AQA A2 CHEMISTRY

TOPIC 4.3

ACIDS AND BASES

BOOKLET OF PAST EXAMINATION QUESTIONS

1. (a) (i) Define the term *Br¤nsted-Lowry acid.*

....

(ii) What is meant by the term *strong* when describing an acid?

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(iii) Give the value of the ionic product of water, K_w, measured at 298K, and state its units.

Value...

Units...

(4)

(b) At 298 K, 25.0 cm^3 of a solution of a strong acid contained 1.50×10^{-3} mol of hydrogen ions.

(i) Calculate the hydrogen ion concentration in this solution and hence its pH.

Hydrogen ion concentration...

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pH...

(ii) Calculate the pH of the solution formed after the addition of 50.0 cm^3 of 0.150 M NaOH to the original 25.0 cm^3 of acid.

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(8)

(c) A solution of a strong acid was found to have a pH of 0.5

(i) Calculate the hydrogen ion concentration in this solution.

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(ii) Calculate the volume of water which must be added to 25.0 cm^3 of this solution to increase its pH from 0.5 to 0.7

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(5)
(Total 17 marks)

2. (a) The pH of a 0.15 M solution of a weak acid, HA, is 2.82 at 300 K.

(i) Write an expression for the acid dissociation constant, K_a, of HA, and determine the value of K_a for this acid at 300 K, stating its units.

Expression for K_a ....

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Value of K_a...

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(ii) The dissociation of HA into its ions in aqueous solution is an endothermic process. How would its pH change if the temperature were increased? Explain your answer.

Effect on pH...

Explanation...

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(8)

(b) Solution **A** contains *n* moles of a different weak acid, HX. The addition of some sodium hydroxide to **A** neutralises one third of the HX present to produce Solution **B.**

(i) In terms of the amount, *n*, how many moles of HX are present in Solution **B**?

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(ii) Determine the ratio $\frac{[HX]}{[X^-]}$ in Solution **B**.

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(iii) Solution **B** has a hydrogen ion concentration of 4.2×10^{-4} mol dm^{-3}. Use this information and your answer to part (b)(ii) to determine the value of the acid dissociation constant of HX.

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(5)

(c) Why is methyl orange **not** suitable as an indicator for the titration of HX with sodium hydroxide?

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(2)

(d) Solution **B** can act as a buffer. Explain what this means and write an equation that shows how Solution **B** acts as a buffer if a little hydrochloric acid is added.

Meaning of buffer...

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Equation...

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(3)
(Total 18 marks)

3. (a) Explain the terms *acid* and *conjugate base* according to the Brønsted-Lowry theory.

Acid ....

Conjugate base ....

(2)

(b) For each of the following reactions, give the formula of the acid and of its conjugate base.

(i) $NH_3 + HBr \rightarrow NH_4^+ + Br^-$

Acid .... *Conjugate base* ....

(1)

(ii) $H_2SO_4 + HNO_3 \rightarrow HSO_4^- + H_2NO_3^+$

Acid .... *Conjugate base* ....

(1)

(c) (i) Write an equation to represent the dissociation of water.

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(1)

(ii) Give the expression for the equilibrium constant, K_c, for the reaction in (c)(i) and use this to derive the expression for the ionic product of water, K_w.

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(3)

(iii) The ionic product of water is 2.92×10^{-14} mol^2 dm^{-6} at 313K. Calculate the pH of water at this temperature.

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(4)

(iv) Given that the pH of water is 7.00 at 298 K, state whether the dissociation of water is endothermic or exothermic. Give a reason for your answer.

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(2)
(Total 14 marks)

4. (a) Give the Brønsted–Lowry definition of a base. State the essential feature of an acid-base reaction in aqueous solution, writing an ionic equation to illustrate your answer.

Definition of a base ....

....

Essential feature ....

....

Equation ....

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(3)

(b) Explain what is meant by the term *weak* when applied to acids and bases.

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(1)

(c) In aqueous solution, the weak acid propanoic acid, $CH_3CH_2COOH(aq)$, produces propanoate ions $CH_3CH_2COO^-(aq)$. Write an expression for the acid dissociation constant, K_a, of propanoic acid and state its units.

Expression for K_a ....

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Units of K_a ....

(2)

(d) (i) Explain what is meant by the term *buffer solution.*

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(ii) Identify two components that could be used to make a buffer solution.

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(iii) Give an example of the use of a buffer solution.

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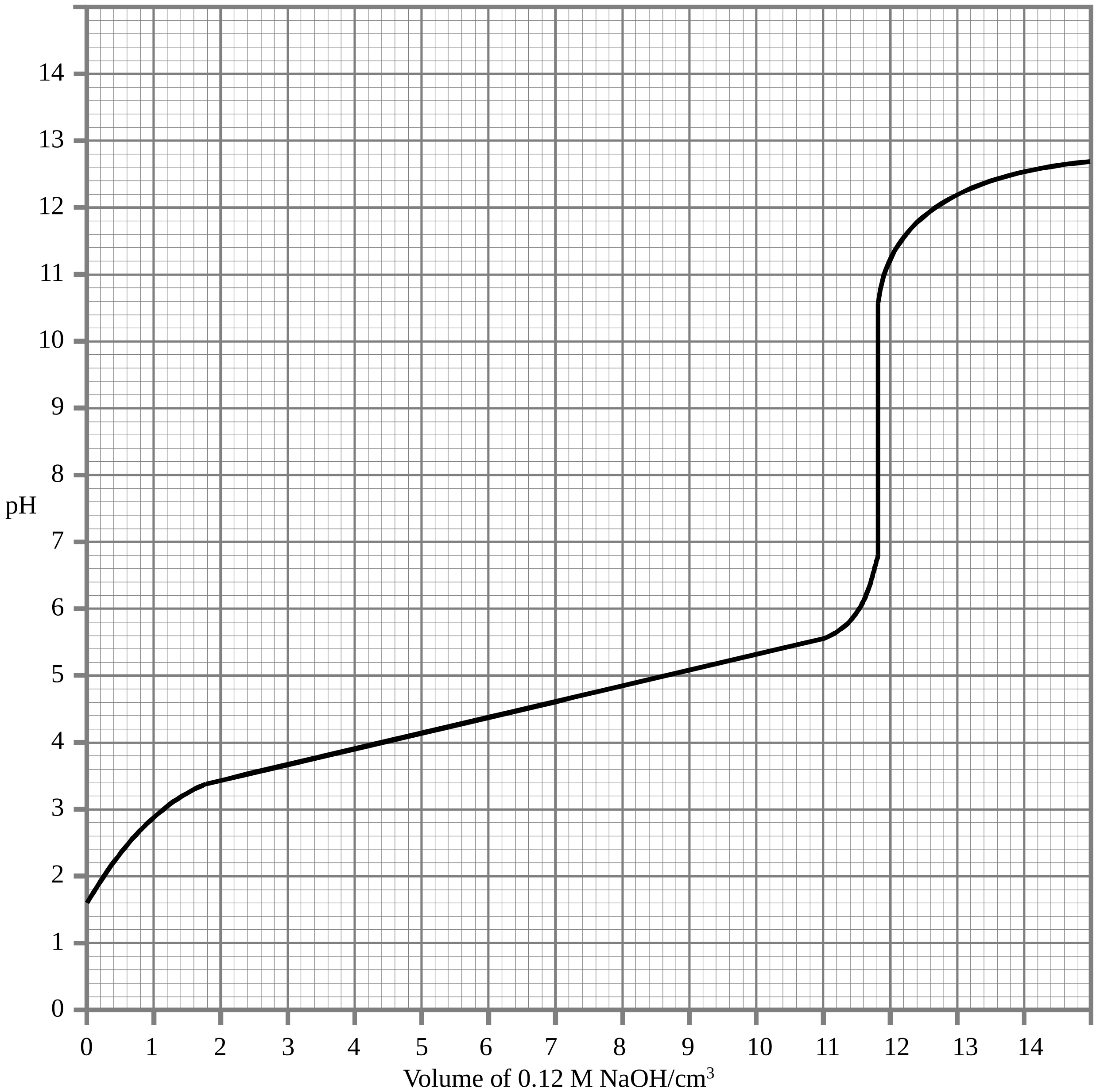
(4)
(Total 10 marks)

5. (a) The graph below shows how the pH changes when 0.12 M NaOH is added to 25.0 cm^3 of a solution of a weak monoprotic acid, HA.



(i) Use the graph to calculate the initial concentration of the acid HA.

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(ii) Write an expression for the dissociation constant. K_a, of the weak acid HA.

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(iii) Determine the volume of sodium hydroxide added when [HA] = [A^-] and use the graph to determine the pH at this point.

Volume of NaOH(aq) added ....

pH ....

(iv) Use your answers to part (a)(ii) and part (a)(iii) to determine the value of K_a for the acid HA.

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(9)

(b) A buffer solution is formed, when approximately half of the original amount of the acid HA(aq) has been neutralised by the base NaOH(aq). Explain how this buffer solution is able to resist change in pH when

(i) a small amount of NaOH(aq) is added,

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(ii) a small amount of HCl(aq) is added.

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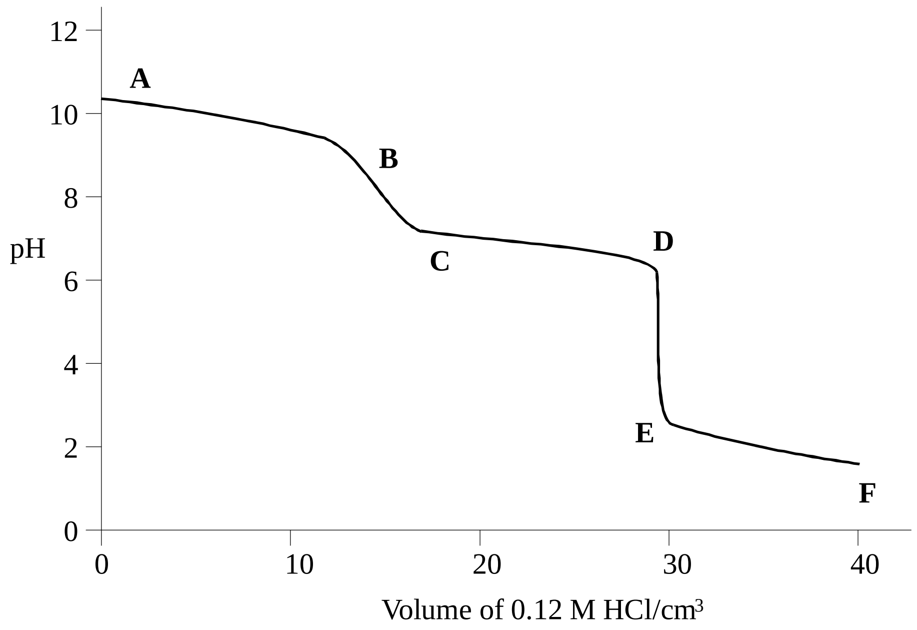
(4)
(Total 13 marks)

6. The graph below shows how the pH changes as 0.12 M HCl(aq) is added to 25.0 cm^3 of a solution of sodium carbonate. There are two end-points. The second end-point is at 30.0 cm^3.



(a) Write equations for the reactions which occur in the solution between point A and point B on the graph and between point C and point D on the graph.

Equation for reaction occurring between A and B.

....

Equation for reaction occurring between C and D.

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(2)

(b) Estimate the minimum volume of hydrochloric acid needed in this experiment for carbon dioxide to be produced from a well-stirred solution of sodium carbonate.

(1)

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(c) Name an indicator which can be used to determine the end-point occurring between points D and E. Explain why this indicator does not change colour between points B and C.

Indicator ....

Explanation ....

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(2)

(d) Use the end-point occurring between points D and E to calculate the concentration of sodium carbonate in the given solution.

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(3)

(e) If the original solution had contained, in addition to sodium carbonate, an equal molar concentration of sodium hydrogen carbonate, at what volumes of hydrochloric acid would the two end-points have been detected?

Volume of HCl(aq) added for first end-point ....

Volume of HCl(aq) added for second end-point ....

(2)
(Total 10 marks)

7. (a) Define the term *Brønsted-Lowry acid.*

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(1)

(b) Write an equation for the reaction between gaseous hydrogen chloride and water. State the role of water in this reaction, using the Brønsted-Lowry definition.

Equation ....

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Role of water ....

(2)

(c) Write an equation for the reaction between gaseous ammonia and water. State the role of water in this reaction, using the Brønsted-Lowry definition.

Equation ....

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Role of water ....

(2)

(d) The ion $H_2NO_3^+$ is formed in the first stage of a reaction between concentrated nitric acid and an excess of concentrated sulphuric acid. In this first stage the two acids react in a 1:1 molar ratio. In the second stage, the $H_2NO_3^+$ ion decomposes to form the nitronium ion, NO_2^+. Write equations for these two reactions and state the role of nitric acid in the first reaction.

Equation for formation of $H_2NO_3^+$ ....

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Role of nitric acid ....

Equation for formation of NO_2^+ ....

(3)

(e) (i) Explain the term *weak acid.*

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....

(ii) Write an expression for the acid dissociation constant, K_a, of HA, a weak monoprotic acid.

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(iii) The value of the acid dissociation constant for the monoprotic acid HX is 144 mol dm^{-3}. What does this suggest about the concentration of undissociated HX in dilute aqueous solution?

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(iv) State whether HX should be classified as a strong acid or a weak acid.
Justify your answer.

Nature of HX ....

Justification ....

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(5)
(Total 13 marks)

8. For acid HA, $K_a = 2.00 \times 10^{-4}$ mol dm^{-3}.

(a) Write an equation for the reaction of HA with NaOH.

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(1)

(b) A solution was formed by adding 15 cm^3 of 0.34 M NaOH to 25 cm^3 of 0.45 M HA.

(i) Calculate the number of moles of A^-(aq) and HA(aq) in this solution. (You should neglect the small number of moles of A^-(aq) formed by ionisation of the remaining HA(aq).)

Number of moles of A^-(aq) ....

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Number of moles of HA(aq) ....

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(ii) Calculate the concentration of A^-(aq) and HA(aq) in this solution.

Number of moles of A^-(aq) ....

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Number of moles of HA(aq) ....

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(iii) Using an expression for K_a, calculate the pH of this solution.

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(9)
(Total 10 marks)

9. (a) Phenol is a *weak acid.* The dissociation of phenol in aqueous solution is represented by the following equation:

$$C_6H_5OH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_6H_5O^-(aq)$$

What is meant by the term *weak acid*?

....

.... **(1)**

(b) (i) Write an expression for the acid dissociation constant, K_a, for phenol.

(1)

(ii) Write an expression linking K_a with pK_a.

.... **(1)**

(iii) The value of the acid dissociation constant, K_a, for phenol is 1×10^{-10} mol dm^{-3}. Calculate the pK_a value of phenol.

.... **(1)**

(iv) Ethanoic acid is a stronger acid than phenol. State whether the p*Ka* value for ethanoic acid will be greater or smaller than that of phenol.

.... **(1)**

(c) The indicator phenolphthalein is a weak acid which can be represented by the formula HIn. It dissociates in solution and has a p*Ka* value of 9.3.

$$\underset{\text{colourless}}{HIn(aq)} \rightleftharpoons H^{+}(aq) + \underset{\text{red}}{In^{-}(aq)}$$

(i) Suggest and explain, with reference to the pK_a value, the pH range of phenolphthalein.

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(2)

(ii) State the colour change that would be observed at the end point in an acid-base titration using phenolphthalein if sodium hydroxide solution were being added from the burette. Explain, in terms of the species present, why this colour is formed.

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(2)

(iii) State why phenolphthalein is unsuitable for a titration between a strong acid and a weak base.

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(1)
(Total 10 marks)

10. (a) (i) Define the term *pH.*

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(ii) Write an expression for the dissociation constant K_a for the weak acid HX.

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(iii) For HX, $K_a = 4.25 \times 10^{-5}$ mol dm$^{-3.}$ Calculate the pH of a 0.45 M solution of this acid.

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(6)

(b) In a 0.25 M solution, a different acid HY is 95% dissociated.

(i) Calculate the pH of this solution.

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(ii) Calculate the value of K_a for the acid HY.

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(6)
(Total 12 marks)

11. (a) Define pK_a

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(1)

(b) Calculate the pH of a 0.52M aqueous solution of the weak monoprotic (monobasic) acid HX (pK_a = 3.72).

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(4)

(c) Write an expression for the acid dissociation constant K_a for HX. Use this to show that the pH of any sample of HX is 3.72 when half of the acid has been neutralised by a solution of sodium hydroxide.

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(3)

(d) Explain why indicators cannot be used to determine the end-point of a titration between a weak acid and a weak base.

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(2)
(Total 10 marks)

12. The value of the acid dissociation constant, K_a, for ethanoic acid is 1.74×10^{-5} mol dm^{-3} at 298 K.

(a) (i) Write an expression for K_a for ethanoic acid.

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(ii) Calculate the pH at 298 K of a 0.220 mol dm^{-3} solution of ethanoic acid.

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(5)

(b) A sample of the 0.220 mol dm^{-3} solution of ethanoic acid was titrated against sodium hydroxide solution.

(i) Calculate the volume of a 0.150 mol dm^{-3} solution of sodium hydroxide required to neutralise 25.0 cm^{3} of the ethanoic acid solution.

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(ii) From the list below, select the best indicator for this titration and explain your choice.

Name of indicator	**pH range**
bromophenol blue	3.0 – 4.6
methyl red	4.2 – 6.3
bromothymol blue	6.0 – 7.6
thymol blue	8.0 – 9.6

Indicator ....

Explanation ....

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(5)

(c) A buffer solution is formed when 2.00 g of sodium hydroxide are added to 1.00 dm^{3} of a 0.220 mol dm^{-3} solution of ethanoic acid.

Calculate the pH at 298 K of this buffer solution.

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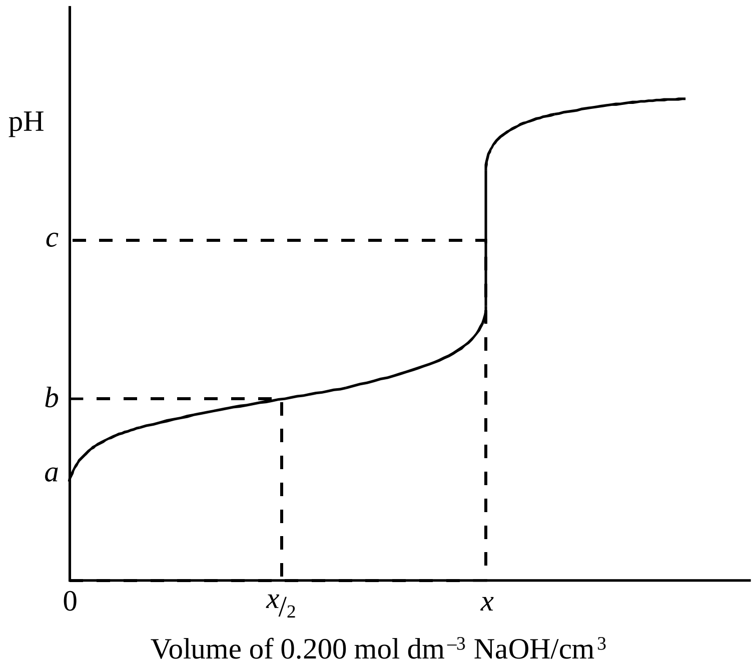
(6)

(Total 16 marks)

13. The sketch below shows the change in pH when a 0.200 mol dm^{-3} solution of sodium hydroxide is added from a burette to 25.0 cm3 of a 0.150 mol dm–3 solution of the weak acid HA at 25 °C.



(a) The volume of sodium hydroxide solution added at the equivalence point is x cm^3. Calculate the value of x.

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(2)

(b) (i) Define the term pH.

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(ii) The pH at the equivalence point is c. Suggest a value for c.

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(iii) Identify a suitable indicator for detecting the equivalence point of the titration.

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(3)

(c) The value of K_c for the weak acid HA at 25 °C is 2.75×10^{-5} mol dm^{-3}.

(i) Explain the term *weak* as applied to the acid HA.

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(ii) Write an expression for K_a for the acid HA.

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(iii) Calculate the pH of the 0.150 mol dm^{-3} solution of acid HA before any sodium hydroxide is added, i.e. the pH at point *a*.

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(5)

(d) Calculate the pH of the solution formed when $x/2$ cm^3 of the 0.200 mol dm^{-3} solution of sodium hydroxide are added to 25.0 cm^3 of the 0.150 mol dm^{-3} solution of HA, i.e. the pH at point *b*.

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(3)
(Total 13 marks)

14. (a) At 50°C, the ionic product of water, K_w, has the value 5.48×10^{-14} mol^2 dm^{-6}.

(i) Define the term K_w

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(ii) Define the term *pH*

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(iii) Calculate the pH of pure water at 50 °C. Explain why pure water at 50 °C is still neutral even though its pH is not 7.

Calculation ....

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Explanation ....

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(5)

(b) At 25°C, K_w has the value 1.00×10^{-14} mol^2 dm^{-6}. Calculate the pH at 25 °C of

(i) a 0.150 mol dm^{-3} solution of sodium hydroxide,

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(ii) the solution formed when 35.0 cm^3 of this solution of sodium hydroxide is mixed with 40.0 cm^3 of a 0.120 mol dm^{-3} solution of hydrochloric acid.

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(8)

(c) In a 0.150 mol dm^{-3} solution of a weak acid HX at 25 °C, 1.80% of the acid molecules are dissociated into ions.

(i) Write an expression for K_a for the acid HX.

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(ii) Calculate the value of K_a for the acid HX at this temperature and state its units.

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(5)
(Total 18 marks)

15. (a) By reference to the forces between molecules, explain why ammonia is very soluble in water.

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(2)

(b) Aqueous solutions of ammonia have a pH greater than 7.

(i) Write an equation for the reaction of ammonia with water.

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(ii) Explain why the pH of a solution containing 1.0 mol dm^{-3} of ammonia is less than 14 at 298 K.

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(3)

(c) An ammonium ion in aqueous solution can behave as a Brønsted–Lowry acid. State what is meant by the term *Brønsted–Lowry acid*.

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(1)

(d) State what is meant by the term *buffer solution*. Identify a reagent which could be added to a solution of ammonia in order to form a buffer solution.

Buffer solution ....

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Reagent ....

(3)

(e) An acidic buffer solution is obtained when sodium ethanoate is dissolved in aqueous ethanoic acid.

(i) Calculate the pH of the buffer solution formed at 298 K when 0.125 mol of sodium ethanoate is dissolved in 250 cm^3 of a 1.00 mol dm^{-3} solution of ethanoic acid. The acid dissociation constant, K_a, for ethanoic acid is 1.70×10^{-5} mol dm^{-3} at 298 K.

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(ii) Write an ionic equation for the reaction which occurs when a small volume of dilute hydrochloric acid is added to this buffer solution.

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(5)
(Total 14 marks)

16. The value of the acid dissociation constant, K_a, for the weak acid HA, at 298 K, is 1.45×10^{-4} mol dm^{-3}.

(a) Write an expression for the term K_a for the weak acid HA.

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(1)

(b) Calculate the pH of a 0.250 mol dm^{-3} solution of HA at 298 K.

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(4)

(c) A mixture of the acid HA and the sodium salt of this acid, NaA, can be used to prepare a buffer solution.

(i) State and explain the effect on the pH of this buffer solution when a small amount of hydrochloric acid is added.

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(ii) The concentration of HA in a buffer solution is 0.250 mol dm^{-3}. Calculate the concentration of A^- in this buffer solution when the pH is 3.59

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(6)
(Total 11 marks)

17. (a) The pH of a 0.120 mol dm^{-3} solution of the weak monoprotic acid, HX, is 2.56 at 298 K.

(i) Write an expression for the term *pH*.

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(ii) Write an expression for the dissociation constant, K_a, for the weak acid HX and calculate its value at 298 K.

Expression for K_a ....

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Calculation ....

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(5)

(b) (i) Write an expression for the ionic product of water, K_w, and give its value at 298 K.

Expression for K_w ....

Value of K_w ....

(ii) Hence, calculate the pH of a 0.0450 mol dm^{-3} solution of sodium hydroxide at 298 K.

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(4)

(c) A titration curve is plotted showing the change in pH as a 0.0450 mol dm^{-3} solution of sodium hydroxide is added to 25.0 cm^3 of a solution of ethanedioic acid, $H_2C_2O_4$
The titration curve obtained has two equivalence points (end points).

(i) Write an equation for the reaction which is completed at the **first** equivalence point.

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(ii) When the **second** equivalence point is reached, a total of 41.6 cm^3 of 0.0450 mol dm^{-3} sodium hydroxide has been added.
Calculate the concentration of the ethanedioic acid solution.

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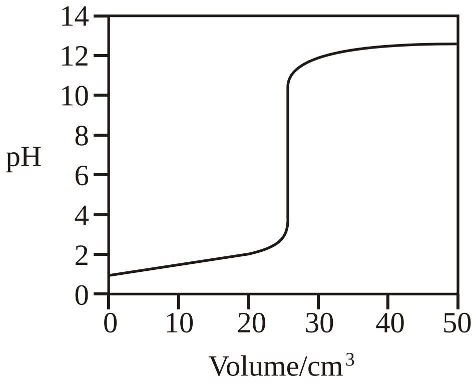
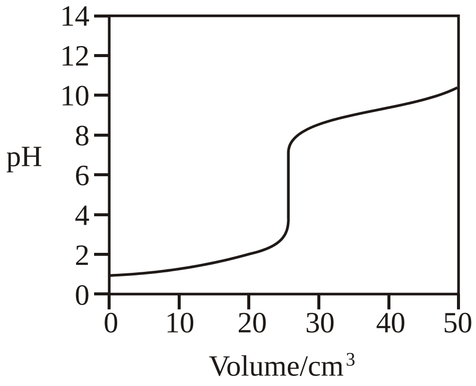
(4)
(Total 13 marks)

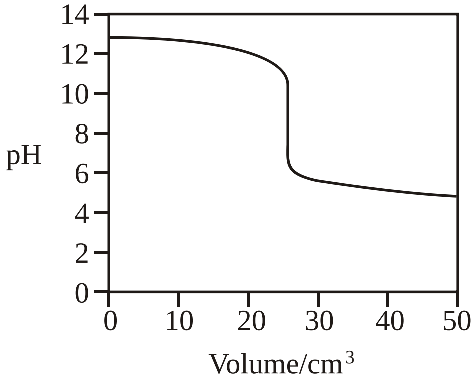
18. (a) Titration curves labelled **A**, **B**, **C** and **D** for combinations of different acids and bases are shown below. All solutions have a concentration of 0.1 mol dm^{-3}.



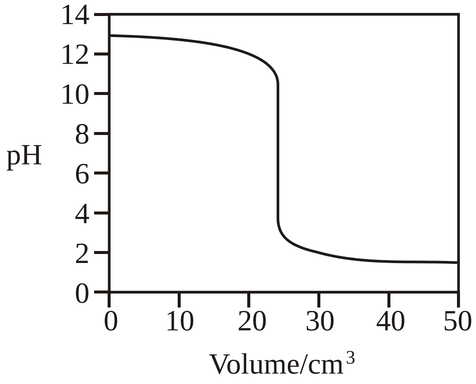
A



B



C



D

(i) Select from **A**, **B**, **C** and **D** the curve produced by the addition of

ammonia to 25 cm^3 of hydrochloric acid ....

ethanoic acid to 25 cm^3 of sodium hydroxide ....

sodium hydroxide to 25 cm^3 of hydrochloric acid ....

(ii) A table of acid–base indicators and the pH ranges over which they change colour is shown below.

Indicator	**pH range**
Thymol blue	1.2 – 2.8
Bromophenol blue	3.0 – 4.6
Methyl red	4.2 – 6.3
Cresolphthalein	8.2 – 9.8
Thymolphthalein	9.3 – 10.5

Select from the table an indicator which could be used in the titration which produces curve **A** but not in the titration which produces curve **B**.

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(4)

(b) (i) Write an expression for the term *pH*.

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(ii) A solution of potassium hydroxide has a pH of 11.90 at 25°C. Calculate the concentration of potassium hydroxide in the solution.

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(4)

(c) The acid dissociation constant, K_a, for propanoic acid has the value of 1.35×10^{-5} mol dm^{-3} at 25°C.

$$K_a = \frac{[H^+][CH_3CH_2COO^-]}{[CH_3CH_2COOH]}$$

In each of the calculations below, give your answer to 2 decimal places.

(i) Calculate the pH of a 0.117 mol dm^{-3} aqueous solution of propanoic acid.

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(ii) Calculate the pH of a mixture formed by adding 25 cm^3 of a 0.117 mol dm^{-3} aqueous solution of sodium propanoate to 25 cm^3 of a 0.117 mol dm^{-3} aqueous solution of propanoic acid.

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(5)

(Total 13 marks)

19. (a) A sample of hydrochloric acid has a pH of 2.34
Write an expression for pH and calculate the concentration of this acid.

pH ....

Concentration ....

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(2)

(b) A 0.150 mol dm^{-3} solution of a weak acid, HX, also has a pH of 2.34

(i) Write an expression for the acid dissociation constant, K_a, for the acid HX.

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....

(ii) Calculate the value of K_a for this acid and state its units.

Calculation ....

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....

Units ....

(iii) Calculate the value of pK_a for the acid HX. Give your answer to two decimal places.

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(5)

(c) A 30.0 cm^3 sample of a 0.480 mol dm^{-3} solution of potassium hydroxide was partially neutralised by the addition of 18.0 cm^3 of a 0.350 mol dm^{-3} solution of sulphuric acid.

(i) Calculate the initial number of moles of potassium hydroxide.

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(ii) Calculate the number of moles of sulphuric acid added.

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(iii) Calculate the number of moles of potassium hydroxide remaining in excess in the solution formed.

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(iv) Calculate the concentration of hydroxide ions in the solution formed.

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(v) Hence calculate the pH of the solution formed. Give your answer to two decimal places.

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(6)
(Total 13 marks)

20. In this question, give all pH values to 2 decimal places.

(a) (i) Write expressions for the ionic product of water, K_w, and for pH.

K_W = ...

pH = ...

(ii) At 318 K, the value of K_w is 4.02×10^{-14} mol^2 dm^{-6} and hence the pH of pure water is 6.70
State why pure water is not acidic at 318 K.

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(iii) Calculate the number of moles of sodium hydroxide in 2.00 cm^3 of 0.500 mol dm^{-3} aqueous sodium hydroxide.

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(iv) Use the value of K_w given above and your answer to part (a)(iii) to calculate the pH of the solution formed when 2.00 cm^3 of 0.500 mol dm^{-3} aqueous sodium hydroxide are added to 998 cm^3 of pure water at 318 K.

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(6)

(b) At 298 K, the acid dissociation constant, K_a, for propanoic acid, CH_3CH_2COOH, has the value 1.35×10^{-5}mol dm^{-3}.

(i) Write an expression for K_a for propanoic acid.

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(ii) Calculate the pH of 0.125 mol dm^{-3} aqueous propanoic acid at 298 K.

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(4)

(c) Sodium hydroxide reacts with propanoic acid as shown in the following equation.

$$NaOH + CH_3CH_2COOH \rightarrow CH_3CH_2COONa + H_2O$$

A buffer solution is formed when sodium hydroxide is added to an excess of aqueous propanoic acid.

(i) Calculate the number of moles of propanoic acid in 50.0 cm^3 of 0.125 mol dm^{-3} aqueous propanoic acid.

....

....

(ii) Use your answers to part (a)(iii) and part (c)(i) to calculate the number of moles of propanoic acid in the buffer solution formed when 2.00 cm^3 of 0.500 mol dm^{-3} aqueous sodium hydroxide are added to 50.0 cm^3 of 0.125 mol dm^{-3} aqueous propanoic acid.

....

....

....

(iii) Hence calculate the pH of this buffer solution at 298 K.

....

....

....

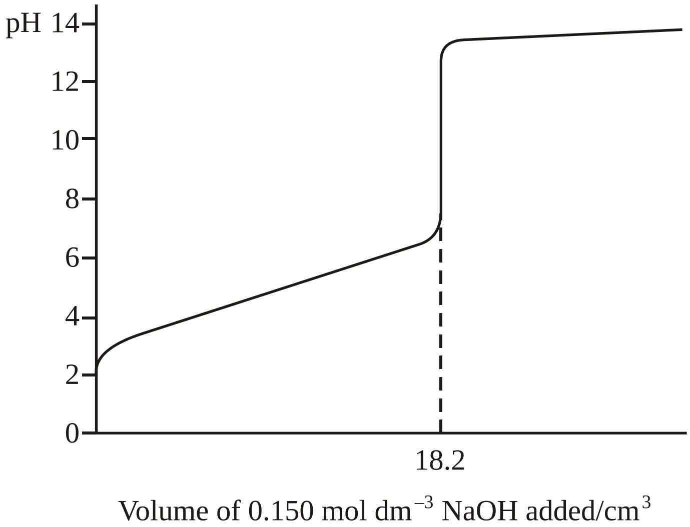
....

....

....

(6)
(Total 16 marks)

21. The pH curve shown below was obtained when a 0.150 mol dm^{-3} solution of sodium hydroxide was added to 25.0 cm^3 of an aqueous solution of a weak monoprotic acid, HA.



(a) Use the information given to calculate the concentration of the acid.

....

....

....

(2)

(b) (i) Write an expression for the acid dissociation constant, K_a, for HA.

....

(ii) Write an expression for pK_a

....

(iii) Using your answers to parts (b)(i) and (b)(ii), show that when sufficient sodium hydroxide has been added to neutralise half of the acid,

pH of the solution = pK_a for the acid HA

....

....

....

(4)

(c) Explain why dilution with a small volume of water does not affect the pH of a buffer solution.

....

....

(2)

(d) (i) Calculate the change in pH when 0.250 mol dm^{-3} hydrochloric acid is diluted with water to produce 0.150 mol dm^{-3} hydrochloric acid.

....

....

....

(ii) Calculate the volume of water which must be added to 30.0 cm^3 of 0.250 mol dm^{-3} hydrochloric acid in order to reduce its concentration to 0.150 mol dm^{-3}.

....

....

....

(4)
(Total 12 marks)

22. (a) Give the Brønsted–Lowry definition of an acid.

....

(1)

(b) (i) Explain the term *weak* when applied to an acid or a base.

....

....

(ii) Give an example of a weak base and write an equation involving this weak base to illustrate the explanation you gave in part (i) above.

Example. ....

....

Equation. ....

....

(3)

(c) In aqueous solution, the weak acid methanoic acid, HCOOH, produces aqueous methanoate ions, $HCOO^{-}(aq)$.

(i) Write an equation, including state symbols, for the formation of methanoate ions and H_3O^+ ions in an aqueous solution of methanoic acid.

....

....

(ii) Identify one substance that acts as a Brønsted–Lowry base in the forward direction, and another in the reverse direction, of the equation you have written in part (i) above.

Base in forward direction. ....

Base in reverse direction. ....

(iii) Write an expression for the acid dissociation constant, K_a, of methanoic acid.

....

....

(4)

(d) (i) A buffer solution can be defined as a solution which resists change in pH under three different circumstances. Identify these circumstances.

Circumstance 1. ....

....

Circumstance 2. ....

....

Circumstance 3. ....

....

(ii) What would you add to methanoic acid in order to make a buffer solution?

....

....

(iii) Apply your knowledge of equilibrium behaviour to the equation you wrote in your answer to part (c)(i) to suggest how this buffer solution is able to resist an increase in pH.

....

....

....

....

(7)

(Total 15 marks)

23. (a) Write an equation, including state symbols, for the reaction of gaseous hydrogen chloride with water.

....

(1)

(b) Calculate the pH of a 1.26M solution of HCl.

....

....

(2)

(c) Suggest a value for the pH of a 1.26M solution of sodium chloride. Explain your answer.

pH. ....

Explanation. ....

....

(2)

(d) Barium hydroxide, $Ba(OH)_2$, dissociates completely in water according to the equation

$$Ba(OH)_2(s) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$$

45.0cm^3 of 1.37M barium hydroxide were added to 95.0cm^3 of l.26M hydrochloric acid.

(i) Calculate the number of moles of H^+ ions in 95.0 cm^3 of I .26M hydrochloric acid.

....

....

(ii) Calculate the number of moles of OH^- ions in 45.0 cm^3 of 1.37 M barium hydroxide.

....

....

(iii) Calculate the pH at 25 °C of the solution formed after mixing the acid and the base.

....

....

....

....

....

....

....

(8)
(Total 13 marks)

24. HA is a weak monoprotic acid which can be used to make buffer solutions. It dissociates in water according to the equation

$$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$$

(a) Explain what is meant by the term *weak* when applied to an acid.

....

....

(1)

(b) A solution containing equal concentrations of undissociated HA and the anion A^- acts as a buffer solution. This solution is able to resist changes in pH. Use the equation above and your knowledge of equilibrium behaviour to suggest how the buffer solution is able to resist:

(i) a decrease in pH when a small amount of strong acid is added;

....

....

....

....

(ii) an increase in pH when a small amount of strong base is added.

....

....

....

....

(4)

(c) (i) Write an expression for the acid dissociation constant, K_a, of the acid HA

....

....

....

....

(ii) Rearrange your expression for K_a to give an equation for the hydrogen ion concentration in the acid HA. Use this equation to suggest how the buffer solution is able to resist changes in pH on dilution.

Equation for [H^+]..

....

....

Resistance to change in pH on dilution..

....

....

(4)
(Total 9 marks)

25. (a) An acid HA has $pK_a = 4.20$

(i) Define the term pK_a

....

(ii) Calculate the value of the dissociation constant, K_a, for the acid HA and state its units.

....

....

(iii) Calculate the pH of a 0.830 M solution of the acid HA.

....

....

....

....

....

(7)

(b) A different acid, HX, has $K_a = 5.25 \times 10^{-5}$ mol dm^{-3}. A solution was formed by mixing 10.5 cm^3 of 0.800 M NaOH with 25.0 cm^3 of 0.920 M HX.

(i) Calculate the number of moles of X^- ions present in the solution formed.
(Ignore any X^- ions formed by dissociation of the excess acid HX)

....

....

(ii) Calculate the number of moles of HX which remain unreacted.

....

....

(iii) Calculate the concentrations of both X^- and HX and use these to determine the pH of the solution formed.

Concentration of X^- ....

....

Concentration of HX ....

....

pH of solution ....

....

....

....

(9)

(c) State qualitatively how the pH of the solution formed in part (b) changes when a small volume of dilute hydrochloric acid is added. Use appropriate equations to explain your answer.

Change in pH ....

Explanation ....

....

....

(3)
(Total 19 marks)

26. (a) State what is meant by the term *monoprotic acid* and give one example

Monoprotic acid..

Example...

(2)

(b) (i) Define pH.

....

(ii) What is the hydrogen ion concentration in a solution which has pH = - 0.20?

....

(2)

(c) Calculate the pH of the solution formed when 35 cm^3 of 0.12 M NaOH are added to 25 cm^3 of 0.15M HCl at 25°C.

....

....

....

....

....

(7)
(Total 11 marks)

27. All solutions in parts (a) to (d) below are maintained at 25 °C.

(a) Write equations to show the reaction of HCl(g) and KOH(s) with water.

Equation for HCl(g) ....

....

Equation for KOH(s) ....

....

(2)

(b) Write an expression to define the ionic product, K_w, of water.

....

(1)

(c) Calculate the pH of a 0.0160 M KOH solution and estimate the pH of a 0.100 M KCl solution. Give your reasoning.

(At 25 °C, the value of K_w is 1.00×10^{-14} mol^2 dm^{-6})

pH of 0.0160 M KOH ....

....

....

....

....

pH of 0.100 M KCl ....

....

(7)

(d) A mass, m, of solid KOH is added to 755 cm^3 of 0.0120 M HCl. The pH after this addition is 11.60, measured at 25 °C. The volume of the resulting solution is still 755 cm^3.

(i) Calculate the number of moles of OH^- ions needed to neutralise exactly the H^+ ions present in the 755 cm3 of 0.0120 M HCl.

....

....

....

(ii) Calculate the number of moles of OH^- ions in excess when the pH is 11.60

....

....

....

....

(iii) Use these results to calculate the total number of moles of KOH added.

....

....

(iv) Hence deduce the value of *m*.

....

....

....

(8)
(Total 18 marks)

28. (a) The hydrogen halides all react with water to form acids. Hydrogen fluoride forms a weak acid while the others all form strong acids. Write equations to show the reactions that occur when hydrogen fluoride and hydrogen chloride are dissolved in water.

Hydrogen fluoride and water ....

Hydrogen chloride and water ....

(2)

(b) (i) Define the term pH..

(1)

(ii) Calculate the pH of an aqueous solution of hydrochloric acid containing 0.050 mol dm^{-3}.

(1)

(c) (i) Write an expression for the dissociation constant, K_a, for hydrofluoric acid.

(1)

(ii) Calculate the pH of an aqueous solution of hydrofluoric acid of concentration 0.050 mol dm^{-3} at 298K, given that $K_a = 5.6 \times 10^{-4}$ mol dm^{-3} at 298 K.

(3)

(d) When hydrogen fluoride is dissolved in pure nitric acid, a reaction takes place that can be represented by the equation:

$$HNO_3 + HF \rightarrow H_2NO_3^+ + F^-$$

State, with a reason, which reactant acts as a Brønsted-Lowry acid in this reaction and give the formula of its conjugate base.

....

....

....

(3)
(Total 11 marks)

29. (a) When dissolved in water, ethanoic acid acts as a weak Brønsted-Lowry acid.

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$

Explain the terms:

(i) *Brønsted-Lowry acid*; ....

....

(1)

(ii) *weak acid.* ....

....

(1)

(b) (i) Write an expression for the acid dissociation constant, K_a, for ethanoic acid.

(1)

(ii) Calculate the pK_a value of aqueous ethanoic acid.
K_a (ethanoic acid) = 1.70×10^{-5} mol dm^{-3} at 25°C

(1)

(c) The pH ranges over which two indicators used in acid-base titrations change colour are given in table below.

Indicator	**pH range**
methyl orange	3.1 – 4.4
phenolphthalein	8.3 – 10.0

In a titration aqueous sodium hydroxide is run into a conical flask containing aqueous ethanoic acid.

(i) State which indicator should be used and explain your answer.

Indicator ....

Explanation ....

....

....

(3)

(ii) State the colour change seen in the conical flask at the end point.

From ....*to* ....

(1)

(iii) Write an equation for the reaction between ethanoic acid and sodium hydroxide in aqueous solution.

....

(1)

(d) You are supplied with aqueous sodium hydroxide and aqueous ethanoic acid, each solution having a concentration of 0.10 mol dm^{-3}. State briefly how you would prepare a buffer solution with a pH equal to the pK_a of ethanoic acid.

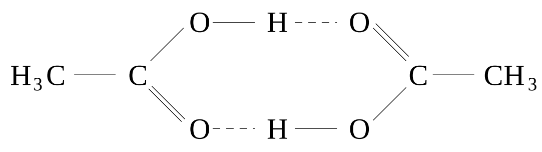
....

....

....

(2)

(e) In the gaseous state, some ethanoic acid molecules are dimerised as shown in diagram below. The broken lines represent hydrogen bonds.



Explain how hydrogen bonds are formed between ethanoic acid molecules.

....

....

....

....

(3)
(Total 14 marks)

30. (a) State what is meant by the term *weak acid* and give one example.

Weak acid...

Example...

(2)

(b) Write an expression for the dissociation constant, K_a, of the weak acid HA and state the units of K_a.

Expression...

....

....

Units...

(2)

(c) When water is cooled, the pH increases but the water remains neutral.

(i) Explain why the pH increases.

....

....

....

(ii) Explain why water remains neutral.

....

(4)

(d) State the characteristic property of a buffer solution.

....

....

(2)
(Total 10 marks)

31. Solution **S** is 0.16 M hydrochloric acid, HCl, a strong acid.
Solution **W** is 0.16 MHX, a weak monoprotic acid. It has a pH of 2.74.
Solution **Z** is 0.12 M barium hydroxide, $Ba(OH)_2$, a strong base.

(a) (i) Explain the terms *weak* and *strong* as applied to acids or bases.

(ii) Determine a value of the acid dissociation constant of the weak acid HX using the expression

$$K_a \approx \frac{[H^+]^2}{c}$$

where *c* is the original concentration of HX. Explain why it is reasonable to use this approximation.

(6)

(b) Show details of **all** calculations in answering this part of the question.

At a temperature of 25°C, 18 cm^3 of solution **S** are titrated with solution **Z**. The titration is repeated using 18 cm^3 of solution **W.**

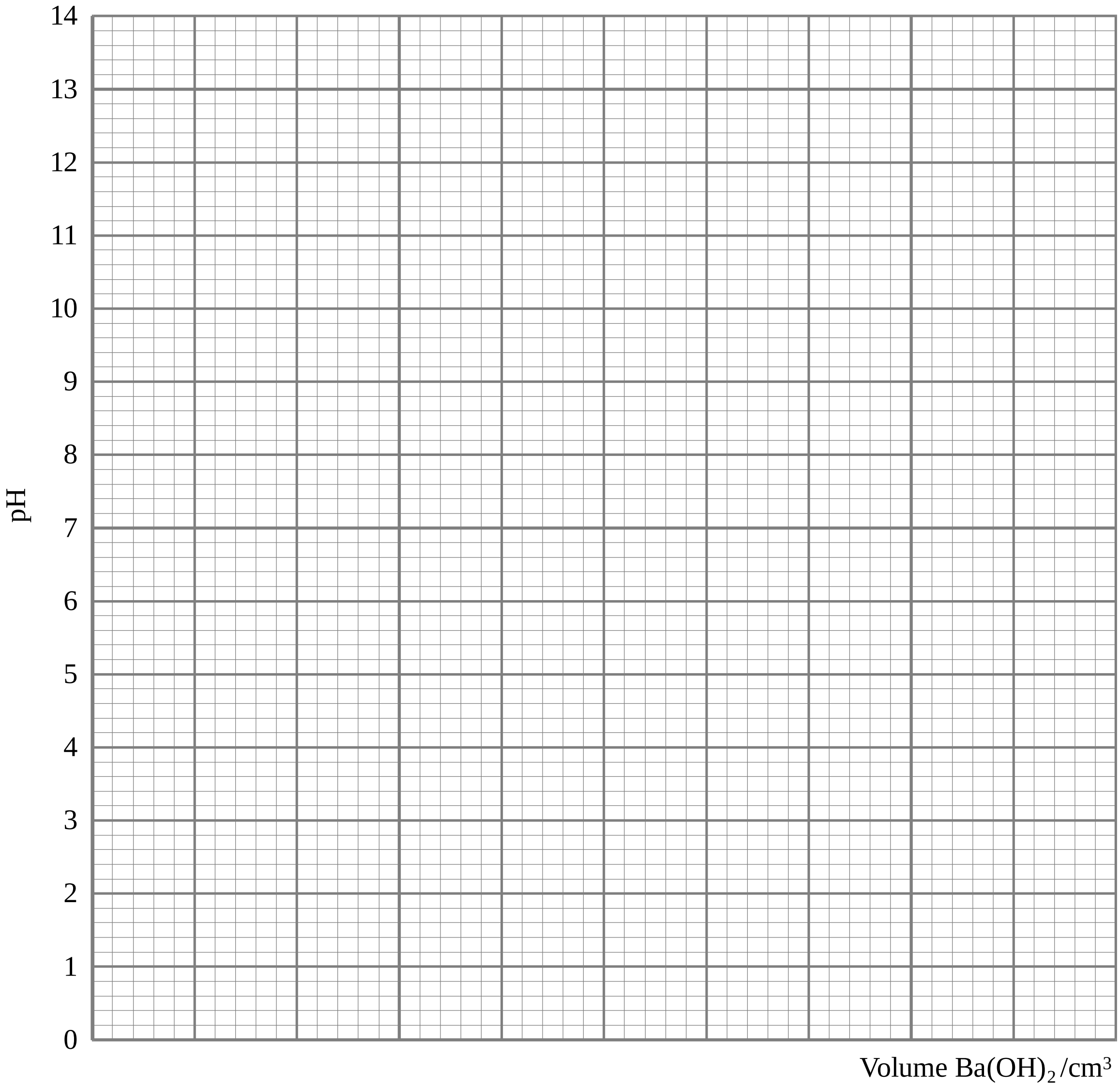
(i) Determine the equivalence volume (end-point) for these titrations and enter your result into the appropriate space in the incomplete table below.
Enter also the half-equivalence volume and the double-equivalence volume in the appropriate space.

(ii) Calculate the missing pH values for each of the two titration solutions and enter these into the table.

(You should assume that, at half-equivalence, [HX]= [X^-] for the weak acid HX, and that, for both acids at double-equivalence, only the alkali that is in excess contributes to the pH of the resulting solution.)

	Start	Half equivalence	Equivalence	Double Equivalence
Volume/cm^3 $Ba(OH)_2$ solution added	0.0			
pH for titration of **S**	0.80		7	
pH for titration of **W**	2.74		8.5	

(iii) Plot these results on the graph below and use the points you have plotted to sketch the complete titration curves for solution **S** and solution **W** titrated with solution **Z**.



(18)

(c) (i) Explain what is meant by the term *buffer solution*. Suggest how solution **W,** when half-neutralised, can behave as a buffer solution.

(ii) State the difference between *acidic* and *basic* buffers. To which of these two types of buffer does a half-neutralised solution of **W** belong? What might you use to make a buffer solution of the other type?

(6)
(Total 30 marks)

32. This question concerns the weak acid, ethanoic acid, for which the acid dissociation constant, K_a, has a value of 1.74×10^{-5} mol dm^{-3} at 25°C.

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

In each of the calculations below, give your answer to 2 decimal places.

(a) Write an expression for the term *pH*. Calculate the pH of a 0.150 mol dm^{-3} solution of ethanoic acid.

(4)

(b) A buffer solution is prepared by mixing a solution of ethanoic acid with a solution of sodium ethanoate.

(i) Explain what is meant by the term *buffer solution*.

(ii) Write an equation for the reaction which occurs when a small amount of hydrochloric acid is added to this buffer solution.

(3)

(c) In a buffer solution, the concentration of ethanoic acid is 0.150 mol dm^{-3} and the concentration of sodium ethanoate is 0.100 mol dm^{-3}.

(i) Calculate the pH of this buffer solution.

(ii) A 10.0 cm^3 portion of 1.00 mol dm^{-3} hydrochloric acid is added to 1000 cm^3 of this buffer solution.
Calculate the number of moles of ethanoic acid and the number of moles of sodium ethanoate in the solution after addition of the hydrochloric acid. Hence, find the pH of this new solution.

(8)
(Total 15 marks)