Q1. (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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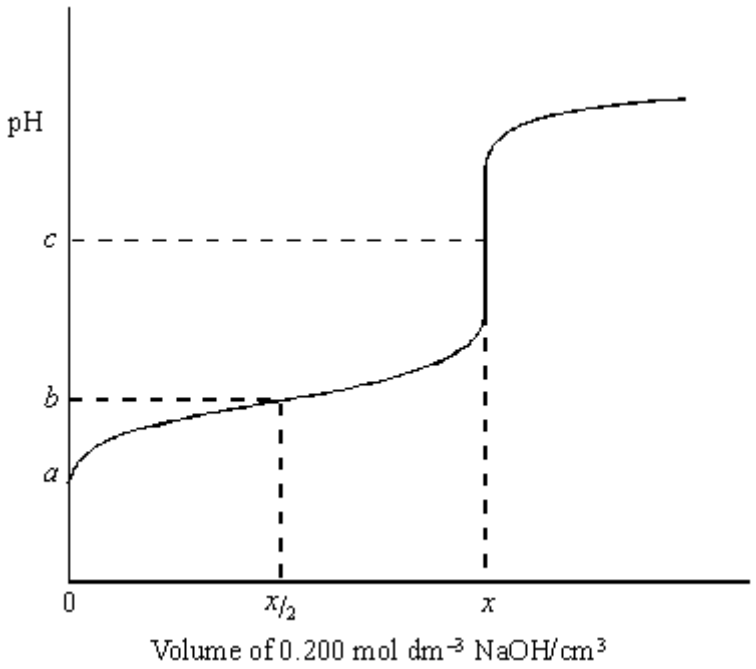
(5)
(Total 14 marks)

Q2. 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 cm^3 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)

Q3. A 0.210 mol dm^{-3} solution of potassium hydroxide was added from a burette to 25.0 cm^3 of a 0.160 mol dm^{-3} solution of ethanoic acid in a conical flask.
Given that the value of the acid dissociation constant, K_a, for ethanoic acid is 1.74×10^{-5} mol dm^{-3}, calculate the pH at 25 °C of the solution in the conical flask at the following three points:

before any potassium hydroxide had been added;

after 8.0 cm^3 of potassium hydroxide solution had been added;

after 40.0 cm^3 of potassium hydroxide solution had been added.

(Total 16 marks)

Q4. 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)