

Questions

Module 6: Acid/Base Reactions

Multiple-choice questions: 1 mark each

1. What did the Brönsted–Lowry definition of acids identify that made it a significant improvement over earlier definitions?
 - (A) Acids contain hydrogen.
 - (B) Acids are proton donors.
 - (C) Acids contain oxygen.
 - (D) Acids are electron-pair acceptors.2002 HSC Q7

2. Which of the following is an acid–base indicator?
 - (A) Methanol
 - (B) Methyl orange
 - (C) Methanoic acid
 - (D) Methyl ethanoate2003 HSC Q1

3. Which is amphiprotic?
 - (A) H_2SO_4
 - (B) NH_4^+
 - (C) HCO_3^-
 - (D) SO_4^{2-}2002 HSC Q6

4. Sodium reacts with water to give hydrogen gas and sodium hydroxide solution.
What volume of gas would be produced from the reaction of 22.99 g of sodium at 25°C and 100 kPa?
 - (A) 11.36 L
 - (B) 12.40 L
 - (C) 22.71 L
 - (D) 24.79 L2013 HSC Q14

5. A solution was obtained by boiling flowers in water. After various substances were added to separate samples of the solution, the colour of each was noted.

<i>Substance added</i>	<i>Colour observed</i>
0.1 mol L ⁻¹ HCl(<i>aq</i>)	Bright pink
0.01 mol L ⁻¹ HCl(<i>aq</i>)	Bright pink
0.001 mol L ⁻¹ HCl(<i>aq</i>)	Pale yellow
Distilled water	Bright yellow
0.001 mol L ⁻¹ NaOH(<i>aq</i>)	Bright yellow
0.01 mol L ⁻¹ NaOH(<i>aq</i>)	Bright yellow

For which of the following titrations would it be appropriate to use this solution as an indicator?

- (A) HCl(*aq*) + NH₃(*aq*)
- (B) HCl(*aq*) + NaOH(*aq*)
- (C) CH₃COOH(*aq*) + NH₃(*aq*)
- (D) CH₃COOH(*aq*) + NaOH(*aq*)

2013 HSC Q19

6. Which pieces of glassware should be used when preparing a primary standard solution?
- (A) Pipette, burette and conical flask
 - (B) Dropper, watch glass and pipette
 - (C) Beaker, filter funnel and volumetric flask
 - (D) Measuring cylinder, stirring rod and conical flask

2012 HSC Q4

7. Which acid / base pair could act as a buffer?
- (A) H₃O⁺ / H₂O
 - (B) H₂O / OH⁻
 - (C) HNO₃ / NO₃⁻
 - (D) H₂PO₄⁻ / HPO₄²⁻

2012 HSC Q8

8. A 25.0 mL sample of a 0.100 mol L^{-1} hydrochloric acid solution completely reacted with 23.4 mL of sodium hydroxide solution.

What volume of the same sodium hydroxide solution would be required to completely react with 25.0 mL of a 0.100 mol L^{-1} acetic acid solution?

- (A) Less than 23.4 mL
- (B) 23.4 mL
- (C) More than 23.4 mL
- (D) Unable to calculate unless the concentration of the sodium hydroxide solution is also known

2013 HSC Q17

9. A pH 3.0 solution of $\text{HCl}(aq)$ is diluted by adding water to produce a pH 5.0 solution.

Which row in the following table correctly identifies an appropriate volume of the original solution and the volume of water added for this dilution?

	<i>Volume of original solution (mL)</i>	<i>Volume of water added (mL)</i>
(A)	100	900
(B)	100	1000
(C)	10	990
(D)	1	1000

2013 HSC Q15

10. Which of the following lists contains ONLY basic substances?

- (A) Oven cleaner, urine, vinegar
- (B) Lemonade, drain cleaner, blood
- (C) Baking soda, ammonia, sea water
- (D) Antacid, dishwashing detergent, lemon juice

2011 HSC Q3

11. Which of the following groups of household substances contains ONLY acidic substances?

- (A) Antacid tablets, baking soda, sugar
- (B) Aspirin, drain cleaner, toothpaste
- (C) Oven cleaner, soda water, milk
- (D) Orange juice, vinegar, wine

12. Which of the following changes take place when 50 mL of water is added to 50 mL of 0.1 mol L^{-1} acetic acid?

<i>pH</i>	<i>Degree of ionisation</i>
(A) Increase	Decrease
(B) Decrease	Increase
(C) Increase	Increase
(D) Decrease	Decrease

2012 HSC Q18

13. What mass of anhydrous sodium carbonate is required to neutralise 100.0 mL of 0.500 mol L^{-1} acetic acid?

- (A) 2.65 g
- (B) 5.30 g
- (C) 10.6 g
- (D) 53.0 g

2012 HSC Q19

14. A household cleaning agent contains a weak base with the formula NaX. 1.00 g of this compound was dissolved in water to give 100.0 mL of solution. A 20.0 mL sample of the solution was titrated with 0.100 mol L^{-1} hydrochloric acid, and required 24.4 mL of the acid for neutralisation.

What is the molar mass of the weak base?

- (A) 82.0 g
- (B) 84.0 g
- (C) 122 g
- (D) 410 g

2011 HSC Q18

15. In a research report a student wrote, ‘Acids are compounds that contain hydrogen and can dissolve in water to release hydrogen ions into solution.’

Who originally stated this theory of acids?

- (A) Arrhenius
- (B) Brönsted–Lowry
- (C) Davy
- (D) Lavoisier

2010 HSC Q8

Use the information provided to answer Questions 16 and 17.

Using 0.100 mol L^{-1} NaOH, a student titrated 25.0 mL of a 0.100 mol L^{-1} weak monoprotic acid, and separately titrated 25.0 mL of a 0.100 mol L^{-1} strong monoprotic acid.

16. Which statement about the volume of base required to reach the equivalence point is correct?
- (A) The weak acid will require the same volume of base as the strong acid.
 - (B) The weak acid will require a larger volume of base than the strong acid.
 - (C) The weak acid will require a smaller volume of base than the strong acid.
 - (D) The volume of base required will depend on the molar mass of the acid used.

2011 HSC Q15

17. Which statement correctly describes the pH at each titration equivalence point?
- (A) The pH of both solutions will be the same.
 - (B) One of the solutions will be neutral while the other will have a pH higher than 7.
 - (C) One of the solutions will be neutral while the other will have a pH lower than 7.
 - (D) One of the solutions will have a pH higher than 7 while the other will have a pH lower than 7.

2011 HSC Q16

18. A group of students produced a red solution by boiling red cabbage leaves in water. When dilute sodium hydroxide was added to the solution, it turned purple. When dilute hydrochloric acid was added to the red solution, no colour change occurred.

Which of these substances, when added, is most likely to cause the red solution to change colour?

- (A) Cleaning solution containing ammonia
- (B) Concentrated hydrochloric acid
- (C) Orange juice
- (D) Vinegar

2001 HSC Q7

- 19.** Citric acid, the predominant acid in lemon juice, is a triprotic acid. A student titrated 25.0 mL samples of lemon juice with 0.550 mol L^{-1} NaOH. The mean titration volume was 29.50 mL. The molar mass of citric acid is $192.12 \text{ g mol}^{-1}$.

What was the concentration of citric acid in the lemon juice?

- (A) 1.04 g L^{-1}
- (B) 41.6 g L^{-1}
- (C) 125 g L^{-1}
- (D) 374 g L^{-1}

2009 HSC Q14

- 20.** According to the Arrhenius theory of acids and bases, an acid is a substance that

- (A) tastes sour.
- (B) is capable of donating a hydrogen ion.
- (C) can accept a pair of electrons to form a co-ordinate covalent bond.
- (D) increases the concentration of hydrogen ions in an aqueous solution.

2008 HSC Q8

- 21.** 20 mL of 0.08 mol L^{-1} HCl is mixed with 30 mL of 0.05 mol L^{-1} NaOH.

What is the pH of the resultant solution?

- (A) 1.1
- (B) 2.7
- (C) 4.0
- (D) 7.0

2008 HSC Q14

- 22.** Acid X is 0.1 mol L^{-1} hydrochloric acid.

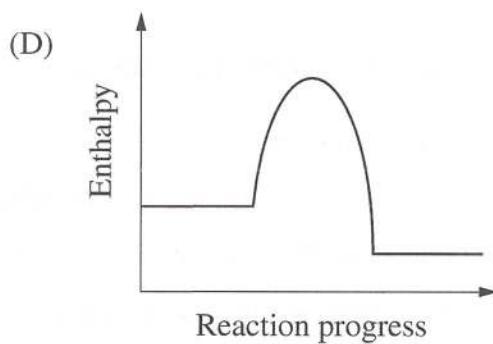
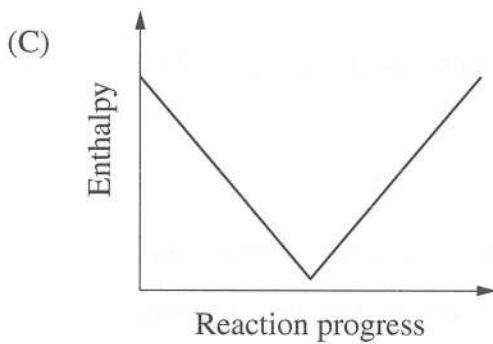
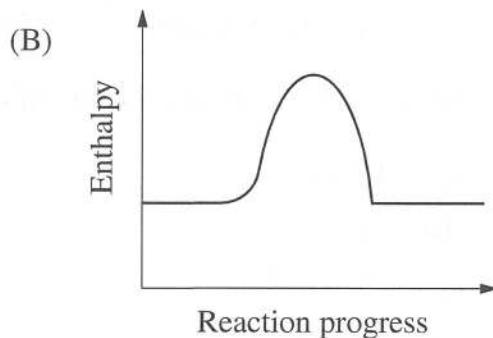
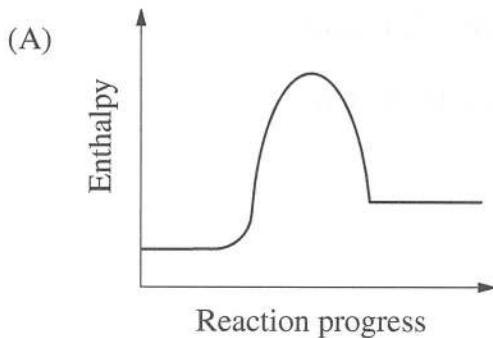
Acid Y is 1.0 mol L^{-1} acetic acid (ethanoic acid).

How does acid X compare with acid Y?

- (A) X is weaker and more dilute than Y.
- (B) X is stronger and more dilute than Y.
- (C) X is weaker and more concentrated than Y.
- (D) X is stronger and more concentrated than Y.

2006 HSC Q8

23. Which graph represents the enthalpy change for an acid-base neutralisation reaction?



2007 HSC Q7

24. A 0.1 mol L^{-1} HCl solution has a pH of 1.0.

What volume of water must be added to 90 mL of this solution to obtain a final pH of 2.0?

- (A) 10 mL
- (B) 180 mL
- (C) 810 mL
- (D) 900 mL

2007 HSC Q10

25. Which of these aqueous solutions would turn the indicator phenolphthalein pink?

- (A) HCl
- (B) NaCl
- (C) NaOH
- (D) CH_3OH

Adapted 2007 HSC Q6

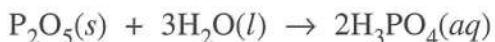
- 26.** Acid X and acid Y are both monoprotic weak acids of equal concentration. Acid X is a stronger acid than acid Y.

Which statement about acid X and acid Y is correct?

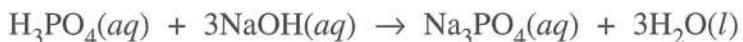
- (A) Acid Y is completely ionised in solution.
- (B) The solution of acid X is less ionised than the solution of acid Y.
- (C) The solution of acid X has a lower pH than the solution of acid Y.
- (D) 1 mole of acid Y requires a greater volume of 1.0 mol L^{-1} NaOH for neutralisation than 1 mole of acid X.

2007 HSC Q8

- 27.** Phosphorus pentoxide reacts with water to form phosphoric acid according to the following equation.



Phosphoric acid reacts with sodium hydroxide according to the following equation.



A student reacted 1.42 g of phosphorus pentoxide with excess water.

What volume of 0.30 mol L^{-1} sodium hydroxide would be required to neutralise all the phosphoric acid produced?

- (A) 0.067 L
- (B) 0.10 L
- (C) 0.20 L
- (D) 5.0 L

2006 HSC Q10

- 28.** In 1884, Svante Arrhenius proposed a definition for acids. His definition was soon accepted as superior to that put forward by earlier chemists.

Why was Arrhenius' definition seen as a major improvement?

- (A) It explained why some acids do not contain oxygen.
- (B) It showed how the solvent can affect the strength of an acid.
- (C) It showed the relationship between pH and the concentration of H^+ ions.
- (D) It could be used to explain why some acids are strong and others are weak.

2006 HSC Q11

29. Which statement best describes the equivalence point in a titration between a strong acid and a strong base?
- (A) The point at which the first sign of a colour change occurs
(B) The point at which equal moles of acid and base have been added together
(C) The point at which equal moles of H^+ ions and OH^- ions have been added together
(D) The point at which the rate of the forward reaction equals the rate of the reverse reaction

2006 HSC Q9

30. What would be the pH of a 0.1 mol L^{-1} solution of sulfuric acid?
- (A) Less than 1.0
(B) Exactly 1.0
(C) Between 1.0 and 7.0
(D) Greater than 7.0

2005 HSC Q8

31. Which of the following pairs would form a buffer solution?
- (A) $\text{HCl}(aq) / \text{Cl}^-(aq)$
(B) $\text{H}_2\text{PO}_4^-(aq) / \text{PO}_4^{3-}(aq)$
(C) $\text{H}_2\text{SO}_4(aq) / \text{HSO}_4^-(aq)$
(D) $\text{CH}_3\text{COOH}(aq) / \text{CH}_3\text{COO}^-(aq)$

2005 HSC Q9

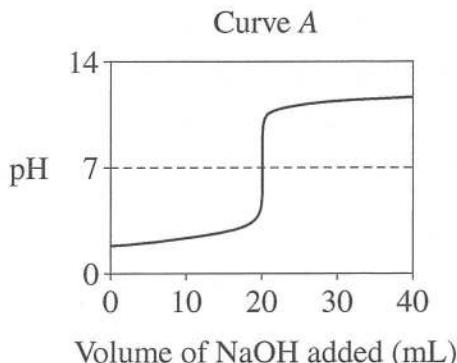
32. A sulfuric acid solution has a concentration of $5 \times 10^{-4} \text{ mol L}^{-1}$.
What is the pH of this solution, assuming the acid is completely ionised?
- (A) 3.0
(B) 3.3
(C) 3.6
(D) 4.0

2003 HSC Q8

33. The hydrogen carbonate ion (HCO_3^-) is amphiprotic. Which statement is correct?
- (A) HCO_3^- is the conjugate base of CO_3^{2-}
(B) H_2CO_3 is the conjugate acid of CO_3^{2-}
(C) CO_3^{2-} is the conjugate base of H_2CO_3
(D) H_2CO_3 is the conjugate acid of HCO_3^-

2000 HSC Q3

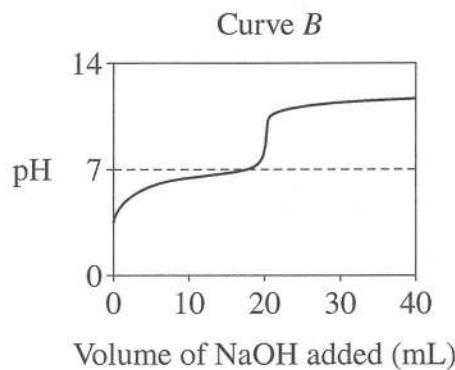
34. A titration was conducted by adding NaOH from a teflon-coated burette to HCl in a conical flask. The pH in the flask was recorded during the titration and Curve A was produced.



The table shows appropriate indicators used to identify the equivalence point in titrations. For NaOH and HCl the appropriate indicator is bromothymol blue.

Indicator	Acidic colour	Range of colour change	Basic colour
Methyl orange	red	3.1 – 4.4	yellow
Methyl red	red	4.4 – 6.2	yellow
Bromothymol blue	yellow	6.0 – 7.6	blue
Cresolphthalein	colourless	8.1 – 9.7	red
Alizarin yellow	yellow	10.1 – 12.0	red

A second titration was conducted by adding NaOH to a different acid. The pH in the flask was recorded during the titration and Curve B was produced.



What is the appropriate indicator for Curve B using the table?

- (A) Methyl orange
- (B) Methyl red
- (C) Cresolphthalein
- (D) Alizarin yellow

2005 HSC Q10

35. In a titration of a strong base with a strong acid, the following procedure was used:

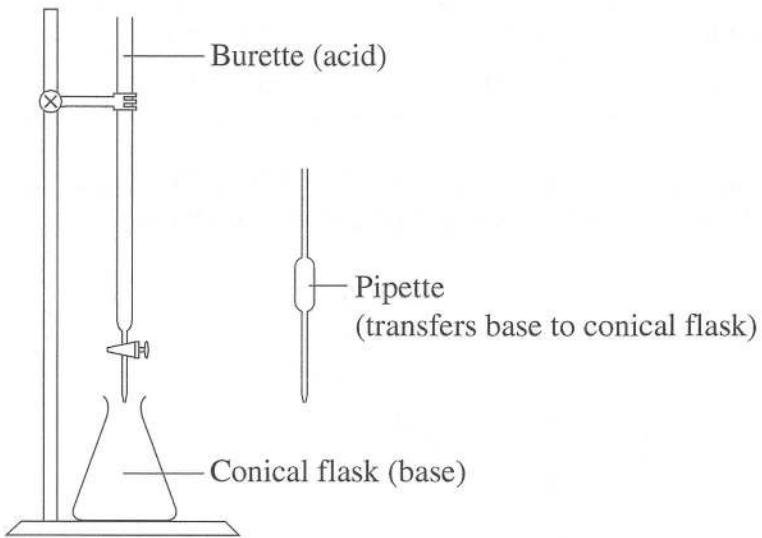
1. A burette was rinsed with water and then filled with the standard acid.
2. A pipette was rinsed with some base solution.
3. A conical flask was rinsed with some base solution.
4. A pipette was used to transfer a measured volume of base solution into the conical flask.
5. Indicator was added to the base sample and it was titrated to the endpoint with the acid.

Which statement is correct?

- (A) The calculated base concentration will be correct.
- (B) The calculated base concentration will be too low.
- (C) The calculated base concentration will be too high.
- (D) No definite conclusion can be reached about the base concentration.

2003 HSC Q14

36. In a titration, an acid of known concentration is placed in a burette and reacted with a base that has been pipetted into a conical flask.

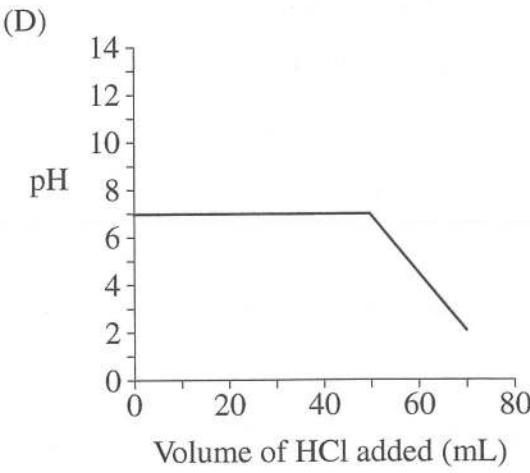
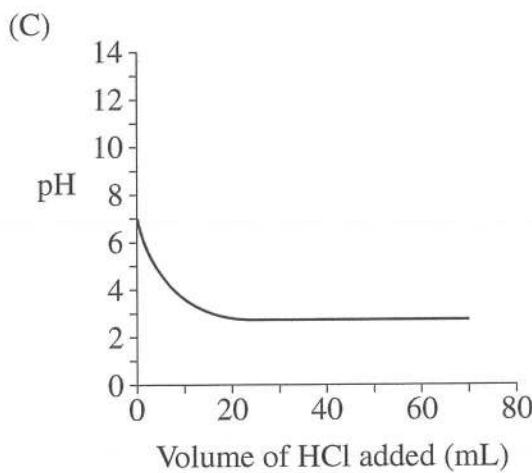
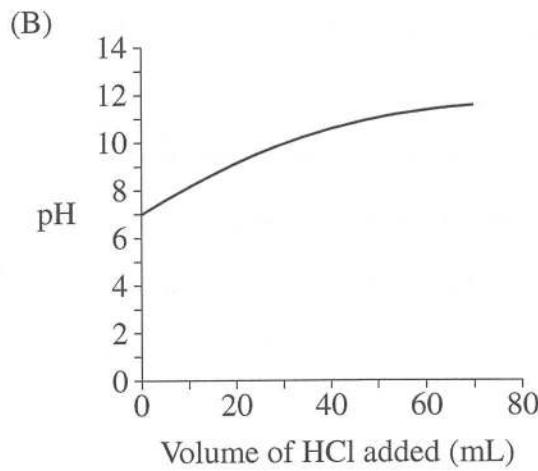
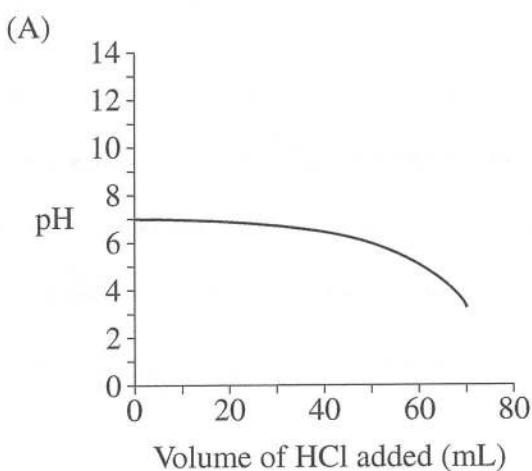


What should each piece of glassware be rinsed with immediately before the titration?

	<i>Burette</i>	<i>Pipette</i>	<i>Conical flask</i>
(A)	acid	base	water
(B)	water	water	water
(C)	acid	base	base
(D)	water	water	base

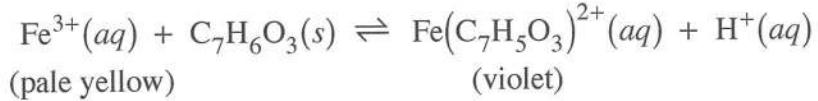
2002 HSC Q8

37. Which of the following graphs shows how pH will vary when dilute HCl is added to 100 mL of dilute natural buffer solution with an initial pH of 7.0?



2003 HSC Q15

38. Iron(III) forms a variety of coloured complex ions. One such example is that produced by the reaction with salicylic acid ($C_7H_6O_3$):



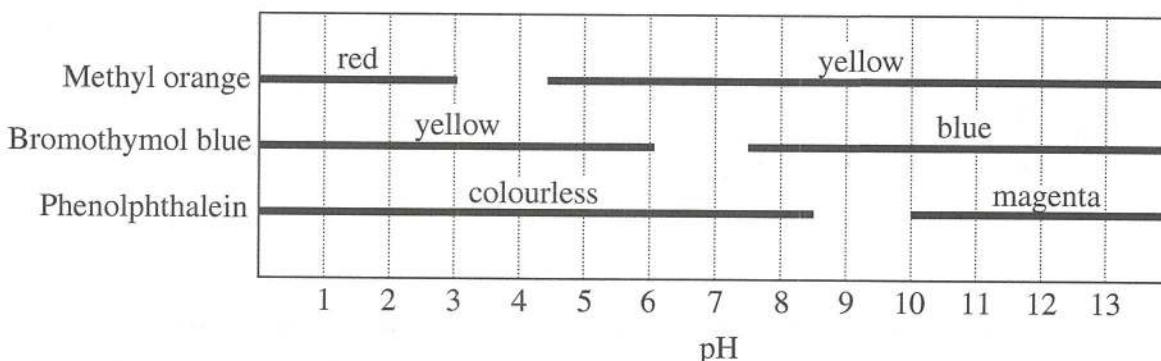
The yield of the violet complex, $\text{Fe}(\text{C}_7\text{H}_5\text{O}_3)^{2+}$, may be decreased by

- (A) adding $\text{C}_7\text{H}_6\text{O}_3(s)$ to the solution.
- (B) adding Fe^{3+} ions to the solution.
- (C) lowering the pH of the solution.
- (D) raising the pH of the solution.

1997 HSC Q10

Refer to the following information to answer Questions 39 and 40.

The graph shows the colour ranges of the acid–base indicators methyl orange, bromothymol blue and phenolphthalein.



- 39.** A solution is yellow in methyl orange, blue in bromothymol blue and colourless in phenolphthalein.

What is the pH range of the solution?

- (A) 4.5 to 6.0
- (B) 6.0 to 7.5
- (C) 7.5 to 8.5
- (D) 8.5 to 10.0

2001 HSC Q6

- 40.** Methyl orange, bromothymol blue and phenolphthalein indicators were mixed together to form a solution.

Over what pH range would the solution be yellow?

- (A) 0 – 14
- (B) 3 – 4.5
- (C) 3 – 7.5
- (D) 4.5 – 6

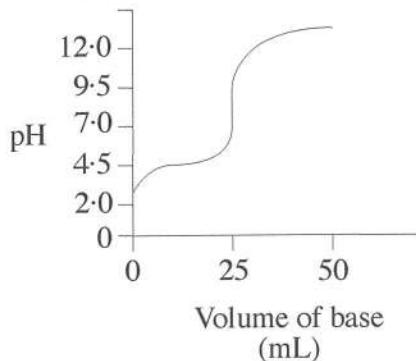
Adapted 2012 HSC Q7

- 41.** What is the conjugate base of HSO_4^- ?

- (A) SO_3^{2-}
- (B) SO_4^{2-}
- (C) H_2SO_4
- (D) HSO_3^-

2009 HSC Q7

42. Which of these is a Brönsted–Lowry definition for a base?
- A substance that donates a proton (H_3O^+) in a chemical reaction.
 - A substance that accepts a proton (H^+) in a chemical reaction.
 - A substance that donates a proton (H^+) in a chemical reaction.
 - A substance that can act as both a proton (H^+) donor and acceptor in a reaction.
43. A pH meter was used to trace the progress of an acid–base titration. 50 mL of $0\cdot100 \text{ mol L}^{-1}$ sodium hydroxide solution was added in 1 mL aliquots (samples) to 50 mL of an acid. The pH of the acid–base mixture was measured after each 1 mL aliquot was added. The results were plotted on the graph below.



Using the information above, it is reasonable to conclude that the acid was

- ethanoic (acetic) acid.
- hydrochloric acid.
- nitric acid.
- sulfuric acid.

1996 HSC Q8

44. Which of the following solutions has the lowest pH?

- 0·1M ethanoic (acetic) acid.
- 0·1M hydrochloric acid.
- 0·2M sodium hydroxide.
- 0·2M nitric acid.

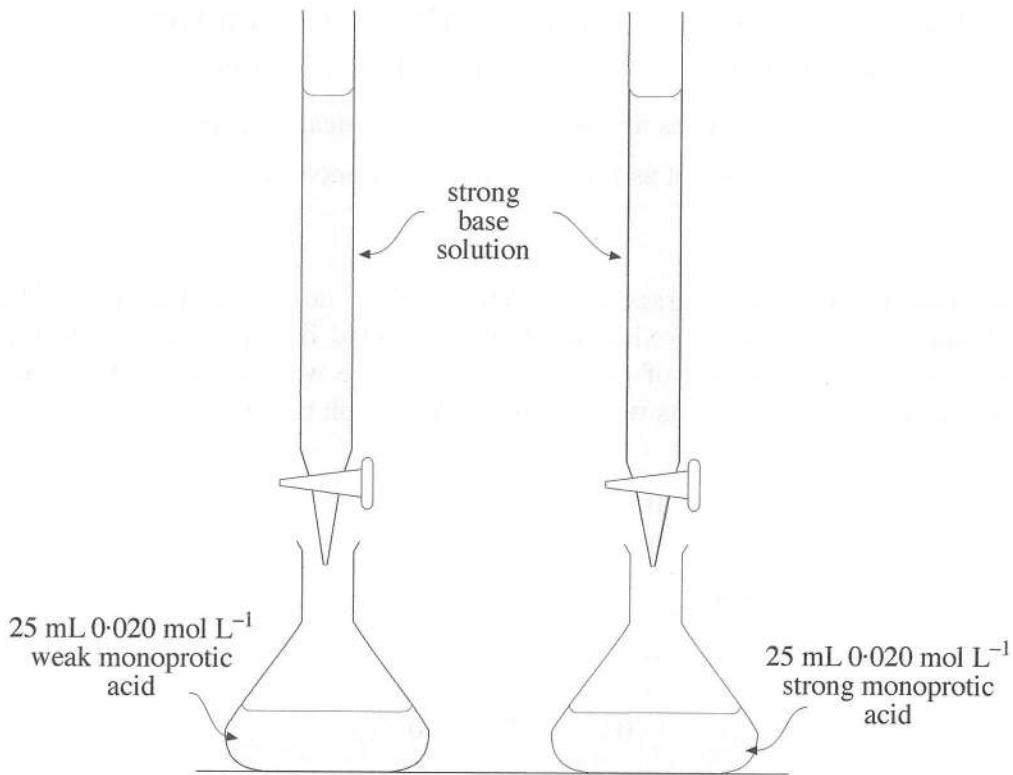
1995 HSC Q2

45. For a given weak acid, HA, the numerical value of K_a

- will change with the pH.
- will change with the temperature.
- cannot be less than 10^{-7} .
- cannot be greater than 10^{-7} .

1996 HSC Q2

46. Two titrations are carried out as shown below.



Which of the following statements is true?

- (A) The volume of base required to reach the equivalence point will depend upon the particular acid used.
- (B) The weak acid will require the same amount of base as the strong acid to reach the equivalence point.
- (C) The weak acid will require less base than the strong acid to reach the equivalence point.
- (D) The weak acid will require more base than the strong acid to reach the equivalence point.

1995 HSC Q10

47. The Brønsted–Lowry theory applies in both aqueous and non-aqueous systems.

The following reactions may take place in solvents other than water.

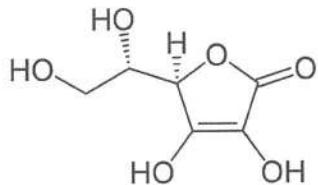
Which is NOT a Brønsted–Lowry reaction?

- (A) $\text{NH}_4^+ + \text{NH}_2^- \rightleftharpoons 2\text{NH}_3$
- (B) $\text{CO}_2 + \text{OH}^- \rightleftharpoons \text{HCO}_3^-$
- (C) $\text{HClO}_4 + \text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COOH}_2^+ + \text{ClO}_4^-$
- (D) $\text{CH}_3\text{CH}_2\text{O}^- + \text{CH}_3\text{NH}_3^+ \rightleftharpoons \text{CH}_3\text{CH}_2\text{OH} + \text{CH}_3\text{NH}_2$

1998 HSC Q9

48. The Aboriginal and Torres Strait Islander Peoples realised that various fruits, such as the quandong, had beneficial effects, e.g. for treating colds and to maintain healthy skin. By performing volumetric analysis such as titration today, chemists have determined the nature and concentration of various substances (acids and bases) contained within the fruits and plants used.

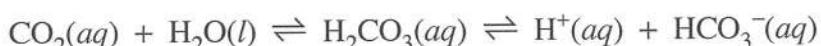
The quandong plant has been shown to have a high level of ascorbic acid (vitamin C). Ascorbic acid ($\text{HC}_6\text{H}_7\text{O}_6$) has the following structure:



How can the Aboriginal and Torres Strait Islander Peoples use of plants such as the quandong be explained?

- (A) Their medicinal use of this plant showed an in-depth knowledge of its properties.
 - (B) They understood the acidic properties of the quandong fruit.
 - (C) They analysed the content of the fruit and found it contained ascorbic acid.
 - (D) They found that the quandong plant had beneficial effects by trial and error.
49. The pH of blood is critical. If the pH falls below 7·4, the condition called acidosis results.

The following equilibrium system is involved in maintaining blood pH.



Acidosis may be reduced by

- (A) introducing more water from cells to shift equilibrium to the right.
- (B) adding an enzyme to catalyse this reaction.
- (C) introducing NaCl into the blood to change pH levels.
- (D) rapid breathing to reduce CO_2 levels in the blood.

1996 HSC Q14

50. Which of the following solutions would have the highest pH?

- (A) $0\cdot10 \text{ mol L}^{-1} \text{ HNO}_3$
- (B) $0\cdot10 \text{ mol L}^{-1} \text{ H}_2\text{SO}_4$
- (C) $0\cdot10 \text{ mol L}^{-1} \text{ HCOOH}$
- (D) $0\cdot10 \text{ mol L}^{-1} \text{ NaCl}$

1997 HSC Q4

51. Which of the following indicators would best identify the end-point in the titration of $0\cdot10 \text{ mol L}^{-1}$ lactic acid ($K_a = 1\cdot4 \times 10^{-4}$ at 25°C) with $0\cdot10 \text{ mol L}^{-1}$ KOH?

	<i>Indicator</i>	pH range
(A)	Methyl orange	3·1–4·4
(B)	Bromophenol blue	3·0–4·6
(C)	Methyl red	4·4–6·2
(D)	Phenol red	7·2–8·8

1997 HSC Q9

52. Equal volumes of four different acids are titrated with the same base at 25°C .

Information about these acids is given below.

<i>Acid</i>	<i>Concentration</i> (mol L^{-1})	pH
HCl	0·1	1·0
H_3PO_4	0·1	1·6
CH_3COOH	0·1	2·9
HCN	0·1	5·1

The acid requiring the greatest volume of base for complete reaction is

- (A) HCl
- (B) H_3PO_4
- (C) CH_3COOH
- (D) HCN

1997 HSC Q14

53. The pH of $0\cdot1 \text{ mol L}^{-1}$ hydrochloric acid (HCl) is about 1·0 and the pH of $0\cdot1 \text{ mol L}^{-1}$ phosphoric acid (H_3PO_4) is about 1·6.

The pH of the phosphoric acid solution is higher because

- (A) HCl dissociates more fully than H_3PO_4 in water.
- (B) H_3PO_4 has more hydrogen atoms.
- (C) HCl is a weaker acid than H_3PO_4 .
- (D) H_3PO_4 is amphiprotic in water.

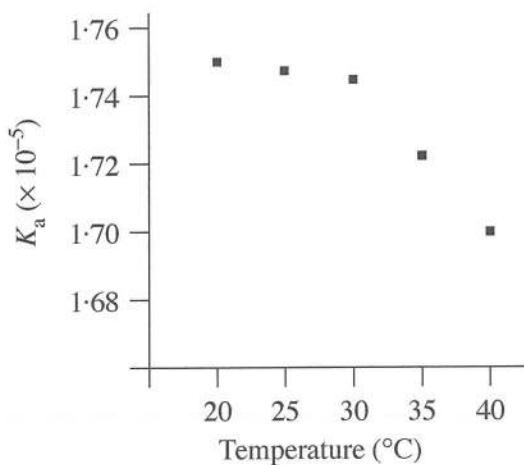
1999 HSC Q6

54. The function of a buffer in a natural system such as blood is to resist changes in

- (A) salt concentration.
- (B) reaction rate.
- (C) temperature.
- (D) pH.

1998 HSC Q13

55. The K_a of ethanoic acid changes with temperature as shown in the graph below.



The graph indicates that within the temperature range shown, as the temperature increases the acid becomes

- (A) more ionised.
- (B) less ionised.
- (C) more concentrated.
- (D) less concentrated.

1998 HSC Q14

56. A student wishes to determine the concentration of a weak acid by titration with a base of known concentration.

What is the most important property of the base?

- (A) It should be weak.
- (B) It should be strong.
- (C) It should be a primary standard.
- (D) It should not react with carbon dioxide.

2000 HSC Q6

57. $0\cdot100 \text{ mol L}^{-1}$ solutions of four different monoprotic acids were prepared. Their pH was measured at 298 K and recorded in the table below.

<i>Acids</i>	pH
<i>A</i>	1·0
<i>B</i>	2·2
<i>C</i>	4·0
<i>D</i>	5·2

The strongest acid is

- (A) acid *A*.
- (B) acid *B*.
- (C) acid *C*.
- (D) acid *D*.

1996 HSC Q10

58. At 25°C , the pH of a $0\cdot050 \text{ mol L}^{-1}$ Ba(OH)_2 solution is

- (A) 1·0
- (B) 1·3
- (C) 12·7
- (D) 13·0

1999 HSC Q14

59. Which of the following solutions has the lowest pH?

- (A) $0\cdot08 \text{ mol L}^{-1}$ sulfuric acid
- (B) $0\cdot08 \text{ mol L}^{-1}$ hydrochloric acid
- (C) $0\cdot20 \text{ mol L}^{-1}$ acetic (ethanoic) acid
- (D) $0\cdot20 \text{ mol L}^{-1}$ nitric acid

2000 HSC Q10

60. 50 mL of a $0\cdot200 \text{ mol L}^{-1}$ solution of sodium hydroxide is diluted to 2·0 L with distilled water.

What is the pH of the diluted solution?

- (A) 11·7
- (B) 12·3
- (C) 12·7
- (D) 13·3

2000 HSC Q14

Short-answer questions

Question 61 (3 marks)

A 0.001 mol L^{-1} solution of hydrochloric acid and a 0.056 mol L^{-1} solution of ethanoic acid both have a pH of 3.0. 3

Why do both solutions have the same pH?

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2010 HSC Q21

Question 62 (4 marks)

A gas is produced when 10.0 g of zinc is placed in 0.50 L of 0.20 mol L^{-1} hydrochloric acid. 4

Calculate the volume of gas produced at 25°C and 100 kPa. Include a balanced chemical equation in your answer.

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Adapted 2010 HSC Q26

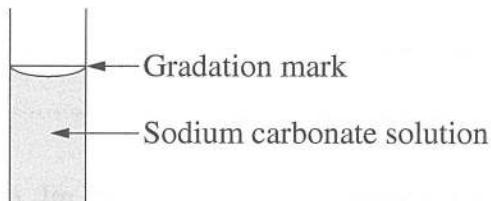
Question 66 (5 marks)

A student attempted to determine the concentration of a hydrochloric acid solution. The following steps were performed.

Step 1. A conical flask was rinsed with water.

Step 2. A 25.0 mL pipette was rinsed with water.

Step 3. The student filled the pipette with a standard sodium carbonate solution to the level shown in the diagram.



Step 4. The standard sodium carbonate solution in the pipette was transferred to the conical flask. The student ensured that all of the sodium carbonate solution was transferred to the conical flask by blowing through the pipette. Three drops of an appropriate indicator were added to the conical flask.

Step 5. A burette was rinsed with the hydrochloric acid solution and then filled with the acid. The student then carried out a titration to determine the concentration of the hydrochloric acid solution.

In steps 2, 3 and 4 above the student did not follow acceptable procedures.

- (a) Identify the mistake the student made in step 4 and propose a change that would improve the validity of the result. 2

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Question 66 continues

Question 66 (continued)

- (b) Explain the effect of the mistakes made in steps 2 and 3 on the calculation of the concentration of the hydrochloric acid solution. 3

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End of Question 66

2013 HSC Q28

Question 67 (4 marks)

- (a) Justify the continued use of the Arrhenius definition of acids and bases, despite the development of the more sophisticated Brönsted–Lowry definition. 3

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- (b) Why does the neutralisation of any strong acid in an aqueous solution by any strong base always result in a heat of reaction of approximately -57 kJ mol^{-1} ? 1

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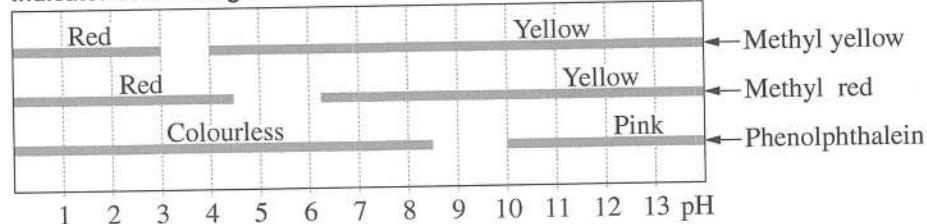
2011 HSC Q29

Question 68 (5 marks)

To determine the pH of garden soil, a sample was first saturated with distilled water in a petri dish. Barium sulfate powder was added to the surface of the sample, and drops of the three indicators listed below were added to separate parts of the sample. The colours observed are shown in the table.

Experimental results

Indicator	Methyl yellow	Methyl red	Phenolphthalein
Colour observed	Yellow	Red	Colourless

Indicator colour ranges**Plant response**

Plant	soil pH range for optimal growth
Carrot	5.5 – 6.8
Chrysanthemum	6.0 – 6.3
Hydrangea Blue	4.0 – 5.0
Hydrangea White	6.5 – 8.0
Potato	5.0 – 5.7

- (a) Why is barium sulfate powder added when testing soil pH? 1

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- (b) Using the information given, select the plant that will grow well at the current soil pH, and justify your selection. 2

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- (c) Outline the method you would use to test a natural indicator that has been prepared in the school laboratory. 2

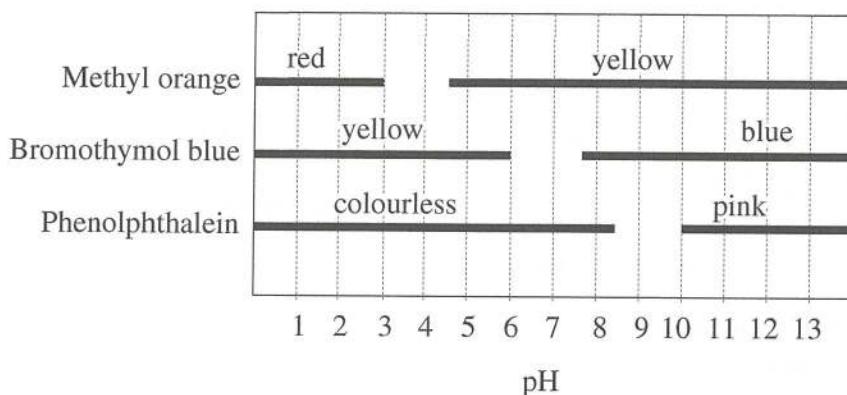
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2011 HSC Q32

Question 69 (5 marks)

Correct swimming pool maintenance requires regular monitoring of the pH level of the water.

- (a) Select the best indicator from the graph to check that the pH of swimming pool water lies within the correct range of 7.0 – 7.6. Justify your choice. 3



- (b) Another part of swimming pool maintenance is adjusting chlorine levels in the pool. ‘Liquid chlorine’ is a solution of sodium hypochlorite (NaOCl) which can be used to do this. Upon addition of sodium hypochlorite to the pool, the following equilibrium reaction occurs: 2



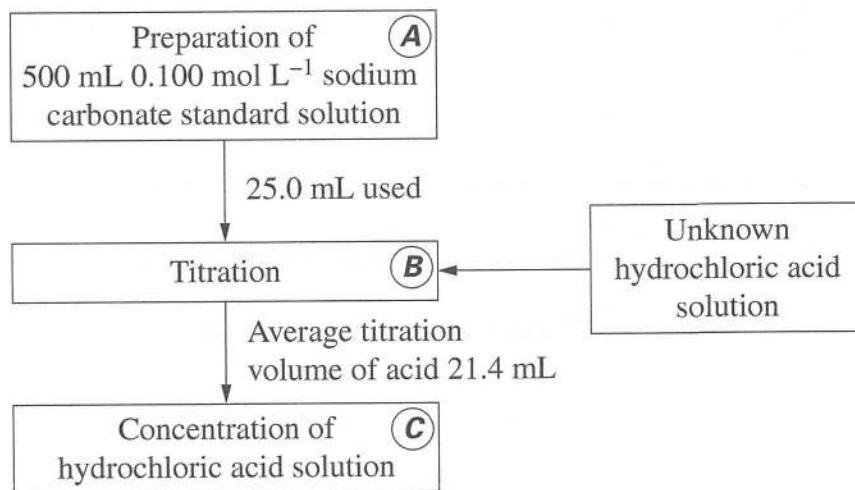
Explain how the addition of sodium hypochlorite will affect the pH of the water in the pool.

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2006 HSC Q23(a), (b)(ii)

Question 70 (8 marks)

The flowchart shown outlines the sequence of steps used to determine the concentration of an unknown hydrochloric acid solution.



Describe steps **A**, **B** and **C** including correct techniques, equipment and appropriate calculations. Determine the concentration of the hydrochloric acid.

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Question 70 continues

Question 70 (continued)

End of Question 70

2010 HSC Q28

Question 71 (7 marks)

The nitrogen content of bread was determined using the following procedure:

- A sample of bread weighing 2.80 g was analysed.
- The nitrogen in the sample was converted into ammonia.
- The ammonia was collected in 50.0 mL of 0.125 mol L⁻¹ hydrochloric acid. All of the ammonia was neutralised, leaving an excess of hydrochloric acid.
- The excess hydrochloric acid was titrated with 23.30 mL of 0.116 mol L⁻¹ sodium hydroxide solution.

- (a) Write balanced equations for the TWO reactions involving hydrochloric acid. 2

Question 71 continues

Question 71 (continued)

- (b) Calculate the moles of excess hydrochloric acid.

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- (c) Calculate the moles of ammonia.

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- (d) Calculate the percentage by mass of nitrogen in the bread.

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End of Question 71

2009 HSC Q22

Question 72 (4 marks)

- Explain how a buffer works with reference to a specific example in a natural system.

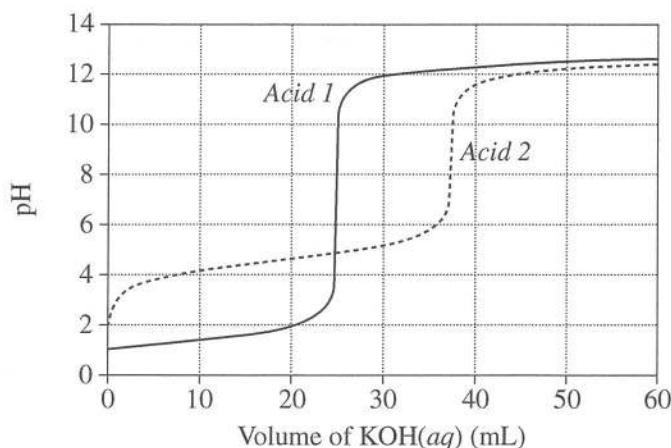
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2008 HSC Q26

Question 73 (6 marks)

The graph shows changes in pH for the titrations of equal volumes of solutions of two monoprotic acids, *Acid 1* and *Acid 2*.



- (a) Explain the differences between *Acid 1* and *Acid 2* in terms of their relative strengths and concentrations. 3

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- (b) Name the salt produced by the reaction of an acid of the same type as *Acid 2* with KOH(aq). 1

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- (c) Calculate the concentration of hydrogen ions when 20 mL of KOH(aq) has been added to *Acid 1*. 1

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- (d) Why would phenolphthalein be a suitable indicator for both titrations? 1

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Question 74 (3 marks)

Propanoic acid gives Swiss cheese its characteristic ‘sweet-nutty’ flavour. The concentration of this acid is important in the quality of the cheese. The pH of a $0\cdot200 \text{ mol L}^{-1}$ solution of propanoic acid is 2·78.

3

Calculate the value of its K_a under the same conditions. Show all working.

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Adapted 1999 HSC Q20

Question 75 (4 marks)

(a) Classify these salts as forming acidic, basic or neutral solutions.

2

<i>Salt</i>	<i>Classification of solution</i>
Ammonium chloride	
Sodium ethanoate	
Sodium chloride	
Ammonium nitrate	

(b) From the table, choose a salt that forms an acidic or basic solution, and justify its classification. Include an equation to illustrate your answer.

2

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2008 HSC Q27

Question 76 (6 marks)

A standard solution was prepared by dissolving 1.314 g of sodium carbonate in water. The solution was made up to a final volume of 250.0 mL.

- (a) Calculate the concentration of the sodium carbonate solution. 2

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This solution was used to determine the concentration of a solution of hydrochloric acid. Four 25.00 mL samples of the acid were titrated with the sodium carbonate solution. The average titration volume required to reach the end point was 23.45 mL.

- (b) Write a balanced equation for the titration reaction. 1

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2008 HSC Q28

Question 77 (2 marks)

Calculate the mass of solid sodium hydrogen carbonate required to make 250 mL of 0.12 mol L^{-1} solution. 2

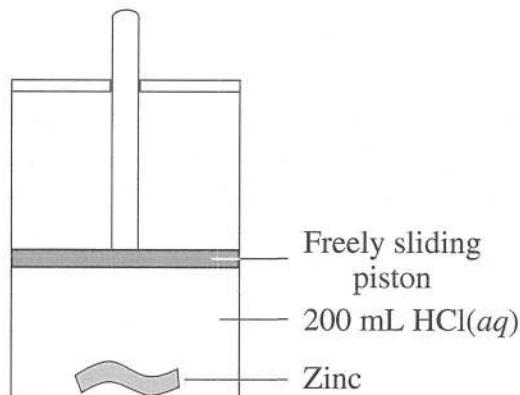
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2004 HSC Q16(b)

Question 78 (7 marks)

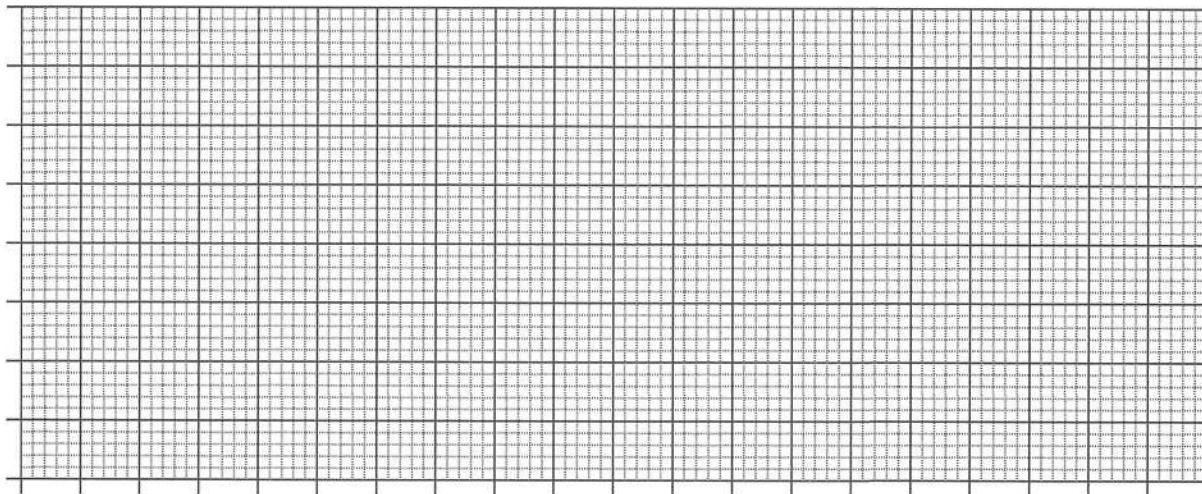
A student carried out a first-hand investigation to identify the relationship between the mass of a metal used in a reaction, and the volume of gas produced. The first-hand investigation was carried out at 25°C, and 100 kPa. In each experiment, 200 mL of hydrochloric acid was added to some zinc, and the volume of gas produced was recorded. The diagram shows the equipment used, and the table contains the student's results.

<i>Mass of zinc (g)</i>	<i>Volume of gas collected at 25°C and 100 kPa (mL)</i>
0.12	45
0.33	125
0.56	115
0.83	315
0.96	365
1.22	380
1.64	380
1.93	380



- (a) Graph the results from these experiments.

3



Question 78 continues

Question 78 (continued)

- (b) Predict the volume of gas that would be produced in this experiment if 3.00 g of zinc had been used. Justify your answer. 2

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- (c) Calculate the theoretical volume of gas produced at 25°C and 100 kPa by the reaction of 0.56 g of zinc with the 200 mL of hydrochloric acid. 2

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End of Question 78

2005 HSC Q22

Question 79 (5 marks)**Red cabbage indicator chart**

Colour	red	violet	purple	blue	green	yellow								
pH	1	2	3	4	5	6	7	8	9	10	11	12	13	14

- (a) State what colour the red cabbage indicator would be in a 0.005 mol L^{-1} solution of H_2SO_4 . Show your working. 1

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- (b) Using the red cabbage indicator, what colour would the solution be if 10 mL of 0.005 mol L^{-1} H_2SO_4 was diluted to 100 mL? 1

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Question 79 continues

Question 79 (continued)

- (c) What volume of 0.005 mol L^{-1} KOH is required to neutralise 15 mL of the diluted solution of H_2SO_4 ? 3

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End of Question 79*2007 HSC Q21***Question 80 (4 marks)**

- (a) Calculate the pH of a 0.2 mol L^{-1} solution of hydrochloric acid. 1

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- (b) Calculate the pH after 20 mL of 0.01 mol L^{-1} sodium hydroxide is added to 50 mL of 0.2 mol L^{-1} hydrochloric acid. Include a balanced chemical equation in your answer. 3

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2006 HSC Q17

Question 81 (5 marks)

An antacid tablet is known to contain calcium carbonate (CaCO_3). To determine the mass of calcium carbonate in the tablet, the following procedure was used.

- The tablet was crushed and then placed in a beaker.
- A pipette was used to add 25.0 mL of 0.600 mol L^{-1} hydrochloric acid to the crushed tablet in the beaker.
- Once the reaction between the calcium carbonate and hydrochloric acid had stopped, phenolphthalein indicator was added to the reaction mixture.
- A teflon-coated burette was then used to add 0.100 mol L^{-1} sodium hydroxide to the beaker to neutralise the excess hydrochloric acid.
- The phenolphthalein changed from colourless to pink after 14.2 mL of the sodium hydroxide solution had been added.

- (a) Write a balanced chemical equation for the reaction that occurred between the calcium carbonate in the tablet and the hydrochloric acid. 1

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- (b) How many moles of hydrochloric acid were added to the tablet? 1

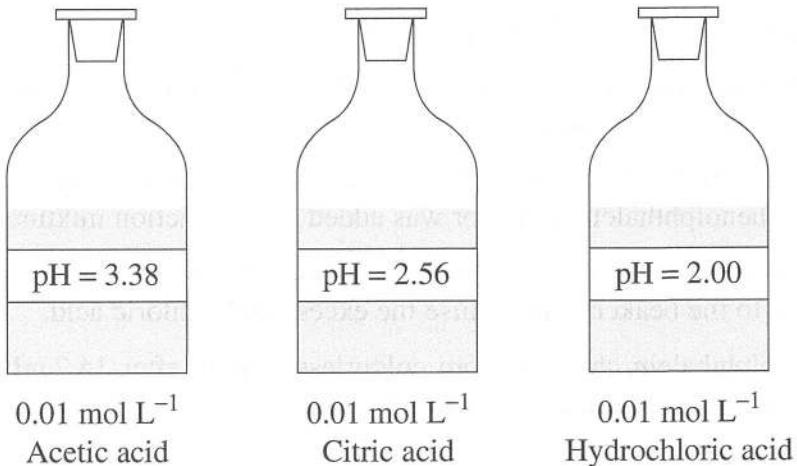
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- (c) Calculate the mass of calcium carbonate in the original antacid tablet. 3

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Question 82 (3 marks)

The diagram shows three reagent bottles containing acids.



- (a) Calculate the pH after 10.0 mL of 0.01 mol L^{-1} hydrochloric acid solution is diluted by the addition of 90.0 mL of distilled water. 1

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- (b) Explain the difference in pH between the three acids in the diagram. 2

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2004 HSC Q24(a) & (c)

Question 83 (4 marks)

25.0 mL of 0.12 mol L⁻¹ standard barium hydroxide solution was titrated with nitric acid. The results are recorded in the table.

<i>Titration</i>	<i>Volume of nitric acid used (mL)</i>
1	20.4
2	18.1
3	18.2
4	18.1

- (a) Write a balanced chemical equation for the reaction of barium hydroxide with nitric acid. 1

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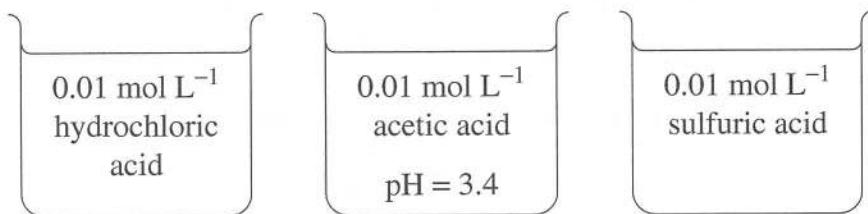
- (b) Calculate the concentration of the nitric acid. 3

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2003 HSC Q23

Question 84 (5 marks)

Solutions of hydrochloric acid, acetic acid and sulfuric acid were prepared. Each of the solutions had the same concentration (0.01 mol L^{-1}). The pH of the acetic acid solution was 3.4.



- (a) Calculate the pH of the hydrochloric acid solution. 1

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- (b) Compare the pH of the sulfuric acid solution to the pH of the hydrochloric acid solution. Justify your answer. (No calculations are necessary.) 2

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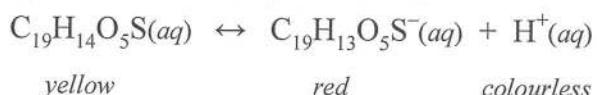
- (c) Explain why the acetic acid solution has a higher pH than the hydrochloric acid solution. 2

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2002 HSC Q22

Question 85 (5 marks)

When phenol red ($C_{19}H_{14}O_5S$) is dissolved in water, the following equilibrium is set up.



In aqueous solution, the molecule $C_{19}H_{14}O_5S$ is yellow, the ion $C_{19}H_{13}O_5S^-$ is red, and the ion H^+ is colourless.

Phenol red is a weak acid with $K_a = 1.3 \times 10^{-8}$.

- (a) Calculate the concentration of hydrogen ions in a 0.0030 mol L^{-1} solution of phenol red. Show working in your answer. 3

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- (b) Phenol red can be used as an acid/base indicator. 2

What colour would a solution of phenol red be if a few drops of it were placed in a 0.1 mol L^{-1} solution of hydrochloric acid? Explain your answer.

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Adapted 1993 HSC Q El 3(b)(i) & (ii)

Question 86 (3 marks)

The K_a of acetic acid (ethanoic acid, CH_3COOH) is 1.6×10^{-5} ($pK_a = 4.8$).

What is the pH of a 0.10 L^{-1} acetic acid solution? 3

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Adapted 1988 HSC Q El 3 (c)(i)

Question 87 (8 marks)

The acidity of wine is due mainly to potassium tartrate (cream of tartar), a weak monoprotic acid of molar mass 188 g mol⁻¹. Three 50 mL samples of wine were titrated with 0·010 M NaOH. The results of these titrations were as follows.

	<i>First titration</i>	<i>Second titration</i>	<i>Third titration</i>
Final reading	9·8 mL NaOH	10·1 mL	10·2 mL
Initial reading	0·0 mL NaOH	0·0 mL	0·1 mL
Volume used	9·8 mL NaOH	10·1 mL	10·1 mL

- (a) What would you use to rinse your pipette before beginning the experiment?

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- (b) Calculate the moles of NaOH used in this titration.

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- (c) What is the molar concentration of potassium tartrate in this wine?

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- (d) What is the concentration of potassium tartrate in g L⁻¹?

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Adapted 1998 HSC Q31

Answers

Module 6: Acid/Base Reactions

Multiple choice

- | | | | | | | | |
|-------|-------|-------|-------|-------|-------|-------|-------|
| 1. B | 2. B | 3. C | 4. B | 5. A | 6. C | 7. D | 8. B |
| 9. C | 10. C | 11. D | 12. C | 13. A | 14. A | 15. A | 16. A |
| 17. B | 18. A | 19. B | 20. D | 21. B | 22. B | 23. D | 24. C |
| 25. C | 26. C | 27. C | 28. D | 29. C | 30. A | 31. D | 32. A |
| 33. D | 34. C | 35. C | 36. A | 37. A | 38. C | 39. C | 40. D |
| 41. B | 42. B | 43. A | 44. D | 45. B | 46. B | 47. B | 48. D |
| 49. D | 50. D | 51. D | 52. B | 53. A | 54. D | 55. B | 56. B |
| 57. A | 58. D | 59. D | 60. A | | | | |

Explanations

- B** Brønsted and Lowry each independently proposed a theory about how acids and bases behave. The Brønsted–Lowry definitions are: an acid is a proton donor, a base is a proton acceptor. So (B) is the answer. (A) is incorrect as it had been known for a long time that acids contained hydrogen, without a full appreciation of the role of the hydrogen. (C) is incorrect, as several acids (e.g. hydrochloric acid, HCl) do not contain oxygen. (D) is incorrect as it is the Lewis definition of an acid, not the Brønsted–Lowry definition.
- B** Methyl orange is the only acid/base indicator, so (B) is the answer. Methanol is an alkanol, methanoic acid an alkanoic acid, and methyl ethanoate an ester.
- C** An amphiprotic species can either accept or donate a proton (i.e. it can act either as a base or as an acid). (A) and (B) are incorrect as they are solely proton donors. (D) is incorrect as it is solely a proton acceptor. (C) is the answer as it is amphiprotic, as shown by these two reactions:
 - HCO_3^- as a proton acceptor: $\text{HCO}_3^-(aq) + \text{H}_3\text{O}^+(aq) \rightarrow \text{H}_2\text{CO}_3(aq) + \text{H}_2\text{O}(l)$
 - HCO_3^- as a proton donor: $\text{HCO}_3^-(aq) + \text{OH}^-(aq) \rightarrow \text{CO}_3^{2-}(aq) + \text{H}_2\text{O}(l)$

- 4. B** $2\text{Na}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{NaOH}(aq) + \text{H}_2(g)$
 So, 1 mol $\text{Na}(s)$ releases 0.5 mol of $\text{H}_2(g)$.
 Molar mass, $M(\text{Na}(s)) = 22.99 \text{ g mol}^{-1}$
 So volume of gas produced = $0.5 \times 24.79 = 12.395$ litres (at 25°C and 100 kPa).
 So (B) is the answer.
- 5. A** The indicator is pink at quite acidic pH values of $\text{pH} < 3$. At $\text{pH} = 3$, the indicator turns pale yellow and is bright yellow at $\text{pH} = 7$ or more. So it would be a suitable indicator for the titration of a strong acid (HCl) and a weak base (NH_3) with an end point at low pH, as given in (A). Both (B) and (C) are incorrect as the end point is near to $\text{pH} = 7$ for acids and bases of similar strength. (D) is incorrect as the end point for the titration of a weak acid and a strong base is at a higher pH.
- 6. C** A standard solution requires the solid anhydrous mass to be weighed in a beaker, dissolved in water, then fully transferred into a volumetric flask using a filter funnel. The solution is then made up to the correct volume. So (C) is the answer. (A) is incorrect as this equipment would be used for a titration. (B) is incorrect as this equipment cannot be used to make any type of solution. A measuring cylinder is not sufficiently accurate for preparing a standard solution, so (D) is incorrect.
- 7. D** A buffer maintains a narrow pH range via a weak Brønsted-Lowry acid (or base) and its conjugate base (or acid), which will also be weak. Only (D) shows such a weak conjugate pair, so it is the answer. (A) and (C) are incorrect as they each involve a strong acid. (B) is incorrect as it involves a strong base.
- 8. B** In monoprotic acid solutions of equal volume and equal concentration, the quantity of H^+ ions available for a neutralisation reaction is the same whether it is a strong acid like hydrochloric acid or a weak acid like acetic acid. Therefore, the same amount of sodium hydroxide solution will be required to react in each case, i.e. 23.4 mL. So (B) is answer.
- 9. C** $[\text{H}^+(aq)]$ in a pH 3.0 solution is $10^{-3} \text{ mol L}^{-1}$ and $[\text{H}^+(aq)]$ in a pH 5.0 solution is $10^{-5} \text{ mol L}^{-1}$. So the dilution factor is 100. Only (C) has a dilution factor of 100 (from 10 mL to 1000 mL). So (C) is the answer. The dilution factor for (A) is 10 (from 100 to 1000 mL), for (B) it is 11 (from 100 to 1100 mL) and for (D) it is 1001 (from 1 to 1001 mL). So (A), (B) and (D) are incorrect.
- 10. C** Baking soda, ammonia and sea water are all basic, so (C) is the answer. Vinegar, lemonade and lemon juice are all acidic, so (A), (B) and (D) are all incorrect.

- 11. D** Orange juice contains citric acid, vinegar is dilute acetic (ethanoic) acid, and wine contains principally tartaric and malic acid with lesser amounts of other acids such as citric acid. So (D) is the answer as it contains three acidic substances. Antacid tablets in (A) are basic (to reduce gastric acidity). Drain cleaner in (B) and oven cleaner in (C) are also basic, so (A), (B) and (C) are incorrect as they each contain a non-acidic substance.
- 12. C** $\text{H}_2\text{O}(l) + \text{CH}_3\text{COOH} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CH}_3\text{COO}^-$
 When 50 mL water is added, it halves the $[\text{CH}_3\text{COOH}]$. So the equilibrium shifts to the right to favour the products and the degree of ionisation increases. Although H_3O^+ ions are produced, the overall $[\text{H}_3\text{O}^+]$ is less than it was before the water was added because the volume has been doubled. Since the $[\text{H}_3\text{O}^+]$ is less, the solution is less acidic, i.e. the pH has increased. So (C) is the answer.
- 13. A** Molar mass $\text{Na}_2\text{CO}_3 = (2 \times 22.99) + 12.01 + (3 \times 16.00) = 105.99 \text{ g mol}^{-1}$
 $\text{Na}_2\text{CO}_3(s) + 2\text{CH}_3\text{COOH}(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + 2\text{CH}_3\text{COO.Na}(aq)$
 From this equation, $n_{\text{sodium carbonate}} = \frac{1}{2} \times n_{\text{acetic acid}}$ for neutralisation
 $= \frac{1}{2} \times \frac{100.0}{1000} \times 0.500 = 0.025 \text{ mol}$
 Mass sodium carbonate $= 0.025 \times 105.99 = 2.65 \text{ g}$... as in (A).
- 14. A** $n(\text{HCl}) = \frac{24.4}{1000} \times 0.1 = n(\text{NaX})$ in 20 mL sample (contains 0.2 g NaX)
 $\therefore M(\text{NaX}) = \frac{\text{mass}}{\text{no. moles}} = \frac{0.2 \times 1000}{24.4 \times 0.01} = 81.97 \approx 82.0 \text{ g mol}^{-1}$... as in (A).
- 15. A** Arrhenius' acid theory (about 1884) is the one quoted, so (A) is the answer.
 Lavoisier (about 1776) stated that acids were compounds containing oxygen, so (D) is incorrect. Davy (about 1810) disproved Lavoisier by proving some acids such as HCl and H₂S did not contain oxygen, so (C) is incorrect. Brønsted-Lowry is not a person, but the name for a later theory of acids developed independently by Brønsted in Denmark and Lowry in England. This theory defines acids in terms of proton transfer. So (B) is incorrect.
- 16. A** At the equivalence point, $n(\text{H}^+) = n(\text{OH}^-)$. Since the volume and concentration of both acids is the same, and since the concentration of NaOH is also the same, equal volumes of acid and base will be required in each case to reach equivalence. So (A) is the answer. The volume of base required depends on the total H⁺ present and not on whether the acid is strong or weak, so (B) and (C) are incorrect. It is the concentration and not the molar mass that determines the volume required, so (D) is incorrect.

- 17. B** At the equivalence point, anions of the weak acid will combine with H⁺ ions from the solution releasing more OH⁻ into the solution, so increasing the pH above 7, whereas the solution formed by the strong acid will have pH = 7. So (B) is the only possible answer.

- 18. A** The solution from the cabbage leaves acts as an indicator. It remains red for acids (such as hydrochloric acid) and turns purple with bases (such as sodium hydroxide). (A) is the answer, as the ammonia solution is basic. (B), (C) and (D) are incorrect, since concentrated hydrochloric acid, orange juice (containing citric acid) and vinegar (containing acetic acid) are all acidic.

19. B

$$n(\text{NaOH}) = 0.550 \text{ mol L}^{-1} \times 29.5 \times 10^{-3} \text{ L} = 0.016225 \text{ mol}$$

$$n(\text{citric acid}) = \frac{0.016225}{3} = 5.408 \times 10^{-3} \text{ mol in 25.0 mL (as citric acid is triprotic)}$$

$$\therefore [\text{citric acid}] = \frac{1000}{25.0} \times 5.408 \times 10^{-3} \text{ mol L}^{-1}$$

$$= \frac{1000}{25.0} \times 5.408 \times 10^{-3} \times 192.12 = 41.559 \approx 41.6 \text{ g L}^{-1} \dots \text{as in (B).}$$

- 20. D** Arrhenius acids are in aqueous solution, and are defined as a substance that, when dissolved in water, increases the concentration of H⁺ ions, so (D) is the answer. (B) is incorrect as there is no mention of water, and it is the Brønsted-Lowry definition of an acid. Acids have a sour taste, but this was not a part of the Arrhenius theory, so (A) is incorrect. Electron pair transfer is involved in the Lewis acid theory, not the Arrhenius theory of acids, so (C) is incorrect.

21. B

$$\text{NaOH}(aq) + \text{HCl}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)$$

$$n(\text{HCl}) = 0.08 \times \frac{20}{1000} = 1.6 \times 10^{-3} \text{ mol}$$

$$n(\text{NaOH}) = 0.05 \times \frac{30}{1000} = 1.5 \times 10^{-3} \text{ mol}$$

$$n \text{ excess HCl} = 0.1 \times 10^{-3} \text{ mol in 50 mL of solution}$$

$$[\text{H}^+] = \frac{1000}{50} \times 0.1 \times 10^{-3} = 2.0 \times 10^{-3} \text{ mol L}^{-1}$$

$$\therefore \text{pH} = -\log 2.0 \times 10^{-3} = 2.7 \quad \dots \text{as in (B).}$$

- 22. B** The acetic acid solution is 10× the concentration of the hydrochloric acid solution, so acid X is *more dilute* than acid Y, as in (A) or (B). Acid strength/weakness relates to how much the acid ionises in water. HCl is a strong acid, having nearly 100% ionisation. Acetic acid is a weaker acid that dissociates in an equilibrium reaction. So acid X is *stronger* than acid Y, as in (B) or (D). So (B) is the answer.

- 23. D** Neutralisation reactions are exothermic so the enthalpy of the products is less than that of the reactants – so (D) is the answer as it shows this change in enthalpy, and (A) and (B) are both incorrect. Enthalpy change diagrams show the enthalpy of the reactants as a horizontal line on the left of the graph and the enthalpy of the products as a horizontal line on the right of the graph. So (C) is incorrect.
- 24. C** HCl is a strong acid and is therefore completely ionised. Thus 0.1 mol L^{-1} HCl has pH=1 and 0.01 mol L^{-1} HCl has pH = 2. As $c_1V_1 = c_2V_2$ (where c = molarity and V = volume), so $0.1 \times 90 = 0.01 \times V_2$ \therefore Final volume, $V_2 = 900 \text{ mL}$, and so 810 mL must be added to the 90 mL to obtain a pH of 2.0, as in (C).
- 25. C** Phenolphthalein is an indicator – it is colourless in solutions below pH 8 and is pink/crimson in solutions above pH 9. So phenolphthalein would be pink in NaOH as it is an alkali, so (C) is the answer. It would be colourless in HCl, an acid, and in NaCl, a salt. So (B) and (C) are incorrect. (D) is incorrect as methanol is neutral, so phenolphthalein would be colourless in it.
- 26. C** Acids X and Y are weak, so neither will be fully ionised. So (A) is incorrect. Since acid X is stronger than acid Y , it will be more ionised than acid Y , so (B) is incorrect. So, acid X will have a greater $[\text{H}^+]$ than acid Y and so its pH will be lower, as in (C). One mole of any monoprotic acid requires the same amount of NaOH for neutralisation whether it is strong or weak, so (D) is incorrect.
- 27. C** Molar mass (P_2O_5) = $(2 \times 30.97) + (5 \times 16.00) = 141.94 \text{ g mol}^{-1}$
 $n(\text{P}_2\text{O}_5) = \frac{1.42}{141.94} = 0.01000 \text{ mol}$
Molar ratio of $\text{P}_2\text{O}_5:\text{H}_3\text{PO}_4 = 1:2$ $\therefore n(\text{H}_3\text{PO}_4) = 2 \times 0.0100 = 0.0200 \text{ mol}$
Molar ratio of $\text{H}_3\text{PO}_4:\text{NaOH} = 1:3$ $\therefore n(\text{NaOH}) = 3 \times 0.0200 = 0.0600 \text{ mol}$
volume (NaOH) = $\frac{0.06}{0.30} = 0.20 \text{ L}$... so (C) is the answer.
- 28. D** Arrhenius proposal was that acids and bases ionise in water to respectively produce H^+ and OH^- , and also that the strength of an acid/base depended on how readily they ionised in water. So (D) is the answer. Arrhenius' definition of acids was unrelated to the presence of oxygen, so (A) is incorrect. Arrhenius was only concerned with water as the solvent, so (B) is incorrect. The pH scale was first proposed by Sörensen many years after Arrhenius, so (C) is incorrect.

- 29. C** The neutralisation reaction is: $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$, as in (C). The colour change of an indicator is used to signal the equivalence point, but an indicator is not necessary for an equivalence point to be reached, so (A) is incorrect. Equal mol of acid and base are only sometimes present at equivalence, so (B) is incorrect. The rate of forward and reverse reactions is not relevant to describing the equivalence point, so (D) is incorrect.
- 30. A** Sulfuric acid will ionise almost completely in a dilute 0.1 mol L^{-1} solution:
 $\text{H}_2\text{SO}_4(aq) \rightarrow \text{SO}_4^{2-}(aq) + 2\text{H}^+(aq)$ For a solution of $0.1 \text{ mol L}^{-1} \text{ H}_2\text{SO}_4$... the
 $[\text{H}^+] \approx 0.2 \text{ mol L}^{-1} \quad \therefore \text{pH} \approx -\log_{10}(0.2) \approx 0.7$ This is < 1.0 , as in (A).
- 31. D** A buffer solution contains a weak acid and its conjugate base. The only acid in the list that is a weak acid is CH_3COOH . So (D) is the answer.
- 32. A** Sulfuric acid is diprotic so $[\text{H}^+] = 2 \times [\text{H}_2\text{SO}_4]$
 $[\text{H}^+] = 2 \times (5 \times 10^{-4}) = 1 \times 10^{-3} \text{ mol L}^{-1} \quad \therefore \text{pH} = -\log_{10}(1 \times 10^{-3}) = 3.0$
- 33. D** When HCO_3^- acts as a base and accepts a proton it becomes H_2CO_3 . So H_2CO_3 is its conjugate acid, as in (D). When HCO_3^- acts as an acid and donates a proton, it becomes CO_3^{2-} . So CO_3^{2-} is the conjugate base of HCO_3^- . (A), (B) and (C) are incorrect as they have incorrect combinations.
- 34. C** In the second titration, the equivalence point is around pH 8.5–9.0. The only indicator that is appropriate for this pH range and which starts and finishes changing colour on the vertical part of the graph is cresolphthalein, as in (C).
- 35. C** Actions 2, 4 and 5 do not affect the moles of acid or base involved. Action 1 makes the acid more dilute than expected, so it will require a greater volume of acid to neutralise the base. Thus, the base concentration will come out too high. Action 3 puts more moles of base in the flask, so a greater volume of acid will have to be added in the titration. The calculation will again produce a value for the base concentration that is higher than it really is. So (C) is the answer.

- 36. A** The titration flask must be washed with water, because although water would dilute the solutions coming into the flask, their volume has already been measured. You must not wash the flask with the acid or base, as that would change the quantity of acid or base in the flask by an unknown amount. Burettes and pipettes should only be rinsed with a little of the solution that is to be used in them, as rinsing with water would dilute the solution that is used in the titration. So the correct combination is in (A), and (B), (C) and (D) are all incorrect.
- 37. A** A buffer will neutralise most (but not all) of the added acid. So, the pH will still fall, but not drastically – until all the buffer is used up. Then the pH will begin to drop as the solution becomes acidic. Only the graph in (A) shows this behaviour.
- 38. C** Lowering the pH would increase $[H^+]$ and the equilibrium would shift towards reactants to use up some of the additional H^+ thus reducing the yield of the violet complex. So (C) is the answer. Both (A) and (B) are incorrect since adding additional amounts of either reactant will increase rather than decrease the yield of the violet complex. (D) is incorrect since raising the pH would reduce $[H^+]$ and this would increase the yield of products.
- 39. C** The pH range occurs where the three data overlap. The yellow bar for methyl orange, the blue bar for bromothymol blue and the colourless bar for phenolphthalein only overlap each other from pH 7.5 to 8.5, as in (C).
- 40. D** The solution will only be yellow when both methyl orange and bromothymol blue are yellow. This occurs from pH 4.5–6, as in (D). Phenolphthalein is colourless in this pH range and so will not affect the colour. Within the range pH 3–4.5, methyl orange will give a reddish hue and within the range pH 6–7.5, bromothymol blue will give a greenish hue, so (B) and (C) are incorrect. (A) is incorrect as the pH range 0–14 will give the whole range of colours.
- 41. B** The hydrogen sulfate ion, HSO_4^- is amphoteric. When it acts as an acid, it produces H^+ and sulfate ions: $HSO_4^- \leftrightarrow H^+ + SO_4^{2-}$. So SO_4^{2-} is the conjugate base, as in (B). When it acts as a base, the conjugate acid is sulfuric acid, H_2SO_4 , so (C) is incorrect. The substances shown in (A) and (D) are conjugate base and acid of sulfur in a different oxidation state so (A) and (D) are incorrect.
- 42. B** Acids are proton (H^+) donors, while bases are proton (H^+) acceptors. The symbol for a hydrogen ion (proton) can be either H^+ or H_3O^+ . So (B) is the answer.

- 43. A** The pure acid has $\text{pH} \approx 3$, as shown on graph where base = 0. At the equivalence point, $\text{pH} \approx 8$. \therefore It is a weak acid. Ethanoic acid is a weak acid, as in (A). So (A) is the answer. (B), (C) and (D) all name strong acids whose solution will have $\text{pH} \leq 1$ and so are incorrect.
- 44. D** The most acidic solution will have the lowest pH, i.e the highest $[\text{H}^+]$. (A) is incorrect as ethanoic acid is a weak acid, $pK_a << 5$. (C) is incorrect as the solution is alkaline, $\text{pH} > 7$. (B) and (D) are both strong acids and fully ionised. So $[\text{H}^+]$ in (D) is twice the $[\text{H}^+]$ in (B). Hence (D) has the lowest pH. So (D) is the answer.
- 45. B** $K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$ The numerical value of the acid dissociation constant, K_a is a characteristic of each particular acid. The K_a value is usually given at 25°C since it varies with temperature as in (B). (A) is incorrect because the equilibrium constant does not vary with the concentration of the reactants and pH is a measure of $[\text{H}^+]$. Both (C) and (D) are incorrect as K_a values have a range from $>> 1$ for strong acids to well below 10^{-7} for very weak acids.
- 46. B** Both the strong and weak acids are present in equal quantities and have the same concentration. So, at the equivalence point, equal numbers of moles of acid and base have neutralised one another, as in (B). The strong acid is completely ionised, so all its H^+ are present in solution initially. A weak acid is only ionised to a very slight extent, and so as it reacts, and H^+ are removed from the solution shifting the weak acid's equilibrium, this results in the release of more H^+ until the equivalence point is reached. So (A) is incorrect as it will not depend on the particular monoprotic acid used, and (C) and (D) are incorrect since each contradict (B).
- [Note: A solution of a weak monoprotic acid contains very few H^+ ions, whereas a solution of a strong monoprotic acid contains a large number of H^+ ions, as it is ionised completely and the weak acid is only ionised to a very slight extent. So the initial pH of the strong acid is lower. While the number of moles at equivalence is the same, the pH at equivalence will be higher for the weak acid since the salt formed from the weak acid hydrolyses in water producing OH^- ions.]
- 47. B** Brønsted–Lowry reactions involve the transfer of a single H^+ from one species to another, as in (A), (C) and (D). This is not the case in (B) where the two species combine, so the answer is (B).
- 48. D** Aboriginal and Torres Strait Islander Peoples did not use any chemical analysis techniques, nor did they have an understanding of the chemistry of acids/bases. But rather, they acquired knowledge through trial and error. So (D) is the answer.

- 49. D** Acidosis occurs when the blood pH is too low, i.e. $[H^+]$ too high. To overcome this, the equilibrium must move to the left, so (A) is incorrect. (B) is incorrect because a catalyst does not change the position of equilibrium. NaCl is the salt of a strong base and a strong acid and will not change the pH, so (C) is also incorrect. If $[CO_2(aq)]$ is reduced, the equilibrium will shift to the left and $[H^+]$ will decrease increasing the pH. So (D) is the answer.
- 50. D** Solutions in (A) and (B) contain strong acids, so $pH \approx 1$. Solution in (C) contains a weak acid, so pH will be about 3. Solution in (D) contains the salt of a strong base and a strong acid, so its $pH \approx 7$. So it is the most alkaline and has the highest pH. Hence (D) is the answer.
- 51. D** Lactic acid is a weak acid and is being titrated with a strong base solution. The end-point will therefore be at $pH > 7$ so indicators (A), (B) and (C) are unsuitable. (D) is the answer as the phenol red range occurs above $pH = 7$.
- 52. B** The concentration of the acid and number of acid hydrogens determines the volume of base required. The pH of the acid solution is irrelevant. The concentrations of all four acids is 0.1 mol L^{-1} . However, H_3PO_4 is a triprotic acid and so requires three times the amount of base. So (B) is the answer. The other acids (HCl, CH_3COOH and HCN) are monoprotic and so each will require a lesser amount of base than H_3PO_4 . So (A), (C) and (D) are incorrect.
- 53. A** At $pH = 1$, $[H^+] = 0.1 \text{ mol L}^{-1}$. So 0.1 mol L^{-1} HCl has completely dissociated if it has $pH = 1$. If H_3PO_4 completely dissociated, it would have $[H^+] = 0.3 \text{ mol L}^{-1}$ and a $pH = -\log_{10}[H^+] \approx 0.5$. The higher pH of 1.6 indicates that H_3PO_4 is only partially dissociated. So (A) is the answer. H_3PO_4 does have more hydrogen atoms as in (B), but this is irrelevant to the question. (C) and (D) are both incorrect as HCl is a stronger acid than H_3PO_4 and H_3PO_4 is not amphiprotic.
- 54. D** A buffer system involves competing equilibria with a common anion so that changes in $[H^+]$ or $[OH^-]$ result in only minimal change in pH, as in (D).



The graph shows that as the temperature increases, the value of K_a decreases. So there are less ions as in (B), rather than more as in (A). So (B) is the answer. The equilibrium constant indicates the proportion of acid present in ionic or molecular form regardless of concentration. So (C) and (D) are incorrect.

- 56. B** To accurately determine the end point close to the equivalence point of the neutralisation reaction, the student would need to use a strong base, as in (B) ... not a weak base, as in (A). So (B) is the answer.

It is not necessary for the strong base to be a primary standard, nor for it to be unreactive to CO_2 . So (C) and (D) are incorrect. A strong base, e.g. NaOH would be suitable. While it is a strong base, it is not a primary standard, nor is it stable in the atmosphere and may react with $\text{CO}_2(g)$ from the air. A strong base can have its concentration determined separately at the time it is to be used by the student.

- 57. A** A strong acid is completely ionised, so all its H^+ are present in solution initially. A weak acid is only ionised to a very slight extent. The lower the pH of monoprotic acid solutions of equal concentration, the greater is the $[\text{H}^+]$ and therefore the greater the degree of dissociation of the acid. Acid A has the lowest pH, so it is the strongest acid. Hence (A) is the answer.

- 58. D** $\text{Ba(OH)}_2 \rightleftharpoons \text{Ba}^{2+} + 2\text{OH}^-$ So if $[\text{Ba(OH)}_2(aq)] = 0.050 \text{ mol L}^{-1}$
then $[\text{OH}^-(aq)] = 0.100 \text{ mol L}^{-1} = 10^{-1} \text{ mol L}^{-1}$
Since $[\text{H}^+] [\text{OH}^-] = 10^{-14}$ $\therefore [\text{H}^+] = 10^{-13}$ and $\text{pH} = 13.0$, as in (D).

- 59. D** Sulfuric acid, hydrochloric acid and nitric acid are all strong acids. So, they will all be completely ionised in dilute solutions, such as 0.08 or 0.20 mol L⁻¹. Hence their $[\text{H}^+]$ will be 0.16, 0.08 and 0.20 mol L⁻¹ respectively (sulfuric acid is diprotic so its $[\text{H}^+]$ is double that of hydrochloric acid). Acetic acid, being a weak acid, only partially ionises and will have a very much lower $[\text{H}^+]$ than the others. The solution with the highest $[\text{H}^+]$ will have the lowest pH. So (D) is the answer.

- 60. A** Original 50 mL sample is diluted to 2.0 L

$$c_1 V_1 = c_2 V_2, \text{ so } c_2 = \frac{c_1 V_1}{V_2}$$

$$c_2 = \frac{0.200 \times 0.05}{2} = 5 \times 10^{-3} \text{ mol L}^{-1}$$

$$\text{So } [\text{NaOH}] = 5 \times 10^{-3} \text{ mol L}^{-1}$$

$$\therefore [\text{OH}^-] = 5 \times 10^{-3} \text{ mol L}^{-1}$$

$$\text{pOH} = -\log_{10}[\text{OH}^-] = -\log_{10}(5 \times 10^{-3}) = 2.30$$

$$\therefore \text{pH} = 14 - \text{pOH} = 14 - 2.30 = 11.7 \dots \text{as in (A).}$$

Short-answer questions

- 61.** Hydrochloric acid (HCl) is a strong acid, so it is fully dissociated (ionised).
 \therefore in 0.001 mol L^{-1} HCl(*aq*), $[\text{H}^+]$ will be 0.001 and $\text{pH} = -\log_{10}(0.001) = 3$
 Acetic (ethanoic) acid (CH_3COOH) is a weak acid, so hardly dissociates (ionises) at all.
 So $[\text{CH}_3\text{COOH}]$ will remain close to 0.056 mol L^{-1} and $[\text{H}^+]$ will be 0.001
 \therefore both solutions have the same pH.
- 62.** $\text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)$
 \therefore 1 mol Zn reacts with 2 mol HCl and produces 1 mol H_2
 $n(\text{Zn}) = \frac{10.0}{65.41} = 0.1529 \text{ mol}$ $n(\text{HCl}) = 0.50 \times 0.20 = 0.1 \text{ mol}$
 So there is an excess of Zn metal and HCl is the limiting reactant.
 \therefore 0.1 mol HCl will produce $0.05 \text{ mol } \text{H}_2 = 0.05 \times 24.79 = 1.24 \text{ L}$ at 25°C and 100 kPa
- 63.** An indicator is an aqueous solution of a weak acid that changes colour over a narrow pH range and is a weak acid/conjugate base pair at equilibrium:

$$\begin{array}{ccc} \text{H}^+(aq) + \text{Ind}^-(aq) & \leftrightarrow & \text{HInd}(aq) \\ \text{green} & & \text{red} \end{array}$$
Hind is a weak acid. When the pH is low, the equilibrium will shift to the right to form more of the red weak acid. When the pH is high, it will form more of the green Ind^- . Since this solution is red in acid and green in alkali, it can be used as an indicator.
- 64.** $n(\text{H}^+) = \text{mol H}^+ = 0.120 \times \frac{75.00}{1000} = 0.009$
 $n(\text{OH}^-) = \text{mol OH}^- = 0.200 \times \frac{25.00}{1000} = 0.005$
 $\therefore \text{H}^+$ is in excess. Excess $\text{H}^+ = 0.009 - 0.005 = 0.004$
 New volume of solution = $75.00 + 25.00 = 100.00 \text{ mL}$
 So new $[\text{H}^+] = 0.004 \text{ mol / } 100 \text{ mL} = 0.04 \text{ mol L}^{-1}$ (using $c = n/V$)
 $\therefore \text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}0.04 = 1.398 \approx 1.4$
- 65.** (a) $2\text{NaOH}(aq) + 2\text{HCl}(aq) \rightarrow 2\text{NaCl}(aq) + \text{H}_2\text{O}(l)$
 As $n(\text{NaOH}) = n(\text{HCl})$, $25.75 \times 10^{-3} \times [\text{OH}^-] = 25.00 \times 10^{-3} \times 0.1034$
 $\text{Molarity NaOH} = [\text{OH}^-] = \frac{25.00 \times 10^{-3} \times 0.1034}{25.75 \times 10^{-3}} = 0.100388 \approx 0.1004 \text{ mol L}^{-1}$

(b) (i) Molar mass, $M(C_9H_8O_4) = (9 \times 12.01) + (8 \times 1.008) + (4 \times 16.00)$
 $= 180.154 \text{ g mol}^{-1}$

Average volume for titration = $\frac{16.60 + 16.55 + 16.55}{3} = 16.55 \text{ mL}$

$n(\text{aspirin}) \text{ in } 35 \text{ mL sample} = 0.1004 \times \frac{16.55}{1000} \text{ mol}$

Mass aspirin in 35 mL sample = $0.1004 \times \frac{16.55}{1000} \times 180.154 = 0.2993 \text{ g}$

∴ average mass of aspirin per tablet = 299 mg

(ii) Ethanol is less polar than water, so it is a better solvent for aspirin.

66. (a) The student should not have blown through the pipette. To improve the validity, the end of the pipette should be allowed touch the inside surface of the flask, to draw out the liquid.
- (b) In Step 2 – the pipette should have been rinsed out with the standard sodium carbonate solution, instead of being rinsed with water as this would dilute the standard solution.
- In Step 3 – the base of the meniscus of the sodium carbonate solution in the pipette should be level with the graduation mark, otherwise less sodium carbonate solution will be delivered.
- Both of these errors mean that less acid is required to reach the titration end-point. So the calculated acid concentration will be higher than the actual value.
67. (a) Arrhenius defined acids as substances that produce hydrogen ions in aqueous solutions, while bases produce hydroxide ions in aqueous solution. The Brønsted-Lowry definition of acids as proton donors and bases as proton acceptors allowed the behaviour of many other substances and their reactions to be explained, e.g. buffers and non-aqueous solvents. However, in everyday situations, most acid base reactions take place in aqueous solution, so the Arrhenius concept gives an adequate explanation of reactions in this case. It also explains acid and base strength in terms of degree of ionisation, as well as the pH system used to express the degree of acidity/alkalinity. This justifies its continued use in many situations.
- (b) All strong acids and strong bases react in the same way in neutralisation. So the reaction taking place in each case is actually the same exothermic reaction and so releases the same amount of energy.

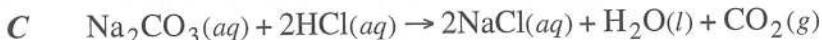
[Note: The reaction is always $\text{H}_3\text{O}^+(aq) + \text{OH}^-(aq) \rightarrow 2\text{H}_2\text{O}(l) \quad \Delta H \approx -57 \text{ kJ mol}^{-1}$]

- 68.** (a) White, insoluble barium sulfate powder absorbs the soil water and the indicator colour can be more readily seen against a white background.
- (b) Hydrangea Blue – these require an optimal pH in the range 4.0–5.0. They are the only plants that will grow optimally at the current soil pH of 4.0–4.5.
- (c) Use a red cabbage solution prepared in the school laboratory. Add 1 mL of this solution into each of 5 test tubes. Then add 5 mL of one of the following to each test tube: distilled water, 0.1 M sodium hydroxide, 0.1 M hydrochloric acid, 0.1 M acetic acid and 0.1 M ammonia solution. Set up a sixth test tube as a control with only 1 mL of red cabbage solution. Observe any colour changes.

[Note: You are not asked for the results. However, the results with red cabbage indicator are: acidic solutions give a red colour, neutral solutions give a purple colour and basic solutions give a greenish-yellow colour.]

- 69.** (a) Bromothymol blue is the best indicator to use – as it undergoes a colour change between pH 6–7.7 and this overlaps the optimal pH of 7.0–7.6 for pool water. The other indicators cannot be used as they do not change colour in the desired pH range, e.g. methyl orange remains yellow for pH = 4.5 and above, while phenolphthalein remains colourless from a low pH up to pH = 8.7
- (b) Addition of NaOCl increases the $[OCl^-]$. This in turn causes the equilibrium to move to the right (Le Châtelier) forming additional OH^- , so the pH of the water increases.
- 70. A** Molar mass, $M(Na_2CO_3) = (2 \times 22.99) + 12.01 + (3 \times 16.00) = 105.99 \text{ g mol}^{-1}$
 Concentration $1\text{M } Na_2CO_3 = 105.99 \text{ g L}^{-1}$
 \therefore concentration $0.100 \text{ M } Na_2CO_3 = 10.599 \text{ g L}^{-1} = 5.2995 \text{ g per 500 mL}$
1. Dry some analytical grade anhydrous Na_2CO_3 by heating until it has constant mass. Allow it to cool in a desiccator.
 2. Accurately weigh a beaker and add 5.2995 g of Na_2CO_3 .
 3. Transfer Na_2CO_3 to a 500 mL volumetric flask ... use distilled water from a wash bottle to ensure all Na_2CO_3 is transferred from the beaker and funnel to the volumetric flask.
 4. Half fill flask with distilled water, stopper it and swirl to dissolve all Na_2CO_3 .
 5. Add distilled water to flask until bottom of meniscus is level with the mark.

- B**
- Rinse a cleaned 25.00 mL pipette with the standard Na_2CO_3 solution and add 25.0 mL of this solution to a cleaned conical flask.
 - Add a few drops of phenolphthalein to the flask.
 - Rinse out a cleaned burette with the acid solution and adjust until the meniscus is at the zero mark of the burette.
 - Titrate the acid against the Na_2CO_3 solution until the indicator starts to become colourless and then continue drop by drop until it is just colourless.
 - Record result and repeat titration at least three times to improve reliability.
 - Eliminate any discrepant results and then calculate the average titration.



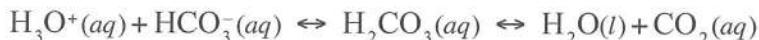
$$n(\text{HCl}) = 2 \text{ mol} \quad n(\text{CO}_3^{2-}) = 1 \text{ mol}$$

$$\therefore [\text{HCl}] = \frac{2 \times 0.100 \times 25.0}{21.4} = 0.2336 \approx 0.234 \text{ M}$$

[Note: You could give an ionic equation instead: $\text{CO}_3^{2-}(aq) + 2\text{H}_3\text{O}^+(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{CO}_2(g)$

- 71.** (a) 1. $\text{NH}_3(g) + \text{HCl}(aq) \rightarrow \text{NH}_4\text{Cl}(aq)$
 2. $\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)$
- (b) n excess $\text{HCl}(aq) \equiv$ moles $\text{NaOH}(aq)$ used $= 23.30 \times 10^{-3} \text{ L} \times 0.116 \text{ mol L}^{-1}$
 $\therefore n$ excess $\text{HCl}(aq) = 2.70 \times 10^{-3} \text{ mol}$
- (c) Original $n \text{ HCl}(aq) = 50.0 \times 10^{-3} \text{ L} \times 0.125 \text{ mol L}^{-1} = 6.25 \times 10^{-3} \text{ mol}$
 n excess $\text{HCl}(aq) = 2.70 \times 10^{-3} \text{ mol}$
 $\therefore n(\text{HCl})$ neutralised by $\text{NH}_3(g) = (6.25 \times 10^{-3}) - (2.70 \times 10^{-3}) \text{ mol}$
 $= 3.55 \times 10^{-3} \text{ mol}$
- From part (a), $n \text{ NH}_3(g) \equiv n \text{ HCl}(aq)$
 $\therefore n \text{ NH}_3(g) = 3.55 \times 10^{-3} \text{ mol}$
- (d) Molar mass, $M(\text{NH}_3) = 14.01 + (3 \times 1.008) = 17.034 \text{ g mol}^{-1}$
 1 mol NH_3 contains 1 mol N atoms
 Mass of nitrogen in sample $= 14.01 \times 3.55 \times 10^{-3} \text{ g}$
 \therefore Percentage nitrogen in bread $= \frac{14.01 \times 3.55 \times 10^{-3}}{2.80} \times \frac{100}{1} = 1.78\%$

72. A buffer consists of a weak acid/base and its conjugate pair. A buffer solution is able to resist a change in pH when a small amount of acid or base is added. An equilibrium is established which, when an acid or base is added, will shift to counteract the change in pH, e.g. in the human body, the maintenance of optimal pH levels is important to ensure enzymes work properly in metabolism. A buffer operates in the blood to maintain a narrow pH around 7.4 and involves this equilibrium:



If CO_2 levels increase, equilibrium shifts to the left to give a lower pH. If more HCO_3^- are formed, equilibrium shifts to the right to produce more CO_2 and pH increases.

[Note: Respiration is the process that produces CO_2 . In healthy people, the normal breathing rate varies to maintain the correct ratio of $[\text{CO}_2\text{(aq)}]$ to $[\text{HCO}_3^{\text{-(}}\text{aq}\text{)}]$, and so helps to maintain the blood at a constant pH level.]

73. (a) Neutralisation of *Acid 1* has an ‘end-point’ at $\text{pH} = 7$ and *Acid 2* has an end-point at $\text{pH} = 8.5$ when titrated with a strong alkali, such as KOH.

Acid 1 is a strong acid (fully ionised). *Acid 2* is a weak acid (partially ionised).

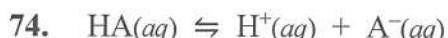
Acid 1 is neutralised by 25 mL of KOH(aq) , whereas *Acid 2* is neutralised by 37.5 mL. So *Acid 2* has a greater concentration in mol L^{-1} than *Acid 1*, because it takes 50% more KOH(aq) to neutralise *Acid 2*.

- (b) potassium acetate

[Note: This salt is produced by acetic acid. But you could have a potassium salt of any other *weak* acid.]

- (c) At 20 mL of KOH(aq) , $\text{pH} = 2 = -\log_{10}[\text{H}^+]$ $\therefore [\text{H}^+] = 10^{-2} = 0.01 \text{ mol L}^{-1}$

- (d) Both titrations have an ‘end-point’ in the colour change range of phenolphthalein.



$$\text{pH} = 2.78 \quad \therefore [\text{H}^+] = 10^{-2.78} = 1.66 \times 10^{-3}$$

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{(1.66 \times 10^{-3})(1.66 \times 10^{-3})}{(0.2 - 1.66 \times 10^{-3})} \approx 1.39 \times 10^{-5}$$

OR By taking the approximation that $(0.200 - 1.66 \times 10^{-3}) \approx (0.200)$ as propanoic acid is weakly ionised, $K_a \approx \frac{(1.66 \times 10^{-3})(1.66 \times 10^{-3})}{(0.200)} \approx 1.38 \times 10^{-5}$

75. (a)

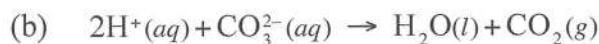
Salt	Classification of solution
Ammonium chloride	acidic
Sodium ethanoate	basic
Sodium chloride	neutral
Ammonium nitrate	acidic

- (b) Sodium ethanoate fully ionises: $\text{NaCH}_3\text{COO}(aq) \rightarrow \text{Na}^+(aq) + \text{CH}_3\text{COO}^-(aq)$
 The sodium ethanoate solution is basic.
 Since ethanoic acid is a weak acid, the ethanoate ions react with water to form undissociated ethanoic acid molecules: $\text{CH}_3\text{COO}^- + \text{H}_2\text{O}(l) \leftrightarrow \text{CH}_3\text{COOH} + \text{OH}^-$
 The surplus OH^- ions raise the pH to make it basic, while the Na^+ ion is neutral.

76. (a) Molar mass, $M(\text{Na}_2\text{CO}_3) = (2 \times 22.99) + 12.01 + (3 \times 16.00) = 105.99 \text{ g mol}^{-1}$

$$n(\text{Na}_2\text{CO}_3) = \frac{1.314}{105.99} \text{ mol per 250.0 mL}$$

$$[\text{Na}_2\text{CO}_3] = \frac{1.314}{105.99} \times 4 = 0.0496 \text{ mol L}^{-1} \quad (\text{to 4 sig figs})$$



- (c) Molar ratio HCl : $\text{Na}_2\text{CO}_3 = 2:1$

$$\text{So } [\text{HCl}] = \frac{23.45}{25.00} \times 2 \times [\text{Na}_2\text{CO}_3] = \frac{23.45}{25.00} \times 2 \times \frac{1.314}{105.99} \times 4 = 0.09303 \\ = 0.093 \text{ mol L}^{-1} \quad (\text{to 4 sig figs})$$

77. $n(\text{NaHCO}_3) = \text{molarity} \times V(\text{in L}) = 0.12 \text{ mol L}^{-1} \times 0.25 \text{ L} = 0.03 \text{ mol}$

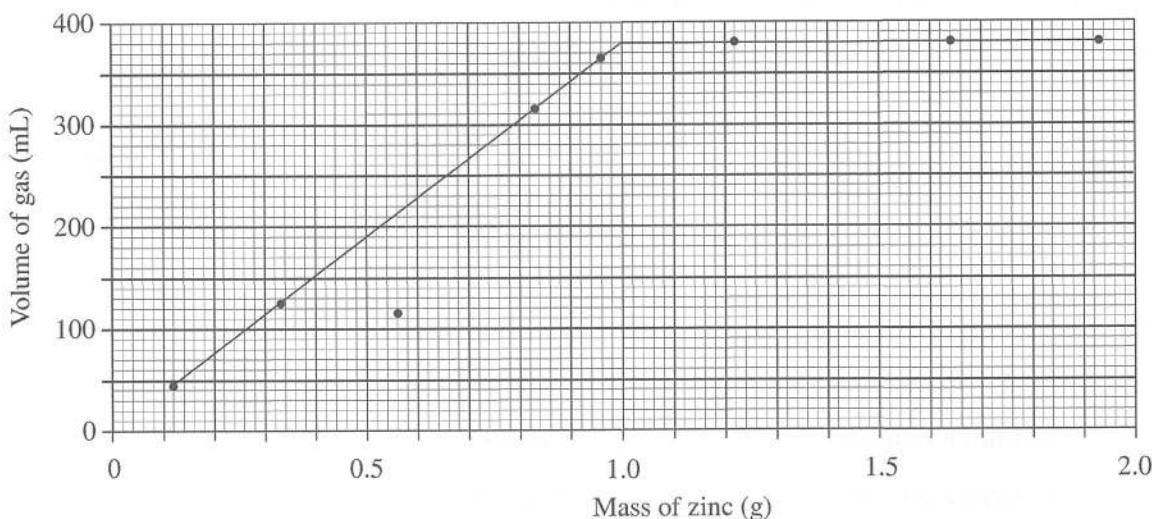
Molar mass, $M(\text{NaHCO}_3) = 22.99 + 1.008 + 12.01 + (3 \times 16.00) = 84.008 \text{ g mol L}^{-1}$

Mass, $m(\text{NaHCO}_3) = M(\text{NaHCO}_3) \times n(\text{NaHCO}_3)$

$$= 84.008 \times 0.03 = 2.52 = 2.5 \text{ g} \quad (\text{to 2 sig figs})$$

78. (a)

Volume of gas produced in an experiment as mass of zinc was increased



[Note: The student's value for 0.56 g has been treated as an outlier.]

- (b) 380 mL of gas would be produced – if more than approx 1 g of zinc is used, it is in excess and HCl is the limiting factor. So increasing the amount of zinc to 3.00 g will not result in any more gas being produced.
- (c) $\text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)$
 Molar mass, $M(\text{Zn}) = 65.41 \text{ g mol}^{-1}$
 $n(\text{Zn}) = \frac{0.56}{65.41} \approx 0.008561 \approx 8.56 \times 10^{-3} \text{ mol}$ and $n(\text{H}_2) = 8.56 \times 10^{-3} \text{ mol}$
 $\therefore \text{volume of H}_2 = 8.56 \times 10^{-3} \times 24.79 = 0.2122024 \text{ L} = 210 \text{ mL}$ (to 2 sig figs)

79. (a) $[\text{H}^+] = 2 \times [\text{H}_2\text{SO}_4] = 2 \times 0.005 = 0.01 \text{ mol L}^{-1}$ [Note: H_2SO_4 is a strong acid and 0.005 mol L^{-1} is a very dilute solution, so assume that the acid will be completely ionised.]
 $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}[0.01] = 2$
 $\therefore \text{red cabbage indicator will be red.}$
- (b) If the H_2SO_4 is diluted $10\times$, $[\text{H}^+] = 0.001 \text{ mol L}^{-1}$ $\therefore \text{pH} = -\log_{10}[0.001] = 3$
 $\therefore \text{red cabbage indicator will be violet.}$
- (c) $\text{H}_2\text{SO}_4(aq) + 2\text{KOH}(aq) \rightarrow \text{K}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}(l)$
 At neutralisation: $n(\text{H}^+) = n(\text{OH}^-)$ and $c_A V_A = c_B V_B$
 So: $[\text{H}^+]_A \times V_A = [\text{OH}^-]_B \times V_B$
 $2 \times [\text{H}_2\text{SO}_4] \times V_A = 0.005 \text{ mol L}^{-1} \times V_B$
 $2 \times 0.0005 \times 15 = 0.005 \times V_B$
 $V_B = \frac{2 \times 0.0005 \text{ mol L}^{-1} \times 15 \text{ mL}}{0.005 \text{ mol L}^{-1}} = 3.0 \text{ mL}$
 $\therefore \text{Volume of KOH needed is } 3.0 \text{ mL}$

80. (a) $[\text{H}_3\text{O}^+] = 0.2 \text{ mol L}^{-1}$
 $\therefore \text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}0.2 = 0.7$
- (b) $\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq)$
 $n(\text{HCl}) = \frac{50}{1000} \times 0.2 = 0.01 \text{ mol}$
 $n(\text{NaOH}) = \frac{20}{1000} \times 0.01 = 0.0002 \text{ mol}$
 $0.0002 \text{ mol NaOH} = 0.0002 \text{ mol HCl}$
 $n(\text{HCl}) \text{ left} = 0.01 - 0.0002 = 0.0098 \text{ mol in } 70 \text{ ml} = 0.0098 \times \frac{1000}{70} = 0.14 \text{ mol L}^{-1}$
 $\therefore \text{Final pH} = -\log_{10}0.14 = 0.85$

- 81.** (a) $\text{CaCO}_3(aq) + 2\text{HCl}(g) \rightarrow \text{CaCl}_2(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)$
- (b) $n(\text{HCl}) = \frac{25.0}{1000} \times 0.600 = 0.015 \text{ mol}$
- (c) $\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)$
 $n(\text{NaOH}) = 0.0142 \times 0.100 = 1.42 \times 10^{-3} \text{ mol}$
 $\text{HCl} : \text{NaOH} = 1:1 \quad \therefore \text{NaOH neutralises } 1.42 \times 10^{-3} \text{ mol HCl}$
 $\therefore n(\text{HCl}) \text{ reacted with CaCO}_3 = 0.015 - 1.42 \times 10^{-3} = 0.01358 \text{ mol}$
 $\text{CaCO}_3 : \text{HCl} = 1:2 \quad \therefore n(\text{CaCO}_3) = \frac{0.1358}{2} = 6.79 \times 10^{-3} \text{ mol}$
 $\text{Molar mass, } M(\text{CaCO}_3) = 40.08 + 12.01 + (16 \times 3) = 100.09 \text{ g L}^{-1}$
 $\therefore \text{mass of CaCO}_3 = 6.79 \times 10^{-3} \times 100.09 = 0.679611 \text{ g} = 0.680 \text{ g (to 3 sig figs)}$

- 82.** (a) $c_1V_1 = c_2V_2$
 $c_2 = 0.01 \text{ mol L}^{-1} \times 10 \text{ mL} / 100 \text{ mL} = 0.001 \text{ mol L}^{-1}$
 $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}0.001 = 3.0$
- (b) The difference in pH is due to the different $[\text{H}^+]$ in each acid solution, which is due to the different degree of ionisation in each acid. pH of HCl is the lowest, and since $\text{pH} = -\log_{10}[\text{H}^+]$, HCl has the greatest $[\text{H}^+]$ and hence the greatest degree of ionisation. The $[\text{H}^+]$ and hence degree of ionisation is lower in citric acid, and lower still in acetic acid, which means that these acids are weaker than HCl, with acetic acid being the weakest of the three acids.

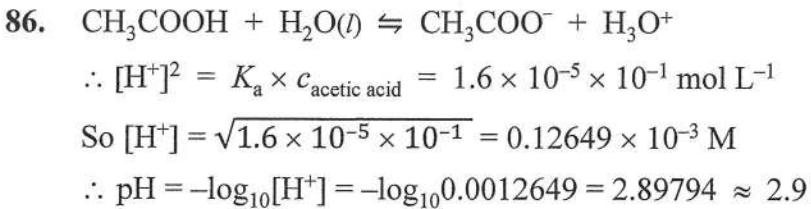
- 83.** (a) $\text{Ba(OH)}_2(aq) + 2\text{HNO}_3(aq) \rightarrow \text{Ba(NO}_3)_2(aq) + 2\text{H}_2\text{O}(l)$
- (b) Average volume of acid used = $(18.1 + 18.2 + 18.1)/3 = 18.13 \text{ mL} = 0.01813 \text{ L}$
 [Note: This is ignoring Titration 1 as it is more than 2 mL above the other three, and so may have been a ‘practice’ or ‘rough’ titration.]
 $n(\text{Ba(OH)}_2) = 0.12 \times 25.0 \times 10^{-3} = 3.0 \times 10^{-3} = 0.003 \text{ mol}$
 $n(\text{HNO}_3) = 2 \times (\text{Ba(OH)}_2) = 2 \times 3.0 \times 10^{-3} = 0.006 \text{ mol}$
 $\therefore [\text{HNO}_3] = \frac{0.006}{0.01813} = 0.3309 = 0.33 \text{ mol L}^{-1}$

- 84.** (a) $\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}0.01 = 2.0$
 [Note: HCl is a strong acid, every HCl molecule ionising to form H^+ . So $[\text{H}^+] = 0.01 \text{ mol L}^{-1}$.]

- (b) Sulfuric acid has a lower pH than HCl, because the $[H^+]$ in H_2SO_4 is higher than in HCl. Since both are strong acids, the first ionisation is complete, and results in identical $[H^+]$. Since H_2SO_4 is diprotic, a second H^+ can break away (although this ionisation is incomplete). This makes the total $[H^+]$ in the sulfuric acid solution higher than in the hydrochloric acid solution which is only monoprotic.
- (c) Acetic acid is a weaker acid than hydrochloric acid. Ionisation is incomplete at equilibrium in acetic acid solution, but complete in hydrochloric acid. Since both acids have the same initial concentration, the $[H^+]$ in acetic acid will be less than in hydrochloric acid. Hence the acetic acid solution has a higher pH.

85. (a) $K_a = \frac{[C_{19}H_{13}O_5S^-][H^+]}{[C_{19}H_{14}O_5S]}$
 Since $[H^+] = [C_{19}H_{13}O_5S^-]$
 $\therefore 1.3 \times 10^{-8} = \frac{[H^+]^2}{0.003}$
 $\therefore [H^+] = 6.2 \times 10^{-6} \text{ mol L}^{-1}$

- (b) Yellow – as HCl is a strong acid, so in a 0.1 mol L^{-1} HCl, $[H^+] = 0.1 \text{ mol L}^{-1}$ and so $\text{pH} = 1$. At this pH, $[H^+]$ is quite high, so the equilibrium will move to the left (Le Châtelier's Principle), causing it to be yellow.



87. (a) The wine.
- (b) Using volume NaOH is 10.1 mL ,
 $\therefore \text{moles NaOH} = 10.1 \times 10^{-3} \times 0.010$
 $= 1.01 \times 10^{-4} \text{ mol}$ [Note: First titration varies by 3% compared to subsequent titrations and so should be ignored as discrepant rather than averaged with the other two results.]
- (c) Potassium tartrate is a monoprotic acid
 So $n(\text{NaOH}) = n(\text{potassium tartrate})$ in 50 mL sample
 $\therefore [\text{potassium tartrate}] = 1.01 \times 10^{-4} \times \frac{1000}{50.0} = 2.02 \times 10^{-3} \text{ mol L}^{-1}$
- (d) $[\text{potassium tartrate}] = 2.02 \times 10^{-3} \text{ mol L}^{-1} \times 188 \text{ g mol}^{-1}$
 $\approx 0.380 \approx 0.38 \text{ g L}^{-1}$

