pH of weak acids, weak bases and salt solutions

2010: 1

An aqueous solution of ammonium chloride (NH_4Cl) has a pH of 4.66.

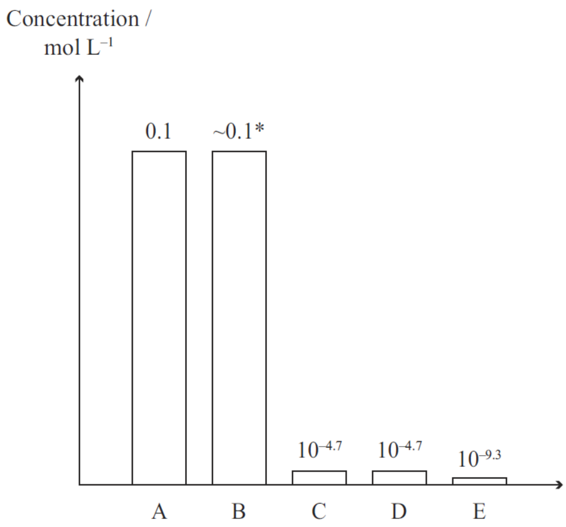
(a) (i) Write the equation for solid ammonium chloride dissolving in water.

(ii) Write the equation for the ammonium ion reacting with water.

(b) Calculate the concentration of the NH_4Cl solution. $pK_a(NH_4^+) = 9.24$ $K_a = 5.75 \times 10^{-10}$.

(c) The bar chart below shows the relative concentrations of the species (excluding water) in a solution of 0.1 mol L^{-1} NH_4Cl. (The bar chart is not drawn to scale.)

Identify the species A to E. Justify your answer.



* ~0.1 means approximately 0.1

2009: 1

Ethanoic acid, CH_3COOH, is a common organic acid. $pK_a\ (CH_3COOH) = 4.76$ $K_a = 1.74 \times 10^{-5}$

(a) (i) Write an equation for the reaction of ethanoic acid with water.

(ii) Write the K_a expression for ethanoic acid.

(b) Calculate the pH of a 0.0500 mol L^{-1} ethanoic acid solution.

(c) Another organic acid is methanoic acid, HCOOH. $pK_a\ (HCOOH) = 3.74$

Account for the fact that 0.0500 mol L^{-1} methanoic acid has a lower pH than 0.0500 mol L^{-1} ethanoic acid.

(d) A solution prepared by dissolving sodium methanoate in water has a pH of 8.65. Determine the concentration of methanoate ions in the solution.

2008: 2

When bromine is added to water, it forms hypobromous acid (HOBr), a weak acid.

(a) (i) Write an equation to show the equilibrium system that is formed with hypobromous acid and water.

(ii) Write the K_a expression for hypobromous acid.

(b) Calculate the pH of a 0.0525 mol L^{-1} hypobromous acid solution. $pK_a(HOBr) = 8.62$

2005: 4

The active ingredient in many sunscreens is para-aminobenzoic acid. It is a weak monoprotic acid and can be represented as HPab, while its conjugate base is Pab^-.

(a) Write an equation for the reactions occurring at equilibrium when HPab is dissolved in water.

(b) Write the expression for Ka(HPab).

A solution of HPab in water was prepared at 25°C and its pH was found to be 3.22.

(c) Calculate the concentration of H_3O^+ in the solution.

The concentration of the HPab solution was determined by titration. A 20.0 mL sample of the HPab solution required 12.0 mL of 0.0500 mol L^{-1} NaOH to reach the equivalence point. The equation for the reaction occurring is

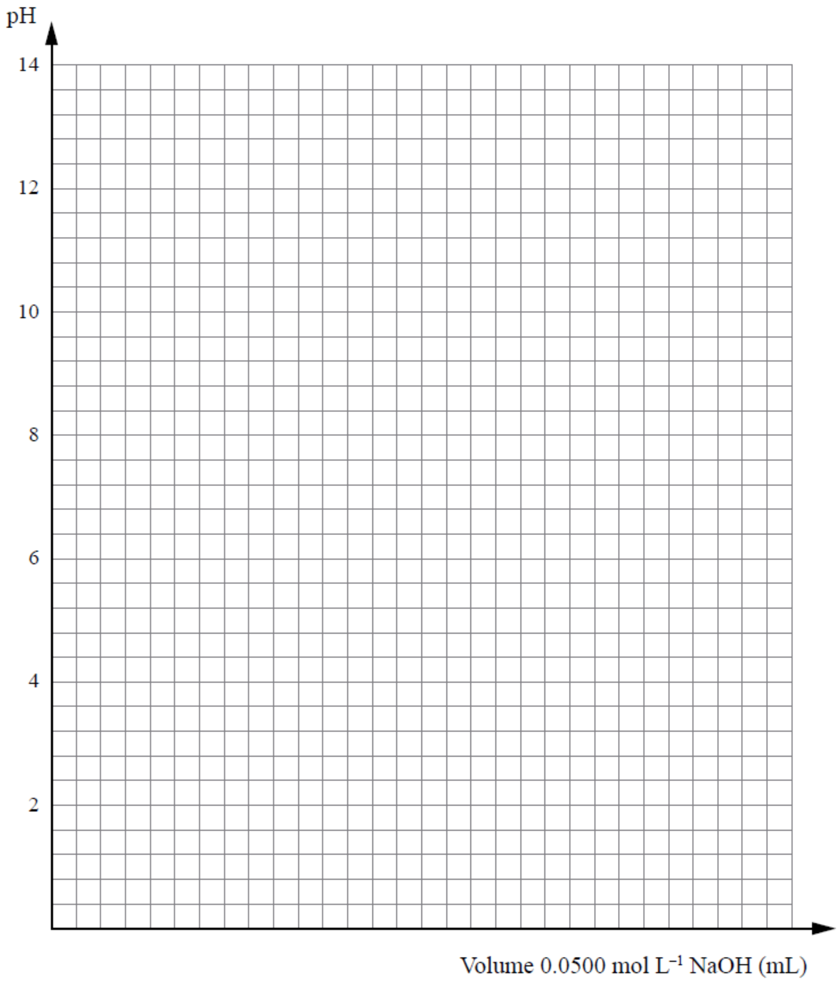
$$HPab + NaOH \rightarrow NaPab + H_2O$$

(d) (i) Calculate the concentration of the HPab solution.

(ii) Using the results from parts (c) and (d)(i), show that pKa(HPab) = 4.92.

(e) Would the pH at the equivalence point of the titration of HPab with NaOH be more than 7, less than 7 or equal to 7? Give reasons and include any relevant equations that support your answer.

(f) Using the information above, sketch a curve showing the change in pH against the volume of sodium hydroxide added to the 20.0 mL HPab solution in the flask.



Species in Solution

2009:4

The following table lists some properties of aqueous solutions of sodium hydroxide, methylamine and methylammonium chloride.

	0.1 mol L^{-1} solutions	pH	Conductivity
A	Sodium hydroxide (NaOH)	13.0	High
B	Methylamine (CH_3NH_2)	11.8	Low
C	Methylammonium chloride (CH_3NH_3Cl)	5.3	High

The solutions above were prepared by adding the compounds to water.

(a) Write equations for the reactions occurring when each of the three compounds are added to water.

NaOH(s)

CH_3NH_2(g)

CH_3NH_3Cl(s)

(b) Justify the differences in the pH and conductivity of the three solutions

2007: 1

(a) (i) For each of the following 0.1 mol L^{-1} solutions, write an equation to show the reaction with water.

CH_3NH_2

NH_4Cl

(ii) List all the species in each of the following 0.1 mol L^{-1} aqueous solutions in order of decreasing concentration. Do not include H_2O.

(b) Explain why aqueous aminomethane, CH_3NH_2, is a weak electrolyte.

2007: 2 (part)

The pH of the solution in the stomach of a patient in hospital is 2.50. As a treatment, the patient is given a small volume of sodium citrate (Na3Cit) solution. Citric acid, H3Cit, is a triprotic acid.

(a) (i) Would the pH of a solution of sodium citrate be less than, equal to or greater than 7?

A calculation is not required.

(ii) Explain your choice, including an appropriate equation in your answer.

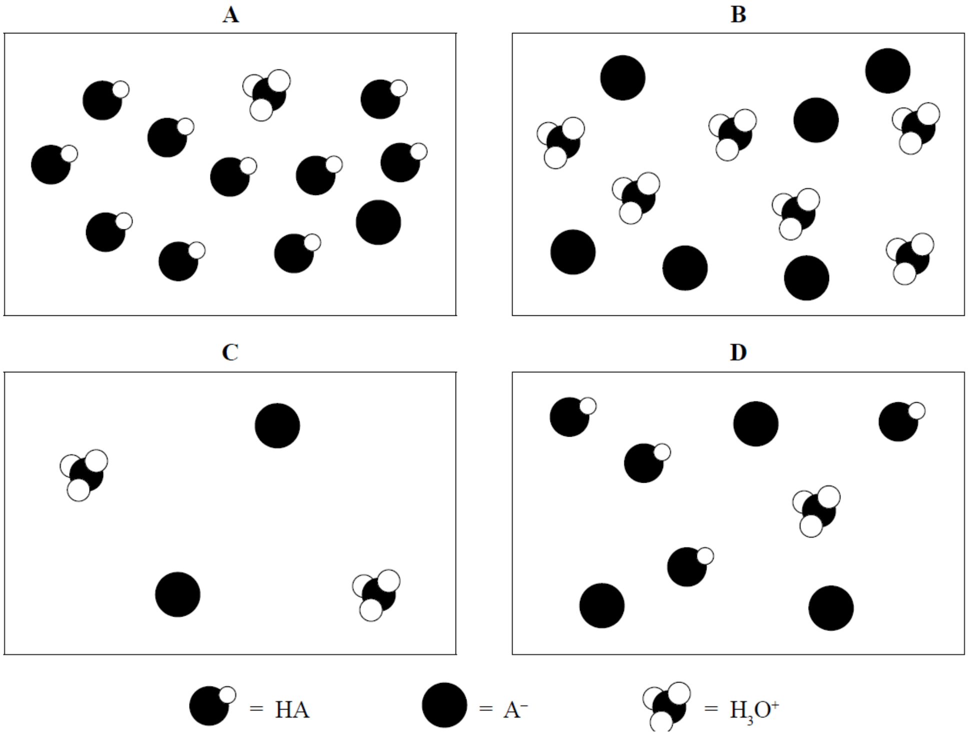
2007: 3

An aqueous ammonia solution has a pH of 10 and when phenolphthalein indicator is added it turns pink. Solid ammonium chloride is added to this solution and the solution turns colourless due to a decrease in pH.

By considering the equilibrium systems, discuss why the pH of the solution decreased. Include a relevant equation in your answer.

2006:1

The boxes below show particle representations of the species (excluding water) in four aqueous solutions.



(a) Choose the box that best illustrates each of the solutions (i)–(iii) below. In each case, give a reason for your answer.

(i) A dilute solution of a strong acid

(ii) A concentrated solution of a weak acid

(iii) A buffer solution

(e) Explain how the **pH** and **buffering properties** of the buffer solution would be affected if it were diluted by a factor of 10.

2005: 1

Arrange the following 0.1 mol L–1 solutions in order of increasing pH.

- NH_3
- NH_4Cl
- HCl
- NaCl
- NaOH

Give reasons for arranging in this order, including equations for any reactions occurring to produce solutions that **do not** have a pH of 7.

Solubility calculations

2010:2

(a) Sufficient Ag_2CrO_4 is dissolved in water to form a saturated solution.
(i) Write the equation for the equilibrium present in a saturated solution of Ag_2CrO_4.
(ii) Write the expression for $Ks(Ag_2CrO_4)$.

(b) Calculate the solubility of Ag_2CrO_4 in a saturated solution in mol L^{-1}. $Ks(Ag_2CrO_4) = 3.00 \times 10^{-12}$

(c) Discuss how the solubility of Ag_2CrO_4 will change if it is dissolved in the following solutions.
No calculations are necessary.
(i) 0.1 mol L^{-1} K_2CrO_4 (ii) 0.1 mol L^{-1} NH_3

2009: 2

Addition of chloride ions to a solution of silver nitrate often results in the formation of a white precipitate of silver chloride (AgCl). $AgCl(s) \rightleftharpoons Ag_+(aq) + Cl^-(aq)$ $K_s(AgCl) = 1.56 \times 10^{-10}$

(a) Calculate the concentration, in mol L^{-1}, of silver ions in a saturated solution of silver chloride at 25°C.

(b) Solid sodium chloride is added to 5.00 L of 0.100 mol L^{-1} silver nitrate solution. Calculate the minimum mass of sodium chloride that would be needed to produce a saturated solution of AgCl.
Assume that there is no change in volume when the sodium chloride is added.
$M(NaCl) = 58.5$ g mol^{-1}.

(c) Discuss reasons for the fact that a precipitate of silver chloride dissolves on the addition of excess aqueous ammonia.

2008: 3

(a) (i) Write an equation for the sparingly soluble salt lead(II) chloride ($PbCl_2$) dissolving in water.
(ii) Write the solubility product expression for lead(II) chloride.

(b) Calculate the solubility, in mol L^{-1}, of $PbCl_2$ in water at 25°C. $Ks(PbCl_2) = 1.60 \times 10^{-5}$ at 25°C

(c) Sea water contains many dissolved salts. The chloride ion concentration in a sample of sea water is 0.440 mol L^{-1}.
Determine whether a precipitate of lead(II) chloride will form when a 1.00 g sample of lead(II) nitrate is added to 500 mL of the sea water. Your answer must be clearly justified.
$M(Pb(NO_3)_2) = 331$ g mol^{-1}.

2008: 4

The K_s of aluminium hydroxide, $Al(OH)_3$, at 25°C, is 3×10^{-34}, indicating that it has very low solubility. The solubility may be altered by changes in pH (due to acidic or basic properties) and formation of complex ions such as the aluminate ion, $[Al(OH)_4]^-$.
Discuss why aluminium hydroxide becomes more soluble in aqueous solutions that have a pH less than 4, or a pH greater than 10.
In your answer include:

- the equation for the reaction that relates to $K_s(Al(OH)_3)$
- equations for the reactions that relate to changes in the solubility of aluminium hydroxide at pH less than 4 or greater than 10
- a discussion of the equilibrium principles involved.

2007: 2 (part)

Magnesium hydroxide (known as milk of magnesia) is another substance that the patient consumed to control the acidity of the solution in the stomach. Ks ($Mg(OH)_2$) = 1.25×10^{-11}

(b) (i) Calculate the solubility of magnesium hydroxide in water in mol L^{-1}.

(ii) What is the concentration of Mg^{2+} in 0.150 mol L^{-1} sodium hydroxide, NaOH, solution?

2006: 2

Sea-water contains appreciable amounts of ions other than Na^+ and Cl^-. One substance that is less soluble than sodium chloride is calcium sulfate. This is precipitated in the first stage of the purification process used to produce table salt (sodium chloride).

K_s ($CaSO_4$) = 2.45×10^{-5}

(a) (i) Write the equation for the equilibrium reaction in a saturated solution of calcium sulfate.

(ii) Calculate the solubility of $CaSO_4$ in water.

Evaporating the sea-water to dryness would produce a mixture of salts including NaCl. However, precipitation of NaCl occurs if concentrated hydrochloric acid is added to a saturated NaCl solution.

(b) Explain why this precipitation occurs.

As part of the process for extracting table salt from sea-water, sodium hydroxide is added to the seawater to precipitate the magnesium ions as magnesium hydroxide. The concentration of Mg^{2+} ions at this stage is 0.555 mol L^{-1}.

(c) Calculate the minimum hydroxide ion concentration and hence the pH of the solution needed for precipitation to occur. Ks ($Mg(OH)_2$) = 7.10×10^{-12}

2005: 2

(a) Describe what is meant by the term '**solubility**'.

(b) The solubility product, K_s, of AgCl has a value of 1.56×10^{-10} at 25°C and this value increases to 2.15×10^{-8} at 100°C.

Explain why K_s is higher at 100°C. Include reference to the relevant equilibrium equation in your answer.

The chloride ion concentration in sea water can be determined by titrating a sample with aqueous silver nitrate ($AgNO_3$) using potassium chromate (K_2CrO_4) as the indicator.

As the silver nitrate is added, a precipitate of silver chloride, (AgCl) forms. When most of the AgCl has precipitated, the Ag^+(aq) concentration becomes high enough for a red precipitate of Ag_2CrO_4 to form.

(c) Show that the solubility of Ag_2CrO_4 in pure water at 25°C is higher than that of AgCl. K_s(AgCl) = 1.56×10^{-10} K_s(Ag_2CrO_4) = 1.30×10^{-12}.

(d) If the concentration of chromate ions is 6.30×10^{-3} mol L^{-1} at the point when the Ag_2CrO_4 starts to precipitate, calculate the concentration of Ag^+ ions in the solution.

Titration curves AND/OR buffer solutions

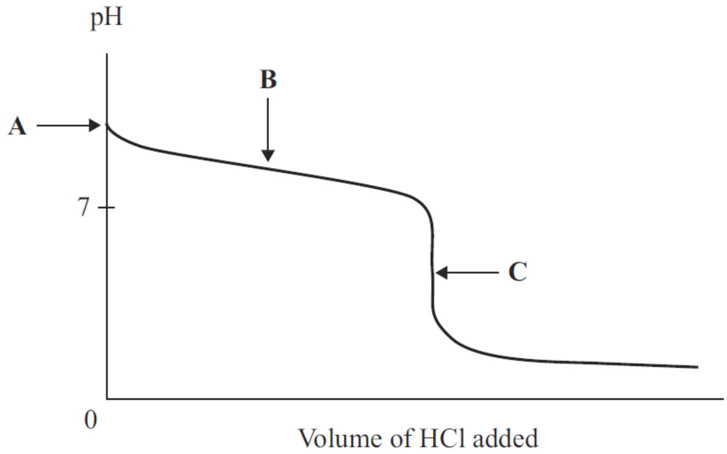
2010: 3

20.00 mL of 0.160 mol L^{-1} ammonia is titrated with 0.230 mol L^{-1} hydrochloric acid.

The equation for the reaction is

$NH_3 + HCl \rightarrow NH_4^+ + Cl^-$ $pK_a(NH_4^+) = 9.24$, $K_a = 5.75 \times 10^{-10}$

The curve for this titration is given below.



(a) Explain, in terms of species present, why the pH at B (half way to the equivalence volume) is 9.24.

(b) Calculate the pH at point A.

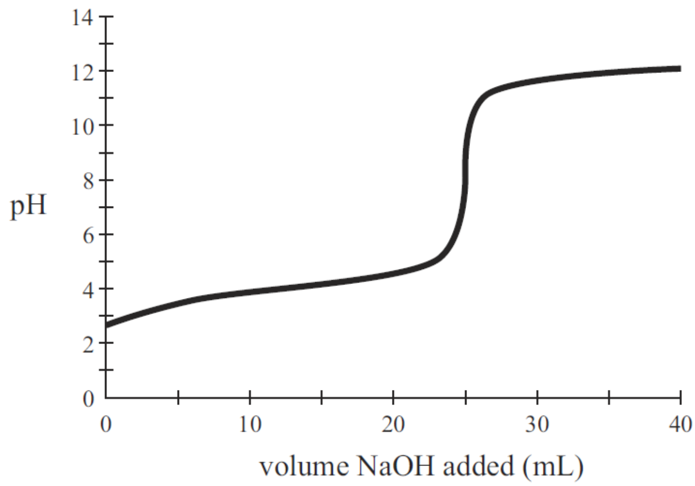
(c) Discuss the pH of the reaction mixture at point **C**, in terms of the species present.

2009: 3

25.0 mL of 0.0500 mol L^{-1} benzoic acid solution (C_6H_5COOH) is titrated with 0.0500 mol L^{-1} sodium hydroxide solution. The equation for the reaction is:

$C_6H_5COOH(aq) + NaOH(aq) \rightarrow C_6H_5COONa(aq) + H_2O(l)$

The titration curve for the reaction is:



(a) Write the formulae of the four chemical species, apart from water and H_3O^+, that are present at the equivalence point.

(b) Explain why the solution in the titration flask has buffering properties after 9.80 mL of the NaOH solution has been added, but not when 25.0 mL has been added.

See NCEA for the answers / assessment schedules 2010:4 means question 4 in 2010!

(c) Some indicators are shown in the table below.

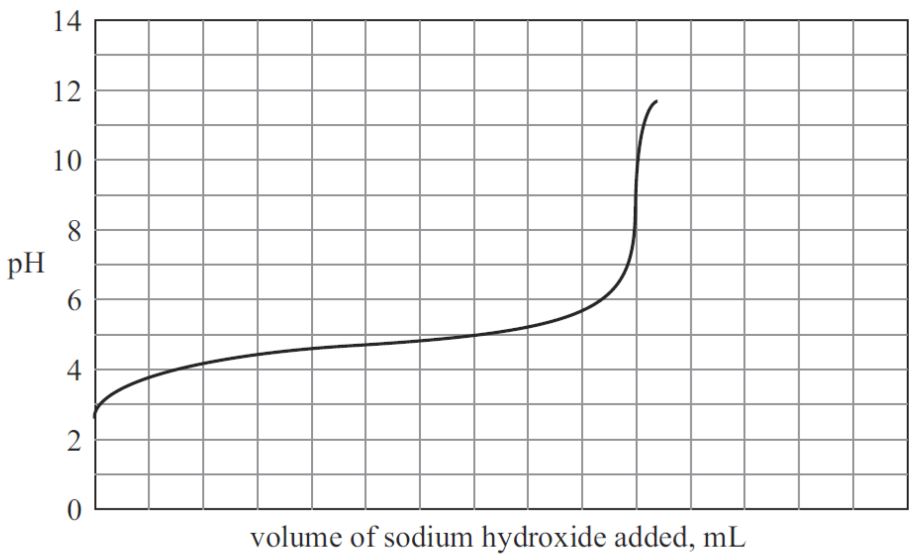
Indicator	pK_a
Methyl orange	3.70
Thymol blue	8.90
Phenolphthalein	9.30

Discuss the suitability of these indicators for this titration.
Your discussion should include:

- identification of the most suitable indicator(s)
- consideration of how indicators are chosen for a titration
- the consequences of choosing an unsuitable indicator.

2008: 1

The following titration curve shows the addition of aqueous 0.100 mol L^{-1} sodium hydroxide to a solution of hydrazoic acid, HN_3. pK_a(HN_3) = 4.72

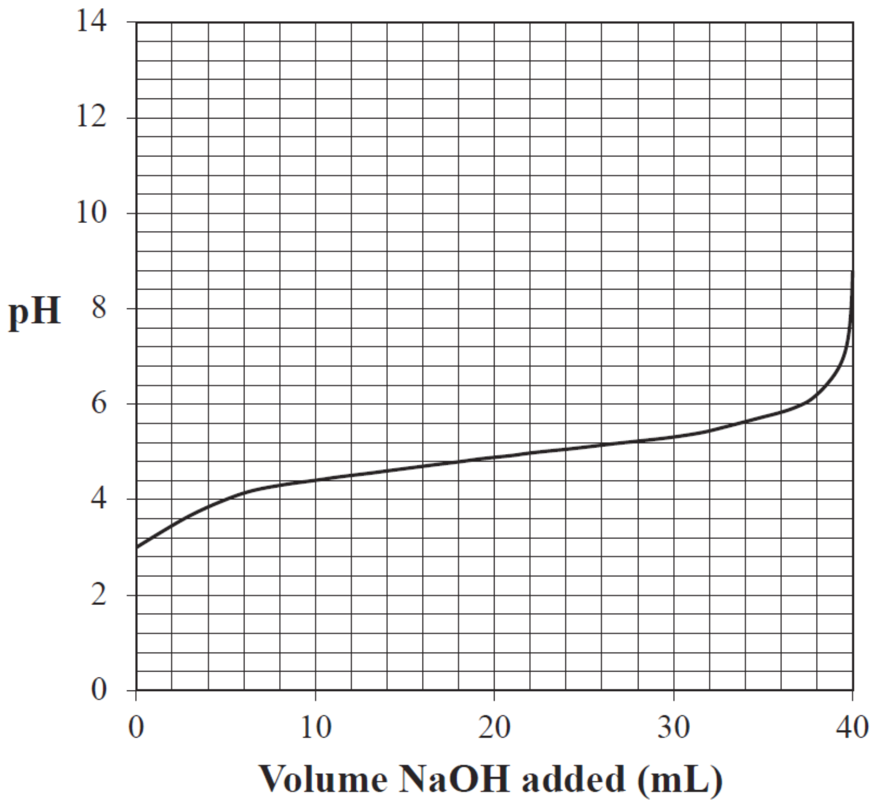


(a) (i) Draw a cross (X) on the titration curve to indicate the pH at the equivalence point of the titration.
(ii) Complete the titration curve to show how the pH changes as more aqueous sodium hydroxide is added.

(b) The initial pH of the hydrazoic acid (HN_3) is 2.60. Calculate the concentration of the HN_3 solution used in the titration.

2007: 4

A 0.160 mol L^{-1} solution of sodium hydroxide is titrated against 50.0 mL of aqueous propanoic acid, HPr. 40.0 mL of the sodium hydroxide solution was required to exactly react with the propanoic acid. The reaction occurring can be represented as: HPr(aq) + NaOH(aq) → NaPr(aq) + H_2O K_a (HPr) = 1.35×10^{-5}



(a) (i) Show that the concentration of the aqueous propanoic acid is 0.128 mol L^{-1}.

(ii) Calculate the pH of the aqueous propanoic acid.

(b) Calculate the pH at the equivalence point.

(c) 35 mL of the sodium hydroxide solution is added to a second 50 mL sample of the same acid to form a buffer solution.

(i) What is the function of a buffer?

(ii) Discuss the ability of the solution formed to act as a buffer. Your answer should include relevant equations.

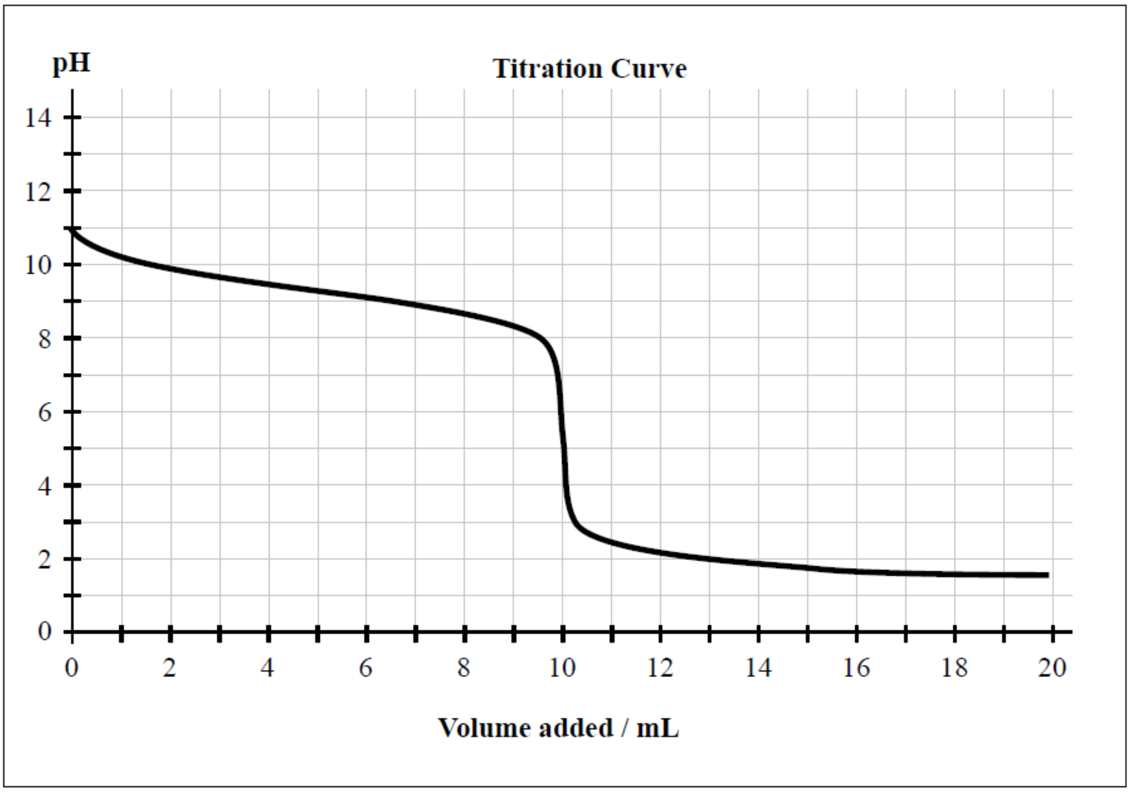
(d) The equivalence point of the titration could also be found using an acid-base indicator. Which of the following indicators would be suitable to use? Explain your choice of indicator.

Indicator	**p*K*a**
Methyl orange	3.7
Bromocresol green	4.7
Methyl red	5.0
Thymol blue	8.9
Phenolphthalein	9.3

2006:3

The graph shows the change in pH when 40.0 mL of 0.0500 mol L^{-1} aqueous NH_3 is titrated with 0.200 mol L^{-1} aqueous HCl.

The equation for the reaction occurring during the titration is: $NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$



(a) Use the curve to determine $pK_a(NH_4^+)$ and hence calculate $K_a(NH_4^+)$.

(b) Explain why the pH at the equivalence point for this titration is less than 7. (Include an equation to support your answer.)

A NH_4^+ / NH_3 buffer solution is prepared with a pH of 9.60.

(c) Use the graph to describe how this buffer solution could be made from 0.0500 mol L^{-1} NH_3 solution and 0.200 mol L^{-1} HCl solutions.

A second titration is carried out – this time 40.0 mL of 0.0500 mol L^{-1} NH_4Cl solution is titrated against 0.200 mol L^{-1} NaOH solution.

(d) Write an equation for the titration reaction.

(e) (i) Show that $[NH_3]$ at the equivalence point is 0.0400 mol L^{-1}.

(ii) Using $K_a(NH_4^+)$ determined in part (a), determine the pH at the equivalence point of the second titration.

2010:4

A buffer solution is made by adding solid sodium methanoate, HCOONa, to an aqueous solution of methanoic acid, HCOOH.

$pK_a(HCOOH) = 3.74$

(a) Describe the function of a buffer solution.

(b) Explain why the solution made with methanoic acid, HCOOH, and sodium methanoate, HCOONa, has the ability to act as a buffer. Your answer should include relevant equations.

(c) Calculate the mass of sodium methanoate that must be added to 100 mL of 0.861 mol L^{-1} methanoic acid to give a solution with a pH of 3.24. Assume there is no volume change on adding the salt.
$pK_a(HCOOH) = 3.74$ $K_a = 1.82 \times 10^{-4}$ M (HCOONa) = 68.0 g mol^{-1}

2008: 5

Two solutions, A and B, were made as described below.

- Solution A: 50 mL of aqueous 1.00 mol L^{-1} ammonium chloride was added to 50 mL of aqueous 1.00 mol L^{-1} ammonia.
- Solution B: 25 mL of aqueous 0.010 mol L^{-1} hydrochloric acid was added to 50 mL of aqueous 0.010 mol L^{-1} ammonia.

(a) (i) Write the K_a expression for NH_4^+.
(ii) Show, by calculation, that the pH of each of the two solutions is 9.24. $pK_a(NH_4^+) = 9.24$

(b) (i) Discuss the abilities of solutions A and B to act as buffers.
(ii) Compare how the pH of each solution would be affected when 1.00 mol L^{-1} sodium hydroxide is added drop-wise to each solution. Calculations are not required, but you should include appropriate equations in your answer.

2005: 3

(a) Explain how a mixture of ethanoic acid (CH_3COOH) and sodium ethanoate (CH_3COONa) can act as a buffer. Include balanced equations for any reactions occurring.

(b) Calculate the concentration of ethanoate ions (CH_3COO^-) in a buffer solution of pH 5.00 if the concentration of CH_3COOH in the buffer is 0.0500 mol L^{-1}. K_a $(CH_3COOH) = 1.76 \times 10^{-5}$ at 25°C.