should include the following:
starting pH = 1;
equivalence point: 25.0 cm^3 of NaOH;
pH at equivalence point = 7;
pH to finish = 12 –13;

[4]
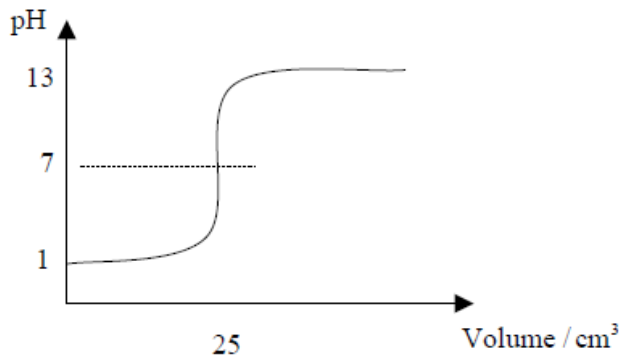
Penalise [1] if profile incorrect.

    (iii) $K_a = 10^{-4.76}/1.74 \times 10^{-5}$;

$$K_a = [H^+]^2 \div [CH_3COOH]/1.74 \times 10^{-5} = \frac{[H^+]^2}{0.100};$$

$[H^+] = 1.32 \times 10^{-3}$ (mol dm^{-3});
starting pH = 2.88;
Accept 3sf.
Award [4] for correct pH.
Allow ECF.
pH at equivalence point: 8 – 9; [5]

(b) (i) HIn is a weak acid;
$HIn \rightleftharpoons H^+ + In^-$ and two colours indicated;

In acid equilibrium moves left or vice versa; [3]

    (ii) phenolphthalein / phenol red / bromothymol blue;
colour change of indicator occurs within the range of pH at equivalence point / on vertical part of graph; [2]

(c) (i) specific examples of weak base and its salt / specific strong acid and weak base; [1]
Name of structure acceptable.
e.g. NH_3 and NH_4Cl.

    (ii) pH changes very little / most acid neutralized by base;
equation from (i); [2]
Any other suitable example.
e.g. $NH_3 + H^+ \rightarrow NH_4^+$ / $NH_4OH + H^+ \rightarrow NH_4^+ + H_2O$.
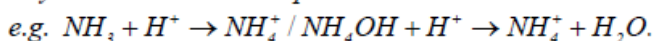
(d) *Brønsted-Lowry acid*
a proton donor;

Lewis acid
electron <u>pair</u> acceptor;

Brønsted-Lowry acid
Any suitable equation;

Lewis acid – BF_3 / $AlCl_3$ /transition metal ions that form complex ion with ligands;
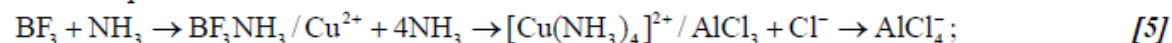
For example
$BF_3 + NH_3 \rightarrow BF_3NH_3$ / $Cu^{2+} + 4NH_3 \rightarrow [Cu(NH_3)_4]^{2+}$ / $AlCl_3 + Cl^- \rightarrow AlCl_4^-$;   *[5]*
Or any suitable equation.

(e) acidic;
$[Al(H_2O)_6]^{3+}$ is (weak) acid due to the formation of H^+ /
$[Al(H_2O)_6]^{3+} \rightleftharpoons [Al(H_2O)_5(OH)]^{2+} + H^+$;   *[2]*

HL B 05w

6

(d) *Brønsted-Lowry acid*
proton donor/*OWTTE*;
CH_3COOH and H_3O^+;

Lewis base
electron pair donor/*OWTTE*;
H_2O and CH_3COO^-;   *[4]*

(e) bromophenol blue is blue **and** phenol red is yellow;
pH of 4.8 is above range of bromophenol blue/bromphenol blue shows its alkaline colour/*OWTTE*;
pH of 4.8 is below range of phenol red/phenol red shows its acidic colour/*OWTTE*;   *[3]*

(f) $K_a = \dfrac{[CH_3COO^-][H^+]}{[CH_3COOH]}$ /rearrangement for $[H^+]$;

$[H^+] = \dfrac{1.74 \times 10^{-5} \times 0.0500}{0.100} = 8.70 \times 10^{-6}$ (mol dm^{-3}) ;

pH $(= -\log[H^+]) = 5.06$;   [3]

OR

$pH = pK_a + \log \dfrac{[CH_3COO^-]}{[CH_3COOH]}$;

$pH = 4.76 + \log\left(\dfrac{0.10}{0.05}\right)$;

$pH = 5.06$;
Accept answer in range 5.0 to 5.1.
ECF from $[H^+]$.
Award [3] for correct final answer.

HL B 05s
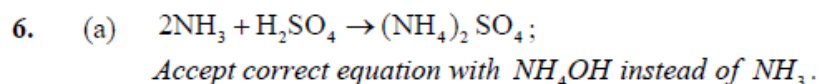
6.   (a) $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$;
Accept correct equation with NH_4OH instead of NH_3.
mol $H_2SO_4 = 0.0201 \times 0.150$;
$2NH_3 = H_2SO_4$ / mol $NH_3 = 6.03 \times 10^{-3}$;
$[NH_3] = 0.241$ (mol dm^{-3}) ;   [4]
Apply −1(SF) if appropriate.
Award [3] for the correct final answer for the concentration calculation.

(b)   bromocresol green;
reaction of weak base and strong acid / *OWTTE*;
pH range of bromocresol green is 3.8 to 5.4 /occurs at pH < 7;   [3]

(c) $K_b = 10^{-4.75} = 1.78 \times 10^{-5}$;

$K_b = \dfrac{[NH_4^+][OH^-]}{[NH_3]}$ / $[OH^-] = \sqrt{K_b[NH_3]}$;

$[OH^-] = \sqrt{1.78 \times 10^{-5} \times 0.121}$;
pOH = 2.83;   [4]
Award [4] for the correct final answer.
Allow ECF, for example any correct conversion of $[OH^-]$ to pOH.

(d)   (i)   a solution which resists change in pH / changes pH very slightly /
keeps pH constant / *OWTTE*;
when <u>small</u> amounts of acid or base are added;
weak acid and its salt / weak acid and its conjugate base;   [3]

(ii)   mol $NH_3 = 0.0050$ and mol HCl $= 0.0025$;
$[NH_4^+] = [NH_3]$;
$[OH^-] = K_b = 1.78 \times 10^{-5}$;
(pOH = 4.75 so) pH = 9.25 (allow 9.2 to 9.3);   [4]
Award [4] for correct final answer.
Accept other valid methods such as Henderson-Hasselbach equation.

(e) (i) proton donor / *OWTTE*;
         suitable equation;   *[2]*

     (ii) electron <u>pair</u> acceptor / *OWTTE*;
          suitable equation;   *[2]*

    (iii) two species whose formulas differ by H^+ / *OWTTE*;
          suitable equation;
          both acid-base pairs correctly identified;   *[3]*

Examples of suitable equations:

$$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$$

  base1   acid2   acid1   base2

$$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$$

$$Cu^{2+} + 4NH_3 \rightarrow Cu(NH_3)_4^{2+}$$

$$NH_2^- + H_2O \rightarrow NH_3 + OH^-$$

$$O^{2-} + H_2O \rightarrow 2OH^-$$
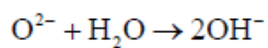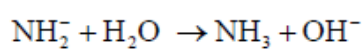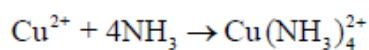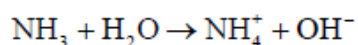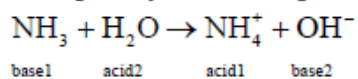
Each equation must use at least one of the species in the question.
Each equation must be balanced.

HL B 03s

8. (a) $pH = -\log_{10}[H^+]$;   [1]

   (b) (i) acidic;
          $Fe(H_2O)_6^{3+}$ is a weak acid / Fe^{3+} reacts with OH^- / equation to show formation
          of HCl or H^+;   [2]
          "FeCl₃ is acidic" is not acceptable.

       (ii) neutral;
           $NaNO_3$ / sodium nitrate is formed from strong base and strong acid / ions do not
           hydrolyse;   [2]

       (iii) alkaline;
            As CO_3^{2-} is weak base / combines with H^+ / equation showing formation of OH^-;   [2]
            Acidic, neutral, alkali mark in each case is independent of reason.

   (c) (i) 8.7 ± 0.7;
           low $[H^+]$ thus small addition of OH^- has great effect / OH^- increases rapidly as
           NaOH is a strong base / logarithmic nature of pH;   [2]

       (ii) volume of NaOH = 8.2 cm³ (exact);
            amount of NaOH $= \dfrac{8.2}{1000} \times 0.1 = 0.00082$ mol;
            $[HA] = \dfrac{0.00082}{0.010} = 0.082$ mol dm⁻³ / 0.082 M;   [3]
            Correct answer [3], units needed for last mark.

       (iii) correct pH reading from graph (2.9) (allow 2.8 or 3.0);
             thus $[H^+] = 1.26 \times 10^{-3}$ (mol dm⁻³);
             $K_a = \dfrac{10^{-2.9} \times 10^{-2.9}}{0.082}$;
             $= 1.9 \times 10^{-5}$ (mol dm⁻³);
             $pK_a = 4.71$ [5]
             Accept 4.7 and allow ECF from (ii).

             If pH given as 2.8, $K_a = 3.06 \times 10^{-5}$ and $pK_a = 4.51$
             If pH given as 3.0, $K_a = 1.22 \times 10^{-5}$ and $pK_a = 4.91$

             If half equivalence method used:
             volume = 4.1 cm³
             $pK_a = 4.75$
             Award [2] out of last [4].

(d) (i) a solution that resists pH change / maintains a (nearly) constant pH;
          when **small** amounts of acid or alkali are added;   [2]

     (ii) M_r of sodium ethanoate;

          moles of sodium ethanoate $= \dfrac{0.25}{82} = (0.0030)$;

          $[CH_3COO^-] = \dfrac{0.0030}{0.2} = 0.015 \,(mol\ dm^{-3})$ *2 sig figs only*;   [3]

          Do not penalize if already penalized in 3 (b) (i) or 6 (c) (iv).

     (iii) $K_a = \dfrac{[H^+]\,[CH_3COO^-]}{[CH_3COOH]}$ *(or with substituted values)*;

          May be assumed from later work.

          $[H^+] = \dfrac{10^{-4.76} \times 0.10}{0.015} = (1.159 \times 10^{-4})$;

          pH = 3.9(4);   [3]

          Allow ECF throughout (d) (ii) and (iii).

HL B 02w

8. (a)   one replaceable hydrogen atom per molecule;   [1]

    (b)   (i)   4.8 / 4.9;
                value of pH at half-equivalence point;   [2]

          (ii) pK_{in} value / pH range within vertical part of graph;
                colour change occurs with addition of small volume of NaOH / end-point
                of indicator corresponds to equivalence point / colour change occurs as pH
                changes rapidly near equivalence point   [2]

    (c)   (i)
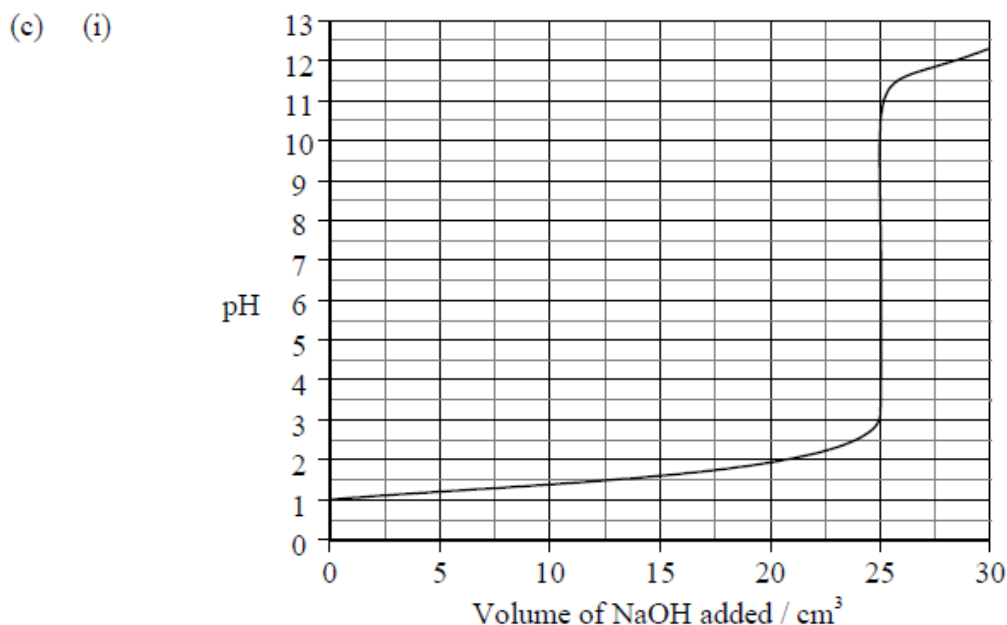
                initial pH = 1; *(accept any value < 3)*
                vertical region from about pH 3-10;
                final pH = 12-13;   [3]

          (ii) *similarities*
                final pH = 12 / same in both because NaOH used;
                same equivalence point / 25 cm^3 alkali needed;

                differences
                initial pH = 1 because $[H^+] = 0.10$ mol dm^{-3} / HCl is strong;
                vertical region longer;   [4]

    (d)   adding H$^+$(aq) shifts equilibrium to the left so HIn colour is seen;
          adding OH$^-$(aq) shifts equilibrium to the right so In$^-$ colour is seen;   [2]

(e) (i) a solution that resists change in pH / whose pH changes slightly;
when a **small** amount of (strong) base or acid is added;

buffer contains (both) a weak acid (HA) and its conjugate base (A^-) / weak base
and its conjugate acid / HA (aq) $\rightleftharpoons$ H^+ (aq) + A^- (aq) ;

if H^+ added, these will react with base of the buffer / H^+ is removed as they
combine with A^- / equilibrium shifts to the left;

so little change in [HA] and [A^-] / H^+ little affected / strong acid H^+
replaced by weak acid HA;   *[5]*

(ii) sodium benzoate $= C_6H_5COONa$

1 mole $= (6\times12)+(5\times1)+12+(2\times16)+23 = 144$ g ;

7.2 g sodium benzoate $= \dfrac{7.2}{144} = 0.05$ moles ;

assume that all of the anion concentration comes from the salt only (as that
from the acid is too small to consider) / assume that the ionization of the
acid is so small it can be ignored / temperature is 298 K;

therefore $6.3\times10^{-5} = \dfrac{[H^+(aq)]\times0.05}{0.02}$;

$[H^+(aq)] = 2.52\times10^{-5}$ mol dm^{-3}; *(accept 2.5×10^{-5} mol dm^{-3})*

pH $= -\log[H^+(aq)] = 4.60$; *(accept 4.6)*   *[6]*

HL B 01s

6. (a) **Bronsted–Lowry**: acid is proton (H^+) donor and base is proton acceptor *[1]*

For example, $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$ *[1]* (*accept other valid example*)

Identification of the acid **and** the base on the left hand side *[1]*

Lewis: acid is electron pair acceptor (electrophile) and base is electron pair donor (nucleophile) *[1]*

For example, $BF_3 + NH_3 \rightarrow BF_3NH_3 / Cu^{2+} + 4NH_3 \rightarrow Cu(NH_3)_4^{2+}$ *[1]*
(*accept other valid example*)

Identification of **both** the acid and the base on the left hand side *[1]*

Bond formed is (dative) covalent / co-ordinate / indication that one species gives bonding electron pair *[1]*   *[7 max]*

(b) Strong acid completely dissociated, weak acid weakly / partially dissociated *[1]*
example of strong **and** weak acid (names or formulas) *[1]*   *[2 max]*

(c) Molecule and anion equilibrium / $HIn \rightleftharpoons H^+ + In^-$ / indicator is a weak acid or weak base *[1]*
Different colours *[1]*
Addition of H^+ or OH^- / acid or alkali causes shift in equilibrium *[1]*
$H^+ \rightarrow$ colour A / OH^- causes colour B *[1]*   *[4 max]*

(d) (i)
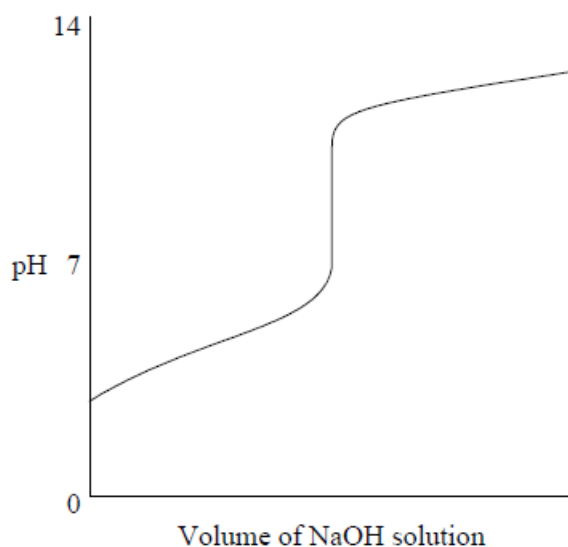
Line starts clearly above $pH = 1$ *[1]*
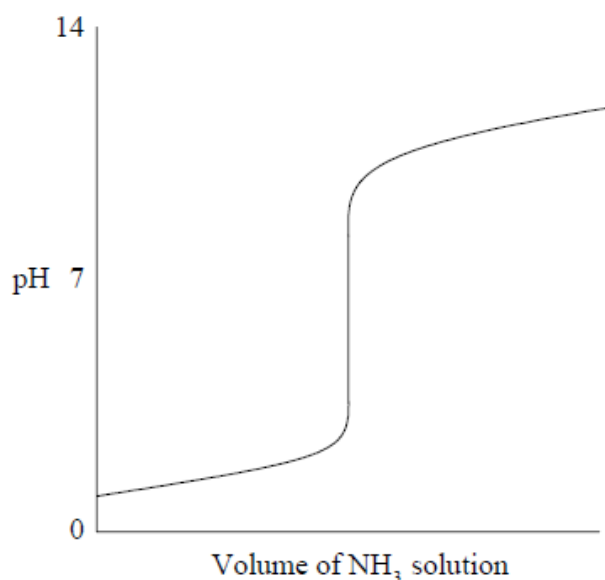More of vertical line in alkaline than acidic region *[1]*
Phenolphthalein *[1]*   *[3 max]*
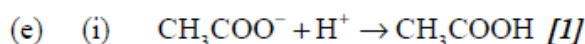
(ii)

Line starts at pH = 1 / same as printed graph *[1]*
More of vertical line in acidic than alkaline region *[1]*
Methyl orange / methyl red / bromophenol blue / bromothymol blue / bromocresol green *[1]*   *[3 max]*

(e) (i) $CH_3COO^- + H^+ \rightarrow CH_3COOH$ *[1]*
Added H^+ ions react with conjugate base / CH_3COO^- *[1]*   *[2 max]*

(ii) $[H^+] = \dfrac{K_a[CH_3COOH]}{[CH_3COO^-]}$ / $pH = pK_a + \log\dfrac{[base]}{[acid]}$ *[1]*

Correct values of CH_3COOH and CH_3COO^- mol or $mol\,dm^{-3}$ *[1]*
Correct value of $[H^+] = 3.48 \times 10^{-5}$ $(mol\,dm^{-3})$ *[1]*
pH = 4.4–4.5 (consequential on $[H^+]$ value) *[1]*   *[4 max]*

HL B 00w
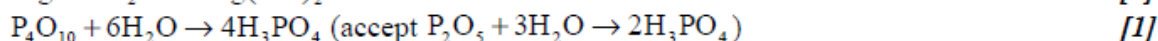
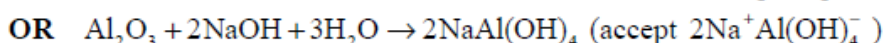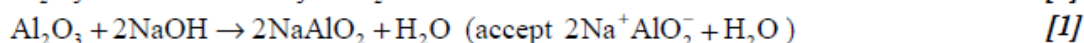(c) Oxides of Na and Mg: basic;   [1]
   oxide of Al: amphoteric;   [1]
   oxide of Si, P and S: acidic.   [1]

   (Award no marks for any one missing; formulas of oxides not asked for.)

   $MgO + H_2O \rightarrow Mg(OH)_2$ [1]
   $P_4O_{10} + 6H_2O \rightarrow 4H_3PO_4$ (accept $P_2O_5 + 3H_2O \rightarrow 2H_3PO_4$)   [1]

   (Accept appropriate acid–base reactions; equations must be balanced to score marks.)

   Slightly acidic from dissolved CO_2 in water forming carbonic acid **OR**
   $CO_2 + H_2O \rightarrow H_2CO_3$ (weak acid) *(No mark if SO_x or NO_x mentioned)*   [1]

(d) (i) $Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$ [1]
         $Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$ (accept $2Na^+AlO_2^- + H_2O$)   [1]
   OR $Al_2O_3 + 2NaOH + 3H_2O \rightarrow 2NaAl(OH)_4$ (accept $2Na^+Al(OH)_4^-$)

   (Accept net ionic equations; equations must be balanced for marks.)

   (ii) $[Fe(H_2O)_6]^{3+} \rightleftharpoons [Fe(H_2O)_5OH]^{3+} + H^+$ [1]
   OR $FeCl_3 + 6H_2O \rightarrow [Fe(H_2O)_6]^{3+} + 3Cl^-$
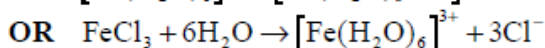
HL B 00w

7. (a) (i)   Weak acid   [1]
         for a strong $0.100 \ mol \ dm^{-3}$, pH would be 1.0
         (thus acid is partially dissociated and weak).   [1]
         pH of final solution closer to 13 means $pOH = 1$ **OR** there is a substantial   [1]
         vertical portion of the pH curve near the end point
         thus a strong base.   [1]

      (ii) $V_b = 0.100 \ mol \ dm^{-3} \times \dfrac{25.0}{22.2}$.   [1]

            $= 0.113 \ mol \ dm^{-3}$ *(accept M as unit).*   [1]

      (iii) $HIn \rightleftharpoons H^+ + In^-$ (generally weak acids)   [1]

           In acid solution, $[H^+]$ will shift equilibrium to the left, indicator is
           predominantly present as HIn which is one colour   [1]

           and in basic solution, H^+ ions are removed by OH^-, shifting equilibrium to the
           right; In^- predominantly present which is another colour.   [1]

      (iv) $K_a = \dfrac{[H^+][In^-]}{[HIn]}$ [1]

           At end point intensity of HIn and In^- is about the same (where colour change
           takes place), **OR** $[HIn] = [In^-]$ [1]
           therefore $K_a = [H^+]$, and $pK_a = pH$.   [1]

      (v) At equivalence point $pH \approx 8$ / between 8 and 9, so use indicator of $pK_a \approx 8$.
           (Accept 9.)   [1]

(b) Na^+ ions are neutral cations from a strong base.   **[1]**

ethanoate ions are basic (from a weak acid)

OR $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$ **OR** hydrolyses to produce weak acid + strong base, therefore basic.   **[1]**

Since ammonium ion is a weak acid **OR** $NH_4^+ \rightleftharpoons NH_3 + H^+$ **[1]**

then the presence of a weak acid and a weak base produces (an approximately) neutral solution.   **[1]**

Alternate answer:

The ions in sodium ethanoate solution are
- CH_3COO^- and Na^+,
- H^+ and OH^-.

Na^+ and OH^- do not combine because NaOH is a strong base. CH_3COO^- and H^+ do combine because CH_3COOH is a weak acid. The solution contains more OH^- than H^+ ions, so is alkaline. The ions in ammonium ethanoate solution are
- CH_3COO^- and NH_4^+,
- H^+ and OH^-.

NH_4OH is a weak base, so the NH_4^+ and OH^- ions combine. As this happens to about the same extent as the combination of CH_3COO^- and H^+, the numbers of H^+ and OH^- ions are about equal, so the solution is approximately neutral.

(c) SO_4^{2-}: neutral anions (from a strong acid).   **[1]**

Al^{3+}: weakly acidic **OR** $Al(H_2O)_6^{3+} \rightleftharpoons \left[Al(H_2O)_5(OH)\right]^{2+} + H^+$ **[1]**

The H^+ reacts with the basic solution, reducing its pH   **[1]**

OR $\left[Al(H_2O)_6\right]^{3+} + OH^- \rightarrow \left[Al(H_2O)_5(OH)\right]^{2+} + H_2O$

(d) (i) $2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$ **[1]**

         (accept $NH_4OH + H_2SO_4 \rightarrow (NH_4)_2SO_4 + H_2O$)

     (ii) $n = cV = 0.4040 \, mol \, dm^{-3} \times 0.02851 \, dm^3 = 0.01152 \, mol$ **[1]**

         mol $NH_3 = 2$ mol acid $= 0.02304$ mol   **[1]**

         $M_r = 17.04$; mass $= 17.04 \, g \, mol^{-1} \times 0.02304 = 0.3926 \, g$ **[1]**

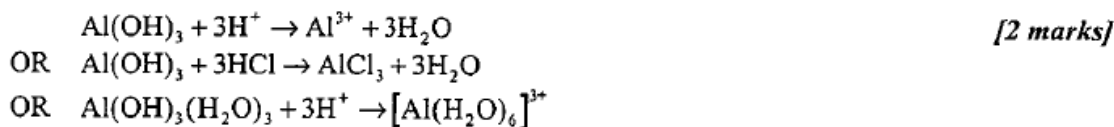   (if $M_r = 17.0$, accept 0.392 g).

         $\% = \dfrac{0.3919}{2.447} \times 100 = 16.04 \, \%$ **[1]**

         (Accept answers within + or − one s.f.; maximum penalty of [1] in question for serious errors in use of significant figures.)

HL B .99s

7. (a) (i) $NaHCO_3$: Brønsted–Lowry base because proton acceptor,   *[1 mark]*
            Lewis base because electron **pair** donor.   *[1 mark]*
            $CaCO_3$: Both   *[1 mark]*

   (ii) $HCO_3^- + H^+ \rightarrow H_2O + CO_2$ *[1 mark]*
   OR $NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$

   Note: If H_2CO_3, no mark given.

        $Al(OH)_3 + 3H^+ \rightarrow Al^{3+} + 3H_2O$ *[2 marks]*
   OR $Al(OH)_3 + 3HCl \rightarrow AlCl_3 + 3H_2O$
   OR $Al(OH)_3(H_2O)_3 + 3H^+ \rightarrow [Al(H_2O)_6]^{3+}$

   Note: *[1 mark]* for correct species, *[1 mark]* for correct balance.

        $CaCO_3 + 2H^+ \rightarrow Ca^{2+} + H_2O + CO_2$ *[1 mark]*
            $+2HCl \rightarrow CaCl_2$

   Note: If H_2CO_3 **again**, do not penalise.

   (iii) $n_{mol}NaHCO_3 = 1\ g \times \dfrac{1\ mol}{84\ g\ mol^{-1}} = 0.012$ *[1 mark]*

        $n_{mol}Al(OH)_3 = 1\ g \times \dfrac{1\ mol}{78\ g\ mol^{-1}} = 0.013$ *[1 mark]*

        $n_{mol}CaCO_3 = 1\ g \times \dfrac{1\ mol}{100\ g\ mol^{-1}} = 0.010$ *[1 mark]*

   $Al(OH)_3$ reacts with 3 mol of H^+ so it is more effective than $CaCO_3$ which reacts with 2 mol of H^+ which is more effective than $NaHCO_3$ which reacts with 1 mol of H^+ *[3 marks]*

   OR $Al(OH)_3$ best   *[1 mark]*
        $CaCO_3$ a further *[2 marks]* if stoichiometry   *[1 mark]*
        $NaHCO_3$ worst   has been used to explain the rest of the   *[1 mark]*
                            order

   Note: If order is wrong look for consequential marking

   (iv) NaOH is a strong alkali   *[1 mark]*
        damages body tissues / corrosive to body   *[1 mark]*
        difficult to store

   Note: Accept other answers on merit

(b) (i) amphoteric/amphiprotic   [1 mark]

(ii) $Zn(OH)_2 + 2H^+ \rightarrow Zn^{2+} + 2H_2O$ *equation 1, balanced 1*   [2 marks]
$Zn(OH)_2 + 2OH^- \rightarrow Zn(OH)_4^{2-}$ (**OR** $ZnO_2^{2-} + 2H_2O$)   [2 marks]

$Al(OH)_3$ / $Pb(OH)_2$ / $Sn(OH)_2$ / Al_2O_3 / $Cr(OH)_3$ / accept other   [1 mark]
suitable oxides/hydroxides (**not** H_2O)

(iii) Electron pair acceptor.   [1 mark]
They have available / empty (*d*) orbitals.   [1 mark]

e.g. $Cu^{2+} + 4NH_3 \rightarrow \left[Cu(NH_3)_4\right]^{2+}$ [2 marks]
choice of base (ligand) 1, formula of suitable complex 1   **Total [25 marks]**

N.B. $\left[Cu(H_2O)_4\right]^{2+}$ or $\left[Cu(H_2O)_6\right]^{2+}$, both are acceptable
Note: Equation does not have to be balanced