Year 12
Chemistry
Acid/Base Reactions

Term 2
Work Book Solutions

MATRIX
EDUCATION

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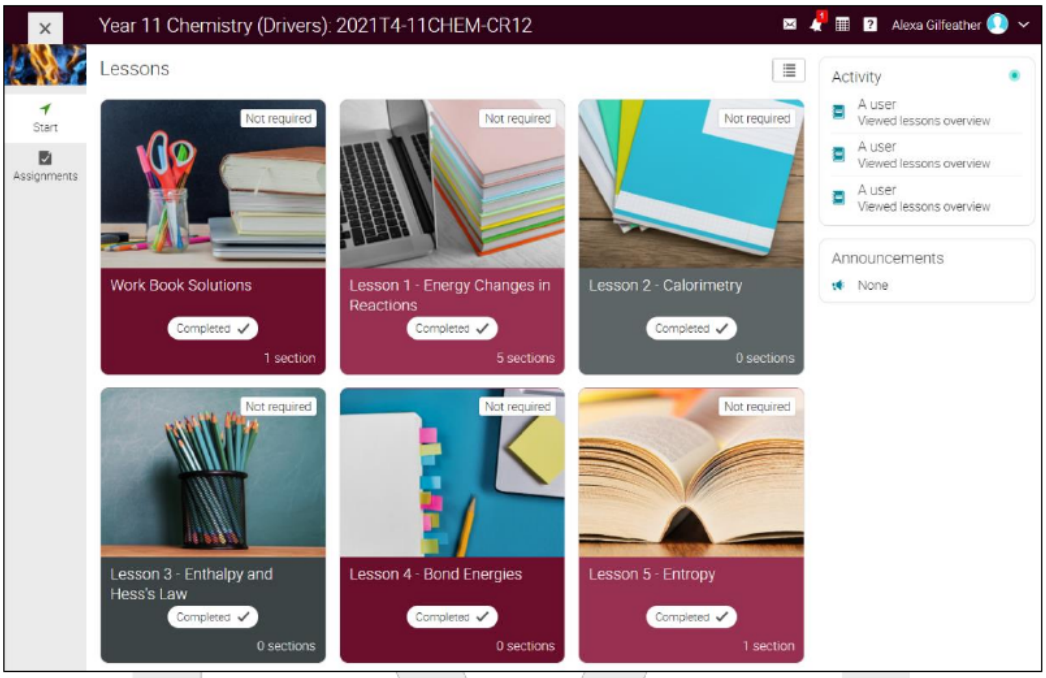

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Acid/base strength and indicators

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Part A: Multiple choice questions

Question 1

Which of the following groups contains ONLY bases?

(a) Na_2O, H_2CO_3, $Ba(OH)_2$, LiOH

(b) Li_2CO_3, NaOH, NH_3, CaO

(c) $NaHCO_3$, CuO, $Mg(OH)_2$, CH_3COOH

(d) KOH, NaCl, Na_2CO_3, MgO

H_2CO_3, CH_3COOH are acids and NaCl is not an acid or base, it is a neutral salt.

Question 2

A student produced a natural indicator from the roots of a green cactus. In order to determine its usefulness as an indicator, the student should test the indicator with:

(a) dilute hydrochloric acid and dilute sodium hydroxide

(b) bromothymol blue, phenolphthalein and methyl orange

(c) a pH probe

(d) a range of household substances

A useful indicator should change colour in acids and bases. Although (d) can theoretically work, it assumes that the range includes both a base and an acid, and that the student knows their acidic/basic nature.

Question 3

Which indicator would best distinguish between rainwater (pH 5.5) and 0.1 mol L^{-1} hydrochloric acid (pH 1)?

(a) Bromothymol blue

(b) Litmus

(c) Phenolphthalein

(d) Methyl orange

Phenolphthalein is colourless for both, bromothymol blue would be yellow for both. Rainwater is in the transition pH range for litmus (pH 4.5-8.3) which could distinguish between the two solutions. However, methyl orange (colour change range at pH 3.1-4.4) gives a much clearer result.

Question 4

Which statement best describes a weak acid solution?

(a) There are no neutral acid molecules present

(b) Only a fraction of the acid molecules is ionised

(c) The total concentration of acid molecules is low

(d) All acid present is ionised to produce hydrogen ions

(a) describes a strong acid, (c) is a dilute solution, (d) describes a strong acid

Question 5

During their chemistry course, students tested a known acid, a known base and water with samples of liquids extracted from various sources. Their results are shown below.

Liquid sample	Colour of liquid	Colour in water	Colour in acid	Colour in base
P	Pink	Pale pink	Pale pink	Orange
Q	Purple	Light purple	Red	Green
R	Green	Green	Purple	Green
S	Yellow	Yellow	Yellow	Yellow

Which liquid sample would be the most effective acid/base indicator to distinguish between acidic, basic and neutral substances?

(a) P

(b) Q

(c) R

(d) S

Acid-base indicators ideally should be different colours in acidic, neutral, and basic solutions.

Sample P is pale pink in both water and an acid.

Sample R is green in water and a base.

Sample S is yellow in everything.

Question 6

Which of the following does not apply to acid-base indicators?

(a) Are usually vegetable pigments

(b) Are used to determine the acidity or basicity of substances

(c) Change colour over their acidity/alkalinity range

(d) Are all weak acids

They can also be weak bases

Question 7

A series of solutions was tested with an electrical conductivity meter. Which of the following solutions would record the highest conductivity reading?

(a) 0.01 mol L^{-1} HCl

(b) 0.1 mol L^{-1} HCl

(c) 0.01 mol L^{-1} CH_3COOH

(d) 0.1 mol L^{-1} CH_3COOH

Conductivity in solution correlates with the concentration of ions in solution. (b) has the largest concentration of ions as it is the most concentrated AND strongest acid.

Question 8

What pH range would a solution be if the solution turns yellow with bromothymol blue, yellow with methyl orange and colourless with phenolphthalein?

(a) 0–14

(b) 3–4.5

(c) 3–7.5

(d) 4.4–6

Yellow with bromothymol blue = pH < 6.0, yellow with methyl orange = pH > 4.4, colourless with phenolphthalein = pH < 8.2. Therefore the narrowest range is pH 4.4–6.0.

Question 9 [2016 HSC Q7]

Which indicator in the table would be best for distinguishing between lemon juice (pH = 2.3) and potato juice (pH = 5.8)?

	Indicator	Colour at different pH	
(a)	Crystal violet	0.2 – yellow	1.8 – blue
(b)	Methyl orange	3.2 – red	4.4 – yellow
(c)	Bromothymol blue	6.0 – yellow	7.6 – blue
(d)	Phenolphthalein	8.2 – colourless	10.0 – pink

This is the only indicator that gives different colours for lemon juice (red) and potato juice (yellow).

Question 10

How does 0.05 mol L^{-1} nitric acid compare to 0.5 mol L^{-1} phosphoric acid?

(a) Phosphoric acid is weaker and more dilute than nitric acid

(b) Phosphoric acid is stronger and more dilute than nitric acid

(c) Phosphoric acid is weaker and more concentrated than nitric acid

(d) Phosphoric acid is stronger and more concentrated than nitric acid

Strength depends on identity of the acid – HNO_3 is strong and H_3PO_4 is weak

Concentration depends on the amount dissolved – 0.05 mol L^{-1} is more dilute than 0.5 mol L^{-1}

Part B: Extended response questions

Question 11

(a) Complete the table by giving the strength (strong or weak) of the following acids and writing the balanced full formula equations for the dissociation of the acids in water to show how each substance produces hydronium ions H_3O^+. 8

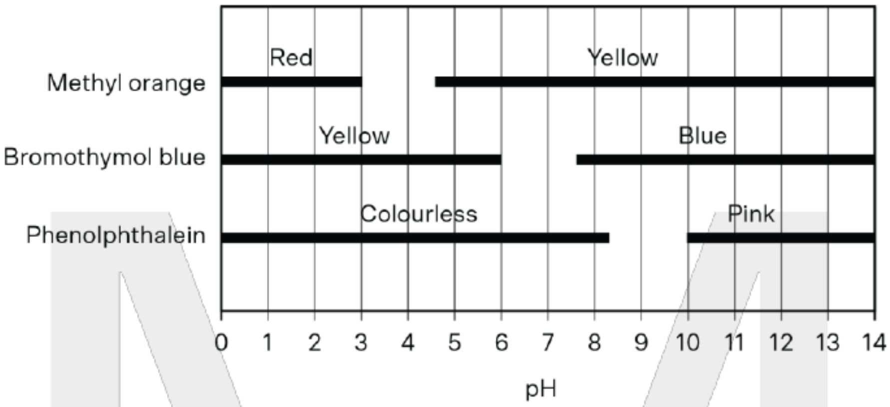
Acid	Strength	Equation
Phosphoric acid	Weak	$H_3PO_{4(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + H_2PO_4^-{}_{(aq)}$
Sulfuric acid	Strong	$H_2SO_{4(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + HSO_4^-{}_{(aq)}$
Sulfurous acid	Weak	$H_2SO_{3(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HSO_3^-{}_{(aq)}$
Hydrofluoric acid	Weak	$HF_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + F^-_{(aq)}$
Nitric acid	Strong	$HNO_{3(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$
Acetic acid	Weak	$CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$
Hydrocyanic acid (HCN)	Weak	$HCN_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CN^-_{(aq)}$
Citric acid ($C_6H_8O_7$)	Weak	$C_6H_8O_{7(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + C_6H_7O_7^-{}_{(aq)}$

(b) Complete the table by giving the strength (strong or weak) of the following bases and writing balanced full formula equations to show how each of the following bases produces hydroxide ions OH^- in water. 8

Base	Strength	Equation
Ammonia	Weak	$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-{}_{(aq)}$
Calcium hydroxide	Strong	$Ca(OH)_{2(s)} \longrightarrow Ca^{2+}{}_{(aq)} + 2OH^-{}_{(aq)}$
Rubidium hydroxide	Strong	$RbOH_{(s)} \longrightarrow Rb^+{}_{(aq)} + OH^-{}_{(aq)}$
Sodium carbonate	Weak	$Na_2CO_{3(s)} \longrightarrow 2Na^+{}_{(aq)} + CO_3{}^{2-}{}_{(aq)}$ $CO_3{}^{2-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HCO_3{}^-{}_{(aq)} + OH^-{}_{(aq)}$
Lithium fluoride	Weak	$LiF_{(s)} \longrightarrow Li^+{}_{(aq)} + F^-{}_{(aq)}$ $F^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HF_{(aq)} + OH^-{}_{(aq)}$
Potassium phosphate	Weak	$K_3PO_{4(s)} \longrightarrow 3K^+{}_{(aq)} + PO_4{}^{3-}{}_{(aq)}$ $PO_4{}^{3-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HPO_4{}^{2-}{}_{(aq)} + OH^-{}_{(aq)}$
Magnesium hydroxide	Strong	$Mg(OH)_{2(s)} \longrightarrow Mg^{2+}{}_{(aq)} + 2OH^-{}_{(aq)}$
Potassium bicarbonate	Weak	$KHCO_{3(s)} \longrightarrow K^+{}_{(aq)} + HCO_3{}^-{}_{(aq)}$ $HCO_3{}^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)} + OH^-{}_{(aq)}$

Question 12

The chart below shows the colour ranges for the indicators methyl orange, bromothymol blue and phenolphthalein.



(a) Identify the possible range of pH values for a solution that will display a yellow colour with both methyl orange and bromothymol blue. 1

4.4–6

(b) To confirm that a solution has a pH range of 3–4, which of the three indicators would be most appropriate? Explain your choice. 3

Methyl orange is the best indicator to use as it changes colour around pH 4 (pH 3.1–4.4).

The other indicators have wider ranges for the colour at pH 4 (bromothymol blue is yellow from pH 0–6, phenolphthalein is colourless for pH 0–8.2).

Question 13 **[2013 HSC Q25]**

An indicator is placed in water. The resulting solution contains the green ion, *Ind*⁻, and the red molecule, *HInd*.

Explain why this solution can be used as an indicator. In your response, include a suitable chemical equation that uses *Ind*⁻ and *HInd*. **4**

Indicators are weak acids (or bases) and partially dissociate into their ions in aqueous solution:

$$HInd_{(s)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + Ind^-_{(aq)}$$

This solution can be used as an indicator because the two sides of the above equilibrium produce two different colours, depending on the pH of the solution it is put into. In high pH (low H_3O^+ concentration), the equilibrium will lie to the right hand side and the solution will be green, but in low pH (high H_3O^+ concentration), the equilibrium will lie to the left hand side and the solution will be red.

Question 14

Compare the properties of acids and bases. **4**

Acids	Bases
Taste sour	Taste bitter
Produce H^+ when dissolved in water	Produce OH^- when dissolved in water
$pH < 7$	$pH > 7$
Corrosive	Corrosive
Turn blue litmus indicator red	Turn red litmus indicator blue
Neutralise bases	Neutralise acids
Conducts electricity	Conducts electricity

Question 15

Acidic solutions can be described as strong or weak and concentrated or dilute.

(a) Explain the difference between strong and weak acidic solutions. Provide an example of each in your answer. 2

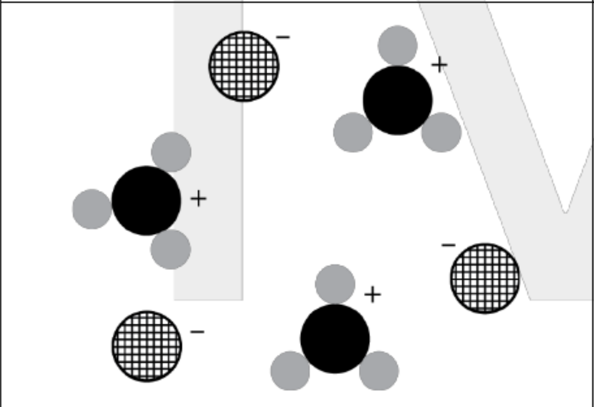
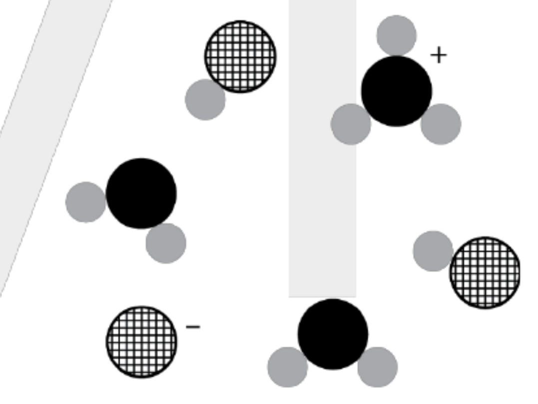
A strong acid completely ionises in water. Nitric acid is an example of a strong acid.

$HNO_{3(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + NO_3^-{}_{(aq)}$

A weak acid will only be partially ionised in water. Acetic acid is an example of a weak acid.

$CH_3COOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + CH_3COO^-{}_{(aq)}$

(b) Illustrate the difference between the ionisation of a strong acid and a weak acid, using at least three acid molecules and three water molecules in each diagram. 2

Strong acid	Weak acid
	

(c) Explain the difference between concentrated and dilute acidic solutions. 2

A concentrated solution has a large amount of solute dissolved per unit volume (e.g., 10 mol L^{-1}), while a dilute solution has a small amount of solute dissolved per unit volume (e.g., 0.1 mol L^{-1}).

Question 16

You have carried out a first-hand investigation to prepare and test a natural indicator with a range of household substances.

(a) Outline the procedure used to prepare the natural indicator. 2

Heat 3 tablespoons of chopped red cabbage in a beaker containing 500 mL water for 30 min.

Strain cabbage through a cloth and collect the filtrate in a 1 L beaker.

Allow to cool before use.

(b) Outline the procedure used to test the natural indicator. 2

Add 5 mL red cabbage indicator into three separate test tubes, labelled A, B and C.

Add 3 mL solution of lemon juice to test tube A, 3 mL solution of cleaning ammonia to test tube B and 3 mL of tap water to test tube C.

Record the colour changes.

(c) Construct a table to show the results obtained from testing this indicator. 2

Substance	Colour of indicator
Lemon juice	Red
Cleaning ammonia	Blue
Tap water	Purple

Question 17

A student performed a first-hand investigation using indicators to determine whether three colourless solutions were acidic or basic. The indicators used are shown in the table.

Indicator	Colour change	pH range
Methyl orange	Red to yellow	3.2–4.4
Methyl red	Red to yellow	4.8–6.0
Thymol blue	Yellow to blue	8.0–9.6
Alizarin	Red to purple	11.0–12.4

Samples of each solution were tested with the indicators. The colours of the resulting solutions are shown in the table.

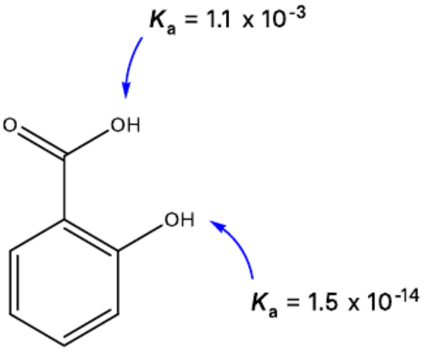
Indicator added	Colour of solution A	Colour of solution B	Colour of solution C
Methyl orange	Yellow	Yellow	Yellow
Methyl red	Yellow	Yellow	Yellow
Thymol blue	Blue	Blue	Yellow
Alizarin	Purple	Red	Red

The student concluded that all three solutions were basic. Assess the validity of this conclusion. 3

The student's assessment of solutions A and B are correct because both turned thymol blue to a blue colour. This indicates a pH greater than 8.0, which is basic. The assessment of solution C may be incorrect. The solution turned thymol blue yellow indicating pH of less than 8.0. The solution also turned methyl red yellow, suggesting solution has a pH between 6 and 8. Thus solution C could be either basic or acidic (or neutral), hence the student's conclusion is not valid.

Question 18

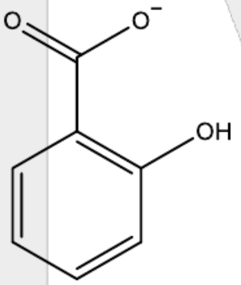
The diprotic acid salicylic acid is found in medications. It can be derivatised to make aspirin. Its structure is shown below, with the K_a values for the dissociation of the two H^+ ions:



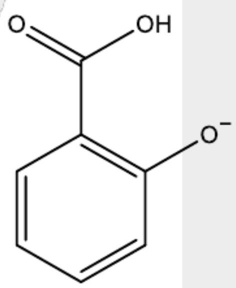
(a) Account for the classification of salicylic acid as a weak acid. **1**

The K_a values are very small. This means equilibrium lies further towards the unionised side (reactant side) and salicylic acid only partially ionises in water.

(b) In an aqueous solution of salicylic acid, which ion will be present in a greater concentration: Ion P or Ion Q? Give reasons for your answer. **2**



Ion P



Ion Q

Ion P will be present in a greater concentration. The K_a value for the dissociation of salicylic acid to H^+ and ion P is larger than the K_a value for the dissociation to give ion Q.

Question 19

Citric acid is a triprotic acid found in foods, including citrus fruits like oranges and lemons.

The K_a value for each dissociation step is given in the table below.

Acid	Formula	K_{a1}	K_{a2}	K_{a3}
Citric	$C_6H_8O_7$	7.4×10^{-4}	1.7×10^{-5}	4.1×10^{-7}

(a) Explain the difference between monoprotic and polyprotic acids. 2

A monoprotic acid is capable of donating one H^+ ion (e.g., HCl), while a polyprotic acid is capable of donating more than one H^+ ion (e.g., H_2SO_4).

(b) Write the equation for each dissociation step and their corresponding acid dissociation constant expressions. 3

$$C_6H_8O_{7(aq)} \rightleftharpoons H^+_{(aq)} + C_6H_7O_7^-{}_{(aq)} \qquad K_{a1} = \frac{[H^+][C_6H_7O_7^-]}{[C_6H_8O_7]}$$

$$C_6H_7O_7^-{}_{(aq)} \rightleftharpoons H^+_{(aq)} + C_6H_6O_7^{2-}{}_{(aq)} \qquad K_{a2} = \frac{[H^+][C_6H_6O_7^{2-}]}{[C_6H_7O_7^-]}$$

$$C_6H_6O_7^{2-}{}_{(aq)} \rightleftharpoons H^+_{(aq)} + C_6H_5O_7^{3-}{}_{(aq)} \qquad K_{a3} = \frac{[H^+][C_6H_5O_7^{3-}]}{[C_6H_6O_7^{2-}]}$$

(c) What can be said about the tendency of subsequent hydrogens to dissociate? 1

Subsequent hydrogens have a lower tendency to dissociate.

(d) Although citric acid is triprotic, the concentration of hydrogen ions in a 0.1 M solution of citric acid is not 0.3 M. Explain why this is the case. 2

Citric acid is a weak acid, which means not all the molecules of citric acid in a solution will ionise to donate all three H^+ ions when dissolved in water.

Question 20

The table below lists some weak bases and their K_b values.

Base	Formula	K_b
Fluoride ion	F^-	2.9×10^{-11}
Ammonia	NH_3	1.8×10^{-5}
Acetate ion	CH_3COO^-	5.6×10^{-10}
Carbonate ion	CO_3^{2-}	1.8×10^{-4}

(a) Order the bases in order of increasing strength. Explain your answer. 3

K_b is the equilibrium constant for the reaction of a base in water to produce hydroxide ions:

$$B_{(aq)} + H_2O_{(l)} \rightleftharpoons BH^+_{(aq)} + OH^-_{(aq)}$$

A larger K_b means the equilibrium lies further towards the right side, producing more hydroxide ions and is considered a stronger base. Therefore, the order is:

$F^- < CH_3COO^- < NH_3 < CO_3^{2-}$

(b) Using an appropriate balanced chemical equation, explain why an aqueous solution containing acetate ions will be basic. 2

Acetate ions will react with water to produce OH^-, which causes the solution to be basic:

$$CH_3COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}$$

(c) Hence write the equilibrium expression for K_b for the acetate ion. 1

$$K_b = \frac{[OH^-][CH_3COOH]}{[CH_3COO^-]}$$

Question 21

Phenolphthalein is an acid-base indicator that is commonly used in laboratories. It is a solution that exists in equilibrium as shown:

$$C_{20}H_{14}O_{4(aq)} \rightleftharpoons 2H^{+}_{(aq)} + C_{20}H_{12}O_4{}^{2-}{}_{(aq)}$$

(a) Phenolphthalein is colourless in acidic solutions and pink in basic solutions. Identify the colour of $C_{20}H_{14}O_4$ molecule and $C_{20}H_{12}O_4^{2-}$ ion. Explain your answer. 2

$C_{20}H_{14}O_4$ is colourless and $C_{20}H_{12}O_4^{2-}$ is pink. In acidic solution there is an excess of H^+ which disturbs the equilibrium. According to Le Chatelier's principle, equilibrium will shift to the left to remove H^+, minimising the disturbance. This produces more $C_{20}H_{14}O_4$. Since in acidic solutions the indicator turns colourless, $C_{20}H_{14}O_4$ must be colourless.

(b) Classify $C_{20}H_{14}O_4$ as a strong acid, weak acid, strong base, or weak base. Explain your answer with reference to the equation above. 2

As shown in the equation above, $C_{20}H_{14}O_4$ ionises reversibly to produce hydrogen ions. Hence it is an acid. Since $C_{20}H_{14}O_4$ only ionises partially, it is a weak acid.

Part C: Additional practice questions

The questions in this part are intended to be used for revision purposes. However, from time to time your teacher may prescribe some of these questions for homework.

Question 22

Equal volumes of methyl orange, bromothymol blue and phenolphthalein indicator were mixed. What colour would be observed at pH 5.5?

(a) Blue

(b) Green

(c) Yellow

(d) Orange

At pH 5.5, methyl orange and bromothymol blue would both be yellow, while phenolphthalein would be colourless. Hence overall, the solution would be yellow.

Question 23

The ionisation of an acid HA can be represented by the general equation:

$$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$$

Which row represents a concentrated solution of a weak acid?

	Concentration of HA (M)	Concentration of H^+ (M)	Concentration of A^- (M)
(a)	1×10^{-6}	3.0	3.0
(b)	0.001	0.5	0.5
(c)	0.7	0.2	0.2
(d)	3.5	0.3	0.3

A solution of over 1 M is regarded as concentrated. A small concentration of H^+ and A^- will be present in a solution of a weak acid.

Question 24

Which of these statements best describes the difference between weak and dilute acid solutions?

(a) A weak acid is not fully ionised and a dilute acid contains a small amount of solute.

(b) A weak acid contains a small amount of solute and a dilute acid is not fully ionised.

(c) A weak acid is fully ionised and a dilute acid is not fully ionised.

(d) A weak acid is fully ionised and a dilute acid contains a small amount of solute.

Question 25

Phenolphthalein is a commonly used indicator. It is colourless in solutions of pH 8 or less and pink in more basic solutions. What colours would a solution of lemon juice with a pH of 2 and a solution of ammonia with a pH of 11 appear if phenolphthalein were added to each?

	Lemon juice (pH = 2)	Ammonia (pH = 11)
(a)	Colourless	Pink
(b)	Colourless	Colourless
(c)	Pink	Pink
(d)	Pink	Colourless

Lemon juice is acidic (pH < 8) hence colourless. Ammonia is basic (pH > 8) hence pink.

Question 26

The pH of a sample was tested using several indicators. The results are shown below:

Indicator	Colour
Bromothymol blue	Yellow
Methyl orange	Yellow
Phenolphthalein	Colourless

What is the most accurate estimate of the pH of the sample?

(a) Between 3.0 and 4.4

(b) Between 4.4 and 6.0

(c) Between 6.0 and 7.0

(d) Between 7.0 and 8.2

Bromothymol blue being yellow indicates pH < 6.0

Methyl orange being yellow indicates pH > 4.5

Phenolphthalein being colourless indicates pH < 8.2

Question 27

Which of the following groups contains only weak bases?

(a) $NaHCO_3$, CaO, $Mg(OH)_2$, $NaOH$

(b) Li_2CO_3, $NaOH$, NH_3, MgO

(c) $NaHCO_3$, $Fe(OH)_2$, $Al(OH)_3$, $NaCH_3COO$

(d) KOH, NH_4Cl, Na_2CO_3, CuO

Strong bases are Group 1 and 2 metal hydroxides. Other bases are weak bases.

Question 28

Which of the following groups contains only polyprotic acids?

(a) H_2SO_3, H_2O, H_3PO_4

(b) H_2SO_4, H_2CO_3, $C_6H_8O_7$

(c) NH_4^+, CH_3COOH, HSO_4^-

(d) $HCOOH$, $H_2PO_4^-$, NH_3

A polyprotic acid is capable of donating multiple hydrogen ions.

Question 29

Citric acid is a weaker acid than sulfuric acid, even though citric acid is triprotic. Which of the following best explains the above statement?

(a) Citric acid ionises more completely than sulfuric acid.

(b) Sulfuric acid reacts completely with a base, citric acid will only react partially with a base.

(c) Sulfuric acid is diprotic and therefore ionises more easily.

(d) Citric acid ionises less completely than sulfuric acid.

Being triprotic means that each citric acid molecule has 3 ionisable hydrogens. However, the number of H^+ ions each acid molecule produces is unrelated to its strength. A strong acid will completely ionise in water whereas a weak acid will only partially ionise in water.

Question 30

List the following acids in order of decreasing strength. Explain your answer. 3

Acid	K_a
Formic	1.77×10^{-4}
Hydrocyanic	6.17×10^{-10}
Benzoic	6.46×10^{-5}

A stronger acid is more ionised in water, i.e. the equilibrium lies further to the right:

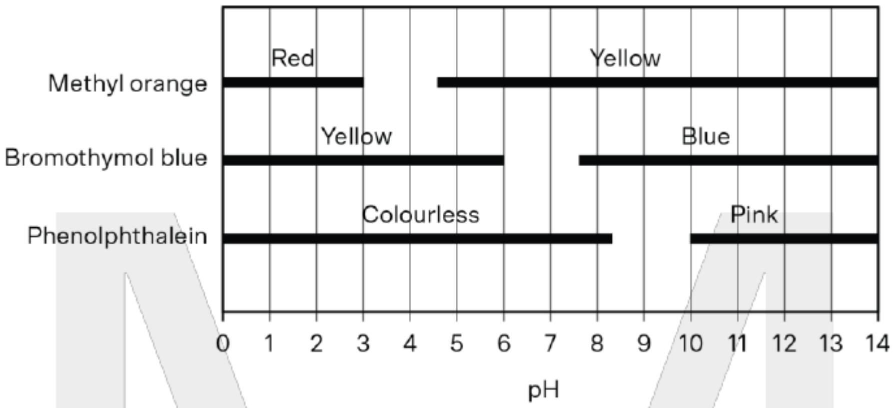
$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$

Hence the equilibrium constant for the dissociation equation K_a will be larger for a stronger acid.

Therefore the order is formic > benzoic > hydrocyanic.

Question 31

The chart below shows the colour ranges for the indicators methyl orange, bromothymol blue and phenolphthalein.



The pH level of the water in swimming pools needs to lie within the range of 7.0–7.6 to ensure the safety of swimmers. Which of the three indicators is the most appropriate for monitoring the pH of swimming pools? Justify your choice. 3

Bromothymol blue is the best choice.

Methyl orange would remain yellow if the pH were to rise a little above 7.6 or fall below 7.0. Phenolphthalein would remain colourless under these conditions. Therefore neither of these indicators allows you to determine if the pH is too high or too low.

On the other hand, bromothymol blue would be green at the correct pH, blue if the pH were too high and yellow-green or yellow if the pH were too low.

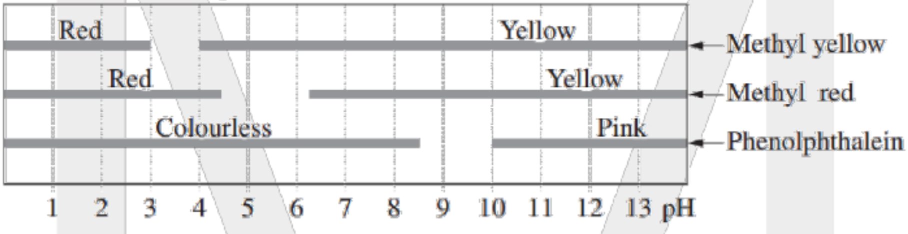
Question 32 [2011 HSC Q32]

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

Barium sulfate is an insoluble white powder that is used as a background to provide contrast. It absorbs the soil moisture and therefore will show the indicator colour.

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

From the data provided, the soil pH is observed to be 4–4.5. The only plants that will grow optimally at this pH will be hydrangeas.

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

Test a small amount of the natural indicator made with several solutions of known pH covering the range from 1–14 and note the colour; for example, various concentrations of HCl and NaOH, acetic acid, Na_2CO_3 solution, ammonia solution and water.

Question 33

The ionisation of hydrosulfuric acid (H_2S) and hydrochloric acid (HCl) are shown below.

$$H_2S_{(aq)} \rightleftharpoons 2H^+_{(aq)} + S^{2-}_{(aq)}$$

$$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

(a) Which of these two acids is the stronger acid? Explain your answer. 2

A strong acid completely ionises in water whereas a weak acid will only be partially ionised in water. Since the ionisation of hydrochloric acid is represented with an irreversible arrow whereas the ionisation of hydrosulfuric acid is shown with a reversible arrow, hydrochloric acid is the stronger acid.

(b) The electrical conductivity of 1.0 mol L^{-1} solutions of each acid was measured. Predict which acid solution would be the best conductor of electricity and justify your answer. 2

Hydrochloric acid is a strong acid which completely ionises into H^+ and Cl^- in water while hydrosulfuric acid is a weak acid, which only partially ionises. Therefore for the same concentration of acid, hydrochloric acid will produce a greater number of ions and hence will have a higher conductivity than hydrosulfuric acid.

Question 34

Sulfurous acid, H_2SO_3, is a weak diprotic acid.

(a) Explain the relationship between the K_a of an acid and its strength. 1

The stronger an acid, the further the equilibrium lies towards the right in the dissociation equation hence the larger the K_a.

(b) Write equations for the dissociation of each hydrogen ion and their corresponding K_a expressions. 2

$H_2SO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HSO_3^-{}_{(aq)}$ $\quad K_{a1} = \frac{[H^+][HSO_3^-]}{[H_2SO_3]}$

$HSO_3^-{}_{(aq)} \rightleftharpoons H^+_{(aq)} + SO_3^{2-}{}_{(aq)}$ $\quad K_{a2} = \frac{[H^+][SO_3^{2-}]}{[HSO_3^-]}$

(c) Explain the relationship between the K_a of an acid and the electrical conductivity of its aqueous solution. 2

The larger the K_a, the further the equilibrium lies towards the right in the dissociation equation hence the higher the concentration of ions in solution. A higher concentration will mean greater conductivity as the ions can carry an electrical current.

Question 35

(a) Define the term acid-base indicator. **1**

An acid-base indicator is a substance which can distinguish between acidic solutions (pH < 7) and basic solutions (pH > 7) by a change in colour.

(b) A natural indicator contained a molecule $C_{16}H_{12}O_6$ that can ionise in water to establish the following equilibrium:

$$C_{16}H_{12}O_{6(aq)} + H^+_{(aq)} \rightleftharpoons C_{16}H_{13}O_6{}^+_{(aq)}$$

Explain what condition must be fulfilled for this solution to be used as an acid-base indicator. **2**

$C_{16}H_{12}O_6$ and $C_{16}H_{13}O_6{}^+$ must be different colours, so that it will display different colours as the pH changes. $C_{16}H_{12}O_6$ will dominate at high pH, and $C_{16}H_{13}O_6{}^+$ will dominate at low pH.

Question 36

Outline a procedure you have performed in school to prepare and test a natural indicator. **4**

Heat 3 tablespoons of chopped red cabbage in a beaker containing 500 mL water for 30 min.

Strain cabbage through a cloth and collect the filtrate in a 1 L beaker. Allow to cool before use.

Add 5 mL of prepared indicator to several solutions of known pH covering the range from 1–14; for example, various concentrations of HCl and NaOH, acetic acid, Na_2CO_3 solution, ammonia solution and water. Record the colour changes.

Question 37

A group of year 12 chemistry students conducted an investigation to determine the relative strengths of hydrochloric acid, hydrosulfuric acid and nitrous acid. The students used 3 indicators to determine the pH range of each solution. The indicators used are given in the table below.

Indicator	Colour at low pH	Colour at high pH	pH of colour change range
Martius yellow	Colourless	Yellow	2.0–3.2
Methyl orange	Red	Yellow	3.1–4.4
Methyl red	Red	Yellow	4.8–6.0

The colours of the resulting solutions are shown in the table.

Indicator	Colour of solution 0.0010 M HCl	Colour of solution 0.0010 M H_2S	Colour of solution 1.00 M HNO_2
Martius yellow	Yellow	Yellow	Colourless
Methyl orange	Red	Yellow	Red
Methyl red	Red	Orange	Red

(a) Determine the pH range of each acid solution. **2**

HCl: pH ranges from 2.0–3.1

H_2S: pH ranges from 4.8–6.0

HNO_2: pH < 2.0

(b) What is the difference between a strong acid and a weak acid? **1**

A strong acid completely ionises in water whereas a weak acid will only be partially ionised in water.

(c) The students concluded that nitrous acid is the strongest acid, followed by hydrochloric acid and then hydrosulfuric acid. Assess the validity of this conclusion drawn from the experimental data. 3

The student's assessment of hydrochloric acid being a stronger acid than hydrosulfuric acid is correct. A stronger acid will be more ionised, causing increased $[H^+]$ with the same acid concentration. Since $pH = -\log_{10}[H^+]$, pH will be lower for higher $[H^+]$.

However, the assessment of nitrous acid being the strongest acid may be incorrect. The pH of an acidic solution is dependent on both the **strength** of the acid and **concentration** of the acidic solution. Since the concentration of the nitrous acid solution is greater than the concentrations of hydrochloric acid and hydrosulfuric acid, the lower pH may not have resulted from greater extent of ionisation. Hence the student's conclusion is not valid.

Question 38 **[Extension]**

Describe an everyday use of indicators and explain why indicators must be used in this situation. 2

Indicators can be used for soil testing. A sample of soil is mixed with universal indicator, and barium sulfate is sprinkled on top so the colour can be seen. The colour is compared to a pH colour chart to determine the pH. This is a convenient way to assess soil pH, which is important as it determines which plants can be grown in the soil, or if the soil pH needs adjustment to improve plant growth.

Year 12
Chemistry
Acid/Base Reactions

Work Book 2
Neutralisation and acid/base theories

MATRIX
EDUCATION

Part A: Multiple choice questions

Question 1

A piece of magnesium metal is added to a solution of sulfuric acid. What are the products of this reaction?

(a) Magnesium sulfate and water

(b) Magnesium sulfate and oxygen

(c) Magnesium sulfate and hydrogen

(d) Magnesium sulfate, carbon dioxide and water

Acid + reactive metal → salt + hydrogen

Question 2

According to the Arrhenius theory of acids, an acid is a substance that:

(a) is able to donate a hydrogen ion

(b) has a pH less than 7

(c) increases the concentration of hydrogen ions in an aqueous solution

(d) reacts with active metals to produce hydrogen gas

Arrhenius' theory only encompasses aqueous solution, so (a) is incorrect.

Question 3

Which of the following solutions would release the greatest amount of heat when mixed together?

	Acid	Base
(a)	100 mL 0.050 M sulfuric acid	100 mL 0.250 M sodium hydroxide
(b)	200 mL of 0.010 M nitric acid	50 mL of 0.150 M potassium hydroxide
(c)	150 mL of 0.100 M hydrobromic acid	150 mL of 0.050 M barium hydroxide
(d)	200 mL of 0.200 M hydrochloric acid	200 mL of 0.025 M lithium hydroxide

The net ionic equation for neutralisation is $H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$

The combination of acid and base producing the largest moles of water will generate the greatest amount of heat.

(a) $n(H_2O) = 0.005$ mol

(b) $n(H_2O) = 0.002$ mol

(c) $n(H_2O) = 0.015$ mol

(d) $n(H_2O) = 0.005$ mol

Question 4

According to Lavoisier's ideas about acids, which of the following compounds would be acidic in water?

(a) NO_2
(b) NH_3
(c) HCl
(d) HBr

Lavoisier said that acids must contain oxygen.

Question 5

Which statement best represents Liebig's definition of an acid?

(a) Acids ionise in solution to form hydrogen ions

(b) Acids contain oxygen

(c) Acids are proton donors

(d) Acids contain replaceable hydrogen

Question 6

Which of the following substances is not amphiprotic?

(a) HCO_3^-

(b) CH_3COO^-

(c) NH_3

(d) HPO_4^{2-}

The acetate ion has no acidic hydrogens that can dissociate.

Question 7

Which of the following are Brønsted-Lowry acid-base reactions?

(i) $CO_{2(g)} + H_2O_{(l)} \rightleftharpoons HCO_3^-{}_{(aq)} + H^+_{(aq)}$

(ii) $H_2SO_{4(l)} + H_2O_{(l)} \longrightarrow HSO_4^-{}_{(aq)} + H_3O^+_{(aq)}$

(iii) $CaO_{(s)} + CO_{2(g)} \longrightarrow CaCO_{3(s)}$

(iv) $NH_{3(g)} + HCl_{(g)} \longrightarrow NH_4Cl_{(s)}$

(a) All of the reactions shown

(b) (i) and (iii) only

(c) (ii) and (iv) only

(d) (iii) and (iv) only

(ii) and (iv) involve proton transfer

(i) produces a proton but does not itself involve proton transfer

(iii) does not have H^+

Question 8

In the following reversible reaction:

$$N_2H_5^+{}_{(aq)} + SCN^-{}_{(aq)} \rightleftharpoons HSCN_{(aq)} + N_2H_{4(aq)}$$

(a) $N_2H_5^+$ acts as an acid

(b) SCN^- acts as an acid

(c) HSCN acts as a base

(d) N_2H_4 acts as an acid

$N_2H_5^+$ donates H^+ in the forward reaction, HSCN donates H^+ in the reverse reaction.

SCN^- accepts H^+ in the forward reaction, N_2H_4 accepts H^+ in the reverse reaction.

Question 9

In an experiment, a stoichiometric amount of 2.00 mol L^{-1} sodium hydroxide solution was added to 100 mL of 2.00 mol L^{-1} nitric acid in a calorimeter. The solution was initially at a temperature of 20.5 °C and rose to a maximum temperature of 29.5 °C. Assuming the heat capacities of both solutions is 4.18×10^3 J kg^{-1} K^{-1} and the densities of both solutions is 1.00 g mL^{-1}, what is the enthalpy of neutralisation for this reaction?

(a) 37.6 kJ mol^{-1}

(b) −37.6 kJ mol^{-1}

(c) 18.8 kJ mol^{-1}

(d) −18.8 kJ mol^{-1}

$q = mc\Delta T = 200 \times 10^{-3} \times 4.18 \times 10^3 \times 9 = 7524$ J

$\Delta H = -q / n = -7.524 / (2 \times 0.1) = -37.62$ kJ mol^{-1}

Question 10

15.0 mL of 0.500 M hydrobromic acid was mixed with 10.0 mL of 0.560 M sodium hydroxide solution. The standard enthalpy of neutralisation is −57 kJ mol^{-1}. If the initial temperature of the solution was 18 °C, what was the final temperature of the solution?

Assume that the heat generated by this reaction only enters the solution and the density of the solution is 1.0 g mL^{-1}.

(a) 25 °C

(b) 21 °C

(c) 15 °C

(d) 18 °C

$n(H^+) = 7.5 \times 10^{-3}$ mol $n(OH^-) = 5.6 \times 10^{-3}$ mol

$n(H_2O) = 5.6 \times 10^{-3}$ mol

$-q = \Delta H \times n = -57 \times 5.6 \times 10^{-3} = -0.3192$ kJ $= -319.2$ J

$q = mc\Delta T = 25 \times 10^{-3} \times 4.18 \times 10^{3} \times \Delta T = 319.2$ J

$\Delta T = 319.2 / (25 \times 10^{-3} \times 4.18 \times 10^{3}) = 3.0545$

$T_f = 3.0545 + 18 = 21.0545$ °C

Part B: Extended response questions

Question 11

Write balanced full formula and net ionic equations for the chemical reactions that occur when the following pairs of substances are combined. Assume all acids and salts are aqueous. **12**

(a) Sulfuric acid + lithium metal

Full formula: $H_2SO_{4(aq)} + 2Li_{(s)} \longrightarrow Li_2SO_{4(aq)} + H_{2(g)}$

Net ionic: $2H^+_{(aq)} + 2Li_{(s)} \longrightarrow 2Li^+_{(aq)} + H_{2(g)}$

(b) Hydrochloric acid + sodium carbonate

Full formula: $2HCl_{(aq)} + Na_2CO_{3(aq)} \longrightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

Net ionic: $2H^+_{(aq)} + CO_3{}^{2-}{}_{(aq)} \longrightarrow H_2O_{(l)} + CO_{2(g)}$

(c) Nitric acid + calcium hydroxide

Full formula: $2HNO_{3(aq)} + Ca(OH)_{2(aq)} \longrightarrow Ca(NO_3)_{2(aq)} + 2H_2O_{(l)}$

Net ionic: $H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$

(d) Acetic acid + calcium oxide

Full formula: $2CH_3COOH_{(aq)} + CaO_{(aq)} \longrightarrow Ca(CH_3COO)_{2(aq)} + H_2O_{(l)}$

Net ionic: $2H^+_{(aq)} + O^{2-}{}_{(aq)} \longrightarrow H_2O_{(l)}$

(e) Phosphoric acid + potassium hydroxide

Full formula: $H_3PO_{4(aq)} + 3KOH_{(aq)} \longrightarrow K_3PO_{4(aq)} + 3H_2O_{(l)}$

Net ionic: $H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$

(f) Sulfuric acid + sodium hydrogen carbonate

Full formula: $H_2SO_{4(aq)} + 2NaHCO_{3(aq)} \longrightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)} + 2CO_{2(g)}$

Net ionic: $H^+_{(aq)} + HCO_3{}^-{}_{(aq)} \longrightarrow H_2O_{(l)} + CO_{2(g)}$

Question 12

A student performed a first-hand investigation to determine the enthalpy of neutralisation. The student reacted 25.0 mL of 0.50 mol L^{-1} hydrochloric acid with an excess of potassium hydroxide solution in a polystyrene cup calorimeter. The final mass of the solution was 34.4 g and a temperature rise of 4.3 °C was measured.

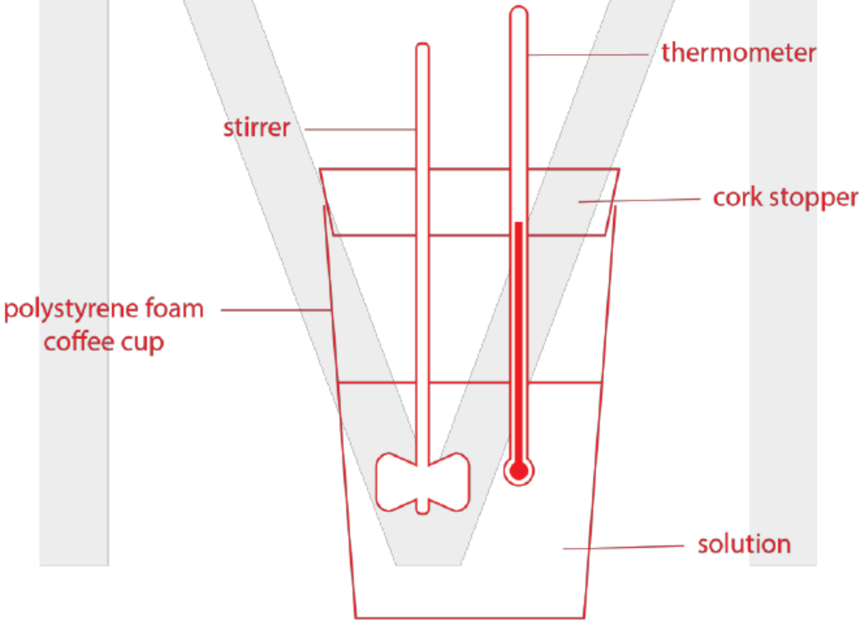
(a) Write a balanced equation for the neutralisation reaction. 1

$HCl_{(aq)} + KOH_{(aq)} \longrightarrow KCl_{(aq)} + H_2O_{(l)}$

(b) What does enthalpy of neutralisation mean? 1

The enthalpy change when an acid and a base react to produce 1 mole of water

(c) Draw a diagram of the apparatus used for this investigation. 2



(d) Use the data above to calculate the measured enthalpy of neutralisation. Assume that the heat capacity of the solution is 4.18×10^3 J kg^{-1} K^{-1}. 2

$q = mc\Delta T = 34.4 \times 10^{-3} \times 4.18 \times 10^3 \times 4.3 = 618.3056$ J

$\Delta H = -q / n = -0.6183056 / (0.5 \times 0.025) = -49.464$ kJ $mol^{-1} = -49$ kJ mol^{-1} (2 s.f.)

(e) Assess the validity of this experiment and suggest a way to improve the validity. 2

This experiment is valid as the enthalpy of neutralisation can be determined thus addressing the aim. However, the validity could be improved if heat loss was further reduced e.g., if the polystyrene cup was more insulated.

(f) The literature value for the enthalpy of neutralisation is −57.62 kJ mol^{-1}. Provide two reasons why the experimentally determined value differs from the literature value. 2

Heat loss to the surroundings, uneven distribution of heat, heat capacity of solution is assumed to be the same as the heat capacity of water.

(g) Suggest how the accuracy of the experiment could be improved. 2

Reduce heat loss to surrounding by insulating polystyrene coffee cup, use a more accurate thermometer, use more accurate scales, use a larger ΔT.

(h) In another calorimetry experiment, the student followed the same procedure but reacted 25.0 mL of 0.50 mol L^{-1} nitric acid with an excess of sodium hydroxide solution instead. The experimental enthalpy of neutralisation was calculated to be −48 kJ mol^{-1}. Using a chemical equation, explain why the enthalpy of neutralisation is approximately the same although the acid and base used was different. 2

The net ionic equation is the same, sodium and potassium cations and the chloride and nitrate anions are spectator ions.

$$H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$$

Therefore, the bonds broken and formed are the same, resulting in the same enthalpy of neutralisation.

(i) If the student repeated the experiment again, but with 25.0 mL of 0.50 mol L^{-1} acetic acid and an excess of sodium hydroxide solution, predict how the experimental enthalpy of neutralisation would differ compared to the previous experiment. Explain your answer. 2

Acetic acid is a weak acid that does not fully ionised in water, hence during neutralisation, energy from the reaction will be absorbed to break the covalent bond in the intact acid to form H$^+$ ions. The enthalpy change will reflect the exothermic reaction between H$^+$ and OH$^-$ and the endothermic bond breaking process. Hence overall ΔH will be less exothermic.

Question 13

One of the advantages of Brønsted-Lowry theory is that it allows many more substances to be defined as acids and bases.

(a) Explain why metal carbonates, such as sodium carbonate, are considered as Brønsted-Lowry bases but not Arrhenius bases. Include relevant balanced chemical equations in your answer. 3

Arrhenius bases contain hydroxide ions (OH^-), therefore metal carbonates are not Arrhenius bases. However, when metal carbonates dissolve, they produce carbonate ions which can accept protons from water, hence acting as a Brønsted-Lowry base:

$$Na_2CO_{3(s)} \longrightarrow 2Na^+_{(aq)} + CO_3{}^{2-}{}_{(aq)}$$

$$CO_3{}^{2-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HCO_3{}^-{}_{(aq)} + OH^-_{(aq)}$$

(b) The Brønsted-Lowry definition is so broad that sodium dihydrogen phosphate (NaH_2PO_4) can be classified as both a Brønsted-Lowry acid and Brønsted-Lowry base.

(i) What is the name given to these types of substances? 1

Amphiprotic substances

(ii) Write TWO ionic equations to show the behaviour of $H_2PO_4^-$ in acidic and basic environments. 2

Basic environment: $H_2PO_4{}^-{}_{(aq)} + OH^-_{(aq)} \longrightarrow HPO_4{}^{2-}{}_{(aq)} + H_2O_{(l)}$

Acidic environment: $H_2PO_4{}^-{}_{(aq)} + H_3O^+_{(aq)} \longrightarrow H_3PO_{4(aq)} + H_2O_{(l)}$

(c) Consider the reaction between ammonia and phosphine gases:

$$NH_{3(g)} + PH_{3(g)} \rightleftharpoons PH_4NH_{2(s)}$$

Explain why this would be considered a Brønsted-Lowry acid-base reaction, but not an acid-base reaction according to Arrhenius. 2

Brønsted-Lowry acid-base reactions require the transfer of a proton. A proton is transferred from NH_3 to PH_3, producing NH_2^- and PH_4^+. However, this is not an Arrhenius acid-base reaction. Firstly, it does not occur in aqueous solution. Secondly, there are no hydroxide ions involved.

(d) The Brønsted–Lowry theory is useful to explain many acid–base reactions but has its limitations. Outline one limitation of the Brønsted-Lowry theory. 2

The Brønsted–Lowry theory cannot explain the acidic behaviour of some substances that do not contain hydrogen and hence should not be able to donate a proton. These include many non-metal oxides, for example CO_2, SO_2 and SO_3, and substances like $AlCl_3$ and BF_3 which produce acidic solutions and can react with bases.

Question 14 **[2011 HSC Q29]**

Justify the continued use of the Arrhenius definition of acids and bases, despite the development of the more sophisticated Brønsted-Lowry definition. 3

Arrhenius defined an acid as a substance that ionises in water to produce hydrogen ions (H^+). He defined a base as a substance that dissociates in water to produce hydroxide ions (OH^-). Alternatively, Brønsted-Lowry defines the acid-base reaction as a transfer of a proton. They defined acids as proton donors and bases as proton acceptors. The Arrhenius definition continues despite the fact that it only applies to aqueous solutions. This is because Arrhenius's theory is encompassed by the Brønsted-Lowry definition, most acid-base reactions occur in aqueous solutions, and the theory demonstrates the formation of salt and water during a neutralisation reaction.

Question 15

A laboratory safety handbook states:

“To minimise risk, large acid spills should be neutralised with limestone ($CaCO_3$) before mopping up.”

Assess the suitability of this recommended method. **4**

Limestone is calcium carbonate, a weak base. It is a good choice for neutralising an acid spill because:

1. It is a weak base, which means it is less caustic if too much is added.

2. It produces carbon dioxide gas on reaction with acid. The lack of effervescence is a sign that neutralisation is complete, which means it is harder to add too much.

It is a better choice than a strong base like sodium hydroxide, which is caustic if too much is added and does not produce gas when it reacts with acid (so there is no indication as to when neutralisation is complete).

However, limestone is not amphiprotic so would not be as ideal as a substance like sodium hydrogen carbonate which is amphiprotic. Amphiprotic substances can prevent pH from rising too rapidly if too much is added. Limestone can be corrosive if too large an excess is added to the spill.

Question 16

A student carried out a calorimetric experiment using the following method:

1. 50 mL of 1.50 M sulfuric acid was added to a polystyrene cup calorimeter.
2. A thermometer was placed in the solution and its initial temperature was recorded.
3. 100 mL of 2.00 M sodium hydroxide solution was added to the calorimeter and the lid was quickly replaced.
4. The mixture was stirred and the maximum temperature reached was recorded.

The initial temperature of the solution was measured to be 18.9 °C.

(a) If the molar enthalpy of neutralisation was determined to be −51 kJ mol^{-1}, calculate the maximum temperature reached by the solution. Assume the density of the solution is 1.00 g mL^{-1} and specific heat capacity of the solution is 4.18×10^3 J kg^{-1} K^{-1}. 3

$H_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

$n(H_2SO_4) = 0.0750$ mol $\quad$ $n(NaOH) = 0.200$ mol

Therefore, H_2SO_4 is limiting and 0.150 mol of water is produced

$q = -\Delta H \times n = -(-51) \times 0.150 = 7.65 \text{ kJ} = 7650 \text{ J}$

$q = mc\Delta T = 150 \times 10^{-3} \times 4.18 \times 10^3 \times \Delta T = 7650 \text{ J}$

$\Delta T = 7650 / (150 \times 10^{-3} \times 4.18 \times 10^3) = 12.201$

$T_f = 12.201 + 18.9 = 31$ °C (2 s.f.)

(b) Explain how the final temperature would differ if solid sodium hydroxide was used instead of aqueous sodium hydroxide solution. 2

Solid sodium hydroxide needs to dissolve before it reacts. This is an exothermic process. Therefore more heat would be released and the final temperature would be higher.

(c) In another calorimetric experiment, the student used the following method:

1. 50 mL of 1.50 M sulfuric acid was added to a polystyrene cup calorimeter containing 100 mL of deionised water.
2. A thermometer was placed in the solution and its initial temperature was recorded.
3. 8.0 g of solid sodium hydroxide was added to the calorimeter and the lid was quickly replaced.
4. The mixture was stirred and the maximum temperature reached was recorded.

The enthalpy of solution for sodium hydroxide is −44.5 kJ mol^{-1}. Determine the final temperature of the solution if the initial temperature of the solution was also 18.9 °C. Assume the density of the sulfuric acid solution is 1.00 g mL^{-1} and specific heat capacity of the solution is 4.18×10^3 J $kg^{-1}K^{-1}$. 5

$NaOH_{(s)} \longrightarrow NaOH_{(aq)}$

$n(NaOH) = 8.0 / (22.99 + 16.00 + 1.008) = 0.20001$ mol

$q_1 = -\Delta H \times n = -(-44.5) \times 0.2001 = 8.900445$ kJ $= 8900.445$ J

$H_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

$n(H_2SO_4) = 0.0750$ mol $n(NaOH) = 0.20001$ mol

Therefore, H_2SO_4 is limiting and $n(H_2O) = 0.150$ mol

$q_2 = -\Delta H \times n = -51 \times 0.150 = 7.65$ kJ $=7650$ J

$q_{total} = q_1 + q_2 = 8900.445 + 7650 = 16550.445$ J

$q_{total} = mc\Delta T = 158 \times 10^{-3} \times 4.18 \times 10^3 \times \Delta T = 16550.445$ J

$T_f = (16550.445 / (158 \times 10^{-3} \times 4.18 \times 10^3)) + 18.9 = 44$ °C

Question 17

Different theories of acids and bases were developed by Lavoisier, Davy, Liebig, Arrhenius and Brønsted-Lowry. Sulfuric acid, H_2SO_4, was classified as an acid by all of these scientists. Explain how each of their theories predict that H_2SO_4 is an acid. Support your answer by using equations where appropriate. 6

Lavoisier claimed that all acids **contain oxygen**. Sulfuric acid (H_2SO_4) contains oxygen, therefore is an acid according to Lavoisier's definition.

Davy claimed that all acids **contain hydrogen**. Sulfuric acid (H_2SO_4) contains hydrogen, therefore is an acid according to Davy's definition.

Liebig stated that active metals would **displace hydrogen** from acids (because all acids contain hydrogen). Magnesium will displace hydrogen from sulfuric acid:

$$H_2SO_{4(aq)} + Mg_{(s)} \longrightarrow MgSO_{4(aq)} + H_{2(g)}$$

Thus sulfuric acid is an acid according to Liebig's definition

Arrhenius stated that acids will **produce hydrogen ions** in solution and neutralise bases.

$$H_2SO_{4(aq)} \longrightarrow H^+_{(aq)} + HSO_4^-{}_{(aq)}$$

$$2NaOH_{(aq)} + H_2SO_{4(aq)} \longrightarrow 2H_2O_{(l)} + Na_2SO_{4(aq)}$$

Thus H_2SO_4 is an acid according to Arrhenius's definition.

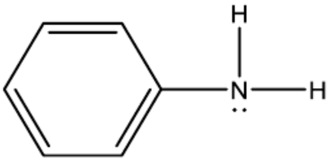
Brønsted-Lowry stated that acids are capable of **donating a proton** and bases are capable of accepting protons.

$$H_2SO_{4(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + HSO_4^-{}_{(aq)}$$

Here, H_2SO_4 is acting as an acid and water is acting as a base.

Question 18

Aniline, $C_6H_5NH_2$, has the structure shown below. It is amphiprotic.



(a) What is an amphiprotic compound? 1

Amphiprotic compounds can donate and accept a proton.

(b) Using equations, show the amphiprotic nature of $C_6H_5NH_2$. 2

Donating a proton: $C_6H_5NH_{2(aq)} + OH^-_{(aq)} \rightleftharpoons C_6H_5NH^-_{(aq)} + H_2O_{(l)}$

Accepting a proton: $C_6H_5NH_{2(aq)} + H_3O^+_{(aq)} \rightleftharpoons C_6H_5NH_3{}^+_{(aq)} + H_2O_{(l)}$

(c) Describe a test that could be carried out in a school laboratory to show that aniline acts as a base in water. 1

Dissolve aniline in water, measure the pH of the solution. If a pH > 7 is measured, it is basic.

(Can also use litmus indicator)

Question 19

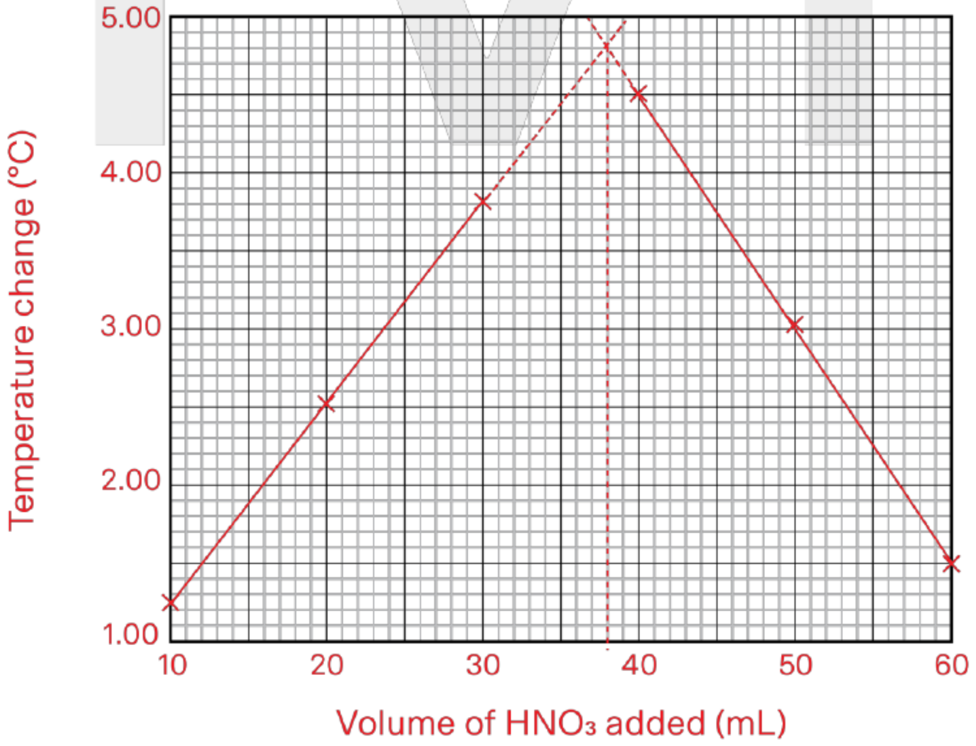
A student carried out an experiment to determine the concentration of a nitric acid solution and the enthalpy of neutralisation when nitric acid is reacted with aqueous sodium hydroxide.

In the experiment, the student added various volumes of nitric acid to 1.00 mol L^{-1} sodium hydroxide solution. The increase in temperature of the mixture was measured after each addition.

The results are given in the table below.

Trial	Volume of nitric acid added (mL)	Volume of sodium hydroxide added (mL)	Temperature change (°C)
1	10.0	60.0	1.25
2	20.0	50.0	2.51
3	30.0	40.0	3.80
4	40.0	30.0	4.50
5	50.0	20.0	3.03
6	60.0	10.0	1.50

(a) Graph the data in the table on the graph below. 3



(b) Determine the concentration of the nitric acid solution. 3

Maximum temperature change occurs when stoichiometric quantities of reactants are mixed. The maximum temperature change occurs when 38.0 mL of HNO_3 is added.

Total volume of solution = 70.0 mL, thus V(NaOH) = 70.0 – 38.0 mL = 32.0 mL

$HNO_{3(aq)} + NaOH_{(aq)} \longrightarrow NaNO_{3(aq)} + H_2O_{(l)}$

$n(NaOH) = c \times V = 1.00 \times 0.032 = 0.032$ mol = $n(HNO_3)$

$c(HNO_3) = n/V = 0.032 / 0.038 = 0.842$ mol L^{-1} (3 s.f.)

(c) Assuming the density of all the solutions are 1.00 g mL^{-1} and specific heat capacity of the resulting solution is the same specific heat capacity as water, calculate the molar enthalpy of neutralisation in kJ mol^{-1} using the data from trial 2. 2

In trial 2, HNO_3 is the limiting reactant, so $n(H_2O) = n(HNO_3)$

$q = mc\Delta T = 70 \times 10^{-3} \times 4.18 \times 10^3 \times 2.51 = 734.426$ J = 0.734426 kJ

$\Delta H = -q / n = -0.734426 / (0.842 \times 0.02) = -43.6$ kJ mol^{-1} (3 s.f.)

(d) The accepted literature value of this enthalpy change is −57.6 kJ mol^{-1}. Calculate the percentage error and provide a reason why the experimentally determined value differs from the literature value. 2

% error = ((57.6 – 43.6) /57.6) × 100 = 24.3% (3.s.f)

Heat loss to the surroundings, uneven distribution of heat, heat capacity of solution is assumed to be the same as the heat capacity of water.

Part C: Additional practice questions

The questions in this part are intended to be used for revision purposes. However, from time to time your teacher may prescribe some of these questions for homework.

Question 20

Outline the definition of an acid that was proposed by each of the following scientists:

(a) Davy **1**

Acids contain hydrogen

(b) Arrhenius **1**

Acids produce H^+ in water

(c) Brønsted and Lowry **1**

Acids are proton donors

(d) Lavoisier **1**

Acids contain oxygen

(e) Liebig **1**

Acids contain replaceable hydrogen

Question 21

Write balanced full formula equations for the following chemical reactions. 5

(a) Sulfuric acid neutralises a sodium hydroxide solution.

$H_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

(b) Acetic acid reacts with sodium carbonate solution and produces a gas that turns limewater milky.

$2CH_3COOH_{(aq)} + Na_2CO_{3(aq)} \longrightarrow 2NaCH_3COO_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

(c) A piece of zinc metal is placed in hydrobromic acid, and bubbles are observed.

$2HBr_{(aq)} + Zn_{(s)} \longrightarrow ZnBr_{2(aq)} + H_{2(g)}$

(d) Nitric acid reacts with magnesium oxide powder.

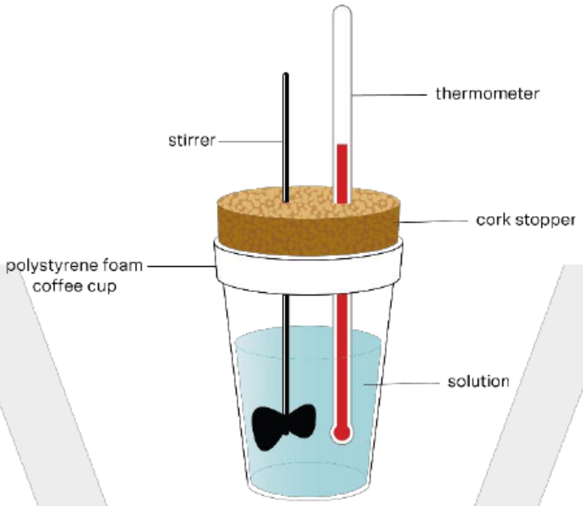
$2HNO_{3(aq)} + MgO_{(s)} \longrightarrow Mg(NO_3)_{2(aq)} + H_2O_{(l)}$

(e) Magnesium hydroxide solution is added slowly to a solution of phosphoric acid to produce solid magnesium phosphate.

$2H_3PO_{4(aq)} + 3Mg(OH)_{2(aq)} \longrightarrow Mg_3(PO_4)_{2(s)} + 6H_2O_{(l)}$

Question 22

The apparatus below was used in an investigation to determine the enthalpy of neutralisation.



The results are shown below:

Volume of 3.0 mol L^{-1} nitric acid = 20.0 mL

Volume of 2.0 mol L^{-1} potassium hydroxide = 35.0 mL

Initial temperature of solution = 17.6 °C

Final temperature of solution = 31.2 °C

(a) What is the name of the apparatus used? **1**

Coffee cup calorimeter

(b) Outline a suitable procedure that could have been used in this investigation. **2**

20 mL of 3.0 M nitric acid was added to the coffee cup calorimeter.

A thermometer was placed in the solution and its initial temperature was recorded.

35 mL of 2.0 M potassium hydroxide solution was added to the calorimeter and the lid was quickly replaced.

The mixture was stirred and the maximum temperature reached was recorded.

(c) Write a balanced chemical equation for this neutralisation reaction. 1

$HNO_{3(aq)} + KOH_{(aq)} \longrightarrow KNO_{3(aq)} + H_2O_{(l)}$

(d) Calculate the enthalpy of neutralisation using the data provided. Assume that the heat capacities of the solutions are 4.18×10^3 J kg^{-1} K^{-1} and the densities of the solutions are 1.00 g mL^{-1}. 2

$q = mc\Delta T = 55 \times 10^{-3} \times 4.18 \times 10^3 \times 13.6 = 3126.64$ J

$n(HNO_3) = 3.0 \times 0.0200 = 0.06$ mol → 0.06 mol H_2O

$n(KOH) = 2.0 \times 0.0350 = 0.07$ mol → 0.07 mol H_2O

Therefore, HNO_3 is limiting, 0.06 mol H_2O produced

$\Delta H = -q / n = -3.12664 / 0.06 = -52.111$ kJ $mol^{-1} = -52$ kJ mol^{-1} (2 s.f.)

(e) Assess the reliability of this experiment. 1

The experiment was only performed once, so the experiment's reliability is unknown.

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

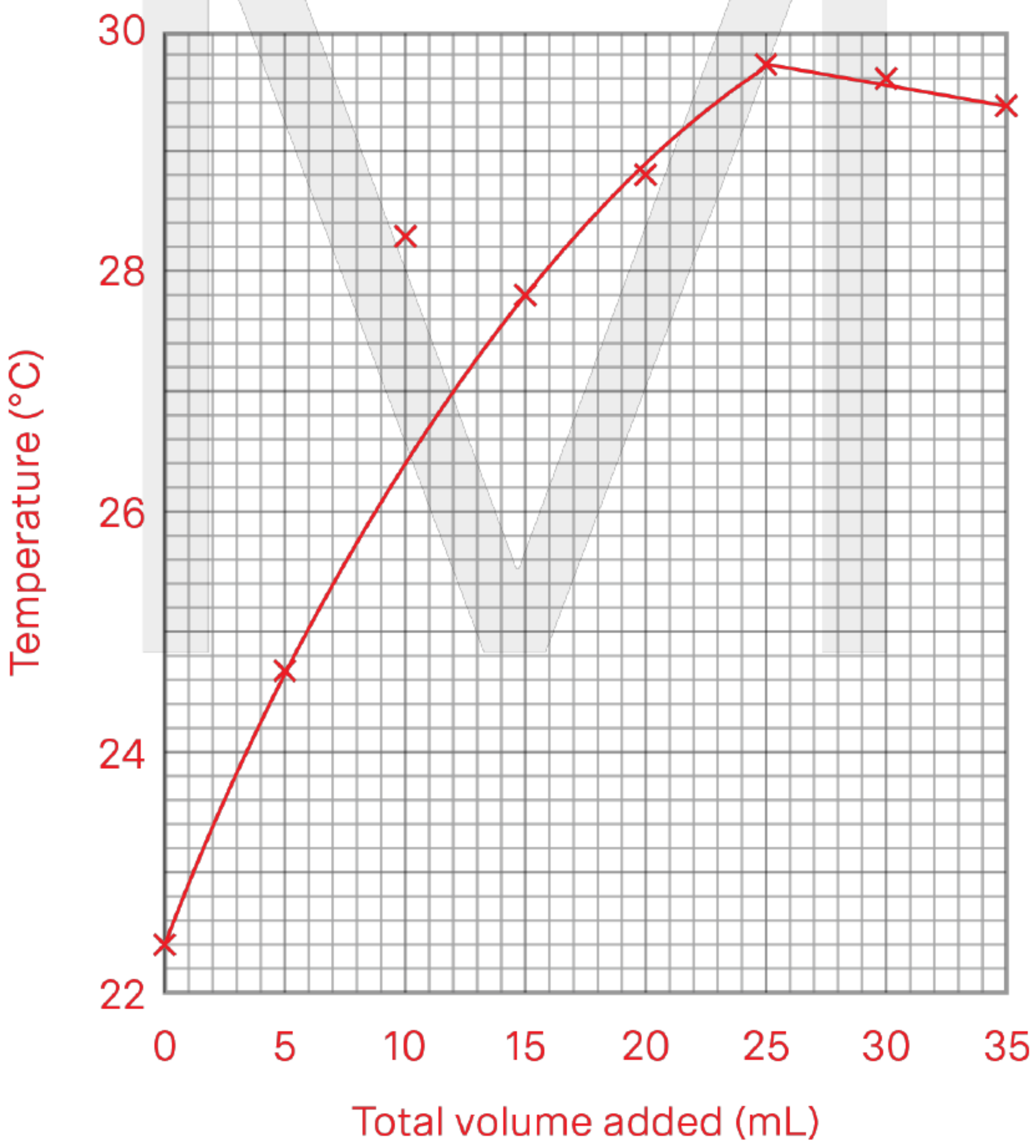
The net ionic equation is always the same, therefore the bonds broken and formed are the same, resulting in the same enthalpy of neutralisation.

Question 23

Tim placed 30.0 mL of 1.10 M hydrochloric acid in a beaker. He added 5.0 mL volumes of 1.32 M sodium hydroxide solution to the beaker and recorded the temperature of the solution, mixing thoroughly after each addition. He recorded the following results:

Total volume added (mL)	0.0	5.0	10.0	15.0	20.0	25.0	30.0	35.0
Temperature (°C)	22.4	24.7	28.3	27.8	28.8	29.7	29.6	29.4

(a) Graph the results on the grid below. 3



(b) Explain why the temperature stops increasing after 25.0 mL of NaOH has been added. **1**

All of the HCl has been neutralised (n(NaOH) = 1.32 x 0.025 = 0.33 = n(HCl)), so the reaction is complete.

(c) Using your graph, determine the temperature change involved in the neutralisation reaction and hence calculate the heat energy produced by the reaction.
Assume the resulting solution has a specific heat capacity of 4.18×10^3 J kg^{-1} K^{-1} and density of all solutions are 1.00 g mL^{-1}. **2**

$\Delta T = 29.7 - 22.4 = 7.3\,°C$

$q = mc\Delta T = 55 \times 10^{-3} \times 4.18 \times 10^3 \times 7.3 = 1678.28$ J = 1.7 kJ (2 s.f.)

(d) Calculate the enthalpy of this reaction in kJ mol^{-1} of HCl. **1**

$\Delta H = -q / n = -1.67828 / (1.32 \times 0.025) = -50.857$ kJ mol^{-1} = -51 kJ mol^{-1} (2 s.f.)

(e) The theoretical enthalpy of this reaction is −57.62 kJ mol^{-1}. Provide a reason why the value calculated from this experiment might be different and suggest a way to improve the validity of this experiment. **2**

Heat loss to the surroundings, improve the insulation of the calorimeter

Question 24

Sodium hydrogen carbonate ($NaHCO_3$) is an amphiprotic compound.

(a) What is an amphiprotic compound? 1

Amphiprotic compounds can donate and accept a proton.

(b) Write TWO ionic equations showing the amphiprotic behaviour of HCO_3^-. 2

Donating a proton: $HCO_3^-{}_{(aq)} + OH^-{}_{(aq)} \longrightarrow CO_3^{2-}{}_{(aq)} + H_2O_{(l)}$

Accepting a proton: $HCO_3^-{}_{(aq)} + H_3O^+{}_{(aq)} \longrightarrow H_2CO_{3(aq)} + H_2O_{(l)}$

(c) Give the name and formula of another amphiprotic compound. 1

Dihydrogen phosphate ion $H_2PO_4^-$

Other answers possible

Question 25

With reference to the reaction of ammonia with water, explain why the Brønsted-Lowry definition of acids and bases is an improvement over earlier ideas. Include a relevant equation in your answer. 3

According to Brønsted-Lowry theory, an acid is a substance which donates protons whereas a base is a substance which accepts protons.

The reaction between ammonia and water is:

$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$

Ammonia has accepted a proton from water to become an ammonium ion, therefore ammonia is acting as a base and water is acting as an acid.

However, according to the earlier Arrhenius theory of acids and bases, a base is a substance which produces OH^- in solution. Since ammonia does not contain OH in its formula, ammonia is not classified as a base, and its basic behaviour in water could not be explained. Similarly, water would not be classified as an Arrhenius acid.

Question 26

Arrhenius put forward an early definition of an acid as any substance which produces hydrogen ions when dissolved in water, and a base as any substance which produces hydroxide ions in water. Brønsted and Lowry extended these definitions in their acid-base theory. They defined an acid as a substance which donates a proton and a base as a substance which accepts a proton in a chemical reaction.

(a) Write an equation to show how the HSO_4^- ion can act as an Arrhenius acid. 1

$HSO_4^-{}_{(aq)} \rightleftharpoons SO_4^{2-}{}_{(aq)} + H^+{}_{(aq)}$

(b) Write an equation to show how the HSO_4^- ion can act as a Brønsted-Lowry base. 1

$HSO_4^-{}_{(aq)} + H_3O^+{}_{(aq)} \rightleftharpoons H_2SO_{4(aq)} + H_2O_{(l)}$

(c) Use an example to explain why the Brønsted-Lowry definition allows many more compounds to be classified as bases. 2

It does not require the base to contain hydroxide ions (OH^-) which was a problem associated with Arrhenius' definition.

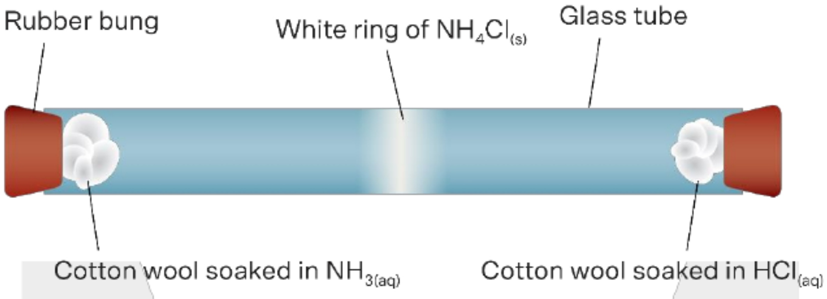
e.g., $NH_{3(aq)} + H^+{}_{(aq)} \rightleftharpoons NH_4^+{}_{(aq)}$

(d) Outline a limitation of the Brønsted-Lowry definition. 2

The Brønsted–Lowry theory cannot explain the acidic behaviour of some substances that do not contain hydrogen and hence should not be able to donate a proton. These include many non-metal oxides, for example CO_2, SO_2 and SO_3, and substances like $AlCl_3$ and BF_3 which produce acidic solutions and can react with bases.

Question 27

A reaction is performed inside a glass tube as shown.



(a) Write an equation to represent the reaction occurring in the set-up in the diagram. **1**

$HCl_{(g)} + NH_{3(g)} \longrightarrow NH_4Cl_{(s)}$

(b) Explain why this would not be an acid-base reaction according to Arrhenius but would be considered a Brønsted-Lowry acid-base reaction. **3**

According to the Arrhenius theory of acids and bases, a base is a substance which produces OH^- in water. Since ammonia does not contain OH in its formula, ammonia is not classified as a base. The reaction also does not occur in aqueous solution. Hence is not an acid-base reaction according to Arrhenius.

Brønsted-Lowry defines the acid-base reaction as a transfer of a proton. This reaction involves a proton transfer: $HCl_{(g)} + NH_{3(g)} \longrightarrow NH_4^+{}_{(s)} + Cl^-{}_{(s)}$

Question 28

In a calorimetric investigation, 25.0 mL of 1.50 mol L^{-1} hydrochloric acid was reacted with 25.0 mL of 1.50 mol L^{-1} sodium hydroxide solution in a calorimeter.

Use the information below to predict the maximum temperature that will be measured.

Initial temperature of solution	23.2 °C
Mass of solution	53 g
Mass of calorimeter	122 g
Specific heat capacity of solution	4.18×10^3 J kg^{-1} K^{-1}
Specific heat capacity of calorimeter	0.61×10^3 J kg^{-1} K^{-1}
Enthalpy of neutralisation	−57 kJ mol^{-1}

Assume that the heat generated by this reaction only enters the solution and the calorimeter and that the calorimeter underwent the same temperature change as the solution. 3

$n(H_2O) = 0.0375$ mol

$q_{total} = -\Delta H \times n = -(-57) \times 0.0375 = 2.1375$ kJ $= 2137.5$ J

$q_{total} = q_{calorimeter} + q_{solution}$

$q_{calorimeter} = 122 \times 10^{-3} \times 0.61 \times 10^3 \times \Delta T = 74.42 \times \Delta T$

$q_{solution} = 53 \times 10^{-3} \times 4.18 \times 10^3 \times \Delta T = 221.54 \times \Delta T$

$2137.5 = (74.42 \times \Delta T) + (221.54 \times \Delta T)$

$\Delta T = 2137.5 / (74.42 + 221.54) = 7.222$ °C

$T_f = 7.222 + 23.2 = 30.422$ °C $= 30$ °C (2 s.f.)

Question 29 **[EXTENSION]**

Explain why aluminium oxide is classified as amphoteric but not amphiprotic. Refer to an example of a chemical species that is both amphoteric and amphiprotic in your answer. 3

Aluminium oxide is amphoteric as it reacts with both acids and bases, but it is not amphiprotic as it is incapable of donating protons (it has no hydrogen in its structure).

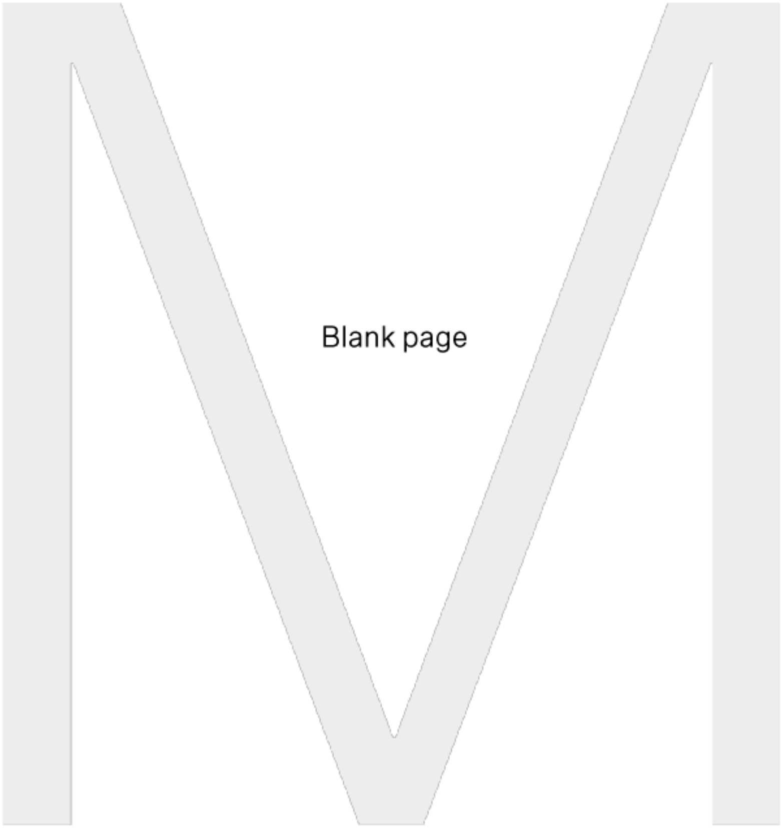
On the other hand, the hydrogen carbonate ion HCO_3^- is both amphiprotic and amphoteric, as it can donate or accept protons (forming H_2CO_3 and CO_3^{2-} respectively), and can react with both acids and bases.

Question 30 **[EXTENSION]**

Explain the difference between amphiprotic and amphoteric. Refer to specific substances in your answer. 3

Amphiprotic describes a substance that can either donate or accept a proton e.g., HCO_3^-, H_2O.

Amphoteric describes a substance that can react with both acids and bases, e.g., Al_2O_3, ZnO. Amphiprotic compounds are a subset of amphoteric compounds.



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Year 12
Chemistry
Acid/Base Reactions

Work Book 3
Applying Brønsted-Lowry theory

MATRIX
EDUCATION

Part A: Multiple choice questions

Question 1

Which salt forms an acidic solution in water?

(a) $MgCl_2$

(b) Na_2CO_3

(c) KCl

(d) NH_4NO_3

NH_4NO_3 will dissociate in water to give NH_4^+ and NO_3^-. NH_4^+ is the conjugate acid of a weak base (NH_3), hence it is a relatively strong acid.

Question 2

A 3.0 mol L^{-1} solution of a monoprotic acid has a pH of 1.2. This solution is:

(a) concentrated and strong

(b) concentrated and weak

(c) dilute and strong

(d) dilute and weak

Concentrated means > 1 M. Strong means completely ionised; if it were strong, it would have a pH of $-\log_{10}(3.0) = -0.48$. Since its pH is higher, it is incompletely ionised (weak).

Question 3

Which of the following does NOT correctly show the formulae of an amphiprotic species with both its conjugate acid and its conjugate base?

	Conjugate acid	Amphiprotic species	Conjugate base
(a)	H_3PO_4	$H_2PO_4^-$	HPO_4^{2-}
(b)	H_2S	HS^-	S^{2-}
(c)	H_2SO_4	HSO_4^-	SO_4^{2-}
(d)	OH^-	H_2O	H_3O^+

An amphiprotic species is capable of both accepting and donating a proton

The conjugate acid must contain 1 more H^+, while the conjugate base must contain 1 less H^+

Question 4

An acetic acid solution has a pH of 2.1. A sulfuric acid solution of the same pH must have:

(a) the same proportion of molecules ionising.

(b) a higher concentration of acid molecules.

(c) a lower concentration of acid molecules.

(d) a higher concentration of hydrogen ions.

Sulfuric acid is a strong acid, so less is required to achieve the same $[H^+]$, which is directly related to pH.

Question 5 [2010 HSC Q18]

Chromate and dichromate ions form an equilibrium according to the following equation.

$$2CrO_4^{2-}{}_{(aq)} + 2H^+{}_{(aq)} \rightleftharpoons Cr_2O_7^{2-}{}_{(aq)} + H_2O_{(l)}$$

Which solution would increase the concentration of the chromate ion (CrO_4^{2-}) when added to the equilibrium mixture?

(a) Sodium nitrate

(b) Sodium chloride

(c) Sodium acetate

(d) Ammonium chloride

Removal of H^+ will shift the equilibrium to the left to increase $[CrO_4^{2-}]$. (c) is basic (due to the acetate ion) and will remove H^+. (a) and (b) are both neutral salts, (d) is acidic (from the NH_4^+).

Question 6

A pH meter was used to measure the pH of several lakes. A severely polluted lake was measured to have a pH of 6.0, while a soda lake was found to have a pH of 9.0. Compared to the polluted lake, the concentration of hydroxide ions in the soda lake is:

(a) 3 times greater

(b) 3 times less

(c) 1000 times greater

(d) 1000 times less

A pH increase of 3 units corresponds to an increase in $[OH^-]$ of 10^3, since $pH = -\log_{10}[H^+]$.

Question 7

What is the pH of a 0.020 mol L^{-1} solution of barium hydroxide?

(a) 1.40

(b) 12.30

(c) 12.60

(d) It has no pH because it is a base

$[OH^-] = 0.02 \times 2 = 0.04$, pOH = 1.3979, pH = 12.60

Question 8

A hydrochloric acid solution and a citric acid solution have the same pH. Which statement is correct regarding the two solutions?

(a) Both solutions contain the same number of intact acid molecules.

(b) Both solutions contain the same number of ions.

(c) The hydrochloric acid solution has a higher concentration of hydrogen ions.

(d) The citric acid has a higher concentration of intact molecules than hydrochloric acid.

(a) is incorrect – since HCl is strong and citric acid is weak, citric acid dissociates to a smaller extent and will have more intact acid molecules.

(b) is incorrect – since citric acid is triprotic, there may be less total ions in the citric acid solution.

(c) is incorrect – same pH means the same $[H^+]$.

Question 9

Which of the following is **NOT** a conjugate acid/base pair?

(a) $HOCl/OCl^-$

(b) H^+/OH^-

(c) HPO_4^{2-}/PO_4^{3-}

(d) OH^-/O^{2-}

A conjugate pair consists of two species that differ by 1 H^+.

Question 10

The K_a values for three weak acids are given in the table below.

Acid	K_a
Hydrofluoric	7.2×10^{-4}
Hydrocyanic	6.2×10^{-10}
Nitrous	4.5×10^{-4}

Which of the following correctly lists the bases in order of increasing strength?

(a) $F^- < NO_2^- < CN^-$

(b) $F^- < CN^- < NO_2^-$

(c) $CN^- < NO_2^- < F^-$

(d) $NO_2^- < F^- < CN^-$

Conjugate bases have inverse strength to their acids. Stronger acids have larger K_a values.

Part B: Extended response questions

Question 11

(a) Complete the following table of conjugate pairs. 10

Acid	Conjugate base
H_2O	OH^-
CH_3COOH	CH_3COO^-
H_2SO_4	HSO_4^-
HCl	Cl^-
$H_2PO_4^-$	HPO_4^{2-}
H_2CO_3	HCO_3^-
HCN	CN^-
HBr	Br^-
$HClO_4$	ClO_4^-
NH_4^+	NH_3

(b) Would the conjugate base of acetic acid or hydrochloric acid be a stronger base? Explain your answer. 2

The conjugate base of acetic acid is a stronger base than the conjugate base of hydrochloric acid.

Hydrochloric acid is a stronger acid than acetic acid. A stronger acid has a stronger tendency to give up H^+, thus its conjugate base has a weaker tendency to accept a proton.

(c) Using relevant equations, explain why sodium cyanide dissolves in water to produce a basic solution. 2

Dissolving: $NaCN_{(s)} \longrightarrow Na^+_{(aq)} + CN^-_{(aq)}$

Reacting with water: $CN^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HCN_{(aq)} + OH^-_{(aq)}$

NaCN dissolves and produces CN^- which is basic as it is the conjugate base of a weak acid. It reacts with water to produce $OH^-_{(aq)}$, increasing the pH and forming a basic solution.

Question 12

In the following equilibrium reaction:

$$HF_{(aq)} + NH_{3(aq)} \rightleftharpoons NH_4{}^+{}_{(aq)} + F^-{}_{(aq)}$$

(a) Which species act as Brønsted-Lowry acids and Brønsted-Lowry bases? 2

Acids: HF, $NH_4{}^+$

Bases: NH_3, F^-

(b) The equilibrium lies to the right. What does this say about the relative strengths of HF and $NH_4{}^+$ as acids? 2

The forward reaction involves HF donating a proton, whereas the reverse reaction involves $NH_4{}^+$ donating a proton. Since the forward reaction is dominant, this means HF is a stronger proton donor and hence a stronger acid.

Question 13

A sample of hydrochloric acid has a pH of 1.50.

(a) Write an equation for the ionisation of hydrochloric acid. 1

$HCl_{(aq)} \longrightarrow H^+{}_{(aq)} + Cl^-{}_{(aq)}$

(b) Calculate $[H^+]$ in this solution. 1

$[H^+] = 10^{-pH} = 0.032$ M (2 s.f.)

(c) Calculate $[OH^-]$ in this solution. 1

$[OH^-] = 10^{-14} / [H^+] = 3.2 \times 10^{-13}$ M (2 s.f.)

(d) Calculate the pOH of this solution. 1

pOH = 14 – pH = 12.50 (2 s.f.)

Question 14

Consider the following acids and their K_a values.

Acid	K_a
HA	1.6×10^{-4}
HB	3.7×10^{-7}
HC	8.0×10^{-3}

(a) List the three acids in order of decreasing strength. Explain your answer. 2

HC > HA > HB

The stronger an acid, the more it will be dissociated, hence it will have a larger K_a.

(b) Calculate the pK_a of each acid. 2

HA: 3.80

HB: 6.43

HC: 2.10

(c) Give the conjugate base for each acid and calculate its K_b value. 3

A^-: 6.3×10^{-11}

B^-: 2.7×10^{-8}

C^-: 1.3×10^{-12}

(d) Which of the following acids will produce the strongest conjugate base in water? 1

HB

Question 15

The following table shows the colour of universal indicator over a range of pH.

pH	1–4	5	6	7	8	9	10–11
Colour	Red	Orange	Yellow	Green	Blue	Purple	Violet

The results of testing four different solutions are shown in the table below.

Solution	Colour
Barium chloride	Green
Ammonium chloride	Orange
Sodium hydroxide	Violet
Potassium acetate	Purple

Write equations for the dissolution of all four compounds and hence explain why dissolving these substances results in acidic, basic or neutral solutions. **4**

$BaCl_{2(s)} \longrightarrow Ba^{2+}_{(aq)} + 2Cl^{-}_{(aq)}$

Both Ba^{2+} and Cl^{-} ions do not react with water to any appreciable extent hence will result in a neutral solution.

$NH_4Cl_{(s)} \longrightarrow NH_4{}^{+}_{(aq)} + Cl^{-}_{(aq)}$ $NH_4{}^{+}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^{+}_{(aq)}$

NH_4Cl dissolves and produces $NH_4{}^{+}$ which is acidic as it is the conjugate acid of a weak base. It reacts with water to produce $H_3O^{+}_{(aq)}$, therefore forms an acidic solution.

$NaOH_{(s)} \longrightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$

NaOH dissolves and produces OH^{-}, therefore forms a basic solution.

$KCH_3COO_{(s)} \longrightarrow K^{+}_{(aq)} + CH_3COO^{-}_{(aq)}$ $CH_3COO^{-}_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + OH^{-}_{(aq)}$

KCH_3COO dissolves and produces CH_3COO^{-} which is basic as it is the conjugate base of a weak acid. It reacts with water to produce $OH^{-}_{(aq)}$, therefore forms a basic solution.

Question 16

Phosphoric acid is a triprotic acid. The K_a value of each dissociation step is given in the table below.

Acid	Formula	K_{a1}	K_{a2}	K_{a3}
Phosphoric	H_3PO_4	7.5×10^{-3}	6.2×10^{-8}	2.2×10^{-13}

(a) Complete the table. **3**

Acid	Conjugate base	K_b	pK_b
H_3PO_4	$H_2PO_4^-$	1.3×10^{-12}	11.88
$H_2PO_4^-$	HPO_4^{2-}	1.6×10^{-7}	6.79
HPO_4^{2-}	PO_4^{3-}	4.5×10^{-2}	1.34

(b) With the aid of appropriate equations, explain why the dihydrogen phosphate ion $H_2PO_4^-$ is amphiprotic, yet an aqueous solution of KH_2PO_4 has a pH lower than 7. **4**

$H_2PO_4^-$ is amphiprotic as it can accept or donate protons:

Accepting: $H_2PO_4^-{}_{(aq)} + H_3O^+{}_{(aq)} \rightleftharpoons H_3PO_{4(aq)} + H_2O_{(l)}$

Donating: $H_2PO_4^-{}_{(aq)} + OH^-{}_{(aq)} \rightleftharpoons HPO_4^{2-}{}_{(aq)} + H_2O_{(l)}$

Since the K_a of $H_2PO_4^-$ (6.2×10^{-8}) is greater than the K_b of $H_2PO_4^-$ (1.3×10^{-12}), the reaction where $H_2PO_4^-$ reacts as an acid in water proceeds to a greater extent than the reaction where $H_2PO_4^-$ reacts as a base in water.

$H_2PO_4^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HPO_4^{2-}{}_{(aq)} + H_3O^+{}_{(aq)}$ $K_a = 6.2 \times 10^{-8}$

$H_2PO_4^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_2PO_4^-{}_{(aq)} + OH^-{}_{(aq)}$ $K_b = 1.3 \times 10^{-12}$

Therefore, in its reaction with water, $H_2PO_4^-$ will preferentially donate a proton. This increases the concentration of H_3O^+ in the aqueous solution. Since $pH = -\log_{10}[H_3O^+]$, pH will decrease to < 7.

Question 17

Hydrobromic acid is a strong monoprotic acid.

(a) A 1.5 L solution of hydrobromic acid is produced by dissolving 2.6 L of hydrogen bromide gas in water at 25 °C and 100 kPa. Calculate the pH of this solution. 3

n(HBr) = 2.6 / 24.79 = 0.104881 mol

c(HBr) = 0.069921 M = $[H^+]$

pH = 1.1554 = 1.16 (2 s.f.)

(b) Calculate the pOH of this solution. 1

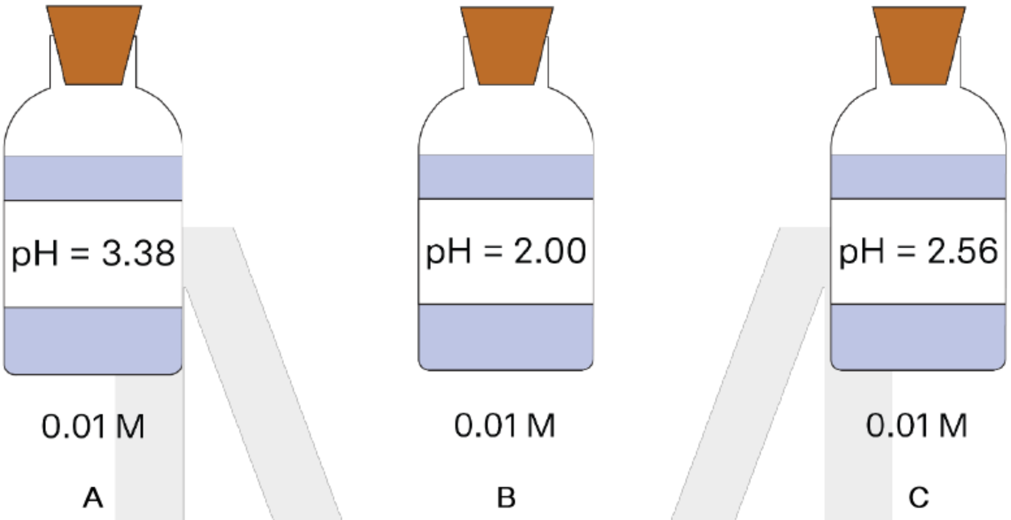
pOH = 14 − pH = 12.84 (2 s.f.)

(c) Predict whether an aqueous solution containing sodium bromide will be acidic, basic or neutral. Explain your answer. 2

Neutral. Sodium ions are neutral. HBr is a strong acid, therefore its conjugate base Br^- will not be a good proton acceptor, hence Br^- will be neutral.

Question 18

The diagram shows three reagent bottles containing acids. The pH was measured to compare their strengths.



(a) List the three acids in order of increasing acid strength. Explain your answer. 2

A < C < B

For the same acid concentration, the stronger an acid, the more dissociated it will be, hence it will have a higher $[H^+]$ and a lower pH since $pH = -\log_{10}[H^+]$.

(b) Calculate the hydrogen ion and hydroxide concentration of each acid solution. 3

Acid	$[H^+]$	$[OH^-]$
A	4.2×10^{-4} M	2.4×10^{-11} M
B	1.0×10^{-2} M	1.0×10^{-12} M
C	2.8×10^{-3} M	3.6×10^{-12} M

(c) Classify each acid as a strong acid or a weak acid. 2

A: weak, B: strong, C: weak

(d) Methyl orange indicator was added to each acid solution. Predict the colour that would be observed in each solution. 2

B and C: red

A: orange

Question 19

A 200 mL solution of calcium hydroxide is prepared by dissolving calcium hydroxide pellets into water at 25 °C until saturation occurs.

(a) Calculate the OH^- concentration in this solution. **3**

$Ca(OH)_{2(s)} \rightleftharpoons Ca^{2+}{}_{(aq)} + 2OH^-{}_{(aq)}$ $K_{sp} = [Ca^{2+}][OH^-]^2 = 5.02 \times 10^{-6}$

Let moles of $Ca(OH)_2$ that dissolves in 1 L = x, thus $[Ca^{2+}] = x$ and $[OH^-] = 2x$

$K_{sp} = (x)(2x)^2 = 5.02 \times 10^{-6}$

$x = 0.0107865$

$[OH^-] = 2x = 0.0215730$ M = 0.0216 M (3 s.f.)

(b) What is the pOH of this solution? **1**

$pOH = -\log_{10}(0.0215730) = 1.666$ (3 s.f.)

(c) What is the pH of this solution? **1**

12.334 (3 s.f.)

(d) Calculate the H^+ concentration in this solution. **1**

4.64×10^{-13} M (3 s.f.)

Question 20

(a) Explain why water is neutral at pH = 7 at 25 °C with reference to the ionisation constant of water. 2

The product of the concentrations of H^+ and OH^- is the ionisation constant of water:

$K_w = [H^+] \times [OH^-] = 10^{-14}$ at 25 °C.

Therefore at 25 °C, $[H^+] = [OH^-] = 10^{-7}$ M and pH = 7. Therefore water is neutral.

(b) The K_w of water at 50 °C is 5.5×10^{-14}. Determine the pH of pure water at 50 °C and thus state how pH varies with temperature. 2

$K_w = [H^+] \times [OH^-] = 5.5 \times 10^{-14}$ at 50 °C.

Therefore $[H^+] = \sqrt{(5.5 \times 10^{-14})} = 2.345 \times 10^{-7}$, $\text{pH} = -\log_{10}(2.345 \times 10^{-7}) = 6.63$.

Thus pH of pure water decreases with increasing temperature.

(c) Justify the classification of water as amphiprotic. Include balanced chemical equations in your answer. 3

Water is amphiprotic as it can both donate and accept a proton.

Donating a proton: $H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$

Accepting a proton: $H_2O_{(l)} + H^+_{(aq)} \rightleftharpoons H_3O^+_{(aq)}$

(d) As the pH of a solution increases by 3, what happens to the following?

(i) concentration of hydrogen ions 1

Decreases 1000-fold

(ii) concentration of hydroxide ions 1

Increases 1000-fold

(iii) pOH 1

Decreases by 3

Question 21 **[2021 HSC Q36]**

The pK_a of sulfurous acid in the following reaction is 1.82.

$$H_2SO_{3(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HSO_3^-{}_{(aq)}$$

The pK_a of hydrogen sulfite in the following reaction is 7.17.

$$HSO_3^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + SO_3^{2-}{}_{(aq)}$$

Calculate the equilibrium constant for the following reaction: 5

$$H_2SO_{3(aq)} + 2H_2O_{(l)} \rightleftharpoons 2H_3O^+_{(aq)} + SO_3^{2-}{}_{(aq)}$$

$$H_2SO_{3(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HSO_3^-{}_{(aq)}$$

$$K_{a1} = \frac{[H_3O^+][HSO_3^-]}{[H_2SO_3]} = 10^{-1.82}$$

$$HSO_3^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + SO_3^{2-}{}_{(aq)}$$

$$K_{a2} = \frac{[H_3O^+][SO_3^{2-}]}{[HSO_3^-]} = 10^{-7.17}$$

$$H_2SO_{3(aq)} + 2H_2O_{(l)} \rightleftharpoons 2H_3O^+_{(aq)} + SO_3^{2-}{}_{(aq)}$$

$$K_{a3} = \frac{[H_3O^+]^2[SO_3^{2-}]}{[H_2SO_3]}$$

$$K_{a1} \times K_{a2} = \frac{[H_3O^+][HSO_3^-]}{[H_2SO_3]} \times \frac{[H_3O^+][SO_3^{2-}]}{[HSO_3^-]} = \frac{[H_3O^+]^2[SO_3^{2-}]}{[H_2SO_3]}$$

Therefore,

$K_{a3} = K_{a1} \times K_{a2} = 10^{-1.82} \times 10^{-7.17} = 1.0 \times 10^{-9}$ (2 s.f.)

Part C: Additional practice questions

The questions in this part are intended to be used for revision purposes. However, from time to time your teacher may prescribe some of these questions for homework.

Question 22

The pH of equimolar acid solutions were measured. Which acid would have the highest pH?

(a) Sulfuric acid

(b) Hydrochloric acid

(c) Methanoic acid

(d) Nitric acid

Methanoic acid is weak, therefore it produces the least H^+.

Question 23

20 g of sodium hydroxide was dissolved in 200 g of distilled water. The solution was then tested with methyl orange, litmus and phenolphthalein. Which row gives the correct results?

	Methyl orange	Litmus	Phenolphthalein
(a)	Yellow	Blue	Pink
(b)	Yellow	Blue	Colourless
(c)	Yellow	Red	Pink
(d)	Red	Red	Pink

n(NaOH) = m / MM = 20 / (22.99 + 16 + 1.008) = 0.5000 mol

c(NaOH) = n / V = 0.5 / 0.2 = 2.5 M

pH = 14 + $\log_{10}(2.5)$ = 14.40, therefore all the colours will be their basic colours.

Question 24

A solution of sodium hydroxide has a pH of 12.35. What is the concentration of hydroxide ions in this solution?

(a) 2.2×10^{-2} M
(b) 2.4×10^{-2} M
(c) 4.5×10^{-13} M
(d) 4.7×10^{-13} M

$[OH^-] = 10^{-(14 - 12.35)}$ M

Question 25

A student uses a pH meter to obtain the pH for solutions of sodium hydroxide and sodium carbonate. The pH reading for each solution is 10.2.

The student correctly concludes that:

(a) both solutions are weak bases
(b) each solution has the same concentration
(c) only the sodium hydroxide contains OH^- ions
(d) both solutions have a $[H_3O^+]$ of $10^{-10.2}$

Sodium hydroxide is a strong base, sodium carbonate is weak. They only have the same pH because more sodium carbonate is required to achieve the same $[OH^-]$.

Question 26

Which compound would form a basic aqueous solution?

(a) Carbon dioxide
(b) Ammonium bromide
(c) Sodium phosphate
(d) Magnesium chloride

Carbon dioxide dissolves to produce carbonic acid.

Ammonium ions are acidic and bromide ions are neutral.

Magnesium and sulfate ions are neutral.

Sodium ions are neutral but phosphate ions are basic.

Question 27

Complete the following table of conjugate pairs. 10

Acid	Conjugate base
HPO_4^{2-}	PO_4^{3-}
NH_3	NH_2^-
H_3O^+	H_2O
HCOOH	$HCOO^-$
H_2S	HS^-
HNO_2	NO_2^-
H_2CO_3	HCO_3^-
HSO_4^-	SO_4^{2-}
$HClO_4$	ClO_4^-
$H_2PO_4^-$	HPO_4^{2-}

Question 28

(a) Write an equation for the ionisation of nitric acid, HNO_3. 1

$HNO_{3(aq)} \longrightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$

(b) What is the pH of a solution of 0.050 mol L^{-1} HNO_3? 1

1.30

(c) Determine the OH^- concentration in a solution of 0.050 mol L^{-1} HNO_3. 1

2.0×10^{-13}

Question 29

(a) For each salt below, classify the nature of solution it would produce in water, and identify the ion causing the pH to change. **8**

Salt	Nature of solution in water (acidic, basic or neutral)	Ion
K_2CO_3	Basic	CO_3^{2-}
$NaSO_3$	Basic	SO_3^{2-}
$MgSO_4$	Neutral	-
NH_4Cl	Acidic	NH_4^+
$NaCH_3COO$	Basic	CH_3COO^-
CaI_2	Neutral	-
$LiNO_2$	Basic	NO_2^-
$AlCl_3$	Acidic	Al^{3+}

(b) For an acidic ion identified above, write an equation to show its acidic nature in aqueous solution. **1**

$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^+{}_{(aq)}$

(c) For a basic ion identified above, write an equation to show its basic nature in aqueous solution. **1**

$CO_3^{2-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HCO_3^-{}_{(aq)} + OH^-{}_{(aq)}$

Question 30 **[2006 HSC Q23(b) modified]**

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

$$OCl^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HOCl_{(aq)} + OH^-_{(aq)}$$

(a) Identify the conjugate acid/base pairs in the above equation. 2

$HOCl$ / OCl^-, H_2O / OH^-

(b) Explain how the addition of sodium hypochlorite will affect the pH of the water in the pool. 2

The addition of OCl^- disturbs the equilibrium. According to Le Chatelier's principle, the equilibrium shifts to the right to remove added OCl^- to minimise the disturbance. Thus the $[OH^-]$ will increase, resulting in a decrease in $[H^+]$. Since $pH = -\log_{10}[H^+]$, the pH rises.

(c) Explain, in terms of Le Chatelier's principle, what effect a reduction in pH will have in the concentration of HOCl at equilibrium. 2

Le Chatelier's principle states that if a system at equilibrium is disturbed then the system responds in such a way as to counteract and minimise the change.

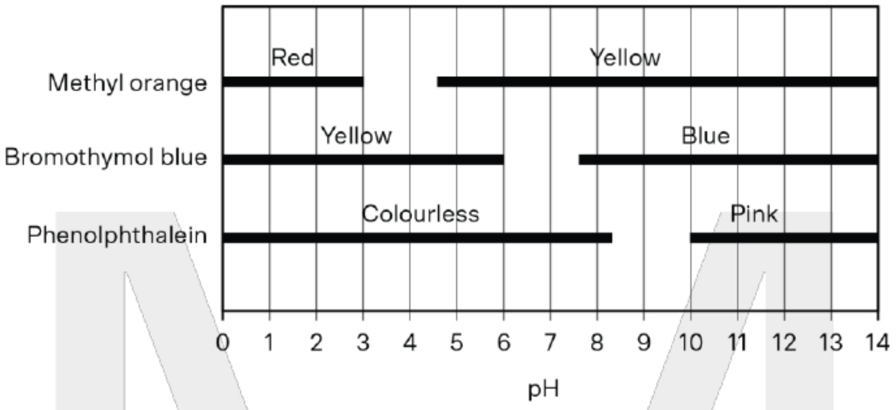
Since $pH = -\log_{10}[H^+]$, reducing the pH of the system means increasing the hydronium ion concentration.

This results in a decrease in concentration of one of the products (OH^-) which results in a shift to the right.

Therefore a reduction in pH increases the concentration of HOCl.

Question 31

The chart below shows the colour ranges for the indicators methyl orange, bromothymol blue and phenolphthalein.



(a) A student tests a solution of HCl using the three indicators. She finds that the solution is orange with methyl orange, yellow with bromothymol blue and colourless with phenolphthalein. What is the most accurate estimate of the HCl concentration in this solution that the student can make? 1

1.0×10^{-3} M to 3.2×10^{-5} M HCl

Explanation: From the methyl orange result, pH is 3.0–4.5, from the bromothymol blue result, pH < 6.0, from the phenolphthalein result, pH < 8.5, therefore the most accurate estimation of pH is 3.0–4.5.

Since $\text{pH} = -\log_{10}[H^+]$, the $[H^+] = 10^{-\text{pH}}$.

Thus, this correspond to 10^{-3} M to $10^{-4.5}$ M HCl = 1.0×10^{-3} M to 3.2×10^{-5} M HCl.

(b) The student wants to determine if a solution is strongly acidic or weakly acidic. Which indicator should she use? Explain. 2

Methyl orange or bromothymol blue are the best choices. Methyl orange will be red when highly acidic and yellow when slightly acidic, while bromothymol blue will be yellow when acidic and green when close to neutral.

Question 32

8.0 L of concentrated hydrochloric acid was spilled in a laboratory accident. The acid in the spill was measured to have a pH of −1.00.

Three substances were considered for use to minimise the damage: solid sodium hydrogen carbonate, powdered limestone (calcium carbonate) and 2 M sodium hydroxide solution.

(a) Calculate the minimum mass of calcium carbonate needed to neutralise the acid. Show numerical working. **3**

$2HCl_{(aq)} + CaCO_{3(s)} \longrightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

$c(HCl) = 10^{-(-1)} = 10$ M

$n(HCl) = c \times V = 8 \times 10 = 80$ mol

$n(CaCO_3) = ½ \times 80$ mol $= 40$ mol (from mole ratio)

$m(CaCO_3) = n \times MM = 40 \times (40.08 + 12.01 + 16 \times 3) = 4003.6$ g $= 4.0$ kg (2 sig. fig.)

(b) Assess the appropriateness of each of the three substances, sodium hydrogen carbonate, powdered limestone and 2 M sodium hydroxide for use in the neutralisation of the spilt acid. **4**

NaOH is the poorest choice to neutralise the spill. It is a strong base, so the neutralisation reaction will be very exothermic. It can also be caustic/dangerous if too much is added, since there is no indication of whether the neutralisation has finished. The fact it is an aqueous solution also means it will increase the size of the spill.

$CaCO_3$ is a better choice to neutralise the spill. Neutralisation with $CaCO_3$ gives off CO_2, and lack of effervescence is an indicator of the end of the neutralisation.

$2HCl_{(aq)} + CaCO_{3(s)} \longrightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

However, if extra $CaCO_3$ is added, the spill may become basic. $NaHCO_3$ is the best choice to neutralise the spill. Not only does it give off CO_2 acting as an indicator for neutralisation similar to $CaCO_3$:

$HCl_{(aq)} + NaHCO_{3(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

It is also amphiprotic. This means that $NaHCO_3$ only acts as a base in an acidic environment. If the environment is too basic, it can instead act as an acid.

Question 33

The autoionisation of water is affected by temperature. The variation in the ionisation constant for water with temperature is given in the table below.

Temperature (°C)	K_w
25	1.0×10^{-14}
40	2.9×10^{-14}
100	51×10^{-14}

(a) Write an equation to show the autoionisation of water and the expression for K_w. **2**

$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$

$K_w = [H^+][OH^-]$

(b) Deduce whether the autoionisation of water is exothermic or endothermic. Explain your reasoning. **2**

Increasing temperature disturbs the equilibrium. According to Le Chatelier's principle, equilibrium shifts in the endothermic direction to remove heat and minimise the disturbance. Since K_w increases with increasing temperature, equilibrium has shifted to the right. Therefore the ionisation of water is endothermic.

(c) State how the pH of pure water varies with temperature. **1**

pH decreases with increasing temperature.

(d) Determine the pH of pure water at 100°C. **2**

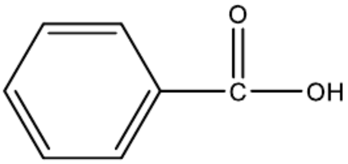
$K_w = [H^+][OH^-] = 51 \times 10^{-14}$ at 100°C

Therefore $[H^+] = \sqrt{(51\times10^{-14})} = 7.1414\times10^{-7}$ M

$pH = -\log_{10}[H^+] = -\log_{10}(7.1414 \times 10^{-7}) = 6.14$ (2 s.f.)

Question 34

Benzoic acid C_6H_5COOH, a food preservative, is a weak monoprotic acid. The structure of benzoic acid is shown below.

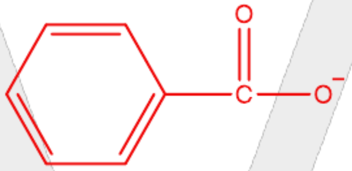


(a) Write the dissociation equation and the corresponding acid dissociation constant (K_a) expression for benzoic acid. 2

$C_6H_5COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + C_6H_5COO^-_{(aq)}$

$$K_a = \frac{[H^+][C_6H_5COO^-]}{[C_6H_5COOH]}$$

(b) Draw the full structural formula of the conjugate base of benzoic acid. 1



(c) Write the dissociation equation and the equilibrium expression for K_b for the conjugate base identified in (b). 2

$C_6H_5COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons C_6H_5COOH_{(aq)} + OH^-_{(aq)}$

$$K_b = \frac{[OH^-][C_6H_5COOH]}{[C_6H_5COO^-]}$$

(d) The K_a of benzoic acid is 6.3×10^{-5}. What is the pK_b of conjugate base of benzoic acid? 1

$K_b = 10^{-14} / K_a = 10^{-14} / 6.3 \times 10^{-5} = 1.5873 \times 10^{-10}$

$pK_b = -\log_{10}(1.5873 \times 10^{-10}) = 9.80$ (2 s.f.)

Question 35

The pK_a values of 3 acids are given in the table below.

Acid	pK_a	K_a	Conjugate base	pK_b	K_b
HCN	9.21	6.2×10^{-10}	CN^-	4.79	1.6×10^{-5}
HCOOH	3.75	1.8×10^{-4}	$HCOO^-$	10.25	5.6×10^{-11}
CH_3COOH	4.74	1.8×10^{-5}	CH_3COO^-	9.26	5.5×10^{-10}

(a) Complete the table above. 4

(b) List the three acids in order of increasing strength. Explain your answer. 2

HCN < CH_3COOH < HCOOH

The stronger an acid, the more dissociated it will be, hence it will have a larger K_a.

Question 36

A 0.010 mol L^{-1} solution of nitric acid and 0.23 mol L^{-1} solution of nitrous acid both have a pH of 2.00.

(a) Explain why both solutions have the same pH. 3

Nitric acid is a strong monoprotic acid and therefore fully ionises in aqueous solution, therefore 0.01 M HCl gives 0.01 M H^+, and since pH = $-\log_{10}[H^+]$, the pH is 2.0. Nitrous acid is a weak acid and does not fully ionise in aqueous solution. Therefore, to reach pH 2.0 or $[H^+] = 0.01$ M, the total concentration of nitrous acid must be higher than that of nitric acid.

(b) Explain why K_a is a better indicator of acid strength than pH. 2

A stronger acid will have a higher H^+ concentration and hence a lower pH, but pH doesn't only depend on acid strength. It also depends on acid concentration, where a higher concentration of acid results in a higher concentration of H^+ ions and thus a lower pH. K_a is independent of concentration of the acid.

Question 37

A 100 mL solution of potassium hydroxide solution is prepared by dissolving 0.55 g of potassium hydroxide pellets into water. Calculate the pH of this solution. 3

n(KOH) = m / MM = 0.55 / (39.1 + 16 + 1.008) = 9.8025×10^{-3} mol

c(KOH) = n / V = 9.8025×10^{-2} M = c(OH^-)

pOH = 1.0087

pH = 12.99 (2 s.f.)

Question 38

Consider the following equilibrium reaction:

$$H_3PO_{4(aq)} + SO_3^{2-}{}_{(aq)} \rightleftharpoons H_2PO_4^-{}_{(aq)} + HSO_3^-{}_{(aq)}$$

(a) Which species act as Brønsted-Lowry acids and Brønsted-Lowry bases? 2

Brønsted-Lowry acids: H_3PO_4, HSO_3^-

Brønsted-Lowry bases: SO_3^{2-}, $H_2PO_4^-$

(b) Using the K_a values given below, predict which side the position of equilibrium lies and explain your prediction. 2

Acid	Formula	K_{a1}	K_{a2}	K_{a3}
Sulfurous	H_2SO_3	1.5×10^{-2}	6.4×10^{-8}	-
Phosphoric	H_3PO_4	7.5×10^{-3}	6.2×10^{-8}	2.2×10^{-13}

The position of equilibrium lies to the right.

K_a is the equilibrium constant for acid dissociation. A larger K_a means the equilibrium lies further towards the ionised side, and the acid is stronger. Thus H_3PO_4 is a stronger acid than HSO_3^-, and therefore it will have a higher tendency to donate a proton.

Year 12
Chemistry
Acid/Base Reactions

Work Book 4
pH calculations

MATRIX
EDUCATION

Part A: Multiple choice questions

Question 1

"Weak bases produce less hydroxide in solution than strong bases."

Ann decides to investigate this hypothesis by measuring the pH of two solutions. Which of the below pairs of solutions would be best to use?

(a) 0.10 mol L^{-1} ammonia and 0.10 mol L^{-1} sodium hydroxide

(b) 1.0 mol L^{-1} sodium carbonate and 1.0 mol L^{-1} ammonia

(c) 0.3 mol L^{-1} sodium carbonate and 0.10 mol L^{-1} sodium hydroxide

(d) 0.10 mol L^{-1} ammonia and 10 mol L^{-1} sodium hydroxide

(b): Both are weak acids.

(c) and (d): All other variables such as concentration should be held the same.

Question 2

The pH of a 0.20 M solution of acetic acid was measured to be close to 2.30. What proportion of acid molecules are ionised?

(a) 1.0%

(b) 2.5%

(c) 15%

(d) 97%

$[H^+] = 10^{-pH} = 5.01 \times 10^{-3}$ M

$((5.01 \times 10^{-3}) / 0.2) \times 100 = 2.5\%$ ionised

Question 3

Which of the following solutions has the highest pH?

	Solution	K_a
(a)	0.1 M Lactic acid	1.4×10^{-4}
(b)	0.5 M Lactic acid	1.4×10^{-4}
(c)	0.1 M Hypochlorous acid	3.0×10^{-8}
(d)	0.5 M Hypochlorous acid	3.0×10^{-8}

Solution with the highest pH will have the smallest [H^+]. Hypochlorous acid is a weaker acid compared to lactic acid since it has a lower K_a, hence will ionise less. A more dilute solution will produce a lower concentration of H^+.

Question 4

In an experiment, a student was given two unknown solids (Solid X and Solid Y). The student dissolved 5 g of each substance in 100 mL of water. The pH of the solutions were then measured at various temperatures. The following data were collected:

Temperature (°C)	pH (Solid X)	pH (Solid Y)
20	2.9	8.1
40	2.9	8.6
60	2.8	9.2
80	2.9	9.9

Which of the following best explains the results?

(a) Both solids X and Y react with water.

(b) Solid X is a weak acid, and solid Y is a strong base.

(c) Solid Y is a weak base, and solid X is a strong acid.

(d) Both solids X and Y are weak acids.

pH of solution Y changes with temperature, hence it exists in an equilibrium (weak base).

Question 5

The pH of a 0.25 M solution of a monoprotic acid is 4.60. Which acid is it most likely to be?

	Solution	pK_a
(a)	Hydrofluoric acid	3.12
(b)	Ethanoic acid	4.77
(c)	Hypobromous acid	8.60
(d)	Hydrocyanic acid	9.20

$K_a = (10^{-4.6})^2 / (0.25 - 10^{-4.6}) = 2.524 \times 10^{-9}$, p$K_a$ = 8.60

Question 6

The K_b of ammonia is 1.8×10^{-5}. What is the pH of a 0.10 mol L^{-1} solution of ammonia?

(a) 7.45×10^{-12}

(b) 1.34×10^{-3}

(c) 2.87

(d) 11.13

$\frac{x^2}{0.1-x} = 1.8 \times 10^{-5}$, $x = 1.34164 \times 10^{-3} = [OH^-]$, pOH = 2.87, pH = 11.13

Question 7

A 0.50 mol L^{-1} solution of hydrazine (NH_2NH_2) has a pOH of 3.19. What is the K_b for this base?

(a) 8.3×10^{-7}

(b) 1.2×10^{-8}

(c) 4.8×10^{-22}

(d) 2.2×10^{-24}

$K_b = (10^{-3.19})^2 / (0.5 - 10^{-3.19}) = 8.3 \times 10^{-7}$

Question 8 **[2021 HSC Q20]**

The trimethylammonium ion, $[(CH_3)_3NH]^+$, is a weak acid. The acid dissociation equation is shown.

$$[(CH_3)_3NH]^+_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + (CH_3)_3N_{(aq)} \qquad K_a = 1.55 \times 10^{-10}$$

At 20 °C, a saturated solution of trimethylammonium chloride, $[(CH_3)_3NH]Cl$, has a pH of 4.46. What is the K_{sp} of trimethylammonium chloride?

(a) 1.26×10^{-9}

(b) 7.76

(c) 60.2

(d) 5.01×10^{10}

pH = 4.46 → $[H_3O^+] = 10^{-4.46} = [(CH_3)_3N]$ (based on stoichiometry)

$[[(CH_3)_3NH]^+] = [H_3O^+][(CH_3)_3N] / K_a = [10^{-4.46}][10^{-4.46}] / 1.55 \times 10^{-10} = 7.7565 \text{ mol L}^{-1}$

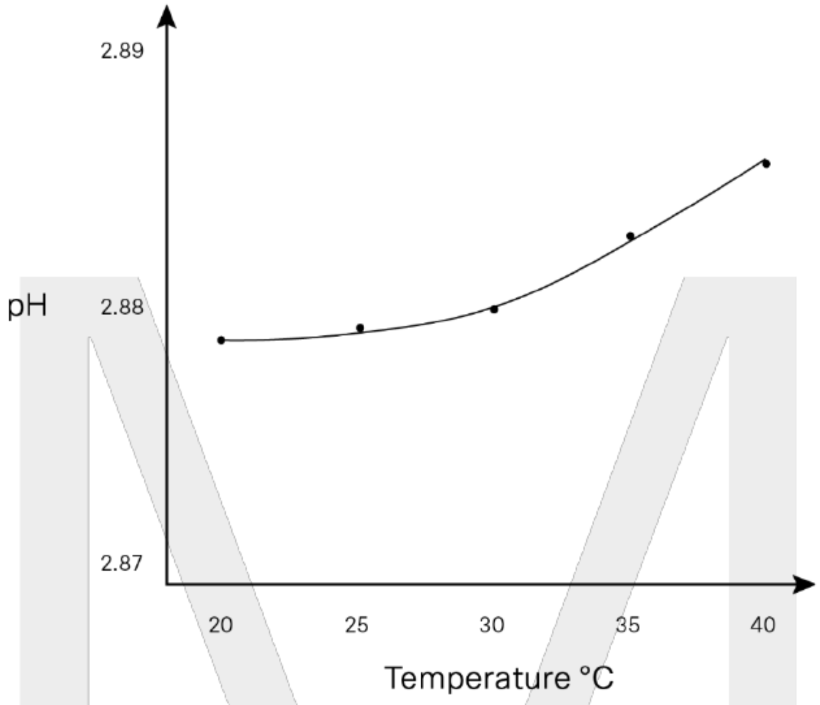
$[(CH_3)_3NH]Cl_{(s)} \leftrightharpoons [(CH_3)_3NH]^+_{(aq)} + Cl^-_{(aq)}$ $K_{sp} = [[(CH_3)_3NH]^+][Cl^-]$

$[Cl^-] = [[(CH_3)_3NH]^+] = 7.7565 \text{ mol L}^{-1}$ (based on stoichiometry)

Therefore, $K_{sp} = (7.7565)(7.7565) = 60.2$

Questions 9 and 10 refer to the following graph.

The change in pH of a 0.10 M ethanoic acid solution is shown below.



Question 9

Is the dissociation of ethanoic acid endothermic or exothermic?

(a) Endothermic

(b) Exothermic

(c) Depends on the temperature

(d) Not enough information

Less H^+ i.e. lower dissociation at higher temperatures. According to Le Chatelier's principle an equilibrium will shift endothermically when heated to minimise disturbance, therefore dissociation is exothermic.

Question 10

As temperature increases, what happens to K_a and percentage ionisation?

	K_a	Percentage ionisation
(a)	Increase	Decrease
(b)	Decrease	Decrease
(c)	Decrease	Increase
(d)	No change	Decrease

Less H^+ i.e. lower dissociation, less products = lower K

Part B: Extended response questions

Question 11

Sodium hypochlorite NaClO dissolves in water to produce a basic solution.

(a) Identify the ion causing the pH to change, and write an equation to show its basic nature in an aqueous solution. 2

Hypochlorite ClO^-

$ClO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HClO_{(aq)} + OH^-_{(aq)}$

(b) Calculate the pH of 0.45 mol L^{-1} sodium hypochlorite solution. The pK_a of hypochlorous acid HClO is 7.52. 4

$pK_b(ClO^-) = 14 - 7.52 = 6.48$

Concentration (M)	ClO^-	HClO	OH^-
Initial	0.45	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.45 - x$	x	x

Assume $[ClO^-]_{eqm}$ = 0.45 M, since K_b is small and dissociation does not change $[ClO^-]_{initial}$ by much.

$\frac{x^2}{0.45} = 10^{-6.48}$

$x = 3.860168 \times 10^{-4} = [OH^-]$

pOH = 3.4133937

pH = 10.5866 = 10.59 (2 s.f.)

Question 12

(a) Write an equation for the dissociation of hydrocyanic acid HCN, and the corresponding acid dissociation constant expression. 2

$HCN_{(aq)} \rightleftharpoons H^+_{(aq)} + CN^-_{(aq)}$ $K_a = \frac{[H^+][CN^-]}{[HCN]}$

(b) 2.0 L of hydrocyanic acid solution was produced by dissolving 4.5 L of hydrogen cyanide gas in water at 25 °C and 100 kPa. The K_a of hydrocyanic acid is 4.9×10^{-10}. Calculate the pH of this solution. 3

$n(HCN) = (100 \times 4.5) / (8.314 \times 298.15) = 0.181538$

$c(HCN) = 0.090769$ M

Concentration (M)	HCN	H^+	CN^-
Initial	0.090769	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.090769 - x$	x	x

Assume $[HCN]_{eqm} = 0.090769$ M, since K_a is small and dissociation does not change $[HCN]_{initial}$ by much.

$$\frac{x^2}{0.090769} = 4.9 \times 10^{-10}$$

$x = 6.66909 \times 10^{-6} = [H^+]$

pH = 5.175933 = 5.18 (2 s.f.)

(c) Write the dissociation equation and the equilibrium expression for K_b for the conjugate base of hydrocyanic acid. 1

$CN^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HCN_{(aq)} + OH^-_{(aq)}$ $K_b = \frac{[HCN][OH^-]}{[CN^-]}$

(d) What is the K_b value for the conjugate base of hydrocyanic acid? 1

$K_b = 10^{-14} / K_a = 2.0 \times 10^{-5}$

Question 13

A student conducted an investigation to determine the difference in strength of two monoprotic acids. The student measured the pH of the two acid solutions using a pH probe. The results are presented in the table below.

	Concentration (mol L^{-1})	pH
Acid 1	0.10	1.59
Acid 2	0.50	1.45

(a) Calculate the K_a of each acid. **4**

$HA_{(aq)} \rightleftharpoons H^+_{(aq)} + A^-_{(aq)}$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Concentration (M)	HA	H^+	A^-
Initial	C	0	0
Change	-10^{-pH}	$+10^{-pH}$	$+10^{-pH}$
Equilibrium	$C - 10^{-pH}$	10^{-pH}	10^{-pH}

$K_a = (10^{-pH})^2 / (C - 10^{-pH})$

For acid 1, $K_a = 8.9 \times 10^{-3}$

For acid 2, $K_a = 2.7 \times 10^{-3}$

(b) Identify which acid is stronger. **1**

Acid 1 is stronger

(c) Explain why K_a is a better indicator of acid strength than pH. **2**

A stronger acid will have a higher H^+ concentration and hence a lower pH, but pH doesn't only depend on acid strength. It also depends on acid concentration, as can be seen in the example where Acid 2 is weaker but gives a lower pH due to the higher concentration used. K_a is independent of concentration.

Question 14

0.658 mL of pure propanoic acid (C_2H_5COOH) was dissolved in water to make a 500 mL solution. The pH of this solution was the same as the pH of 0.677 millimolar HCl. Given that the density of propanoic acid is 0.993 g mL^{-1} and it is monoprotic, determine the K_a of propanoic acid. 3

$m(C_2H_5COOH) = V \times d = 0.658 \times 0.993 = 0.653394$ g

$n(C_2H_5COOH) = m / MM = 8.82035 \times 10^{-3}$ mol

$c(C_2H_5COOH) = n / V = 0.0176407$ M

$pH = -\log_{10}(0.677 / 1000) = 3.1694$

Concentration (M)	C_2H_5COOH	H^+	$C_2H_5COO^-$
Initial	0.0176407	0	0
Change	$-10^{-3.1694}$	$+10^{-3.1694}$	$+10^{-3.1694}$
Equilibrium	$0.0176407 - 10^{-3.1694}$	$10^{-3.1694}$	$10^{-3.1694}$

$K_a = (10^{-3.1694})^2 / (0.0176407 - 10^{-3.1694}) = 2.70197 \times 10^{-5} = 2.70 \times 10^{-5}$ (3 s.f.)

Question 15

Nitrous acid was dissolved in a basic solution. The final solution was measured to have a pH of 4.50 and 96% of nitrous acid was ionised.

(a) If the original concentration of nitrous acid was 0.080 M, what is the concentration of unionised acid? **1**

$0.04 \times 0.08 = 3.2 \times 10^{-3}$ M (2 s.f.)

(b) What is the pK_a of nitrous acid? **3**

Concentration (M)	HNO_2	H^+	NO_2^-
Initial	0.08		0
Change	−0.0768		+0.0768
Equilibrium	3.2×10^{-3}	$10^{-4.5}$	0.0768

$K_a = (0.0768 \times 10^{-4.5}) / (3.2 \times 10^{-3}) = 7.58947 \times 10^{-4}$

pK_a = 3.1198 = 3.12 (2 s.f.)

Question 16

(a) Outline how you could use a pH probe to determine the K_a of benzoic acid, if you have access to pure benzoic acid powder (formula C_6H_5COOH, solubility 3.44 g per L). **4**

Maximum concentration = 3.44 / 122.118 = 0.0282 M

For 250 mL of 0.02 M, n = 0.02 × 0.25 = 5×10^{-3} mol

m = 0.61059 g

1. Weigh out approximately 0.61 g of benzoic acid (record exact mass)

2. Dissolve in distilled water and transfer to a clean 250 mL volumetric flask

3. Add water until the bottom of the meniscus is resting on the mark

4. Stopper and invert to mix

5. Calibrate a pH probe

6. Measure the pH of the benzoic acid solution

7. Calculate the K_a using the calculated concentration of the solution and the measured pH

(b) State two advantages and two disadvantages of the use of indicators over the use of pH probes for determining pH. 2

Advantages: cheaper, easier to use

Disadvantages: less precise (pH range only), destructive

Question 17 **[2024 HSC Q39]**

Water and octan-1-ol do not mix. When an aqueous solution of bromoacetic acid ($BrCH_2COOH$) is shaken with octan-1-ol, an equilibrium system is established between bromoacetic acid dissolved in the octan-1-ol and in the water.

$$BrCH_2COOH_{(aq)} \rightleftharpoons BrCH_2COOH_{(octan\text{-}1\text{-}ol)}$$

The equilibrium constant expression for this system is

$$K_{eq} = \frac{[BrCH_2COOH_{(octan-1-ol)}]}{[BrCH_2COOH_{(aq)}]}$$

An aqueous solution of bromoacetic acid with an initial concentration of 0.1000 mol L^{-1} is shaken with an equal volume of octan-1-ol. Bromoacetic acid does not dissociate in octan-1-ol but does dissociate in water, with $K_a = 1.29 \times 10^{-3}$. When the system has reached equilibrium, the $[H^+]$ is 9.18×10^{-3} mol L^{-1}.

Calculate the equilibrium concentration of aqueous bromoacetic acid and hence, or otherwise, calculate the K_{eq} for the octan-1-ol and water system. **4**

For the ionisation reaction: $BrCH_2COOH_{(aq)} \rightleftharpoons BrCH_2COO^-_{(aq)} + H^+_{(aq)}$ $K_a = \frac{[H^+][BrCH_2COO^-]}{[BrCH_2COOH]}$

$[BrCH_2COO^-_{(aq)}]$ at equilibrium = $[H^+_{(aq)}]$ at equilibrium = 9.18×10^{-3} mol L^{-1}

$K_a = 1.29 \times 10^{-3} = \frac{(9.18 \times 10^{-3})(9.18 \times 10^{-3})}{[BrCH_2COOH]}$

Therefore $[BrCH_2COOH_{(aq)}]$ at equilibrium = $\frac{(9.18 \times 10^{-3})(9.18 \times 10^{-3})}{1.29 \times 10^{-3}}$ = 0.06532744186 mol L^{-1}

Concentration (M)	$BrCH_2COOH_{(aq)}$	$BrCH_2COO^-_{(aq)}$	$H^+_{(aq)}$
Initial	x	0	0
Change	-9.18×10^{-3}	$+9.18 \times 10^{-3}$	$+9.18 \times 10^{-3}$
Equilibrium	0.06532744186	9.18×10^{-3}	9.18×10^{-3}

$x = 0.06532744186 + 9.18 \times 10^{-3} = 0.07450744186$ mol L^{-1}

Suppose we begin with 1 L of 0.1000 mol L^{-1} $BrCH_2COOH$:

$n(BrCH_2COOH)_{total} = 0.1000$ mol

$n(BrCH_2COOH_{(aq)}) = 0.07450744186$ mol

$n(BrCH_2COOH_{(octan\text{-}1\text{-}ol)}) = 0.1000 - 0.07450744186 = 0.02549255814$ mol

Since aqueous solution and octan-1-ol are of equal volume:

$[BrCH_2COOH_{(octan\text{-}1\text{-}ol)}] = 0.02549255814$ mol L^{-1}

$K_{eq} = \frac{0.0254925581}{0.06532744186} = 0.390$ (3sf)

Question 18

The concentration of iron(III) ions in the oceans affect the growth of many aquatic life forms. The availability of iron depends on the solubility of iron(III) hydroxide, which has a $K_{sp} = 2.79 \times 10^{-39}$ at 25 °C.

(a) If the pH of the ocean is 8.179, determine the concentration of iron(III) ions in the ocean. 3

$Fe(OH)_{3(s)} \rightleftharpoons Fe^{3+}_{(aq)} + 3OH^{-}_{(aq)}$ $K_{sp} = [Fe^{3+}][OH^{-}]^3 = 2.79 \times 10^{-39}$

pOH = 14 − 8.179 = 5.821

$[OH^{-}] = 10^{-5.821} = 1.51 \times 10^{-6}$ M

$K_{sp} = [Fe^{3+}] \times (1.51 \times 10^{-6})^3 = 2.79 \times 10^{-39}$

$[Fe^{3+}] = 8.10 \times 10^{-22}$ M (3 s.f.)

(b) An increase in the concentration of carbon dioxide in the atmosphere results in an acidification of the oceans due to the following equilibrium.

$$CO_{2(g)} + H_2O_{(l)} \rightleftharpoons HCO_3^{-}{}_{(aq)} + H_3O^{+}{}_{(aq)}$$

Explain, with reference to Le Chatelier's principle, how the concentration of iron(III) ions will be affected by an increase in the concentration of atmospheric carbon dioxide. 3

Acidification of the oceans due to an increase in the concentration of atmospheric carbon dioxide will result in a decrease in the concentration of hydroxide ions, disturbing the equilibrium. According to Le Chatelier's principle, the equilibrium will shift right to produce more hydroxide ions to minimise the disturbance. Hence, more iron(III) ions will be produced, thus the concentration of iron(III) ions increases.

Question 19

Hydrogen cyanide is a toxic gas that can dissolve in water and form an acidic solution.

$$HCN_{(g)} \rightleftharpoons HCN_{(aq)}$$

$$HCN_{(aq)} + H_2O_{(l)} \rightleftharpoons CN^-_{(aq)} + H_3O^+_{(aq)}$$

(a) The pK_a of hydrocyanic acid is 9.20. If a hydrocyanic acid solution has a pH of 6.50, what is the concentration of hydrocyanic acid? 3

$HCN_{(aq)} \rightleftharpoons H^+_{(aq)} + CN^-_{(aq)}$ $K_a = \frac{[H^+][CN^-]}{[HCN]}$

Concentration (M)	HCN	H^+	CN^-
Initial	C	0	0
Change	$-10^{-6.5}$	$+10^{-6.5}$	$+10^{-6.5}$
Equilibrium	$C - 10^{-6.5}$	$10^{-6.5}$	$10^{-6.5}$

$K_a = (10^{-6.5})^2 / (C - 10^{-6.5}) = 10^{-9.2}$

$c(HCN) = 1.6 \times 10^{-4}$ M

(b) Calculate the percentage ionisation of the HCN solution at this concentration. 1

$\% = 10^{-6.5} / (1.6 \times 10^{-4}) \times 100 = 0.2\%$

(c) Solutions of HCN are used in the gold and silver mining industry. Explain why HCN solutions are kept at pH > 7. 2

At lower pH, there would be excess H_3O^+, disturbing the equilibrium. According to Le Chatelier's principle, equilibrium would shift to the left to remove H_3O^+ and minimise the disturbance. This would increase the amount of HCN, which could evaporate as toxic HCN gas.

Question 20 [2023 HSC Q35]

(a) A 0.2000 mol L^{-1} solution of dichloroacetic acid ($CHCl_2COOH$) has a pH of 1.107.

Dichloroacetic acid is monoprotic.

Calculate the K_a for dichloroacetic acid. 3

NESA sample answer

$[H^+] = 10^{-pH} = 10^{-1.107} = 0.0782$ mol L^{-1}

Concentration (M)	**$CHCl_2COOH_{(aq)}$**	**H^+**	**$CHCl_2COO^-_{(aq)}$**
Initial	0.2000	0	0
Change	−0.0782	+0.0782	+0.0782
Equilibrium	0.1218	0.0782	0.0782

$K_a = [H^+][CHCl_2COO^-] / [CHCl_2COOH] = (0.0782)(0.0782) / (0.1218) = 0.0501$

(b) The following data apply to the ionisation of acetic acid (CH_3COOH) and trichloroacetic acid (CCl_3COOH).

	CH_3COOH	CCl_3COOH
pK_a	4.76	0.51
$\Delta H°$ (kJ mol^{-1})	−0.1	+1.2
$\Delta S°$ (J K^{-1} mol^{-1})	−91.6	−5.8
$-T\Delta S°$ (kJ mol^{-1})	+27.3	+1.7
$\Delta G°$ (kJ mol^{-1})	+27.2	+2.9

Explain the relative strength of these acids with reference to the data. 3

The pK_a of trichloroacetic acid is lower than the pK_a of acetic acid, so trichloroacetic acid is a stronger acid than acetic acid. The major difference between the data for the two acids is the magnitude of the $\Delta S°$ terms. Both are negative and will make an unfavourable contribution to the $\Delta G°$; however, the value for acetic acid is much larger than that for trichloroacetic acid. This means that ionisation of acetic acid is less favourable than it is for trichloroacetic acid, making the latter the stronger acid.

Answers could include: Reference to $-T\Delta S°$ or $T\Delta S°$.

Part C: Additional practice questions

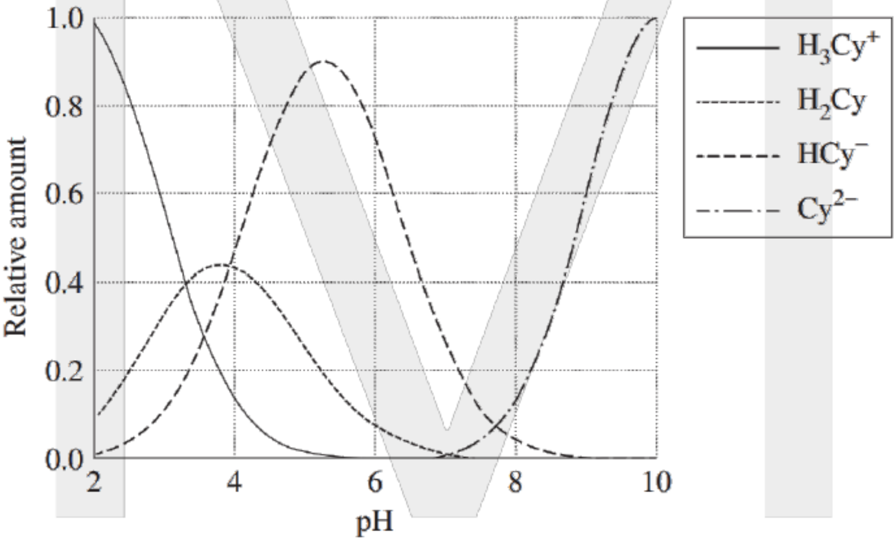
The questions in this part are intended to be used for revision purposes. However, from time to time your teacher may prescribe some of these questions for homework.

Question 21 **[2022 HSC Q20]**

Cyanidin is a plant pigment that may be used as a pH indicator. It has four levels of protonation, each with a different colour, represented by these equilibria:

$$\underset{\text{Red}}{H_3Cy^+} \underset{+H^+}{\overset{-H^+}{\rightleftharpoons}} \underset{\text{Colourless}}{H_2Cy} \underset{+H^+}{\overset{-H^+}{\rightleftharpoons}} \underset{\text{Purple}}{HCy^-} \underset{+H^+}{\overset{-H^+}{\rightleftharpoons}} \underset{\text{Blue}}{Cy^{2-}}$$

The following graph shows the relative amount of each species present at different pH values.



What colour would the indicator be if added to a 0.75 mol L^{-1} solution of hypoiodous acid, HIO (pK_a = 10.64)?

(a) Red

(b) Colourless

(c) Purple

(d) Blue

$HIO_{(aq)} \leftrightharpoons H^+_{(aq)} + IO^-_{(aq)}$

pK_a = 10.64 ⟶ $K_a = 10^{-10.64}$

Let x be the change in concentration of HIO

Concentration (M)	HIO	H^+	IO^-
Initial	0.75	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.75 - x$	x	x

$K_a = [H^+][IO^-]/[HIO] = x^2/(0.75 - x)$

Since K_a is small assume $0.75 - x \approx 0.75$

$K_a = x^2/0.75 = 10^{-10.64} \longrightarrow x = 4.14506 \times 10^{-6} = [H^+]_{equilibrium}$

$pH = -\log_{10}[H^+] = -\log_{10}(4.14506 \times 10^{-6}) = 5.38$

Based on the graph, the major species in the solution at this pH is HCy^-, which is purple.

Question 22

The pH of a 0.10 M solution of a monoprotic acid was measured to be close to 3. What proportion of acid molecules are unionised?

(a) 0%

(b) 1%

(c) 60%

(d) 99%

$[H^+] = 10^{-pH} = 0.001$ M

0.001 / 0.1 = 0.01 = 1% ionised, therefore 99% unionised

Question 23

Which of the following solutions has the highest pH?

	Solution	K_a
(a)	0.1 M Hydrofluoric acid	6.8×10^{-4}
(b)	0.5 M Hydrofluoric acid	6.8×10^{-4}
(c)	0.1 M Chlorous acid	1.1×10^{-2}
(d)	0.5 M Chlorous acid	1.1×10^{-2}

Solution with the highest pH will have the smallest $[H^+]$. Hydrofluoric acid is a weaker acid compared to chlorous acid since it has a lower K_a, hence will ionise less. A more dilute solution will produce a lower concentration of H^+.

Question 24

What proportion of acid molecules are ionised in a 0.050 M solution of methanoic acid? The K_a of methanoic acid is 1.8×10^{-5}.

(a) 1.9%

(b) 2.5%

(c) 4.7%

(d) 5.3%

$\frac{x^2}{0.05} = 1.8 \times 10^{-5}$, $x = 9.4868 \times 10^{-4} = [H^+]$

$\% = (9.4868 \times 10^{-4}) / 0.05 \times 100 = 1.9\%$

Question 25

The pH of a 0.074 M solution of a monoprotic acid was measured to be 2.49. Which acid is it most likely to be?

	Solution	K_a
(a)	Hydrobromic acid	1.0×10^{9}
(b)	Benzoic acid	6.5×10^{-5}
(c)	Lactic acid	1.5×10^{-4}
(d)	Phenol	1.6×10^{-10}

$K_a = (10^{-2.49})^2 / (0.074 - 10^{-2.49}) = 1.48 \times 10^{-4}$

Question 26

The concentration of acetic acid ($pK_a = 4.74$) in vinegar was found to be 1.0 mol L^{-1}. What would the pH of vinegar be?

(a) 4.27×10^{-3}

(b) 1.82×10^{-5}

(c) 2.18

(d) 2.37

$$\frac{x^2}{1.0} = 10^{-4.74}$$

$x = 4.266 \times 10^{-3} = [H^+]$

Question 27

Which list contains only acidic substances?

(a) Wine, vinegar, aspirin

(b) Oven cleaner, rainwater, blood

(c) Baking soda, antacid tablets, laundry detergents

(d) Drain cleaner, milk, coffee

Vinegar contains acetic acid, wine contains a mix of acids. Aspirin is acetylsalicylic acid.

Household cleaning products tend to be basic.

Question 28

A student investigated the equilibrium mixture produced by the dissociation of hypochlorous acid (HOCl) in water. A pH probe was placed in a beaker containing 1.5×10^{-3} M HOCl solution. The pH was measured to be 5.18. What is the K_a for HOCl?

(a) 2.8×10^{-6}

(b) 2.7×10^{-7}

(c) 2.9×10^{-8}

(d) 2.6×10^{-9}

$K_a = (10^{-5.18})^2 / (1.5 \times 10^{-3} - 10^{-5.18})$

Question 29

Butyric acid C_3H_7COOH is a weak monoprotic acid.

(a) Write an equation for the dissociation of butyric acid in water. 1

$C_3H_7COOH_{(aq)} \rightleftharpoons C_3H_7COO^-_{(aq)} + H^+_{(aq)}$

(b) The K_a of butyric acid at 25 °C is 1.5×10^{-5}. What is the pH of 0.10 M butyric acid solution? 3

Concentration (M)	C_3H_7COOH	H^+	$C_3H_7COO^-$
Initial	0.10	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.10 - x$	x	x

Assume $[C_3H_7COOH]_{eqm}$ = 0.10 M, since K_a is small and dissociation does not change $[C_3H_7COOH]_{initial}$ by much.

$\frac{x^2}{0.1} = 1.5 \times 10^{-5}$

$x = 1.2247 \times 10^{-3} = [H^+]$

pH = 2.91195 = 2.91 (2 s.f.)

(c) The pH of 0.10 M hydrochloric acid is 1.0. Why is the pH of 0.10 M butyric acid not 1.0? 2

Butyric acid is a weak acid. This means it will only partially ionise in water and thus 0.1 M butyric acid produces less than 0.1 M H^+. Since $pH = -\log_{10}[H^+]$, pH > 1.0.

(d) What is the percentage dissociation of 0.10 M butyric acid at this temperature? 1

% dissociated = $[H^+]_{eqm} / [HA]_{initial} \times 100 = (1.2247 \times 10^{-3}) / 0.1 \times 100$ = 1.2% (2 s.f.)

Question 30

Formic acid HCOOH is a weak monoprotic acid. 0.024 g of formic acid was dissolved in water to make 0.600 L of solution, and its pH was measured to be 3.40. What is the pK_a for formic acid? 3

$HCOOH_{(aq)} \rightleftharpoons H^+_{(aq)} + HCOO^-_{(aq)}$

n(HCOOH) = 0.00052144 mol

c(HCOOH) = 0.00086907 M

$[HCOO^-]$ and $[H^+] = 10^{-3.4} = 0.0003981 \text{ mol L}^{-1}$

Concentration (M)	HCOOH	$HCOO^-$	H^+
Initial	0.00086907	0	0
Change	−0.0003981	+0.0003981	+0.0003981
Equilibrium	0.00047096	0.0003981	0.0003981

$$K_a = \frac{[H^+][HCOO^-]}{[HCOOH]} = \frac{[0.003981]^2}{[0.00047096]} = 3.365 \times 10^{-4}$$

pK_a = 3.47 (2 s.f.)

Question 31

The ascorbate ion $HC_6H_6O_6^-$ dissolves to produce a basic aqueous solution.

(a) Write the dissociation equation and the equilibrium expression for K_b for the ascorbate ion. 2

$HC_6H_6O_6^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HC_6H_6O_6H_{(aq)} + OH^-{}_{(aq)}$

$$K_b = \frac{[HC_6H_6O_6H][OH^-]}{[HC_6H_6O_6^-]}$$

(b) A 0.40 mol L^{-1} solution of sodium ascorbate has a pH of 8.85. Calculate the K_b of the ascorbate ion. 2

Concentration (M)	$HC_6H_6O_6^-$	$HC_6H_6O_6H$	OH^-
Initial	0.4	0	0
Change	$-10^{-5.15}$	$+10^{-5.15}$	$+10^{-5.15}$
Equilibrium	$0.4 - 10^{-5.15}$	$10^{-5.15}$	$10^{-5.15}$

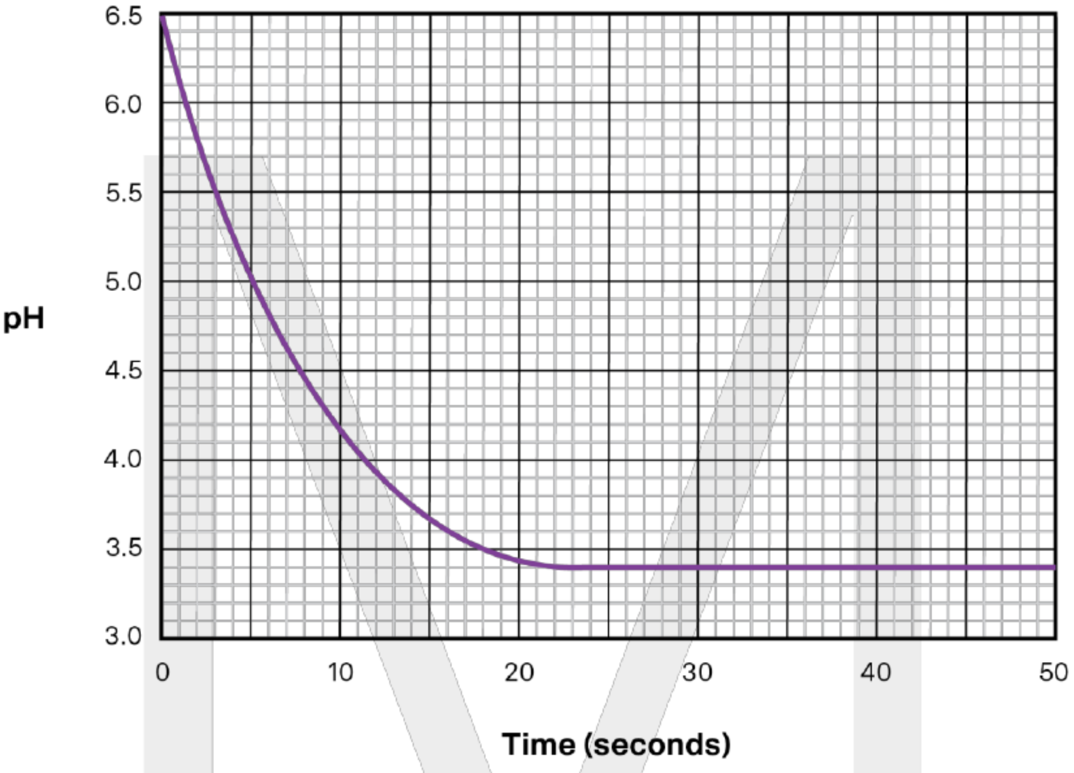
$K_b = (10^{-5.15})^2 / (0.4 - 10^{-5.15}) = 1.253 \times 10^{-10} = 1.3 \times 10^{-10}$ (2 s.f.)

(c) What is the K_a value for the conjugate acid of the ascorbate ion? 1

$K_a = 10^{-14} / K_b = 8.0 \times 10^{-5}$ (2 s.f.)

Question 32

Valeric acid C_4H_9COOH is a monoprotic acid used in the production of many industrial chemicals. 0.306 g of valeric acid was dissolved in water to make 250 mL of solution. The pH of the solution was measured over time:



Calculate the K_a of valeric acid. 3

$n(C_4H_9COOH) = 0.306 / (12.01 \times 5 + 1.008 \times 10 + 16 \times 2) = 2.99618 \times 10^{-3}$ mol

$[C_4H_9COOH]_{initial} = 2.99618 \times 10^{-3} / 0.25 = 0.011984725$ mol L^{-1}

Concentration (M)	C_4H_9COOH	H^+	$C_4H_9COO^-$
Initial	0.011984725	0	0
Change	$-10^{-3.40}$	$+10^{-3.40}$	$+10^{-3.40}$
Equilibrium	$0.011984725 - 10^{-3.40}$	$10^{-3.40}$	$10^{-3.40}$

$$K_a = \frac{[C_4H_2COO^-][H^+]}{[C_4H_2COOH]} = \frac{(10^{-3.4})^2}{(0.011984725 - 10^{-3.4})} = 1.4 \times 10^{-5} \text{ (2 s.f.)}$$

Question 33

The ammonium ion is an acidic ion. It has a K_a value of 5.6×10^{-10}.

(a) Write an equation for the dissociation of the ammonium ion in water. **1**

$NH_4{}^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^+{}_{(aq)}$

(b) What is the pH of a 450 mL solution containing 3.45 g of ammonium chloride? **3**

$n(NH_4Cl) = m / MM = 0.0644956$ mol

$c(NH_4Cl) = n / V = 0.1433236$ M $= c(NH_4{}^+)$

Concentration (M)	**$NH_4{}^+$**	**NH_3**	**H_3O^+**
Initial	0.1433236	0	0
Change	$- x$	$+ x$	$+ x$
Equilibrium	$0.1433236 - x$	x	x

Assume $[NH_4{}^+]_{eqm} = 0.1433$ M, since K_a is small and dissociation does not change $[NH_4{}^+]_{initial}$ by much.

$\frac{x^2}{0.1433} = 5.6 \times 10^{-10}$

$x = 8.958 \times 10^{-6} = [OH^-]$

pH = 5.0478 = 5.05 (2 s.f.)

Question 34

Arsenious acid is a highly toxic triprotic acid. The pH of a 0.055 M solution of arsenious acid is 5.24. Calculate its K_a, assuming that the second and third ionisation produce a negligible amount of H^+ ions. 2

$H_3A_{(aq)} \rightleftharpoons H^+_{(aq)} + H_2A^-_{(aq)}$

Concentration (M)	H_3A	H^+	H_2A^-
Initial	0.055	0	0
Change	$-10^{-5.24}$	$+10^{-5.24}$	$+10^{-5.24}$
Equilibrium	$0.055 - 10^{-5.24}$	$10^{-5.24}$	$10^{-5.24}$

$$K_a = \frac{[H^+][H_2A^-]}{[H_3A]} = \frac{(10^{-5.24})^2}{(0.055 - 10^{-5.24})} = 6.0 \times 10^{-10} \text{ (2 s.f.)}$$

Question 35

Outline a method involving a pH probe that can be used to determine the relative strengths of 3 solutions of 0.10 M weak monoprotic acids. Include an appropriate safety precaution in your answer and explain the expected result. 4

1. Pour 15 mL of each acid into 3 separate, dry test tubes.

2. Calibrate pH probe.

3. Measure the pH of each solution using a pH probe and record results.

A lower pH corresponds to higher $[H^+]$, which means greater dissociation. Since acid strength relates to the extent of dissociation, lower pH means stronger acid.

Safety: Perform experiment near a source of water so splashes can be quickly diluted

Question 36

Phenol, C_6H_5OH, is a weak monoprotic acid commonly used as an antiseptic.

(a) Write an equation for the dissociation of phenol in water. **1**

$C_6H_5OH_{(aq)} \rightleftharpoons C_6H_5O^-_{(aq)} + H^+_{(aq)}$

(b) A 3.5×10^{-3} mol L^{-1} solution of phenol has a pH of 6.22. Calculate the K_a of phenol. **2**

Concentration (M)	C_6H_5OH	H^+	$C_6H_5O^-$
Initial	3.5×10^{-3}	0	0
Change	$-10^{-6.22}$	$+10^{-6.22}$	$+10^{-6.22}$
Equilibrium	$3.5 \times 10^{-3} - 10^{-6.22}$	$10^{-6.22}$	$10^{-6.22}$

$K_a = (10^{-6.22})^2 / (3.5 \times 10^{-3} - 10^{-6.22}) = 1.03754 \times 10^{-10} = 1.0 \times 10^{-10}$ (2 s.f.)

(c) What is the percentage dissociation of phenol in this solution? **1**

$\% = 10^{-6.22} / (3.5 \times 10^{-3}) \times 100 = 0.017\%$

Question 37

(a) Magnesium hydroxide is sparingly soluble in water. What is the molar solubility of magnesium hydroxide in water at 25 °C? 2

$Mg(OH)_{2(s)} \rightleftharpoons Mg^{2+}_{(aq)} + 2OH^{-}_{(aq)}$ $K_{sp} = [Mg^{2+}][OH^{-}]^2 = 5.61 \times 10^{-12}$

Let moles of $Mg(OH)_2$ in 1 L = x, thus $[Mg^{2+}] = x$ and $[OH^{-}] = 2x$

$K_{sp} = (x)(2x)^2 = 5.61 \times 10^{-12}$

$x = 1.11935 \times 10^{-4}$

Solubility of $Mg(OH)_2 = 1.12 \times 10^{-4}$ mol/L (3 s.f.)

(b) What is the pH of a saturated solution of magnesium hydroxide in water at 25 °C? 2

$[OH^{-}] = 2x = 2.2387 \times 10^{-4}$ mol/L

$pH = 14 + \log_{10}(2.2387 \times 10^{-4}) = 10.350$ (3 s.f.)

(c) What is the solubility of magnesium hydroxide in a solution that has a pH of 11.24 at 25 °C? 3

pOH = 2.76, $[OH^{-}] = 10^{-2.76}$ M

Let moles of $Mg(OH)_2$ in 1 L = x

Concentration (M)	$Mg(OH)_2$	Mg^{2+}	$2OH^{-}$
Initial	–	0	$10^{-2.76}$
Change	–	$+x$	$+2x$
Equilibrium	–	x	$10^{-2.76} + 2x$

Assuming the $[OH^{-}]_{eqm}$ does not change very much since K_{sp} is very small:

$K_{sp} = x \times (10^{-2.76})^2 = 5.61 \times 10^{-12}$

$x = 1.8576 \times 10^{-6}$

Thus solubility of $Mg(OH)_2$ is 1.9×10^{-6} mol /L (2 s.f.)

Question 38

(a) Calculate the pH of a 0.070 M solution of sulfuric acid (K_{a1} = very large, $K_{a2} = 1.2 \times 10^{-2}$). **4**

$H_2SO_{4(aq)} \longrightarrow H^+_{(aq)} + HSO_4^-{}_{(aq)}$

$[H^+]$ and $[HSO_4^-] = 0.070$ M

$HSO_4^-{}_{(aq)} \rightleftharpoons H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

Concentration (M)	**HSO_4^-**	**H^+**	**SO_4^{2-}**
Initial	0.070	0.070	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.070 - x$	$0.070 + x$	x

$K_a = \frac{(0.070 + x)(x)}{0.070 - x} = 1.2 \times 10^{-2}$

Solve using quadratic equation: $x = 0.009209561$

Total $[H^+] = 0.070 + 0.009209561 = 0.07921$ M

$pH = -\log_{10}(0.07921) = 1.10$ (2 s.f.)

(b) What is the ratio of HSO_4^- to SO_4^{2-} ions in the solution? **2**

$[HSO_4^-] = 0.06079$ M

$[SO_4^{2-}] = 0.009209561$ M

0.06079 M / 0.009209561 M = 6.6

ratio = 6.6 : 1

Year 12
Chemistry
Acid/Base Reactions

Work Book 5
pH of mixed solutions and buffers

MATRIX
EDUCATION

Part A: Multiple choice questions

Question 1

A small amount of acid is added to a buffer solution. As a result, the pH of the solution will be:

(a) the same

(b) slightly lower

(c) slightly higher

(d) pH 7

A buffer resists change in pH, but does not completely eliminate it.

Question 2

Which acid / base pair could act as a buffer?

(a) HBr / Br^-

(b) H_3O^+ / H_2O

(c) CH_3COOH / CH_3COO^-

(d) $H_2PO_4^-$ / PO_4^{3-}

A buffer requires a weak acid and its conjugate base.

Question 3

Which action would result in a pH increase of 2.0?

(a) Adding 90 mL water to 10 mL of 0.01 mol L^{-1} HCl solution

(b) Adding 90 mL of water to 10 mL of 0.01 mol L^{-1} NaOH solution

(c) Adding 990 mL water to 10 mL of 0.01 mol L^{-1} HCl solution

(d) Adding 990 mL of water to 10 mL of 0.01 mol L^{-1} NaOH solution

pH change of 2 units = 100-fold dilution. Diluting acid = increase pH whereas diluting base = decrease pH.

Question 4

For which of the following combinations would the resultant solution have the lowest pH?

(a) 20.0 mL of 0.100 M NaOH and 10.0 mL of 0.010 M HCl

(b) 20.0 mL of 0.100 M NH_3 and 80.0 mL of 0.050 M NaOH

(c) 50.0 mL of 0.100 M HCl and 80.0 mL of 0.100 M HNO_3

(d) 20.0 mL of 0.100 M HNO_3 and 200 mL of distilled water

(a) and (b) both will end up basic.

(d) starts at pH 1 and is diluted, whereas (c) remains at pH 1.

Question 5

Consider the following:

- (i) HCl and NaOH
- (ii) CH_3COOH and CH_3COONa
- (iii) NH_3 and NH_4Cl

Of these three pairs of compounds, which could be combined to make a buffer?

(a) all of them

(b) i and ii

(c) ii and iii

(d) only iii

A buffer consists of a weak acid and its conjugate base.

Acetic acid/acetate and ammonium/ammonia are both conjugate pairs that could function as buffers.

Question 6

What is the pH of a solution prepared by mixing 60.0 mL of 0.10 M sulfuric acid with 50.0 mL of 0.20 M barium hydroxide solution?

(a) 12.86

(b) 1.14

(c) 7.00

(d) 12.56

$H_2SO_{4(aq)} + Ba(OH)_{2(aq)} \longrightarrow BaSO_{4(s)} + 2H_2O_{(l)}$

$n(H_2SO_4) = c \times V = 0.06 \times 0.1 = 0.006$ mol

$n(Ba(OH)_2) = c \times V = 0.2 \times 0.05 = 0.01$ mol

Excess $n(Ba(OH)_2) = 0.01 - 0.006 = 0.004$ mol $= 0.008$ mol OH^-

Total volume $= 0.06 + 0.05 = 0.11$ L

$[OH^-] = 0.008 / 0.11 = 0.072727$ M

$pH = 14 + \log_{10}(0.072727) = 12.86$

Question 7

A solution of nitric acid was diluted. As a result, the pH changed from 2 to 4. What was the change in the hydrogen ion concentration of the diluted solution?

(a) Halved

(b) Doubled

(c) Decreased by 100 times

(d) Increased by 100 times

pH change of 2 units = 100-fold dilution. Increase pH = diluting hydrogen concentration.

Question 8

Which of the following solutions could buffer the largest amount of excess acid or base?

(a) 0.1 M NH_3 and 0.1 M NH_4^+

(b) 0.5 M NH_3 and 0.5 M NH_4^+

(c) 0.1 M NH_3 and 0.1 M NaOH

(d) 0.5 M NH_3 and 0.5 M NaOH

A buffer requires a weak acid and its conjugate base, so (c) and (d) are not buffers.

(a) and (b) are both buffers however (b) has the higher concentration, hence the equilibrium can continue to shift to absorb excess acid or base for the longest.

Question 9

Given the following ionisation equations of hypothetical acids and bases in water:

$$HL_{(g)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + L^-_{(aq)}$$

$$MOH_{(s)} \longrightarrow M^+_{(aq)} + OH^-_{(aq)}$$

$$HZ_{(g)} + H_2O_{(l)} \rightleftharpoons Z^-_{(aq)} + H_3O^+_{(aq)}$$

$$H_2D^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HD^{2-}_{(aq)} + H_3O^+_{(aq)}$$

Which combinations would produce a buffer?

(a) NaL + HL and MOH + NaCl

(b) NaZ + HL and MOH + NaCl

(c) NaH_2D + Na_2HD and NaZ + HZ

(d) NaH_2D + Na_2HD and MCl + NaOH

A buffer consists of a mixture of a weak acid and its conjugate base. The first two equations shows HL is a strong acid and MOH is a strong base (full arrow).

Question 10

A solution of pH 2 has the necessary changes made to it so that it is now a pH of 5. What change has been made to the concentration of H^+?

(a) It has become less concentrated by a factor of 1000

(b) It has become more concentrated by a factor of 3

(c) It has become more concentrated by a factor of 1000

(d) It has become less concentrated by a factor of 3

pH uses a $\log_{10}$ scale. This means that for every 1 point rise in pH, there is a 10x decrease in H^+ ion concentration.

Question 11

The pK_a values for phosphoric acid are given in the table below.

Acid	Formula	pK_{a1}	pK_{a2}	pK_{a3}
Phosphoric acid	H_3PO_4	2.12	7.21	12.66

Which of the following solutions would produce a buffer that is effective in the range of 7–8?

(a) Equal amounts of H_3PO_4 and NaH_2PO_4

(b) Equal amounts of NaH_2PO_4 and Na_2HPO_4

(c) Equal amounts of Na_2HPO_4 and Na_3PO_4

(d) NaH_2PO_4 only

A buffer is most effective at a pH near its pK_a.

Part B: Extended response questions

Question 12

A 500 mL solution was prepared using 1.20 g of HCl gas. 10.0 mL of this solution was taken and diluted to 150 mL. What is the pH of the final solution? 3

n = 1.20 / (1.008 + 35.45) = 0.0329146 mol

$c_1 = n / V = 0.0329146 / 0.5 = 0.065829$ M

$c_1V_1 = c_2V_2$

$c_2 = c_1V_1 / V_2 = 0.065829 \times 0.01 / 0.15 = 4.39 \times 10^{-3}$ M

$\text{pH} = -\log_{10}(4.39 \times 10^{-3}) = 2.358$ (3 sig. fig.)

Question 13

The dissociation of hydrofluoric acid in water is shown below.

$$HF_{(aq)} + H_2O_{(l)} \rightleftharpoons F^-_{(aq)} + H_3O^+_{(aq)}$$

Predict how equilibrium will shift if a solution of hydrofluoric acid is diluted to one-third of its original concentration. Explain your answer. 2

$$K_a = \frac{[H^+][F^-]}{[HF]}$$

Diluting causes all concentrations to be reduced to 1/3 of the original:

$$Q = \frac{\frac{[H^+]}{3} \times \frac{[F^-]}{3}}{\frac{[HF]}{3}} = K_a \times \frac{1}{3}$$

Therefore $Q < K$, equilibrium will shift to the right to reach a new equilibrium with $Q = K$.

Question 14

One of the buffer systems used to regulate the pH of the blood consists of a mixture of carbonic acid and hydrogen carbonate.

(a) Hydrogen ions are produced during the breakdown of glucose in blood during exercise. Explain, using a balanced equation, the importance of the carbonic acid-hydrogen carbonate buffer in maintaining a stable blood pH during exercise. Refer to Le Chatelier's principle in your answer. 3

Buffers are solutions that resist changes in pH when small quantities of acid or base are added. The following equilibrium buffers the pH of blood:

$H_2CO_{3(aq)} \rightleftharpoons HCO_3^-{}_{(aq)} + H^+_{(aq)}$

When additional H^+ is produced during exercise, the equilibrium is disturbed. According to Le Chatelier's principle, the position of equilibrium shifts to the left to remove H^+ and minimise the disturbance. This keeps the concentration of H^+ in blood stable, and since $pH = -\log_{10}[H^+]$ the pH is also stable.

(b) Explain why hydrochloric acid and sodium chloride cannot be used to make a buffer solution. 2

HCl is a strong acid, therefore it is completely ionised in water and does not exist in equilibrium with Cl^-. Therefore, an equilibrium cannot shift to minimise the disturbance from a change in H^+/OH^- concentration.

Question 15

Calculate the pH of the final solution when the following solutions are mixed:

(a) 55 mL of 0.040 M $Ba(OH)_2$ and 30 mL of 0.50 M KOH 3

$n(Ba(OH)_2) = 0.04 \times 0.055 = 2.2 \times 10^{-3}$ mol

$n(KOH) = 0.5 \times 0.03 = 0.015$ mol

$n(OH^-)_{total} = 2 \times 2.2 \times 10^{-3} + 0.015 = 0.0194$ mol

$[OH^-] = 0.228235$ M

pOH = 0.6416

pH = 13.36 (2 s.f.)

(b) 25.0 mL of HBr with a pH of 1.20 and 30.0 mL of HCl with a pH of 2.40 3

$[H^+] = 10^{-pH}$

$n(H^+)$ from HBr $= 10^{-1.2} \times 0.025 = 1.5774 \times 10^{-3}$ mol

$n(H^+)$ from HCl $= 10^{-2.4} \times 0.03 = 1.19432 \times 10^{-4}$ mol

$n(H^+)_{total} = 1.69683 \times 10^{-3}$ mol

$[H^+] = 0.030851$ M

pH = 1.51 (2 s.f.)

(c) 150 mL of 0.020 M HNO_3, 0.20 g LiOH powder and 100 mL of distilled water 3

$$HNO_{3(aq)} + LiOH_{2(s)} \longrightarrow LiNO_{3(aq)} + H_2O_{(l)}$$

$n(HNO_3) = c \times V = 0.02 \times 0.15 = 3 \times 10^{-3}$ mol

$n(LiOH) = m / MM = 0.2 / (6.941 + 16 + 1.008) = 8.351079 \times 10^{-3}$ mol

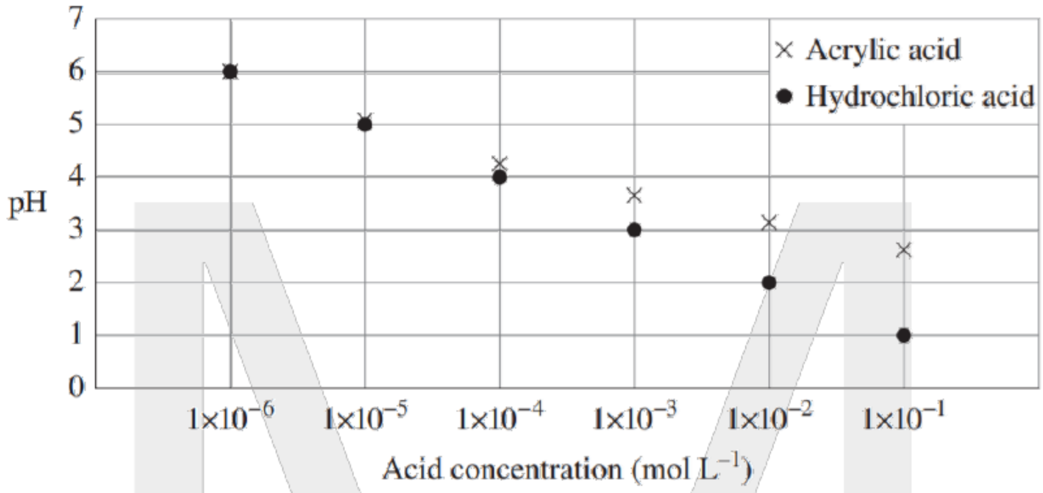
Excess $n(LiOH) = 8.351079 \times 10^{-3} - 3 \times 10^{-3} = 5.351079 \times 10^{-3}$ mol $= n(OH^-)$

$[OH^-] = 5.351079 \times 10^{-3} / 0.250 = 0.0214043$ M

$pH = 14 - (-\log_{10}[OH^-]) = 12.33$ (2 sig. fig)

Question 16 [2020 HSC Q34]

The effect of concentration on the pH of acrylic acid (C_2H_3COOH) and hydrochloric acid (HCl) solutions is shown in the graph. Both of these acids are monoprotic.



Explain the trends in the graph. Include relevant chemical equations in your answer. **4**

NESA sample answer: The trend for both acids is for pH to decrease with increasing concentration as $[H^+]$ increases. The strong acid, HCl, shows a linear trend with pH (the log of [HCl]). This is because HCl fully dissociates at all concentrations, and therefore the $[H^+]$ equals [HCl]. However, acrylic acid is a weak acid. The degree of dissociation is dependent on the concentration of the acid. We can see that the trend of the acrylic acid is not linear. It bends up at higher concentrations. This reflects the proportionally lower fraction of dissociation. This is given by the following equilibrium expression.

$$C_2H_3COOH_{(aq)} \rightleftharpoons C_2H_3COO^-_{(aq)} + H^+_{(aq)}$$

$$K_a = \frac{[C_2H_3COO^-][H^+]}{[C_2H_3COOH]}$$

$$[H^+] = \frac{K_a[C_2H_3COOH]}{[C_2H_3COO^-]}$$

Answers could include:

- Increasing concentration decreases pH for both acids
- The pH of the HCl is lower than that of the C_2H_3COOH
- For C_2H_3COOH the degree of ionisation increases as concentration decreases.

Question 17

A buffer solution can be prepared by combining ammonia with ammonium chloride.

(a) Write the equation for the equilibrium that buffers this system. 1

$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-{}_{(aq)}$

(b) Given the pK_b of ammonia at 298 K is 4.75, calculate the mass of ammonium chloride that needs to be added to 2.5 L of 0.10 mol L^{-1} ammonia solution to form a buffer solution with a pH of 9.20 at 298 K.

Assume the volume of the solution is constant with the addition of ammonium chloride. 5

pOH = 14.00 – 9.20 = 4.80, hence $[OH^-] = 10^{-pOH} = 10^{-4.80}$ M

Concentration (M)	NH_3	NH_4^+	OH^-
Initial	0.10	x	0
Change	$-10^{-4.80}$	$+10^{-4.80}$	$+10^{-4.80}$
Equilibrium	$0.10 - 10^{-4.80}$	$x + 10^{-4.80}$	$10^{-4.80}$

$$K_b = \frac{[NH_4{}^+][OH^-]}{[NH_3]} = \frac{(x + 10^{-4.80})(10^{-4.80})}{(0.10 - 10^{-4.80})} = 10^{-4.75}$$

$[NH_4^+]_{initial} = x = 0.112168$ M

$n(NH_4^+)$ in 2.5 L = 0.112168 × 2.5 = 0.2804205 mol = $n(NH_4Cl)$

$m(NH_4Cl) = 0.2804205 \times (14.01 + 1.008 \times 4 + 35.45) = 15$ g (2 s.f.)

Alternate working:

$pH = pK_a + \log_{10}([A^-] / [HA])$

$9.20 = (14 - 4.75) + \log_{10}(0.100/[NH_4^+])$

$[NH_4^+] = 0.100 / 10^{(9.20 - 9.25)} = 0.1122$ M

$n(NH_4^+)$ in 2.5 L = 0.1122 × 2.5 = 0.2805 mol = $n(NH_4Cl)$

$m(NH_4Cl) = 0.2805 \times (14.01 + 1.008 \times 4 + 35.45) = 15$ g (2 s.f.)

Question 18

(a) 10.0 mL of a solution of HNO_3 at pH 1.0 was diluted to 100.0 mL. What is the pH of the new solution? **1**

10-fold dilution, therefore pH = 2.0

(b) 10.0 mL of a solution of acetic acid at pH 2.13 was diluted to 100.0 mL. What is the pH of the new solution? The pK_a of acetic acid is 4.75. **4**

$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$

Work out original concentration (C):

Concentration (M)	CH_3COOH	H^+	CH_3COO^-
Initial	C	0	0
Change	$- 10^{-2.13}$	$+ 10^{-2.13}$	$+ 10^{-2.13}$
Equilibrium	$C - 10^{-2.13}$	$10^{-2.13}$	$10^{-2.13}$

$(10^{-2.13})^2 / (C - 10^{-2.13}) = 10^{-4.75}$

$C = 10^{-2.13} + (10^{-2.13})^2 / 10^{-4.75} = 3.0977$ M

Dilution: $c_1V_1 = c_2V_2$, $c_2 = 0.30977$ M

pH of diluted solution:

Concentration (M)	CH_3COOH	H^+	CH_3COO^-
Initial	0.30977	0	0
Change	$- x$	$+ x$	$+ x$
Equilibrium	$0.30977 - x$	x	x

Assume $[CH_3COOH]_{eqm} = 0.30977$ M, since K_a is small and dissociation does not change $[CH_3COOH]_{initial}$ by much.

$$\frac{x^2}{0.30977} = 10^{-4.75}$$

$x = 2.347 \times 10^{-3} = [H^+]$, pH = 2.6295 = 2.63 (2 s.f.)

Question 19

A buffer solution is prepared by mixing 100 mL of 0.10 mol L^{-1} sodium dihydrogen phosphate and 80 mL of 0.12 mol L^{-1} of sodium hydrogen phosphate.

(a) What determines the pH of a buffer solution? 1

The pK_a of the acid and the relative concentrations of the acid and its conjugate base.

(b) Calculate the approximate pH of this buffer solution. The K_a of dihydrogen phosphate ion is 6.2×10^{-8}. 2

The approximate pH can be calculated using Henderson–Hasselbalch equation.

$pH = pK_a + \log_{10}([A^-] / [HA]) = -\log_{10}(6.2 \times 10^{-8}) + \log_{10}(0.120/0.100) = 7.29$ (2 s.f.)

(c) Explain how the pH of this solution would be affected by the addition of a small amount of potassium hydroxide solution. Include balanced equations in your answer. 3

The following equilibrium occurs in this buffer solution:

$$H_2PO_4^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HPO_4^{2-}{}_{(aq)} + H_3O^+{}_{(aq)}$$

The addition of OH^- ions will cause a reaction with H_3O^+ ions, reducing their concentration in the equilibrium mixture. This will force the reaction to the right according to Le Chatelier's principle, increasing the H_3O^+ concentration to minimise the disturbance. As H_3O^+ concentration remains relatively the same, the pH remains relatively the same since $pH = -\log_{10}[H_3O^+]$.

(d) Dihydrogen phosphate ion dissociates to give hydrogen phosphate ion. Explain why approximately equal amounts of dihydrogen phosphate and hydrogen phosphate are mixed together to create a buffer and not just dihydrogen phosphate ion only. 2

The dissociation of dihydrogen phosphate ion to hydrogen phosphate ion is minimal since dihydrogen phosphate ion is a weak acid. Hence when only dihydrogen phosphate ion is used, there will not be enough hydrogen phosphate ion (conjugate base) to react with added acid, causing the pH to rapidly decrease. Thus solution will not be a good buffer.

Question 20 **[2020 HSC Q33]**

Excess solid calcium hydroxide is added to a beaker containing 0.100 L of 2.00 mol L^{-1} hydrochloric acid and the mixture is allowed to come to equilibrium.

(a) Show that the amount (in mol) of calcium hydroxide that reacts with the hydrochloric acid is 0.100 mol. 2

NESA sample answer:

$Ca(OH)_{2(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + 2H_2O_{(l)}$

$n(HCl) = 2.00 \text{ mol L}^{-1} \times 0.100 \text{ L} = 0.200 \text{ mol}$

$n(Ca(OH)_2) = n(HCl) / 2 = 0.200 \text{ mol} / 2 = 0.100 \text{ mol}$

(b) It is valid in this instance to make the simplifying assumption that the amount of calcium ions present at equilibrium is equal to the amount generated in the reaction in part (a). Calculate the pH of the resulting solution. 4

NESA sample answer:

$[Ca^{2+}] = 0.100 \text{ mol} / 0.100 \text{ L} = 1.00 \text{ mol L}^{-1}$

$K = 5.02 \times 10^{-6} = [Ca^{2+}][OH^-]^2$

Let $[OH^-] = x$ mol L^{-1}

$\therefore 1.00 \times x^2 = 5.02 \times 10^{-6}$

$x = \sqrt{5.02 \times 10^{-6}} = 2.24 \times 10^{-3} \text{ mol L}^{-1}$

$\therefore [OH^-] = 2.24 \times 10^{-3} \text{ mol L}^{-1}$

$pOH = -\log_{10}(2.24 \times 10^{-3}) = 2.650$

$pH = 14.00 - 2.650 = 11.35$

Question 21

A student prepared 200 mL of 2.0 mol L^{-1} solution of a weak base. The pH of the solution was found to be 9.80.

(a) Calculate the K_b for this weak base. **3**

pOH = 4.2

$B^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HB_{(aq)} + OH^-_{(aq)}$

$$K_b = \frac{[HB][OH^-]}{[B^-]}$$

Concentration (M)	B^-	HB	OH^-
Initial	2	0	0
Change	$-10^{-4.2}$	$+10^{-4.2}$	$+10^{-4.2}$
Equilibrium	$2 - 10^{-4.2}$	$10^{-4.2}$	$10^{-4.2}$

$K_b = (10^{-4.2})^2 / (2 - 10^{-4.2}) = 1.9906 \times 10^{-9} = 2.0 \times 10^{-9}$ (2 s.f.)

(b) 30.0 mL of this basic solution was diluted to 100.0 mL. What is the pH of the new solution? **4**

$c_2 = 0.03 \times 2 / 0.1 = 0.6$ M

Concentration (M)	B^-	HB	OH^-
Initial	0.6	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.6 - x$	x	x

Assume $[B^-]_{eqm}$ = 0.6 M, since K_b is small and dissociation does not change $[B^-]_{initial}$ by much.

$\frac{x^2}{0.6} = 1.9906 \times 10^{-9}$

$x = 3.45595 \times 10^{-5} = [OH^-]$

pOH = 4.4614

pH = 9.54 (2 s.f.)

Part C: Additional practice questions

The questions in this part are intended to be used for revision purposes. However, from time to time your teacher may prescribe some of these questions for homework.

Question 22

Which combination of equimolar solutions would produce the most basic mixture?

(a) Nitric acid and calcium hydroxide

(b) Nitric acid and sodium carbonate

(c) Sulfuric acid and calcium hydroxide

(d) Sulfuric acid and sodium carbonate

Nitric acid is monoprotic, whereas calcium hydroxide contains two hydroxide ions, therefore there will be an excess of hydroxide ions. Although sodium carbonate also reacts with nitric acid in a 1:2 ratio, calcium hydroxide will still result in a higher pH as it is strong.

Question 23

500 mL of 2.0 mol L^{-1} sodium hydroxide solution is added to 500 mL of 2.0 mol L^{-1} sulfuric acid. Which correctly describes the resulting solution?

(a) The solution will be neutral

(b) The solution will contain an excess of hydrogen ions

(c) The solution will contain an excess of hydroxide ions

(d) The pH will fluctuate

$H_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow Na_2SO_{4(s)} + 2H_2O_{(l)}$

Same moles present, therefore, H_2SO_4 is in excess

Question 24

A phosphate buffer is typically made by mixing phosphoric acid (pH = 1.6) and a salt solution. What would be an appropriate salt solution for making a phosphate buffer, and the buffer's estimated effective pH range?

(a) Sodium chloride, pH < 1.6
(b) Sodium chloride, pH = 7.0
(c) Sodium dihydrogen phosphate, pH > 1.6
(d) Sodium dihydrogen phosphate, pH < 1.6

A buffer consists of a weak acid and its conjugate base, so sodium dihydrogen phosphate is required. Since sodium dihydrogen phosphate is the conjugate base, the pH of this buffer must be greater than that of phosphoric acid (1.6).

Question 25

What volume of 10.0 M HCl is required to prepare 500 mL of 0.350 M HCl?

(a) 6.75 mL
(b) 17.5 mL
(c) 21.5 mL
(d) 35.0 mL

$c_1V_1 = c_2V_2$

$V_1 = 0.35 \times 0.5 / 10 = 0.0175$ L = 17.5 mL

Question 26

Calculate the pH of the resultant solution when the following solutions are mixed:

(a) 20.0 mL of 0.100 M NaOH and 80.0 mL of 0.050 M HCl 3

$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

$n(NaOH) = c \times V = 0.1 \times 0.02 = 2 \times 10^{-3}$ mol

$n(HCl) = c \times V = 0.05 \times 0.08 = 4 \times 10^{-3}$ mol

Excess $n(HCl) = 4 \times 10^{-3} - 2 \times 10^{-3} = 2 \times 10^{-3}$ mol

$c(HCl) = n / V_{total} = 2 \times 10^{-3} / 0.1 = 2 \times 10^{-2}\ M = [H^+]$

$pH = -\log_{10}[H^+] = 1.70$ (2 sig. fig)

(b) 45.00 mL of 0.050 M HCl and 0.18 g $Ca(OH)_2$ powder 3

$2HCl_{(aq)} + Ca(OH)_{2(s)} \longrightarrow CaCl_{2(aq)} + 2H_2O_{(l)}$

$n(HCl) = c \times V = 0.05 \times 0.045 = 2.25 \times 10^{-3}$ mol

$n(Ca(OH)_2) = m / MM = 0.18 / (40.08 + 16 \times 2 + 1.008 \times 2) = 2.42928 \times 10^{-3}$ mol

Excess $n(Ca(OH)_2) = 2.42928 \times 10^{-3} - (2.25 \times 10^{-3} / 2) = 1.3048 \times 10^{-3}$ mol

$c(Ca(OH)_2) = n / V_{total} = 1.3048 \times 10^{-3} / 0.045 = 2.8996 \times 10^{-2}\ M$

$[OH^-] = 2.8996 \times 10^{-2} \times 2 = 0.05799\ M$

$pH = 14 - \log_{10}[OH^-] = 12.76$ (2 sig. fig)

(c) 150 mL of 0.0200 M HNO_3 and 100 mL of distilled water 2

$n(HNO_3) = 0.02 \times 0.15 = 3 \times 10^{-3}$ mol

Total V = 0.25 L

$[H^+] = 0.012\ M$

pH = 1.921 (2 s.f.)

(d) 50.0 mL of 0.300 M LiOH and 50.0 mL of 0.0200 M HNO_3 3

$HNO_{3(aq)} + LiOH_{(aq)} \longrightarrow LiNO_{3(aq)} + H_2O_{(l)}$

$n(LiOH) = c \times V = 0.3 \times 0.05 = 0.015$ mol

$n(HNO_3) = c \times V = 0.02 \times 0.05 = 1 \times 10^{-3}$ mol

Excess $n(LiOH) = 0.015 - 1 \times 10^{-3} = 0.014$ mol $= n(OH^-)$

$[OH^-] = n(OH^-)_{excess} / V_{total} = 0.14$ M

$pH = 14 - \log_{10}[OH^-] = 13.146$ (3 sig. fig)

(e) 200 mL of 0.00305 M HNO_3 and 100 mL of 0.0260 M HI 3

$n(HNO_3) = 0.00305 \times 0.2 = 6.1 \times 10^{-4}$ mol

$n(HI) = 0.026 \times 0.1 = 2.6 \times 10^{-3}$ mol

$n(H^+)_{total} = 3.21 \times 10^{-3}$ mol

$[H^+] = n(H^+)_{total} / V_{total} = 0.0107$ M

pH = 1.971 (3 s.f.)

(f) 100 mL of water, 50.0 mL of 0.20 M H_2SO_4 and 1.40 g of KOH 3

$H_2SO_{4(aq)} + 2KOH_{(aq)} \longrightarrow K_2SO_{4(aq)} + 2H_2O_{(l)}$

$n(H_2SO_4) = c \times V = 0.2 \times 0.05 = 0.01$ mol

$n(KOH) = m / MM = 1.4 / (39.1 + 16 + 1.008) = 0.02495$ mol

Excess $n(KOH) = 0.02495 - 0.01 \times 2 = 0.00495$ mol $= n(OH^-)$

$[OH^-] = n(OH^-)_{excess} / V_{total} = 0.0330125$ M

pOH = 1.4813

pH = 12.52 (2 s.f.)

(g) 70.0 mL of 0.050 M $Ca(OH)_2$ and 125.0 mL of 0.130 M HBr 3

$2HBr_{(aq)} + Ca(OH)_{2(s)} \longrightarrow CaBr_{2(aq)} + 2H_2O_{(l)}$

$n(HBr) = 0.13 \times 0.125 = 1.625 \times 10^{-2}$ mol

$n(Ca(OH)_2) = 0.05 \times 0.07 = 3.5 \times 10^{-3}$ mol

Excess $n(HBr) = 1.625 \times 10^{-2} - (3.5 \times 10^{-3} \times 2) = 9.25 \times 10^{-3}$ mol $= n(H^+)$

$[H^+] = n(H^+)_{excess} / V_{total} = 0.0474359$ M

pH = 1.32 (2 sig. fig)

Question 27

Explain how a buffer works. Refer to a natural buffer system in your response. 4

Buffers are solutions that resist changes in pH when small quantities of acid or base are added. The following equilibrium buffers the pH of blood at approximately 7.4:

$H_2CO_{3(aq)} \rightleftharpoons HCO_3^-{}_{(aq)} + H^+_{(aq)}$

When additional H^+ is produced during exercise, the equilibrium is disturbed. According to Le Chatelier's principle, the position of equilibrium shifts to the left to remove H^+ and minimise the disturbance. This keeps the concentration of H^+ in blood stable, and since $pH = -\log_{10}[H^+]$ the pH is also stable.

Question 28

A solution was made in which 50.0 mL of 0.12 M HNO_2 and 50.0 mL of 0.035 M $NaNO_2$ were added together.

(a) Calculate the approximate pH of the final solution using the Henderson–Hasselbalch equation. The K_a of HNO_2 is 7.1×10^{-4}. **2**

The solution contains a weak acid (HNO_2) and its conjugate base (NO_2^-).

$\text{pH} = \text{p}K_a + \log_{10}([A^-] / [HA]) = -\log_{10}(7.1 \times 10^{-4}) + \log_{10}(0.035/0.12) = 2.61$ (2 s.f.)

Note: The concentration of HNO_2 and NO_2^- in the final volume do not need to be determined here as both concentrations will be halved and the ratio of NO_2^- and HNO_2 will be the same.

(b) Determine the actual pH of the final solution. (HINT: You will need to construct an ICE table) **4**

Total V = 2 × original V of each solution, therefore concentrations will be halved:

$c(HNO_2) = 0.12 / 2 = 0.06$ M

$c(NO_2^-) = c(NaNO_2) = 0.035 / 2 = 0.0175$ M

Concentration (M)	HNO_2	H^+	NO_2^-
Initial	0.06	0	0.0175
Change	$-x$	$+x$	$+x$
Equilibrium	$0.06 - x$	x	$0.0175 + x$

$x(0.0175 + x) / (0.06 - x) = 7.1 \times 10^{-4}$

$0.0175x + x^2 = 4.26 \times 10^{-5} - (7.1 \times 10^{-4})x$

$x^2 + 0.01821x - 4.26 \times 10^{-5} = 0$

$x = 2.0977 \times 10^{-3} = [H^+]$

pH = 2.68 (2 s.f.)

Question 29

The conjugate acid/base pair $H_2PO_4^-/HPO_4^{2-}$ plays an important role in maintaining the pH of living cells. Explain how this conjugate acid/base pair works to maintain the pH. 3

$H_2PO_4^-/HPO_4^{2-}$ forms a buffer system in cells, which can neutralise small amounts of additional acids or bases, maintaining a stable pH. This is important as a lot of processes that occur in living systems are pH dependent.

$H_2PO_4^-{}_{(aq)} + H_2O_{(l)} \rightleftharpoons HPO_4^{2-}{}_{(aq)} + H_3O^+{}_{(aq)}$

If acid (H_3O^+) is added, the equilibrium will shift left to remove H_3O^+ from the system and minimise the disturbance (Le Chatelier's principle), keeping pH relatively constant, since $pH = -\log_{10}[H_3O^+]$

If base (OH^-) is added: $H_2PO_4^-{}_{(aq)} + OH^-{}_{(aq)} \longrightarrow HPO_4^{2-}{}_{(aq)} + H_2O_{(l)}$

The added OH^- is removed, again keeping pH relatively constant.

Question 30 [2012 HSC Q28]

A solution was made by mixing 75.00 mL of 0.120 mol L^{-1} hydrochloric acid with 25.00 mL of 0.200 mol L^{-1} sodium hydroxide.

What is the pH of the solution? 3

$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

$n(HCl) = c \times V = 0.120 \times 0.07500 = 0.00900$ mol

$n(NaOH) = c \times V = 0.200 \times 0.02500 = 0.00500$ mol

Excess $n(HCl) = n(HCl)_{initial} - n(HCl)_{reacted} = 0.00900 - 0.00500 = 0.00400$ mol

$n(H^+) =$ excess $n(HCl) = 0.00400$ mol

$[H^+] = n / V = 0.00400 / (0.075 + 0.025) = 0.0400$ M

$pH = -\log_{10}[H^+] = -\log_{10}(0.0400) = 1.398$ (3 sig. fig.)

Question 31

(a) 7.00 mL of a solution of hydrofluoric acid at pH 3.65 was diluted to 65.0 mL. What is the pH of the new solution? The pK_a of hydrofluoric acid is 3.17. **4**

$HF_{(aq)} \rightleftharpoons H^+_{(aq)} + F^-_{(aq)}$

Work out original concentration:

Concentration (M)	HF	H^+	F^-
Initial	C	0	0
Change	$-10^{-3.65}$	$+10^{-3.65}$	$+10^{-3.65}$
Equilibrium	$C - 10^{-3.65}$	$10^{-3.65}$	$10^{-3.65}$

$(10^{-3.65})^2 / (C - 10^{-3.65}) = 10^{-3.17}$

$(10^{-3.65})^2 = 10^{-3.17} \times (C - 10^{-3.65})$

$(10^{-3.65})^2 / 10^{-3.17} = C - 10^{-3.65}$

$C = 10^{-3.65} + (10^{-3.65})^2 / 10^{-3.17} = 2.9800 \times 10^{-4}$ M

Dilution: $c_1V_1 = c_2V_2$, $c_2 = 3.2093 \times 10^{-5}$ M

pH of diluted solution:

Concentration (M)	HF	H^+	F^-
Initial	3.2093×10^{-5}	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$3.2093 \times 10^{-5} - x$	x	x

$\frac{x^2}{3.2093 \times 10^{-5} - x} = 10^{-3.17}$

$x^2 + 10^{-3.17}x - (3.2093 \times 10^{-5} \times 10^{-3.17}) = 0$

$x = 3.0699 \times 10^{-5} = [H^+]$

pH = 4.51 (2 s.f.)

(b) 21.50 mL of a solution of sodium fluoride at pH 8.47 was diluted to 150.0 mL. What is the pH of the new solution? **4**

pOH = 5.53

$\text{p}K_b = 14 - 3.17 = 10.83$

$F^-_{(aq)} + H_2O_{(l)} \rightleftharpoons HF_{(aq)} + OH^-_{(aq)}$

Work out original concentration:

Concentration (M)	**F^-**	**HF**	**OH^-**
Initial	C	0	0
Change	$-10^{-5.53}$	$+10^{-5.53}$	$+10^{-5.53}$
Equilibrium	$C - 10^{-5.53}$	$10^{-5.53}$	$10^{-5.53}$

$(10^{-5.53})^2 / (C - 10^{-5.53}) = 10^{-10.83}$

$C = 10^{-5.53} + (10^{-5.53})^2 / 10^{-10.83} = 0.58847$ M

Dilution: $c_1V_1 = c_2V_2$, $c_2 = 0.0844$ M

pH of diluted solution:

Concentration (M)	**F^-**	**HF**	**OH^-**
Initial	0.0844	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.0844 - x$	x	x

Assume $[F^-]_{eqm} = 0.0844$ M, since K_b is small and dissociation does not change $[F^-]_{initial}$ by much.

$\frac{x^2}{0.0844} = 10^{-10.83}$

$x = 1.11731 \times 10^{-6} = [OH^-]$

pH = 14 − pOH = 8.05 (2 s.f.)

Question 32

A buffer solution was produced by mixing together acetic acid and sodium acetate. The total concentration of acetic acid and acetate ion was 0.250 M and the pH of the buffer solution was measured to be 5.35. The pK_a of acetic acid is 4.74.

(a) Calculate the concentrations of acetic acid and acetate ion in the solution. **4**

$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$

Let the concentration of CH_3COOH at equilibrium be x

Then the concentration of CH_3COO^- is $0.25 - x$

$[H^+] = 10^{-5.35}$ M

$$\frac{10^{-5.35} \times (0.25 - x)}{x} = 10^{-4.74}$$

$x = 0.0492727$ M

$[CH_3COOH] = 0.049$ M (2 s.f.)

$[CH_3COO^-] = 0.25 - 0.0492727 = 0.20$ M (2 s.f.)

Note: Can also solve using the Henderson-Hasselbalch equation.

(b) Explain how this buffer system resists the change in pH when a small amount of acid is added. **3**

$CH_3COOH_{(aq)} \rightleftharpoons H^+_{(aq)} + CH_3COO^-_{(aq)}$

When additional H^+ is added, the equilibrium is disturbed. According to Le Chatelier's principle, the position of equilibrium shifts to the left to remove H^+ and minimise the disturbance. This keeps the concentration of H^+ stable, and since $pH = -\log_{10}[H^+]$ the pH is also stable.

Question 33

The human stomach has a pH range of 1-2 due to the presence of hydrochloric acid.

(a) What is the concentration range of hydrochloric acid in the stomach? 2

pH = $-\log_{10}[H_3O^+]$ and HCl is a strong acid.

Therefore, [HCl] = $[H_3O^+] = 10^{-pH}$

Therefore, the pH range = $10^{-1} - 10^{-2} = 0.1 - 0.01$ M

(b) Explain why the solution in the human stomach is not a buffer solution. 2

HCl is a strong acid, therefore it is completely ionised in water and does not exist in equilibrium with Cl^-. Therefore, an equilibrium cannot shift to minimise the disturbance from a change in H^+/OH^- concentration.

(c) Mylanta is a medication often taken to neutralise excess stomach acid. A tablet contains 200 mg of aluminium hydroxide and 200 mg of magnesium hydroxide. If a patient typically has 200 mL of stomach fluid at a pH =1.00, what will be the resulting pH after the patient consumes one tablet? 4

$n(OH^-)$ from $Mg(OH)_2$ = $2 \times n(Mg(OH)_2) = 2 \times (0.200 / 58.326) = 6.858 \times 10^{-3}$ mol

$n(OH^-)$ from $Al(OH)_3$ = $3 \times n(Mg(OH)_2) = 3 \times (0.200 / 78.004) = 7.692 \times 10^{-3}$ mol

$n(OH^-)$ total = 0.0145 mol

$[H^+] = 10^{-1.00} = 0.100$ M

$n(H^+)$ in 200 mL= $0.100 \times 0.200 = 0.0200$ mol

$$H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)}$$

$n(H^+)$ remaining = $0.200 - 0.0145 = 5.45 \times 10^{-3}$ mol

$[H^+]$ remaining = $5.45 \times 10^{-3} / 0.200 = 0.0273$ mol L^{-1}

pH = $-\log_{10}(0.0273) = 1.56$ (2 s.f.)

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Year 12
Chemistry
Acid/Base Reactions

Work Book 6
Titration calculations

MATRIX
EDUCATION

Part A: Multiple choice questions

Question 1

If 10 mL of 1.0 M HCl was required to neutralise a 20 mL NaOH solution, what was the concentration of NaOH in the solution?

(a) 0.50 M

(b) 1.5 M

(c) 2.0 M

(d) 2.5 M

n(HCl) = c × V = 1.0 × 0.01 = 0.01 mol = n(NaOH)

c(NaOH) = n / V = 0.01 / 0.02 = 0.5 M

Question 2

What mass of sodium carbonate completely neutralises 150.0 mL of 0.850 mol L^{-1} nitrous acid?

(a) 3.38 g

(b) 6.76 g

(c) 13.5 g

(d) 27.0 g

$2HNO_{2(aq)} + Na_2CO_{3(aq)} \longrightarrow 2NaNO_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

$n(HNO_2)$ = c × V = 0.85 × 0.15 = 0.1275 mol

$n(Na_2CO_3) = n(HNO_2)$ / 2 = 0.06375 mol

$m(Na_2CO_3)$ = n × MM = 0.06375 × 105.99 = 6.76 g

Question 3

Equal volumes of equimolar solutions of the following acids were titrated with the same base solution. Which one would require the greatest volume of base to reach the equivalence point?

(a) Sulfuric acid

(b) Citric acid

(c) Carbonic acid

(d) Nitric acid

Citric acid is triprotic thus would have the greatest number of ionisable protons.

Question 4 [2011 HSC Q18]

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 mol^{-1}
(b) 84.0 g mol^{-1}
(c) 122 g mol^{-1}
(d) 410 g mol^{-1}

X^- = monoprotic reaction

n(HCl) = c × V = 0.1 × 0.0244 = 0.00244 mol = n(NaX) in 20 mL

Therefore 100 mL = 0.0122 mol

mm(NaX) = m / n = 1 / 0.0122 = 81.96 g mol^{-1}

Question 5

In volumetric analysis, the properties of the reactants, as well as the nature of the reaction between them, will determine if a back titration is to be used.

Consider the following cases:

(i) The substance being analysed is volatile.
(ii) The substance being analysed is insoluble in water but is soluble in dilute acid.
(iii) The end point of the reaction is difficult to detect.

In which cases would a back titration be more suitable than a simple forward titration?

(a) (i) and (ii) only
(b) (i) and (iii) only
(c) (ii) and (iii) only
(d) (i), (ii) and (iii)

(i) faster to add excess reactant than to titrate, hence less analyte will evaporate

(ii) insolubility can be overcome with a back titration e.g., for the analysis of $CaCO_3$

(iii) If the endpoint is difficult to detect (e.g., a weak acid/weak base titration, covered in Lesson 8), a strong acid or base can be added in excess instead.

Question 6

A 25.00 mL aliquot of 0.200 mol L^{-1} nitric acid solution was titrated with potassium hydroxide.

A titre volume of 34.70 mL was required to reach equivalence point.

If in another experiment, 25.00 mL of 0.200 mol L^{-1} hydrocyanic acid solution was used instead of nitric acid, what volume of the same potassium hydroxide solution would be required to reach equivalence point?

(a) Less than 34.70 mL
(b) More than 34.70 mL
(c) 34.70 mL
(d) Unable to calculate as more information is required

Both acids are monoprotic, so both balanced equations have a 1:1 mole ratio of acid to KOH.

Question 7

25 mL of each of the following solutions is titrated with 0.100 M ammonia:

(i) 0.07 M HF (pH 2.0)
(ii) 0.058 M H_2SO_3 (pH 1.7)
(iii) 0.030 M H_2SO_4 (pH 1.5)
(iv) 0.10 M HCl (pH 1.0)

Which requires the greatest volume of base to reach the equivalence point?

(a) HF
(b) H_2SO_3
(c) H_2SO_4
(d) HCl

H_2SO_3 has the highest concentration of ionisable H^+

Question 8

What volume of 0.020 mol L^{-1} sodium hydroxide solution is required to neutralise 15 mL of 0.010 mol L^{-1} citric acid solution?

(a) 2.5 mL

(b) 22.5 mL

(c) 7.5 mL

(d) 9.0 mL

n(citric) = c × V = 0.01 × 0.015 = 1.5×10^{-4} mol

n(NaOH) = 3 × n(citric) since citric acid is triprotic

n(NaOH) = 4.5×10^{-4} mol

V(NaOH) = n / c = 22.5 mL

Question 9

A sample containing approximately 0.5 g sodium carbonate is to be analysed via a back titration. Which of the following would NOT be suitable as the excess reagent?

(a) 25.0 mL of 0.50 M nitric acid

(b) 50.0 mL of 0.12 M hydrochloric acid

(c) 100.0 mL of 0.050 M sulfuric acid

(d) All of the above are suitable

n(Na_2CO_3) = m / MM = 4.717×10^{-3} mol

$CO_3^{2-}{}_{(aq)} + 2H^+{}_{(aq)} \longrightarrow H_2O_{(l)} + CO_{2(g)}$

Therefore need at least 9.434×10^{-3} mol ionisable H^+

(a) n(H^+) = 0.0125 mol

(b) n(H^+) = 6×10^{-3} mol

(c) n(H^+) = 0.01 mol

Question 10 **[2024 HSC Q20]**

The concentration of ascorbic acid (MM =176.124 g mol^{-1}) in solution A was determined by titration.

- A 25.00 mL sample of solution A was titrated with potassium hydroxide solution.
- 50.00 mg of ascorbic acid was added to a second 25.00 mL sample of solution A, which was titrated in the same way.

Titration volumes for both titrations are given.

Solution	Titre (mL)
25.00 mL solution A	17.50
25.00 mL solution A + 50.00 mg of ascorbic acid	33.10

What is the concentration of ascorbic acid in solution A?

(a) 5.352×10^{-3} mol L^{-1}

(b) 6.004×10^{-3} mol L^{-1}

(c) 1.012×10^{-2} mol L^{-1}

(d) 1.274×10^{-2} mol L^{-1}

n(ascorbic acid) in 50.00 mg = 0.05000 / 176.124 = $2.838908951 \times 10^{-4}$ mol

Ascorbic and KOH react in a 1:1 ratio.

n(KOH) reacting with 50.00 mg of ascorbic acid = $2.838908951 \times 10^{-4}$ mol

V(KOH) required to react with 50.00 mg of ascorbic acid = 33.10 – 17.50 = 15.60 mL

c(KOH) = $2.838908951 \times 10^{-4}/0.01560$ = 0.0181981 mol L^{-1}

n(KOH) reacting with ascorbic acid in solution A = 0.0181981 × 0.01750 = 3.18467×10^{-4} mol

n(ascorbic) in solution A = 3.18467×10^{-4} mol

c(ascorbic acid) = = $3.18467 \times 10^{-4}/0.025$ = 0.01274 mol L^{-1}

Part B: Extended response questions

Question 11

Write balanced full formula equations for the following neutralisation reactions:

(a) Acetic acid + sodium hydroxide **1**

$CH_3COOH_{(aq)} + NaOH_{(aq)} \longrightarrow CH_3COONa_{(aq)} + H_2O_{(l)}$

(b) Sulfuric acid + potassium hydrogen carbonate **1**

$H_2SO_{4(aq)} + 2KHCO_{3(aq)} \longrightarrow K_2SO_{4(aq)} + 2H_2O_{(l)} + 2CO_{2(g)}$

(c) Phosphoric acid + calcium hydroxide **1**

$2H_3PO_{4(aq)} + 3Ca(OH)_{2(aq)} \longrightarrow Ca_3(PO_4)_{2(s)} + 6H_2O_{(l)}$

Question 12

(a) Lactic acid is a weak monoprotic acid. If 20.00 mL of 0.09924 M lithium hydroxide was titrated with lactic acid solution and the volume required for neutralisation was 15.26 mL, find the concentration of lactic acid used. **2**

$HA_{(aq)} + LiOH_{(aq)} \longrightarrow LiA_{(aq)} + H_2O_{(l)}$

$n(LiOH) = 0.09924 \times 0.02 = 1.9848 \times 10^{-3} = n(HA)$

$c(HA) = n / V = 0.1301$ M (4 s.f.)

(b) 25.0 mL of sulfuric acid required 24.3 mL of 0.200 mol L^{-1} potassium hydroxide for neutralisation. Determine the concentration of the sulfuric acid solution. **2**

$H_2SO_{4(aq)} + 2KOH_{(aq)} \longrightarrow K_2SO_{4(aq)} + 2H_2O_{(l)}$

$n(KOH) = 0.2 \times 0.0243 = 4.86 \times 10^{-3}$

$n(H_2SO_4) = 2.43 \times 10^{-3}$

$c(H_2SO_4) = n / V = 0.0972$ M (3 s.f.)

(c) A 40.0 mL aliquot of ammonia solution was titrated with 0.575 M hydrobromic acid. The average titre was 14.85 mL. Calculate the concentration of ammonia in the original solution in mol L^{-1} and g L^{-1}. 3

$HBr_{(aq)} + NH_{3(aq)} \longrightarrow NH_4Br_{(aq)}$

n(HBr) = 0.575 × 0.01485 = 8.53875×10^{-3} = n(NH_3)

c(NH_3) = n / V = 0.21347 M = 0.213 M (3 s.f.)

In 1 L, there are 0.21347 moles NH_3

m(NH_3) in 1 L = n × MM = 0.21347 × (14.01 + 1.008 × 3) = 3.64 g L^{-1} (3 s.f.)

(d) A 0.843 g sample containing potassium carbonate was dissolved in 50.0 mL of distilled water. The resulting solution required 31.40 mL of 0.278 M hydrochloric acid for neutralisation. Calculate the mass percentage of potassium carbonate in the original sample. 3

$2HCl_{(aq)} + K_2CO_{3(aq)} \longrightarrow 2KCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

n(HCl) = 0.278 × 0.0314 = 8.7292×10^{-3}

n(K_2CO_3) = 4.3646×10^{-3}

m(K_2CO_3) = 0.60323 g

% = 0.60323 / 0.843 × 100 = 71.6% by mass (3 s.f.)

(e) 5.00 mL of concentrated acetic acid solution was diluted to 100.0 mL. 25.00 mL aliquots of the diluted solution were titrated against 0.1450 M potassium hydroxide solution. If the average titre was 31.75 mL, what was the concentration of the original acetic acid solution? 3

Solve the titration:

$CH_3COOH_{(aq)} + KOH_{(aq)} \longrightarrow CH_3COOK_{(aq)} + H_2O_{(l)}$

$n(KOH) = 0.145 \times 0.03175 = 4.60375 \times 10^{-3}$ mol $= n(CH_3COOH)$

$c(CH_3COOH) = n / V = 4.60375 \times 10^{-3} / 0.025 = 0.18415$ M

Solve the dilution:

$c_1V_1 = c_2V_2$

$c_1 = 0.18415 \times 0.1 / 0.005 = 3.683$ M $= 3.68$ M (3 s.f.)

(f) 50.0 mL of 0.235 M hydrochloric acid was added to a sample of ammonia. The excess hydrochloric acid required 24.65 mL of 0.114 M sodium hydroxide for neutralisation. How many moles of ammonia were originally present? 4

Reaction 1: $HCl_{(aq)} + NH_{3(aq)} \longrightarrow NH_4Cl_{(aq)}$

Reaction 2 (titration): $HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

$n(HCl)_{total} = 0.235 \times 0.05 = 0.01175$ mol

In Reaction 2 (titration):

$n(NaOH) = 0.114 \times 0.02465 = 2.8101 \times 10^{-3} = n(HCl)_{Rxn\ 2}$

$n(HCl)_{Rxn\ 1} = n(HCl)_{total} - n(HCl)_{Rxn\ 2} = 0.01175 - 2.8101 \times 10^{-3} = 8.9399 \times 10^{-3}$ mol

$n(NH_3) = 8.9399 \times 10^{-3} = 8.94 \times 10^{-3}$ mol (3 s.f.)

Question 13

A toilet cleaning agent contains hydrochloric acid. To determine the exact concentration of acid in the product, a student first took a 10.0 mL aliquot of the toilet cleaning agent and diluted it to 250.0 mL in a volumetric flask. 25.0 mL aliquots of the diluted solution were then titrated against 0.0602 mol L^{-1} sodium carbonate solution. The experiment was repeated several times. The results are shown below:

Run number	Volume (mL)
1	24.2
2	22.1
3	22.2
4	22.5

Calculate the concentration (% w/w) of hydrochloric acid in the toilet cleaning agent (density of cleaning agent = 1.2 g/mL). **4**

Solve titration first:

$2HCl_{(aq)} + Na_2CO_{3(aq)} \longrightarrow 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Average V(Na_2CO_3) = 22.2667 mL (24.2 mL is an outlier)

n(Na_2CO_3) = c × V = 0.0602 × 0.0222667 = $1.34045333 \times 10^{-3}$ mol

n(HCl) = $1.34045333 \times 10^{-3} \times 2 = 2.68090666 \times 10^{-3}$ mol

c(HCl) = n / V = $2.68090666 \times 10^{-3}$ / 0.025 = 0.1072362667 M

Solve dilution:

c_1 = 0. 1072362667 × 0.25 / 0.01 = 2.6809 mol L^{-1}

In 1 L:

Mass of total = V × d = 1200 g

Mass of solute = 2.6809 × (1.008 + 35.45) = 97.74 g

%w/w = 97.74 / 1200 × 100 = 8.15% w/w (3 s.f.)

Question 14

A particular brand of antacid tablets uses aluminium hydroxide as the sole active ingredient. Since aluminium hydroxide is insoluble at neutral pH, it is most easily measured via back titration. The following procedure was used to analyse the tablets:

5. An antacid tablet was crushed in a conical flask
6. 25.0 mL of 0.2011 mol L^{-1} sulfuric acid was added
7. The mixture was gently heated on a steam bath for 5 minutes then cooled
8. Bromothymol blue indicator was added to the flask
9. Sodium hydroxide (0.125 mol L^{-1}) was required to reach the equivalence point.
10. The procedure was repeated 3 times.

The average titre was 18.45 mL. The average antacid tablet weighed 1.33 g.

(a) Write a balanced chemical equation for the reaction between the antacid tablet and the sulfuric acid. **1**

$2Al(OH)_{3(s)} + 3H_2SO_{4(aq)} \longrightarrow Al_2(SO_4)_{3(aq)} + 6H_2O_{(l)}$

(b) How many moles of sulfuric acid were added to each tablet? **1**

$n(H_2SO_4) = c \times V = 0.2011 \times 0.0250 = 5.0275 \times 10^{-3} = 5.03 \times 10^{-3}$ mol (3 s.f.)

(c) Calculate the average mass of aluminium hydroxide in each tablet. **3**

Reaction 1: $2Al(OH)_{3(s)} + 3H_2SO_{4(aq)} \longrightarrow Al_2(SO_4)_{3(aq)} + 6H_2O_{(l)}$

Reaction 2: $H_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

$n(NaOH) = c \times V = 0.125 \times 0.01845 = 2.3063 \times 10^{-3}$ mol

$n(H_2SO_4)_{Rxn\ 2} = 2.3063 \times 10^{-3} / 2 = 1.1531 \times 10^{-3}$ mol

$n(H_2SO_4)_{Rxn\ 1} = n(H_2SO_4)_{total} - n(H_2SO_4)_{Rxn\ 2} = 5.0275 \times 10^{-3} - 1.1531 \times 10^{-3} = 3.8743 \times 10^{-3}$ mol

$n(Al(OH)_3) = 3.\ 874375 \times 10^{-3} \times 2/3 = 2.58291666 \times 10^{-3}$ mol

$m(Al(OH)_3) = n \times MM = 2.58291 \times 10^{-3} \times (26.98 + 16 \times 3 + 1.008 \times 3) = 0.201$ g (3 s.f.)

(d) State an appropriate safety precaution for this procedure. **1**

Wear safety goggles, gloves, lab coat to protect against corrosive acid splashes

Question 15

25.0 mL of a solution of sulfuric acid was titrated with 0.1088 M sodium hydroxide, and the average titre was found to be 23.42 mL. The bottle of sulfuric acid was accidentally left in a sunny spot with the cap off for two weeks, during which time some of the water evaporated. The titration was performed again with the 0.1088 M NaOH solution, and the new titre was found to be 26.91 mL. If there was originally 318.0 mL of sulfuric acid left in the bottle after the first titration, what volume of water evaporated before the second titration? 5

$H_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

Titration 1:

$n(NaOH) = c \times V = 0.1088 \times 0.02342 = 2.548096 \times 10^{-3}$ mol

$n(H_2SO_4) = n(NaOH) / 2 = 2.548096 \times 10^{-3} / 2 = 1.27405 \times 10^{-3}$ mol

$c(H_2SO_4) = 1.27405 \times 10^{-3} / 0.025 = 0.0509619$ M

Titration 2:

$n(NaOH) = c \times V = 0.1088 \times 0.02691 = 2.92781 \times 10^{-3}$ mol

$n(H_2SO_4) = n(NaOH) / 2 = 1.463905 \times 10^{-3}$ mol

$c(H_2SO_4) = 1.463905 \times 10^{-3} / 0.025 = 0.0585562$ M

$c_1V_1 = c_2V_2$

$V_2 = c_1V_1 / c_2 = 0.0509619 \times 0.318 / 0.0585562 = 0.276758 = 276.76$ mL

Volume evaporated $= V_1 - V_2 = 318 - 276.76 = 41.2$ mL (3 sig. fig.)

Question 16

Explain why acid should always be added to water during a dilution, and not the other way around. 3

Diluting concentrated acid is very exothermic. If acid is added to water, the solution formed is more dilute, leading to less hazardous splashes. The heat released is also absorbed better due to water's high heat capacity.

Question 17

To determine the amount of carbon dioxide produced by decomposing leaves, a student collected the gas produced by 4.0 g of crushed leaves over 4 days and absorbed it into 250.0 mL of 0.201 M NaOH. The following reaction occurred:

$$2NaOH_{(aq)} + CO_{2(g)} \longrightarrow Na_2CO_{3(aq)} + H_2O_{(l)}$$

Excess barium chloride was added to precipitate out the carbonate ions, and the excess NaOH was determined by titration with 0.128 M benzoic acid, a monoprotic acid. The average titre was determined to be 28.4 mL.

Calculate the amount of CO_2 produced as g CO_2 per g leaves. **4**

Reaction 1: $2NaOH_{(aq)} + CO_{2(g)} \longrightarrow Na_2CO_{3(aq)} + H_2O_{(l)}$

Reaction 2: NaOH + benzoic acid ⟶ sodium benzoate + water

(We do not know the formula of benzoic acid, but since we are told that it is monoprotic, we know that it reacts with NaOH in a 1:1 ratio)

$n(NaOH)_{total} = n(NaOH)_{Rxn\ 1} + n(NaOH)_{Rxn\ 2}$

$n(NaOH)_{total} = c \times V = 0.201 \times 0.25 = 0.05025$ mol

For Reaction 2 (Titration):

$n(\text{benzoic acid}) = c \times V = 0.128 \times 0.0284 = 3.6352 \times 10^{-3}$ mol $= n(NaOH)_{Rxn\ 2}$

$n(NaOH)_{Rxn\ 1} = n(NaOH)_{total} - n(NaOH)_{Rxn\ 2} = 0.0466148$ mol

$n(CO_2) = n(NaOH)_{Rxn\ 1} / 2 = 0.0233074$ mol

$m(CO_2) = n \times MM = 0.0233074 \times 44.01 = 1.02576$ g

g CO_2 per g leaves = 1.02576 / 4 = 0.26 g CO_2 per g leaves (2 s.f.)

Question 18 **[2020 HSC Q25]**

Citric acid reacts with sodium hydroxide according to the following chemical equation:

$$C_6H_8O_{7(aq)} + 3NaOH_{(aq)} \longrightarrow Na_3C_6H_5O_{7(aq)} + 3H_2O_{(l)}$$

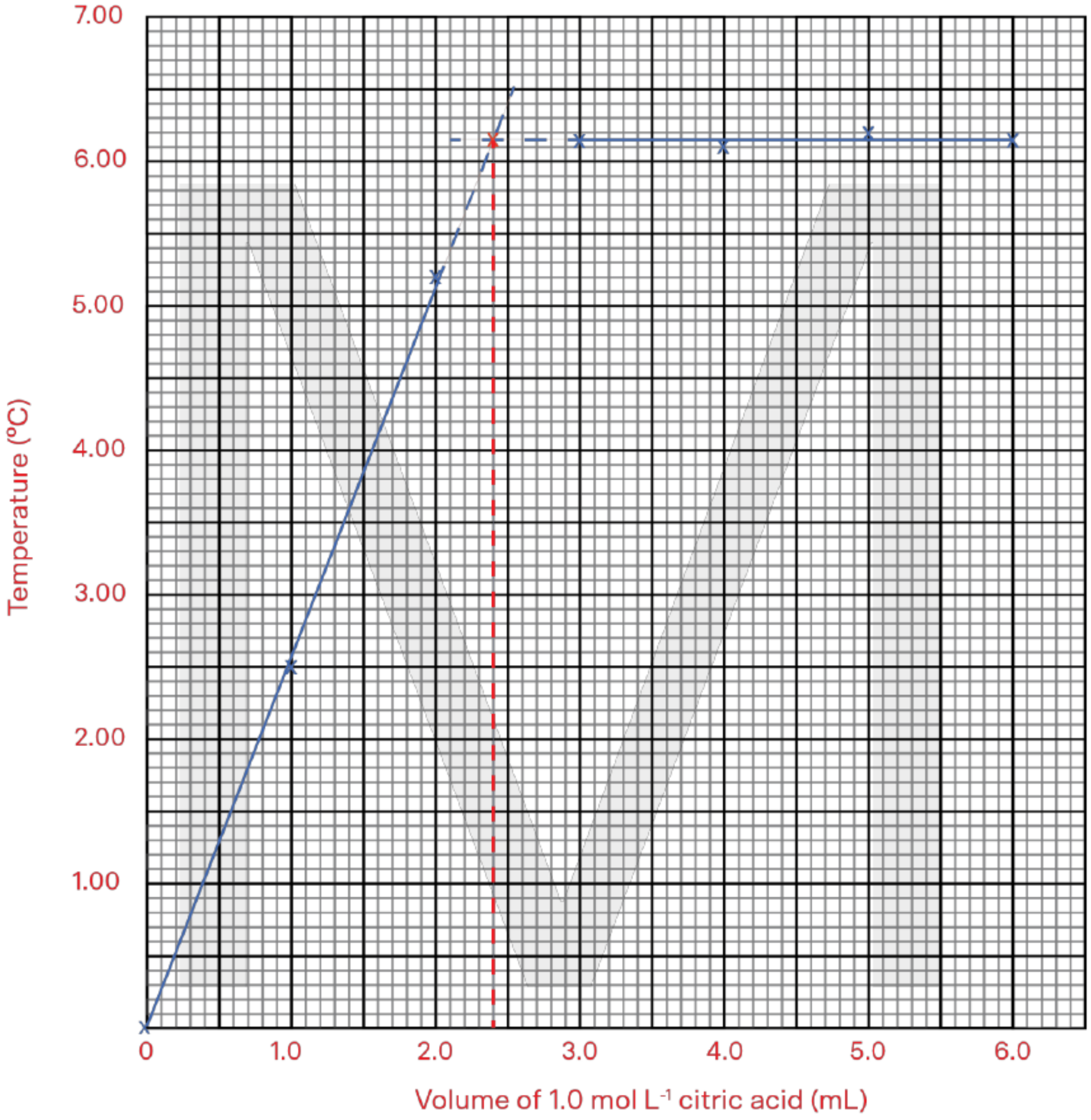
Various volumes of 1.0 mol L^{-1} citric acid solution were mixed with 8.0 mL of a sodium hydroxide solution of unknown concentration and sufficient deionised water added to make the total volume of the resulting solution 14.0 mL. The change in temperature of each solution was measured.

The data are given in the table.

Volume of 1.0 mol L^{-1} citric acid$_{(aq)}$ (mL)	Temperature increase (°C)
0.0	0.00
1.0	2.50
2.0	5.20
3.0	6.15
4.0	6.10
5.0	6.20
6.0	6.15

By graphing the data in the table and performing relevant calculations, determine the concentration of the sodium hydroxide solution. 7

Solution temperature vs volume of citric acid added to a solution of $NaOH_{(aq)}$



n(citric acid) = 1.0 mol L^{-1} × 0.0024 L = 0.0024 mol

n(NaOH) = 3 × n(citric acid) = 3 × 0.0024 mol = 0.0072 mol in 8.0 mL

c(NaOH) = 0.0072 / 0.0080 = 0.90 mol L^{-1} (2 s.f.)

Part C: Additional practice questions

The questions in this part are intended to be used for revision purposes. However, from time to time your teacher may prescribe some of these questions for homework.

Question 19

Which of the following describes the correct procedure for diluting concentrated acids?

(a) Slowly add the concentrated acid to water.

(b) Slowly add water to the concentrated acid.

(c) Slowly add an aqueous weak base to the concentrated acid.

(d) Slowly add an aqueous strong base to the concentrated acid.

Acid should be added carefully to water to avoid spitting as acid dissolution is very exothermic.

Question 20

Calculate the concentration of the solution formed when 4.3 g of Na_2CO_3 is dissolved to make 100 mL of solution.

(a) 0.31 M

(b) 0.41 M

(c) 0.51 M

(d) 0.61 M

$n = m / MM = 4.3 / 105.99 = 0.0405699$ mol, $c = n / V = 0.0405699 / 0.1 = 0.41$ M

Question 21

In the reaction of an acid with a strong base, the common ionic reaction involves ions of:

(a) hydronium and hydroxide

(b) sodium and chloride

(c) hydrogen and hydronium

(d) hydroxide and chloride

Question 22

Compared to the original temperature, the reaction mixture after titrating hydrochloric acid against sodium hydroxide will be:

(a) the same temperature

(b) slightly warmer

(c) slightly cooler

(d) significantly cooler

Strong acid + strong base reactions are exothermic.

Question 23

Which solution requires the largest volume of 0.1 M HCl to neutralise?

(a) 50 mL of 0.1 M $NaHCO_3$

(b) 5 mL of 0.5 M $NaHCO_3$

(c) 50 mL of 0.1 M Na_2CO_3

(d) 5 mL of 0.5 M Na_2CO_3

Na_2CO_3 requires two H^+ to neutralise, as opposed to $NaHCO_3$ which only requires one H^+.

50 mL of 0.1 M = 0.005 mol, 5 mL of 0.5 M = 0.0025 mol.

Question 24

Back titrations are recommended when:

(a) one of the reactants is volatile

(b) the acid or base does not dissolve in water

(c) direct titration would have an endpoint that is difficult to determine.

(d) all of the above.

Question 25

Calculate the mass of calcium hydroxide required to make 100 mL of 0.0050 mol L^{-1} solution.

(a) 0.012 g

(b) 0.029 g

(c) 0.037 g

(d) 0.041 g

$n(Ca(OH)_2) = c \times V = 0.0005$ mol

$m(Ca(OH)_2) = n \times MM = 0.0005 \times (40.08 + 16 \times 2 + 1.008 \times 2) = 0.037$ g

Question 26

Which of the following would require the greatest volume of a NaOH solution for neutralisation?

(a) 20.0 mL of 0.108 M nitric acid

(b) 30.0 mL of 0.050 M sulfuric acid

(c) 0.13 g of acetic acid

(d) 0.095 g of carbonic acid

(a) n(H^+) = 2.16×10^{-3} mol

(b) n(H^+) = 3×10^{-3} mol

(c) n(H^+) = 2.16×10^{-3} mol

(d) n(H^+) = 3.06×10^{-3} mol

Question 27

The solubility of thermally unstable salts such as ammonium chloride may be determined by back titration.

In one experiment a 5.00 mL saturated solution of ammonium chloride, NH_4Cl, at 20 °C, was diluted with distilled water to 250.0 mL in a volumetric flask.

A 20.0 mL aliquot of this solution was added to 10.0 mL of 0.400 M NaOH solution. The solution was heated to drive off the ammonia formed by this reaction.

When the flask had cooled, the excess hydroxide ions were neutralised by 14.7 mL of 0.125 M HCl solution. The molar mass of ammonium chloride is 53.5 g mol^{-1}.

(a) Write an equation for the neutralisation reaction. 1

$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

(b) Determine the moles of NaOH originally added to the ammonium chloride solution. 1

$n(NaOH)_{total} = c \times V = 0.4 \times 0.01 = 4.00 \times 10^{-3}$ mol (3 sig. fig.)

(c) Determine the moles of ammonium chloride in the 20.0 mL aliquot. 2

Titration:

$n(HCl) = c \times V = 0.125 \times 0.0147 = 1.8375 \times 10^{-3}\ mol = n(NaOH)_{excess}$

Original reaction:

$NH_4Cl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + NH_{3(g)} + H_2O_{(l)}$

$n(NaOH)_{reaction} = n(NaOH)_{total} - n(NaOH)_{excess} = (4 - 1.8375) \times 10^{-3} = 2.1625 \times 10^{-3}$ mol

$n(NH_4Cl)$ in aliquot $= n(NaOH)_{reaction} = 2.1625 \times 10^{-3}\ mol = 2.16 \times 10^{-3}$ mol (3 sig. fig.)

(d) Calculate the moles of ammonium chloride in 5.00 mL of the saturated solution. 2

$c(NH_4Cl) = n / V = 2.1625 \times 10^{-3} / 0.02 = 0.108125$ M

Dilution:

$c_1V_1 = c_2V_2$

$c_1 = c_2V_2 / V_1 = 0.108125 \times 0.25 / 0.005 = 5.40625$ M

$n(NH_4Cl) = c \times V = 5.40625 \times 0.005 = 0.02703\ mol = 0.0270$ mol (3 sig. fig.)

(e) Calculate the solubility of ammonium chloride in water at 20 °C, in g L^{-1}. 2

0.02703 mol in 5 mL = 5.40625 mol in 1000 mL

$m(NH_4Cl) = n \times MM = 5.40625 \times (14.01 + 1.008 \times 4 + 35.45) = 289.19$ g

Solubility = 289 g L^{-1} (3 sig. fig.)

Question 28

(a) If 20.00 mL of 0.0548 M calcium hydroxide required 14.75 mL acetic acid to reach the equivalence point, calculate the concentration of the acetic acid solution. 2

$2CH_3COOH_{(aq)} + Ca(OH)_{2(aq)} \longrightarrow Ca(CH_3COO)_{2(aq)} + 2H_2O_{(l)}$

$n(Ca(OH)_2) = 1.096 \times 10^{-3}$

$n(CH_3COOH) = 2.912 \times 10^{-3}$

$c(CH_3COOH) = 0.14861 = 0.149$ M (3 s.f.)

(b) What is the concentration of a sodium hydrogen carbonate solution if 30.00 mL is required to neutralise 29.10 mL of 0.2500 M nitric acid? Give your answer in mol L^{-1} and g mL^{-1}. 3

$HNO_{3(aq)} + NaHCO_{3(aq)} \longrightarrow NaNO_{3(aq)} + H_2O_{(l)} + CO_{2(g)}$

$n(HNO_3) = 7.275 \times 10^{-3} = n(NaHCO_3)$

$c(NaHCO_3) = 0.2425$ M (4 s.f.)

In 1 mL, 0.2425×0.001 mol $NaHCO_3$ present

$m(NaHCO_3) = n \times MM = 0.02037$ g mL^{-1} (4 s.f.)

(c) 3.50 mL of concentrated hydrochloric acid was diluted to 50.0 mL. 15.00 mL aliquots of the diluted solution were titrated against 0.450 M potassium hydroxide solution. If the average titre was 12.55 mL, what was the concentration of the original acid solution as %w/w? The density of the original acid solution was 1.12 g mL^{-1}. 3

Titration:

$HCl_{(aq)} + KOH_{(aq)} \longrightarrow KCl_{(aq)} + H_2O_{(l)}$

$n(KOH) = 5.6475 \times 10^{-3} = n(HCl)$

$c(HCl) = 0.3765$ M

Dilution:

$c_1 = c_2V_2 / V_1 = 5.37857 \text{ mol L}^{-1}$

$\%w/w = m(HCl) / m(solution) \times 100$

m(HCl) In 1 L = $n \times MM = 5.37857 \times (1.008 + 35.45) = 196.092$ g

$m(solution) = V \times d = 1120$ g

%w/w = 17.5% (3 s.f.)

Question 29

A student wants to determine the concentration of ammonia in a bottle of household cleaner. She accurately dilutes a 25.00 mL aliquot of the cleaner to 250.0 mL, then titrates 25.00 mL aliquots of the diluted solution against 0.2240 M HCl. The average titre obtained is 21.85 mL. Calculate the concentration of NH_3 in the cloudy ammonia as %w/w if the density of the original solution is 0.950 g mL^{-1}. **4**

$HCl_{(aq)} + NH_{3(aq)} \longrightarrow NH_4Cl_{(aq)}$

n(HCl) = c × V = 0.2240 × 0.02185 = 4.8944×10^{-3} mol

$n(NH_3)$ = n(HCl) = 4.8944×10^{-3} mol

$c(NH_3)$ = n / V = 4.8944×10^{-3} / 0.025 = 0.195776 M

Dilution:

$c_1 = c_2V_2 / V_1$ = 0.195776 × 0.25 / 0.025 = 1.95776 M

$m(NH_3)$ in 1 L = n × MM = 1.95776 × (14.01 + 1.008 × 3) = 33.348 g

mass of 1 L = 0.950 × 1000 = 950 g

% w/w = 33.348 / 950 × 100 = 3.51% w/w (3 sig. fig.)

Question 30

Lemon juice contains citric acid, a triprotic acid that reacts with sodium hydroxide to form sodium citrate ($Na_3C_6H_5O_7$).

A chemist decides to measure the citric acid content of lemon juice. First, he used 0.0997 mol L^{-1} hydrochloric acid to standardise a sodium hydroxide solution. The concentration of sodium hydroxide was determined to be 0.1006 mol L^{-1}.

A 50.0 mL sample of lemon juice was diluted to 250 mL, and 25 mL aliquots of the diluted juice were titrated against sodium hydroxide. The average titre was 34.75 mL. The density of the lemon juice was 1.00 g mL^{-1}.

(a) Write an equation for the neutralisation reaction between sodium hydroxide and citric acid. **1**

$C_6H_8O_{7(aq)} + 3NaOH_{(aq)} \longrightarrow Na_3C_6H_5O_{7(aq)} + 3H_2O_{(l)}$

(b) Calculate the concentration of citric acid in the undiluted lemon juice (%w/w). **4**

Solve titration first:

$C_6H_8O_{7(aq)} + 3NaOH_{(aq)} \longrightarrow Na_3C_6H_5O_{7(aq)} + 3H_2O_{(l)}$

$n(NaOH) = c \times V = 0.1006 \times 0.03475 = 3.49585 \times 10^{-3}$ mol

$n(C_6H_8O_7) = 3.49585 \times 10^{-3} \times 3 = 0.00116528$ mol

$c(C_6H_8O_7) = n / V = 0.00116528 / 0.025 = 0.0466113$ M

Solve dilution:

$c_1V_1 = c_2V_2$

$c_1 = 0.0466113 \times 0.25 / 0.05 = 0.2330566$ M

%w/w = mass of solute / mass of total × 100

$= 0.2330566 \times (12.01 \times 6 + 1.008 \times 8 + 16 \times 7) / 1000 \times 100 = 4.5\%$ w/w (2 s.f.)

Question 31

A student wished to determine the percentage of calcium carbonate present in a shell found at the beach. The clean dry shell, which weighed 1.306 g, was placed in a beaker and 10.0 mL of 5.00 M hydrochloric acid was added. When the shell had completely dissolved, the resulting solution was transferred to a volumetric flask and the volume made up to 25.0 mL with distilled water. A 10.0 mL sample from this solution required 11.2 mL of 1.00 M sodium hydroxide for complete neutralisation.

(a) Write an equation for the reaction of calcium carbonate with hydrochloric acid. **1**

$CaCO_{3(s)} + 2HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

(b) How many moles of acid remained in the beaker after the reaction with the shell (before the dilution was made)? **2**

n(NaOH) = c × V = 1 × 0.0112 = 0.0112 mol = n(HCl) in 10 mL

n(HCl) in 25 mL = 0.0112 × 0.025 / 0.010 = 0.028 mol = $n(HCl)_{excess}$

(c) What was the percentage of calcium carbonate in the shell? **4**

$n(HCl)_{total}$ = c × V = 5 × 0.01 = 0.05 mol

$n(HCl)_{reacted} = n(HCl)_{total} - n(HCl)_{excess}$ = 0.05 − 0.028 = 0.022 mol

$n(CaCO_3)$ = n(HCl) / 2 = 0.011 mol (from mole ratio)

$m(CaCO_3)$ = n × MM = 1.10099 g

% w/w = component / total × 100 = 1.10099 / 1.306 × 100 = 84.3% (3 sig. fig.)

Question 32

Drain cleaners typically contain up to 50% w/w sodium hydroxide as the active ingredient to effectively break up blocked drains. A quality assurance chemist was tasked with investigating the concentration of the active ingredient in 2 separate brands of drain cleaners.
1.00 mL of each drain cleaner brand was diluted to 250 mL in a volumetric flask. Four 25.00 mL aliquots of each diluted drain cleaner were titrated against a standardised 0.1021 mol L^{-1} hydrochloric acid using phenolphthalein as the indicator. The following results were obtained.

Titration number	Volume of HCl added (mL) to drain cleaner X	Volume of HCl added (mL) to drain cleaner Y
1	10.1	15.2
2	11.2	13.4
3	11.3	13.1
4	11.0	13.5

Calculate the active ingredient concentration in %w/w for each brand of cleaner (density of liquid drain cleaner = 1.1 g/mL). **4**

Titration: $HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

For drain cleaner X: Average titre = (11.2 + 11.3 + 11.0) / 3 = 11.1667 mL (10.1 mL is an outlier)

n(HCl) = c × V = 0.1021 × 0.0111667 = 1.14012 × 10^{-3} mol = n(NaOH)

c(NaOH) = n / V = 1.14012 × 10^{-3} / 0.025 = 0.045605 mol L^{-1}

$c_1 = c_2V_2 / V_1$ = 0. 045605 × 0.25 / 0.001 = 11.4012 mol L^{-1}

m(NaOH) in 1 L = n × MM = 11.4012 × (22.99 + 16 + 1.008) = 456.02 g

m(cleaner) = 1000 × 1.1 = 1100 g

Concentration = (456.02 g / 1100 g) × 100 = 41.5% w/w

For drain cleaner Y, average titre = (13.4 + 13.1 + 13.5) / 3 = 13.3333 mL

Therefore concentration = (41.4545 / 11.1667) × 13.3333 = 49.5% w/w

Year 12
Chemistry
Acid/Base Reactions

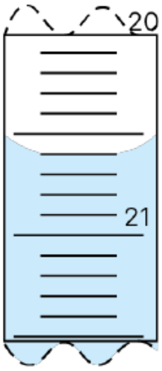
Work Book 7
Titration technique

MATRIX
EDUCATION

Part A: Multiple choice questions

Question 1

The diagram below shows the level of sodium hydroxide solution in a burette at the endpoint of a titration with acetic acid. The starting level of the base in the burette was 0.60 mL.



Which statement is correct?

(a) The base volume used was 20.0 mL and at endpoint pH > 7.

(b) The base volume used was 20.8 mL and at endpoint pH > 7.

(c) The base volume used was 22.0 mL and at endpoint pH > 7.

(d) The base volume used was 19.6 mL and at endpoint pH < 7.

Burette markings start at the top at 0.0 mL. The bottom of the meniscus is at 20.6 mL, therefore the volume used is 20.0 mL. This is a weak acid + strong base titration, hence the salt formed (sodium acetate) is basic, and the endpoint pH should be > 7.

Question 2

Which one of the following is NOT a feature of a good primary standard?

(a) Available in solid form

(b) Have a large molecular mass

(c) Ionises completely in water

(d) Soluble in water

Primary standards are ideally available in solid form (for ease of weighing), have a large molecular mass (for accuracy), and soluble in water (to make an aqueous solution). Complete ionisation is not important.

Question 3

Which of the following should be avoided during a titration?

(a) Rinse all apparatus with distilled water and use immediately.

(b) When filling the burette, use a tissue to wipe off any solution that might have spilled onto the outside of the burette.

(c) Any solution that runs down the inside wall of the conical flask should be washed into the mixture with distilled water from a wash bottle.

(d) All of the above.

Only the conical flask should have distilled water as the final rinse. Both pipette and burette should be rinsed with their filling solutions.

Question 4

When using a burette, it is important to:

(a) rinse it with distilled water before filling

(b) ensure it is clamped vertically

(c) rinse it with the solution in the conical flask

(d) ensure that the last drops are drained by shaking it

The burette should be used vertically to ensure the correct readings are taken.

(a) and (c): It should be rinsed with the filling solution.

(d): If used correctly, the burette should not be empty.

Question 5

In a titration, a 25.00 mL titre of 1.00 M hydrochloric acid neutralised a 20.00 mL aliquot of sodium hydroxide solution. If, in repeating the titration, a student had rinsed one of the pieces of glassware as described below, which statement regarding the titre is correct?

(a) If water was left in the titration flask after final rinsing, the titre would be equal to 25.00 mL

(b) If the final rinsing of the burette is with water, the titre would be less than 25.00 mL

(c) If the final rinsing of the 20.00 mL pipette is with water the titre would be greater than 25.00 mL

(d) If the titration flask had been rinsed with the acid prior to the addition of the aliquot the titre would be greater than 25.00 mL

HCl is in the burette, NaOH is in the pipette/conical flask for this titration.

(b): water would dilute the acid, meaning a larger titre would be required

(c): water would dilute the base, meaning a smaller titre would be required

(d) acid would neutralise some base, meaning a smaller titre would be required.

Question 6

A titration between hydrochloric acid and potassium hydroxide is to be performed. If the acid is to be dispensed from the burette, which of the following indicates the best final rinsing solution for each piece of apparatus?

	Burette	Pipette	Conical flask
(a)	HCl	H_2O	KOH
(b)	H_2O	KOH	H_2O
(c)	H_2O	H_2O	KOH
(d)	HCl	KOH	H_2O

Burette and pipette should be rinsed with the filling solution, conical flask should be rinsed with water.

Question 7 [2013 HSC Q19]

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.

Substance added	Colour observed
0.1 mol L^{-1} $HCl_{(aq)}$	Bright pink
0.01 mol L^{-1} $HCl_{(aq)}$	Bright pink
0.001 mol L^{-1} $HCl_{(aq)}$	Pale yellow
Distilled water	Bright yellow
0.001 mol L^{-1} $NaOH_{(aq)}$	Bright yellow
0.01 mol L^{-1} $NaOH_{(aq)}$	Bright yellow

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

(a) $HCl_{(aq)} + NH_{3(aq)}$

(b) $HCl_{(aq)} + NaOH_{(aq)}$

(c) $CH_3COOH_{(aq)} + NH_{3(aq)}$

(d) $CH_3COOH_{(aq)} + NaOH_{(aq)}$

The indicator changes colour with 0.001 M HCl, i.e. pH 3. Therefore, this indicator can be used for a strong acid/weak base titration.

Question 8

In a titration experiment, a student placed a 25 mL aliquot of 0.50 M hydrochloric acid into the conical flask and 0.50 M sodium carbonate standard solution into the burette. Which of the following correctly identifies the solution that should be used to rinse the glassware immediately before use?

	Conical flask	Burette
(a)	0.50 M HCl	0.50 M Na_2CO_3
(b)	Distilled water	Distilled water
(c)	Distilled water	0.50 M Na_2CO_3
(d)	0.50 M HCl	Distilled water

Question 9

The volume of solution delivered by pipette is the:

(a) aliquot

(b) titre

(c) titrand

(d) analyte

Aliquot is an accurately known volume of liquid. The liquid transferred via pipette is of an accurately known volume.

Question 10

The following procedure was used by a student in the titration of a monoprotic acid with sodium hydroxide:

(i) The burette was rinsed with distilled water and then filled with a standardised NaOH solution.

(ii) The conical flask and pipette were rinsed with the acid solution.

(iii) An aliquot of acid solution was transferred into the conical flask using the pipette.

(iv) An indicator was added to the conical flask and the acid was titrated with the base until a permanent colour change was observed with the indicator.

Which statement is correct?

(a) The calculated concentration of the acid will be too high.

(b) The calculated concentration of the acid will be too low.

(c) The calculated concentration of the acid will be correct.

(d) No conclusion can be reached about the acid concentration.

Incorrect steps: rinsing burette with water only will dilute the base, rinsing conical with acid adds extra acid. Both will cause the calculated acid concentration to be too high.

Part B: Extended response questions

Question 11

Explain why the equivalence point for each neutralisation is acidic, basic or neutral and identify a suitable indicator to use to determine the equivalence point. Include balanced chemical equations in your answer.

(a) Sodium hydroxide + nitric acid 3

At the equivalence point, sodium nitrate is present.

$NaOH_{(aq)} + HNO_{3(aq)} \longrightarrow NaNO_{3(aq)} + H_2O_{(l)}$

$NaNO_3$ dissociates to give Na^+ and NO_3^-. Both Na^+ and NO_3^- ions do not react with water to an appreciable extent hence will result in a neutral solution. Bromothymol blue is a suitable indicator.

(b) Potassium hydroxide + ethanoic acid 3

At the equivalence point, potassium acetate is present.

$KOH_{(aq)} + CH_3COOH_{(aq)} \longrightarrow KCH_3COO_{(aq)} + H_2O_{(l)}$

KCH_3COO dissociates to give CH_3COO^- which is basic as it is the conjugate base of a weak acid. It reacts with water to produce $OH^-_{(aq)}$, therefore forms a basic solution.

$CH_3COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}$

Phenolphthalein is a suitable indicator.

(c) Ammonia + hydrobromic acid 3

At the equivalence point, ammonium bromide is present.

$NH_{3(aq)} + HBr_{(aq)} \longrightarrow NH_4Br_{(aq)}$

NH_4Br dissociates to give NH_4^+ which is acidic as it is the conjugate acid of a weak base. It reacts with water to produce $H_3O^+_{(aq)}$, therefore forms an acidic solution.

$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^+_{(aq)}$

Methyl orange is a suitable indicator.

(d) Sodium carbonate + hydrochloric acid 3

Neutralisation produces carbon dioxide: $Na_2CO_{3(aq)} + 2HCl_{(aq)} \longrightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

Although sodium chloride is a neutral salt, carbon dioxide reacts with water to produce carbonic acid, thus forming an acidic solution.

$$H_2O_{(l)} + CO_{2(g)} \rightleftharpoons H_2CO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HCO_3^-{}_{(aq)}$$

Methyl orange is a suitable indicator.

Question 12

25 mL of 0.10 mol L^{-1} hydrochloric acid is titrated against sodium hydroxide. Identify the final rinsing substance prior to setting up the titration for each piece of equipment:

(a) conical flask 1

distilled water

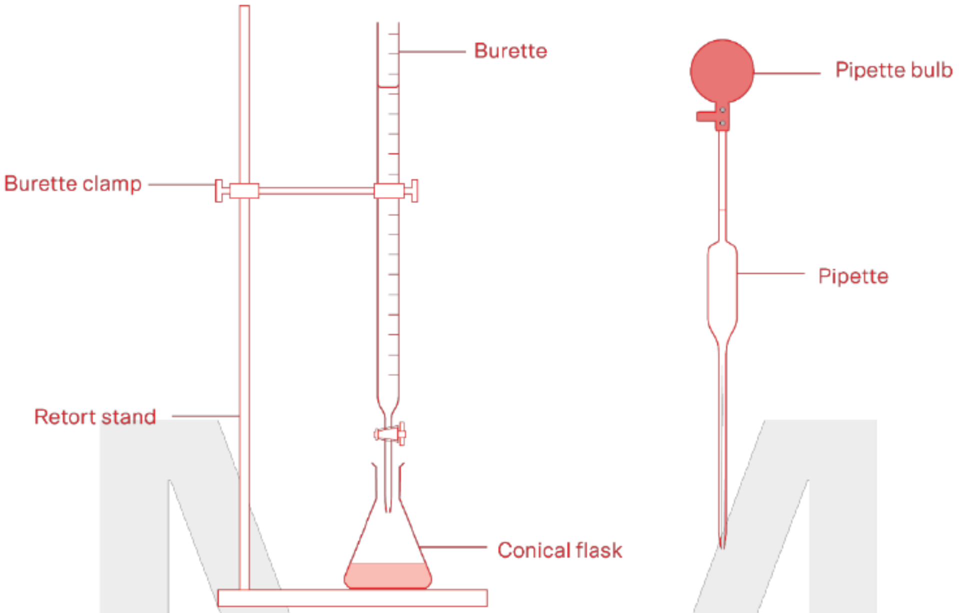
(b) pipette 1

0.1 M HCl

(c) burette 1

NaOH

(d) Draw a labelled diagram of the experimental set-up for this titration experiment. 3


Burette
Pipette bulb
Burette clamp
Pipette
Retort stand
Conical flask

Question 13

Explain the difference between endpoint and equivalence point. 2

Endpoint: the point in a titration when the indicator first permanently changes colour

Equivalence point: the point in a titration when stoichiometric amounts of reactants are present (when reaction is complete)

Question 14

Part of a titration procedure performed by a student is described below.

1. All titration equipment was washed with distilled water and used immediately.
2. The solution of unknown concentration was transferred to a conical flask using a pipette. The pipette was gently tapped against the inside wall of the conical flask to ensure complete transfer.
3. A few drops of the appropriate indicator solution were added to the contents of the flask and swirled gently to mix.
4. Before use, the burette was rinsed with the standardised titrant with the tap open.
5. The burette was clamped so that the zero mark was at eye level. The burette was also confirmed to be vertical by alignment.
6. The burette was filled with the titrant up to the 0 mL mark.
7. The flask containing the unknown solution was placed under the burette.

Suggest changes that would reduce experimental errors and ensure that the student's results were valid. Explain your answer. 3

The procedure contains some steps which reduced the validity of the results.

- The pipette should be rinsed with the filling solution as the final rinse, not water. Doing so dilutes the solution contained in the pipette to an unknown extent, and so the actual number of moles transferred will be lower than recorded.

- The pipette should have a drop of solution remaining in the tip after transfer. It should not be completely transferred. This adds extra moles to the conical flask.

The other steps in the process are correct.

Question 15 **[2022 HSC Q32]**

The concentration of citric acid, a triprotic acid, in a carbonated soft drink was to be determined.

Step 1: A solution of $NaOH_{(aq)}$ was standardised by titrating it against 25.00 mL aliquots of a solution of the monoprotic acid potassium hydrogen phthalate (KHP). The KHP solution was produced by dissolving 4.989 g in enough water to make 100.0 mL of solution. The molar mass of KHP is 204.22 g mol^{-1}.

The results of the standardisation titration are given in the table.

Titration	Volume NaOH (mL)
1	28.60
2	27.40
3	27.20
4	27.60

Step 2: A 75.00 mL bottle of the drink was opened and the contents quantitatively transferred to a beaker. The soft drink was gently heated to remove CO_2.

Step 3: The cooled drink was quantitatively transferred to a 250.0 mL volumetric flask and distilled water was added up to the mark.

Step 4: 25.00 mL samples of the solution were titrated with the $NaOH_{(aq)}$ solution. The average volume of $NaOH_{(aq)}$ used was 13.10 mL.

(a) Calculate the concentration of the triprotic citric acid in the soft drink. 6

For the standardisation of NaOH by titration with KHP:

n(KHP) = m / MM = 4.989 / 204.22 = 0.02442...mol

c(KHP) = n / V = 0.02442... / 0.1 = 0.2442...M

$n(KHP)_{titration}$ = c × V = 0.2442... × 0.025 = 0.006107...mol

$KHP_{(aq)} + NaOH_{(aq)} \longrightarrow NaKP_{(aq)} + H_2O_{(l)}$

n(NaOH) = n(KHP)titration = 0.006107...mol

$V_{average}(NaOH)$ = (0.0274 + 0.0272 + 0.0276) / 3 = 0.0274 L (trial 1 is an outlier)

c(NaOH) = n / V = 0.006107... / 0.0274 = 0.2228...M

For the titration of NaOH against citric acid (H_3A):

n(NaOH) = c × V = 0.2228... × 0.0131 = 0.002919...mol

$H_3A_{(aq)} + 3NaOH_{(aq)} \longrightarrow Na_3A_{(aq)} + 3H_2O_{(l)}$

n(H_3A) = 1/3 × n(NaOH) = 1/3 × 0.002919... = 0.0009733...mol

$c(H_3A)_{dilute}$ = n / V = 0.0009733... / 0.025 = 0.03893...M

$c(H_3A)_{original}$ = $((H_3A)_{dilute} \times V_{dilute}) / V_{original}$ = 0.03893... × 0.250 / 0.075 = 0.1298 M (4 sig. fig.)

(b) Explain how your answer to part (a) would be different if the carbon dioxide was not removed from the soft drink. 2

Carbon dioxide, $CO_{2(g)}$ is sparingly soluble in water. In solution, $CO_{2(aq)}$ can react with water to produce carbonic acid, $H_2CO_{3(aq)}$ according to the following equilibrium:

$CO_{2(g)} \rightleftharpoons CO_{2(aq)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)}$

$H_2CO_{3(aq)}$ can react with $NaOH_{(aq)}$ in a neutralisation reaction:

$H_2CO_{3(aq)} + 2NaOH_{(aq)} \longrightarrow Na_2CO_{3(aq)} + 2H_2O_{(l)}$

If $CO_{2(aq)}$ was not removed from the soft drink before titrating it against $NaOH_{(aq)}$, additional $NaOH_{(aq)}$ titrant would be required to neutralise the $CO_{2(aq)}$ as well as the citric acid, $H_3A_{(aq)}$. This would result in the calculation of a greater n(NaOH), a greater n(H_3A), a greater $c(H_3A)_{dilute}$, and finally a greater $c(H_3A)_{original}$ than in part (a).

Question 16 **[2016 HSC Q29]**

A solution of hydrochloric acid was standardised by titration against a sodium carbonate solution using the following procedure.

- All glassware was rinsed correctly to remove possible contaminants
- Hydrochloric acid was placed in the burette
- 25.0 mL of sodium carbonate solution was pipetted into the conical flask

The titration was performed and the hydrochloric acid was found to be 0.200 mol L^{-1}.

(a) Identify the substance used to rinse the conical flask and justify your answer. 2

Distilled water should be used to rinse the conical flask to ensure there are no moles of any reactants that have not been accounted for in the reaction.

(b) Seashells contain a mixture of carbonate compounds. The standardised hydrochloric acid was used to determine the percentage by mass of carbonate in a seashell using the following procedure.

- A 0.145 g sample of the seashell was placed in a conical flask.
- 50.0 mL of the standardised hydrochloric acid was added to the conical flask.
- At the completion of the reaction, the mixture in the conical flask was titrated with 0.250 mol L^{-1} sodium hydroxide.

The volume of sodium hydroxide used in the titration was 29.5 mL.

Calculate the percentage by mass of carbonate in the sample of the seashell. 4

Reaction 1: $2HCl_{(aq)} + CO_3{}^{2-}{}_{(aq)} \longrightarrow 2Cl^-{}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

Reaction 2 (titration): $HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

$n(NaOH) = c \times V = 0.25 \times 0.0295 = 7.375 \times 10^{-3}$ mol $= n(HCl)_{Rxn\ 2}$

$n(HCl)_{total} = c \times V = 0.2 \times 0.05 = 0.01$ mol

$\therefore n(HCl)_{Rxn\ 1} = n(HCl)_{total} - n(HCl)_{Rxn\ 2} = 0.01 - 7.375 \times 10^{-3} = 2.625 \times 10^{-3}$ mol

$n(CO_3{}^{2-}) = n(HCl)_{Rxn\ 1} / 2 = 1.3125 \times 10^{-3}$ mol

$m(CO_3{}^{2-}) = n \times MM = 1.3125 \times 10^{-3} \times (12.01 + 16 \times 3) = 0.078763$ g

$\%m/m = m(CO_3{}^{2-}) / m(shell) \times 100 = 54.3\%$ (3 s.f.)

Question 17

A student performed an experiment to determine the concentration of acetic acid in Supersour brand vinegar. 25.00 mL of Supersour brand vinegar was diluted accurately to 250.0 mL, then 20.00 mL of this diluted solution was titrated with 0.102 M sodium hydroxide. The average titre obtained was 24.85 mL.

(a) Determine the concentration of the diluted vinegar. **2**

$CH_3COOH_{(aq)} + NaOH_{(aq)} \longrightarrow NaCH_3COO_{(aq)} + H_2O_{(l)}$

$n(NaOH) = c \times V = 0.102 \times 0.02485 = 2.5347 \times 10^{-3}$ mol $= n(CH_3COOH)$

$c(CH_3COOH) = n / V = 2.5347 \times 10^{-3} / 0.020 = 0.127$ M (3 sig. fig.)

(b) Determine the concentration of acetic acid in the original vinegar in %w/v. **2**

$c_1V_1 = c_2V_2$

$c_1 = c_2V_2 / V_1 = 0.127 \times 0.25 / 0.025 = 1.27$ M

In 1 L (1000 mL), $m(CH_3COOH) = 1.27 \times (12.01 \times 2 + 1.008 \times 4 + 16 \times 2) = 76.106$ g

% w/V = 76.1069022 / 1000 × 100 = 7.61% (3 sig. fig.)

(c) The student accidentally rinsed the burette with water before filling it with sodium hydroxide. What would be the effect on the calculated concentration of the vinegar compared to the actual concentration? **1**

Rinsing would dilute the sodium hydroxide, hence a larger titre would be recorded. Hence the calculated concentration of vinegar would be higher than the actual concentration.

(d) Explain why the sodium hydroxide solution must be standardised before use in the titration, for the results to be accurate. **1**

Sodium hydroxide is hygroscopic, hence the concentration cannot be determined from the mass of solid NaOH used to make the solution, and the concentration needs to be determined by standardisation (usually titration against a standard solution)

(e) Suggest a suitable indicator for determining the endpoint of this titration. **1**

Weak acid / strong base = basic endpoint, so phenolphthalein would be suitable.

Question 18

A student wants to standardise a solution of sulfuric acid of unknown concentration. She prepared a 0.1040 M standard solution by dissolving pure sodium carbonate in deionised water and making it up to 500 mL.

(a) Explain why sodium hydroxide was not used as the primary standard. 1

Sodium hydroxide reacts with carbon dioxide and readily absorbs water vapour from the air, so it cannot be accurately weighed hence the concentration of the prepared solution will not be known.

(b) Determine the mass of sodium carbonate used to prepare the solution. 2

n(Na_2CO_3) = c × V = 0.1040 × 0.500 = 0.052 mol

m(Na_2CO_3) = n × MM = 0.052 × (22.99 × 2 + 12.01 + 16 × 3) = 5.51 g (3 s.f.)

(c) Outline the procedure the student should have used to prepare the solution. 3

1. Accurately weigh approximately 5.51 g Na_2CO_3 into a clean beaker.

2. Add distilled water to the Na_2CO_3 to dissolve. Carefully transfer it to a clean 500 mL volumetric flask. Rinse all of the equipment that came into contact with the weighed mass into the flask.

3. Add distilled water until the bottom of the meniscus is resting on the mark on the neck.

4. Stopper the flask and invert several times to ensure the solution is well mixed.

Part C: Additional practice questions

The questions in this part are intended to be used for revision purposes. However, from time to time your teacher may prescribe some of these questions for homework.

Question 19

An indicator changes colour in the pH range 8.3–10.0. This indicator should be used when titrating a known volume of:

(a) a strong acid with a weak base

(b) a weak acid with a weak base

(c) a weak base with a strong acid

(d) a weak acid with a strong base

The indicator colour change range should match the pH of the equivalence point. The salt formed from weak acid + strong base will be basic. Therefore, this indicator can be used for a weak acid/strong base titration.

Question 20

A student used a pipette to transfer a solution to a flask. After doing so, a small amount of solution remained in the tip of the pipette. To deliver the correct volume, the student should:

(a) blow the remaining solution into the flask

(b) touch the inside of the flask with the tip of the pipette

(c) shake the pipette to dislodge the remaining solution

(d) rinse the pipette with a small quantity of distilled water into the flask

The pipette is calibrated to deliver the correct amount if the tip is touched to the inside of the flask for 3 seconds. A small amount of solution should remain in the tip of the pipette.

Question 21

In an acid-base titration, the equivalence point:

(a) occurs when the indicator changes colour permanently.

(b) occurs only when the pH of the solution reaches 7.

(c) is always the same as the endpoint.

(d) occurs when stoichiometric amounts of H^+ from the acid and OH^- from the base are mixed.

This is when reaction is complete.

Question 22

A natural indicator extracted from flower petals changes colour in the pH range of 4.0–5.1. This indicator would be suitable for determining the equivalence point for a titration between:

(a) NaOH and H_2SO_4

(b) NaOH and HF

(c) NH_3 and HF

(d) NH_3 and H_2SO_4

Strong acid + weak base gives an acidic equivalence point.

Question 23

(a) Explain the difference between an *aliquot* and a *titre.* 2

An aliquot is the volume of the solution that is usually transferred by the pipette.

A titre is the volume of titrant added to reach the equivalence point.

(b) Explain why it is important not to store concentrated acid or basic solutions in glassware such as volumetric flasks and burettes. 2

Concentrated acids/bases are corrosive/caustic and if fragile volumetric glassware is broken or spilt, they are very harmful. Concentrated bases can dissolve glass and affect the accuracy of the glassware.

Question 24

Pipettes were traditionally filled by sucking up the required liquid by mouth. Now we use pipette fillers.

(a) Explain the advantage of this change in technique. 1

Much safer as there is no risk of aspirating or ingesting liquid (which can be toxic, corrosive or caustic). More hygienic as no transfer of germs from one person to another.

(b) Why is it incorrect to shake or blow the last drop of liquid from a pipette? 1

It is calibrated in such a way that the last drop is to remain in the pipette for the correct volume to be transferred.

(c) What precautions should you take when using a pipette filler? 1

Don’t suck solution up into pipette filler as the solutions are frequently corrosive and will damage the filler.

Question 25

Mary wanted to determine the amount of citric acid in a sample of lemon juice. The procedure she used is shown below.

1. A standard solution of sulfamic acid, a monoprotic acid, was prepared. 1.912 g of solid sulfamic acid (H_3NSO_3) was made up to 100 mL.
2. This solution was used to standardise a ~0.2 M solution of NaOH. 27.78 mL of NaOH was required to neutralise 25.00 mL of the sulfamic acid solution.
3. Next, a 50.00 mL sample of lemon juice was diluted to 250.0 mL and 25.00 mL aliquots of the diluted juice were titrated with the NaOH solution. The average titre was 23.73 mL.
4. The concentration of citric acid in the lemon juice was calculated as %w/w, assuming that citric acid was the only acid present.

The molar mass of sulfamic acid is 97.104 g mol^{-1}, the molar mass of citric acid is 192.124 g mol^{-1} and the density of lemon juice is 1.031 g mL^{-1}.

(a) Calculate the concentration of the sulfamic acid standard solution in mol L^{-1}. 2

$n(H_3NSO_3)$ = m / MM = 1.912 / (1.008 × 3 + 14.01 + 32.07 + 16 × 3) = 0.0196902 mol

$c(H_3NSO_3)$ = n / V = 0.0196902 / 0.1 = 0.196902 = 0.197 M (3 sig. fig.)

(b) Calculate the concentration of the sodium hydroxide solution in mol L^{-1}. 2

$n(H_3NSO_3)$ = c × V = 0.1969 × 0.025 = 4.92256 × 10^{-3} mol = n(NaOH)

c(NaOH) = n / V = 4.92256 × 10^{-3} / 0.02778 = 0.177198 = 0.177 M (3 sig. fig.)

(c) Calculate the concentration of citric acid in the diluted lemon juice in mol L^{-1}. 2

n(NaOH) = c × V = 0.177198 × 0.02373 = 4.204906 × 10^{-3} mol

From mole ratio, n(citric acid) = n(NaOH) / 3 = 1.401635 × 10^{-3} mol

c(citric acid) = n / V = 1.401635 × 10^{-3} / 0.025 = 0.0560654 = 0.0561 M (3 sig. fig.)

(d) Calculate the concentration of citric acid in the original juice sample in % w/w. 2

$c_1V_1 = c_2V_2$, $c_1 = 0.0560654 \times 0.25 / 0.05 = 0.280327$ mol L^{-1}

m(citric acid) in 1 L = n × MM = 0. 280327 × 192.124 = 53.8857555 g

m(1 litre of juice) = 1.031 × 1000 = 1031 g

%w/w = component / total × 100 = 53.8857555 / 1031 × 100 = 5.22% (3 sig. fig.)

Question 26

A 250 mL standard solution of 0.200 mol L^{-1} solution of sodium carbonate is to be prepared.

(a) Calculate the mass of sodium carbonate required to make this solution. 2

$n(Na_2CO_3) = c \times V = 0.2 \times 0.25 = 0.05$ mol

$m(Na_2CO_3) = n \times MM = 5.2995 = 5.30$ g (3 s.f.)

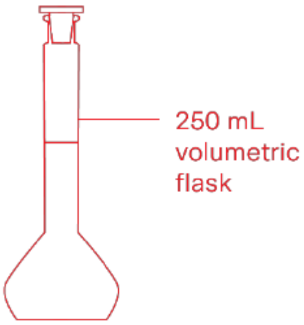
(b) Outline a suitable procedure to make this solution. Include a diagram of the volumetric equipment required. 5

1. Accurately weigh approximately 5.30 g of anhydrous Na_2CO_3 into a clean beaker.

2. Add distilled water to the Na_2CO_3 to dissolve. Carefully transfer it to a clean 250 mL volumetric flask. Rinse all of the equipment that came into contact with the weighed mass into the flask.

3. Add distilled water until the bottom of the meniscus is resting on the mark on the neck.

4. Stopper the flask and invert several times to ensure the solution is well mixed.



Question 27 **[2011 HSC Q26]**

A manufacturer makes lemon cordial by mixing flavouring, sugar syrup and citric acid. The concentration of the citric acid is determined by titration with NaOH.

The sodium hydroxide solution is prepared by dissolving 4.000 g of NaOH pellets in water to give 1.000 L of solution. This solution is standardised by titrating 25.00 mL with a 0.1011 mol L^{-1} standardised solution of HCl. The average titration volume is found to be 24.10 mL.

To analyse the lemon cordial, 50.00 mL of the cordial is diluted to 500.0 mL. Then 25.00 mL of the diluted solution is titrated with the NaOH solution to the phenolphthalein endpoint.

The following data were collected during one of the analysis runs of the lemon cordial.

Titration	#1	#2	#3	#4
Volume (mL)	26.55	27.25	27.30	27.30

(a) Explain why the calculated concentration of the standardised NaOH solution is different from the concentration calculated using the mass given, assuming no human error occurred. 2

c(NaOH) from mass given = n / V = (4.000 / 39.998) / 1.000 L = 0.1000 mol L^{-1}

c(NaOH) from titration:

n(NaOH) = n(HCl) = c × V = 0.1011 × 0.02410 = 0.00243651 mol

c(NaOH) = n / V = 0.002437 / 0.025 = 0.09746 mol L^{-1}

The difference in concentration is due to the NaOH being impure. NaOH is hygroscopic and therefore absorbs water from the air easily. Therefore, the weighed NaOH most likely contained H_2O as well, so less NaOH was in the solution than anticipated.

(b) Determine the concentration of citric acid in the lemon cordial. 4

Average V(NaOH) = 27.283 mL (remove outlier)

n(NaOH) = c × V = 0.09746 × 0.027283 = 0.0026590 mol

n(citric acid) = n(NaOH) / 3 = 0.000886343819 mol

c(dilute citric) = n / V = 0.000886348 / 0.025 = 0.0354539 M

c_1 = 0.03545 × 0.500 / 0.05 = 0.3545 M (4 sig. fig.)

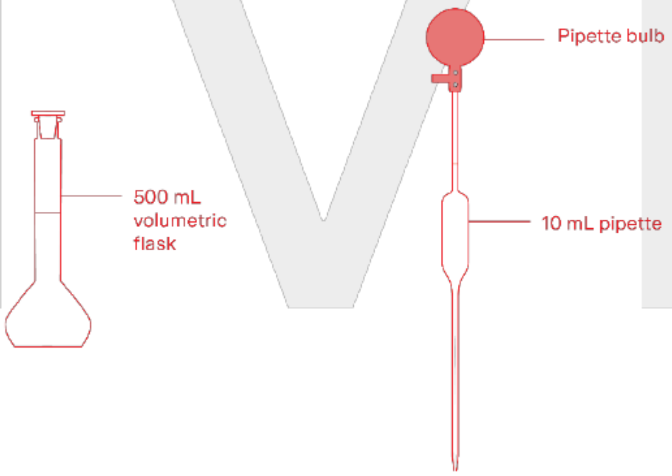
Question 28

In order to increase accuracy and safety, unknown samples are sometimes diluted before titrations. This is particularly important when handling unknown acids or bases as diluted samples can also reduce the risk associated with spillage.

To measure the concentration of sodium hydroxide in a liquid drain cleaner, a student titrated 25.0 mL samples of the drain cleaner that had been diluted by a factor of 50 against 0.1050 mol L^{-1} acetic acid solution. The results are shown below.

Titration number	Volume of acid (mL)
1	28.3
2	29.0
3	29.3
4	29.1

(a) Outline a procedure to dilute the drain cleaner accurately and draw a diagram of the volumetric equipment required. 5



1. Rinse a 10.0 mL pipette with distilled water 3 times, then 3 times with liquid drain cleaner.

2. Fill the pipette with drain cleaner until the meniscus rests on the etched mark.

3. Transfer the aliquot to a 500 mL volumetric flask (rinsed with distilled water 3 times). Touch the tip of the pipette on the side of the flask. Do not shake out the final drop.

4. Fill the flask with distilled water until the bottom of the meniscus is sitting on the mark on the neck of the flask, then stopper the flask and invert to mix.

(b) Which indicator would be suitable for this titration? Explain why. 2

Phenolphthalein: weak acid + strong base produces a basic salt (sodium acetate) which means the equivalence point will be basic. Since phenolphthalein's colour change range is 8.2 – 10.0, it will be suitable.

(c) Calculate the concentration (in mol L^{-1}) of sodium hydroxide in the undiluted drain cleaner. 3

$CH_3COOH_{(aq)} + NaOH_{(aq)} \longrightarrow NaCH_3COO_{(aq)} + H_2O_{(l)}$

Average $V(CH_3COOH) = 29.133$ mL (28.3 is an outlier)

$n(CH_3COOH) = c \times V = 0.1050 \times 0.029133 = 3.059 \times 10^{-3}$ mol $= n(NaOH)$

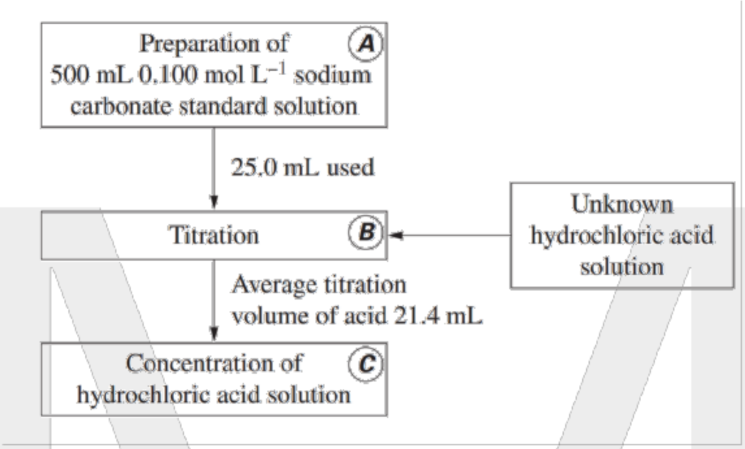
$c(NaOH) = n / V = 3.059 \times 10^{-3} / 0.025 = 0.12236$ M

Since the drain cleaner was diluted by a factor of 50,

Original concentration $= 0.12236 \times 50 = 6.12$ M (3 s.f.)

Question 29 **[2010 HSC Q28]**

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 hydrochloric acid. 8

STEP A: Na_2CO_3 should initially be dried and stored in a desiccator.

m(Na_2CO_3) = 0.1 × 0.5 × 105.99 = 5.30 g

Dried Na_2CO_3, 5.30 g, should be weighed out accurately (> 2 d.p.)

500 mL volumetric flask cleaned and rinsed with distilled water. Dissolve weighed Na_2CO_3, transferred to volumetric flask, rinse all equipment that came into contact with weighed mass into flask. Add distilled water until the bottom of the meniscus is resting on the mark on the neck. Stopper the flask and invert several times to mix.

STEP B:

Clean, rinse and fill a 50 mL burette with the unknown acid and clamp to a retort stand.

Clean a 250 mL conical flask and rinse with distilled water.

Clean a 25 mL pipette and rinse with 0.1 M Na_2CO_3 solution. Fill pipette with Na_2CO_3 solution to mark with bottom of meniscus.

Add pipette volume into conical flask. Add suitable indicator to the conical flask. Place a white tile under the conical flask.

Slowly add acid solution from the burette into the conical flask, while swirling the flask. When the indicator permanently changes colour, record the final volume. This initial titration should be taken as a rough test. Three more titrations should be performed with the average titration being used for calculations.

STEP C:

$2HCl_{(aq)} + Na_2CO_{3(aq)} \longrightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$

$n(Na_2CO_3) = 0.1 \times 0.025 = 0.0025$ mol

$n(HCl) = 2 \times n(Na_2CO_3) = 0.0050$ mol

$c(HCl) = 0.0050 / 0.00214 = 0.234$ mol L^{-1}

Question 30

Standardised sodium hydroxide solution was used by a chemist to determine the concentration of acetic acid in an old storage bottle. The acid concentration was calculated to be 0.1344 M.

(a) Identify a property of sodium hydroxide which makes it unsuitable as a primary standard. 1

It is hygroscopic (absorbs moisture from the air) therefore cannot be weighed accurately.

(b) Outline the safety precautions needed when using sodium hydroxide. 2

Wear safety glasses, lab coat and gloves (as NaOH is corrosive), perform experiments near a source of running water in case of spills.

(c) Calculate the pH of the acetic acid solution, assuming that 1% of the acetic acid molecules are ionised in solution. 2

$[H^+] = 0.1344 \times 0.01 = 0.001344$ M

$pH = -\log_{10}[H^+] = 2.9$ (1 sig. fig.)

Question 31

Permanent hair dyes generally come with 2 separate components, one of which contains ammonia which opens the hair shaft so the dye can diffuse inside the hair fibre. James wants to determine the concentration of ammonia in a hair dye component via a straightforward titration. The following chemicals are available in the laboratory:

Acid solutions	0.1025 M hydrochloric acid	0.1025 M acetic acid
Indicators	Methyl orange	Phenolphthalein

(a) Identify an appropriate acid and indicator to use for this investigation, and justify your choices. 3

Hydrochloric acid – it is strong, therefore can be used to titrate against a weak base using an indicator to determine the endpoint. If weak acetic acid was used, the change in pH at the endpoint would be too gradual to give a sharp colour change.

Methyl orange – hydrochloric acid reacts with ammonia to form ammonium chloride at the equivalence point. It is a weak acid, so methyl orange (colour change range pH 3.1–4.4) will appropriately indicate the equivalence point.

Concerned about the volatility of ammonia, Lily decided to measure the ammonia content using a back titration procedure instead. She took a 15 mL sample of the hair dye and diluted it with 25 mL distilled water. She then added 50 mL of 0.75 M HNO_3 to the diluted dye, and titrated the excess with 0.124 M sodium hydroxide solution. The average titre was 27.45 mL.

(b) Calculate the concentration of ammonia in the dye component in %w/w. 4

Reaction 1: $HNO_{3(aq)} + NH_{3(aq)} \longrightarrow NH_4NO_{3(aq)}$

Reaction 2: $HNO_{3(aq)} + NaOH_{(aq)} \longrightarrow NaNO_{3(aq)} + H_2O_{(l)}$

$n(NaOH) = c \times V = 0.124 \times 0.02745 = 3.4038 \times 10^{-3}$ mol $= n(HNO_3)_{Rxn\,2}$

$n(HNO_3)_{Rxn\,1} = n(HNO_3)_{total} - n(HNO_3)_{Rxn\,2} = (0.75 \times 0.05) - 3.4038 \times 10^{-3} = 0.0340962$ mol

$n(NH_3) = 0.0340962$ mol in 15 mL concentrated hair dye

$m(NH_3) = n \times MM = 0.0340962 \times (14.01 + 1.008 \times 3) = 0.58079$ g

%w/w = 0.58079 / 15 × 100 = 3.87%

Question 32

A scientist wishes to quantitatively determine the concentration of sulfuric acid in a polluted water sample by titrating 25 mL aliquots of the sample against a standardised solution of 0.1120 mol L^{-1} NaOH.

(a) Outline the method the scientist should follow when performing the investigation. Specify the rinsing solutions and an appropriate indicator in your answer. 3

Water sample:

1. Rinse a clean 25 mL pipette 3 times with the polluted water sample.

2. Fill pipette with polluted water until the bottom of the meniscus is resting on the mark.

3. Transfer the aliquot to a clean conical flask that has been rinsed 3 times with distilled water. Do not shake out the last drop.

NaOH:

1. Rinse a clean burette 3 times with the standard NaOH solution. Ensure there are no bubbles trapped in the tap.

2. Fill up the burette with NaOH solution until the bottom of the meniscus is between 0 and 5 mL. Record the actual reading.

Titration:

1. Add 3 drops of bromothymol blue indicator to the conical flask.

2. Add NaOH from the burette to the acid in the conical flask slowly until a permanent colour change occurs (from yellow to green).

3. Record the final reading on the burette (at the bottom of the meniscus).

4. Repeat the entire procedure 3 times.

(b) The following table shows the volume of base used in various attempts. Calculate the concentration of sulfuric acid in the water sample in mol L^{-1}. **2**

Attempt	1	2	3	4
Volume NaOH (mL)	19.20	18.65	18.55	18.65

$2NaOH_{(aq)} + H_2SO_{4(aq)} \longrightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$

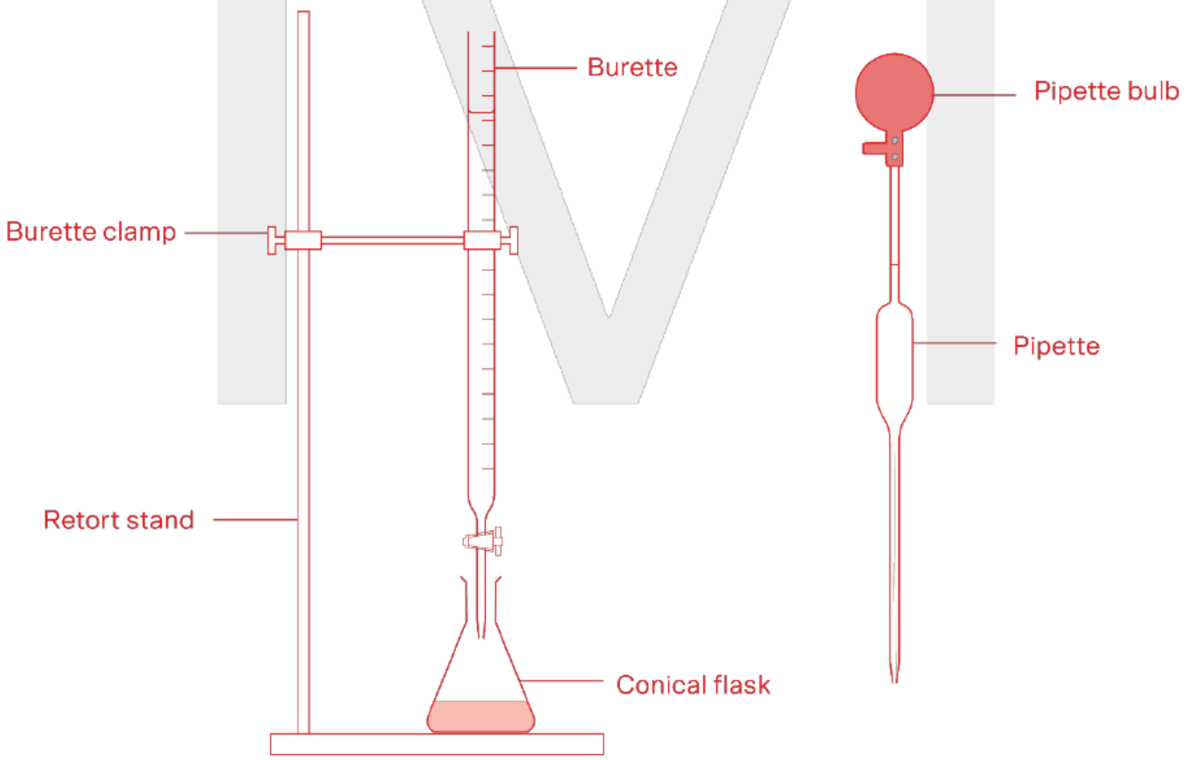
Average V(NaOH) = 18.6167 mL (19.20 mL is an outlier)

n(NaOH) = c × V = 0.1120 × 0.0186167 = 2.08507 × 10^{-3} mol

$n(H_2SO_4)$ = n(NaOH) / 2 = 1.0425333 × 10^{-3} mol

$c(H_2SO_4)$ = n / V = 1.0425333 × 10^{-3} / 0.025 = 0.042 M (2 s.f.)

(c) Draw and label the apparatus used for this titration. **3**



Question 33

To determine the degree of zinc accumulation in a river close to an industrial processing plant, soil sediment was collected from the bottom of the river. The soil was thoroughly dried in an oven to remove excess moisture. 30.0 mL of 0.300 mol L^{-1} nitric acid solution was added to a 4.33 g sample of the dried soil. After all the zinc had reacted, the remaining solution was made up to 500 mL, then titrated in 25.0 mL aliquots with 0.0264 mol L^{-1} sodium hydroxide solution using phenolphthalein as the indicator. The titration results are given below:

Titration attempt	Volume of NaOH (mL)
1	13.9
2	12.7
3	12.8
4	12.9

(a) Determine the moles of nitric acid remaining after reaction with the zinc. 2

$HNO_{3(aq)} + NaOH_{(aq)} \longrightarrow NaNO_{3(aq)} + H_2O_{(l)}$

n(NaOH) = c × V = 0.0264 × 0.0128 = 3.3792×10^{-4} mol = n(HNO_3) in 25.0 mL

Remaining n(HNO_3) = moles in 500 mL = (3.3792×10^{-4} / 25) × 500 = 6.7584×10^{-3} mol

= 6.76×10^{-3} mol (3 s.f.)

(b) Determine the mass percentage (%w/w) of zinc in the soil sample. 3

Reaction 1: $Zn_{(s)} + 2HNO_{3(aq)} \longrightarrow Zn(NO_3)_{2(aq)} + H_{2(g)}$

$n(HNO_3)_{total}$ = c × V = 0.03 × 0.3 = 9×10^{-3} mol

$n(HNO_3)_{Rxn\ 1} = n(HNO_3)_{total} - n(HNO_3)_{remaining} = (9 - 6.7584) \times 10^{-3} = 2.2416 \times 10^{-3}$ mol

n(Zn) = (2.2416×10^{-3}) / 2 = 1.1208×10^{-3} mol

m(Zn) = n × MM = 0.073278 g = 0.0733 g (3 s.f.)

Concentration (%w/w) = m(Zn) / m(soil) × 100 = 1.69% (3 s.f.)

(c) Explain why phenolphthalein is appropriate as the indicator for this titration. 2

The titration is between a strong acid and a strong base, hence has a rapid steep change in pH around the neutral equivalence point. Hence phenolphthalein (colour change range pH 8.2–10.0) will change colour at this point.

(d) Outline an alternative method of determining the mass of zinc in a sample, based on gas volume. 2

A weighed soil sample is reacted with sulfuric acid under an inverted measuring cylinder, initially filled with water. After equalising water levels the volume of hydrogen gas collected is measured and converted to moles, using the molar volume of a gas under collection conditions. As one mole of Zn produces one mole of hydrogen, the mass of Zn can be determined.

Question 34 **[2012 HSC Q30]**

A chemist analysed aspirin tablets for quality control.

The initial step of the analysis was the standardisation of a NaOH solution. Three 25.00 mL samples of a 0.1034 mol L^{-1} solution of standardised HCl were titrated with the NaOH solution. The average volume required for neutralisation was 25.75 mL.

(a) Calculate the molarity of the NaOH solution. 2

$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

n(NaOH) = n(HCl) = 0.1034 × 0.02500 = 0.002585 mol

c(NaOH) = n / V = 0.002585 / 0.02575 = 0.1004 mol L^{-1} (4 sig. fig.)

Three flasks were prepared each containing a mixture of 25 mL of water and 10 mL of ethanol. An aspirin tablet was dissolved in each flask. The aspirin in each solution was titrated with the standardised NaOH solution according to the following equation:

$$C_9H_8O_{4(aq)} + NaOH_{(aq)} \longrightarrow C_9H_7O_4Na_{(aq)} + H_2O_{(l)}$$

The following titration results were obtained.

Tablet	Volume (mL)
1	16.60
2	16.50
3	16.55

(b) Calculate the average mass (mg) of aspirin per tablet. 3

n(aspirin) = n(NaOH) = c × V = 0.1004 × 0.01655 = 0.001661 mol

m(aspirin) = n × MM = 0.001661 × (9 × 12.01 + 8 × 1.008 + 4 × 16.00) = 0.2993 g

Average mass of aspirin per tablet is 299.3 mg (4 sig. fig.)

(c) Why was it necessary to include the ethanol in the mixture? 1

Ethanol is necessary to improve the solubility of aspirin

Year 12
Chemistry
Acid/Base Reactions

Work Book 8
Applications of titration

MATRIX
EDUCATION

Part A: Multiple choice questions

Question 1

During the titration of an aliquot of a strong acid with a strong base whose concentration is known:

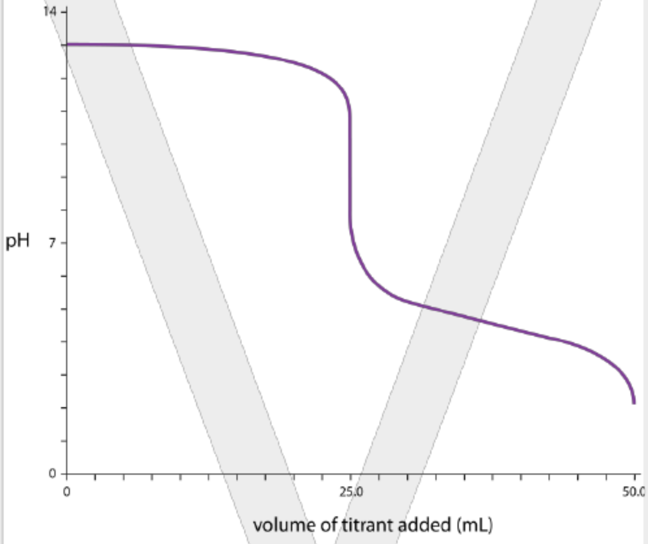
(a) there is a steady increase in pH

(b) there is a sharp increase in pH around the end point

(c) there is a steady decrease in pH

(d) there is a sharp decrease in pH around the end point

Question 2



Which combination of acid and base could produce the titration curve above?

(a) Sodium hydroxide and hydrochloric acid

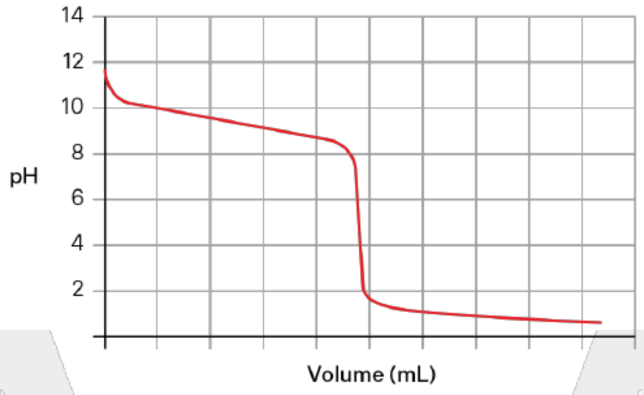
(b) Potassium hydroxide and sulfuric acid

(c) Ammonia and hydrochloric acid

(d) Potassium hydroxide and acetic acid

The pH of the equivalence point (most vertical point) is basic therefore weak acid + strong base.

Questions 3 and 4 refer to the figure below.



Question 3

Which titration set-up produced the titration curve above?

	Solution in conical flask	Solution in burette
(a)	HCl	NH_3
(b)	NH_3	HCl
(c)	CH_3COOH	NaOH
(d)	NaOH	CH_3COOH

The curve shape is for a weak base/strong acid titration, and begins at high pH.

Question 4

The table shows the indicators used to identify the equivalence point in titrations.

Indicator	Range of colour change
Crystal violet	0.2–1.8
Methyl red	4.4–6.2
Cresolphthalein	8.1–9.7
Alizarin yellow	10.1–12.0

Which indicator is the most suitable for determining the equivalence point of this titration?

(a) Crystal violet

(b) Methyl red

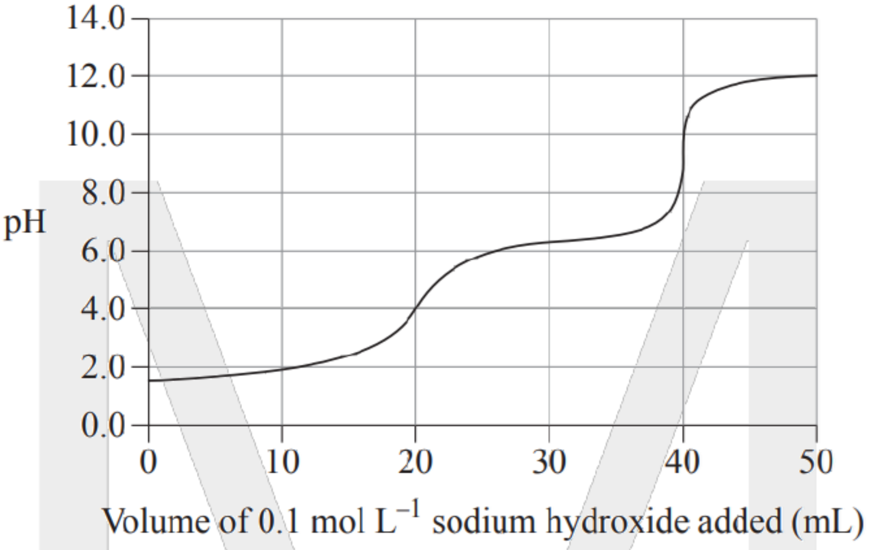
(c) Cresolphthalein

(d) Alizarin yellow

Equivalence point is the most vertical part of the graph; it is at pH ~5. The indicator which changes colour at this pH is methyl red.

Question 5 [2024 HSC Q17]

20 mL of a 0.1 mol L^{-1} solution of an acid is titrated against a 0.1 mol L^{-1} solution of sodium hydroxide. A graph of pH against the volume of sodium hydroxide for this experiment is shown.



Which of the following acids was used in the titration?

	Acid	pK_{a1}	pK_{a2}
(a)	1	4.76	-
(b)	2	Strong	-
(c)	3	1.91	6.30
(d)	4	4.11	9.61

The titration curve depicts two equivalence points, hence the acid is diprotic and will have two pK_a values.

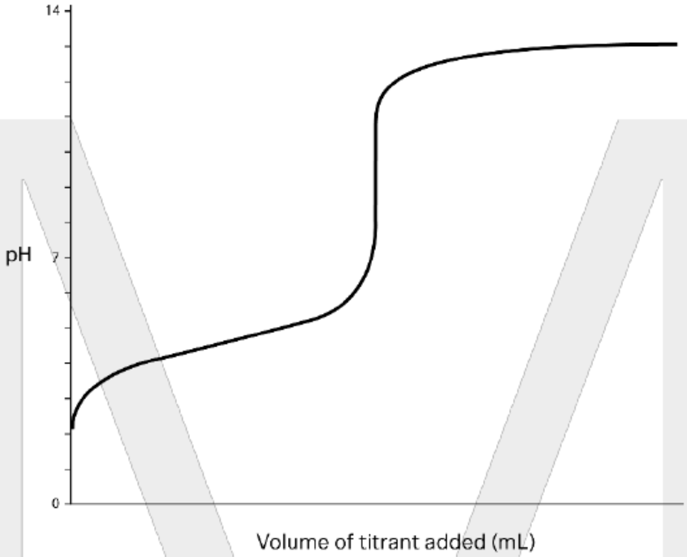
The pH at the half-equivalence points gives the pK_a values for each dissociation. The first equivalence point occurs at 20 mL, hence the first half-equivalence point occurs at 10 mL. The pH here is 1.91, so pK_a for the first dissociation is 1.91.

The second equivalence point occurs after another 20 mL (total 40 mL) of NaOH is added, so the second half-equivalence point occurs after another 10 mL has been added (total 30 mL). The pH here is 6.30, so pK_a for the second dissociation is 6.30.

Questions 6 – 8 refer to the following information.

In an investigation, a student titrated 20.0 mL of a 0.100 mol L^{-1} strong monoprotic acid with 0.100 mol L^{-1} KOH. Then in another titration, the student titrated 20.0 mL of another 0.100 mol L^{-1} monoprotic acid with 0.100 mol L^{-1} KOH.

The graph shows the changes in pH during the second titration:



Question 6

Which of the following substances could be the acid used in the second titration?

(a) Hydrochloric acid

(b) Hydrocyanic acid

(c) Nitric acid

(d) Citric acid

The equivalence point (the most vertical part) is basic, which corresponds to a strong base/weak acid titration (since the weak acid produces its moderately strong conjugate base upon neutralisation). According to the question, the acid used in the second titration is monoprotic.

Question 7

Which pH range should an indicator have to be used in this titration?

(a) 3.1–4.4

(b) 5.0–8.0

(c) 6.0–7.6

(d) 8.3–10.0

Equivalence point is the most vertical part of the graph; it is at pH ~9. The indicator which changes colour at this pH is (d).

Question 8

Which statement regarding the volume of KOH required to reach the equivalence point is correct?

(a) The strong acid will require a larger volume of KOH for neutralisation than the second acid.

(b) The strong acid will require a smaller volume of KOH for neutralisation than the second acid.

(c) The strong acid will require the same volume of KOH for neutralisation as the second acid.

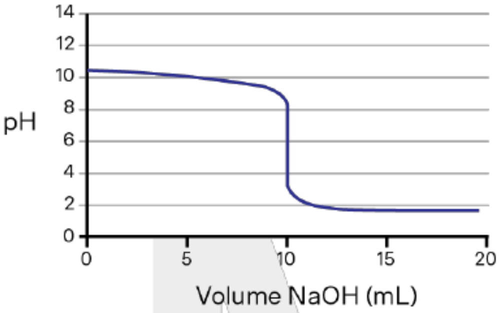
(d) The volume of KOH cannot be determined as the identities of the acid are unknown.

The amount of base required only depends on the number of ionisable protons. Both acids are monoprotic.

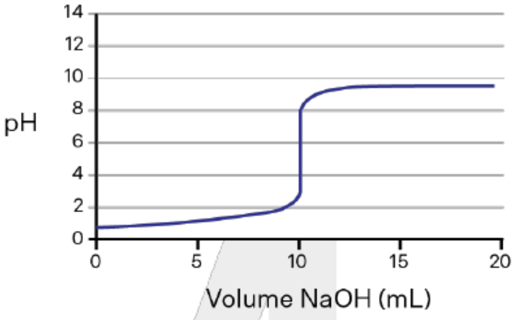
Question 9

Which titration curve best represents the change in pH as 0.100 M NaOH solution is added to a 10.0 mL aliquot of 0.100 M HCl solution?

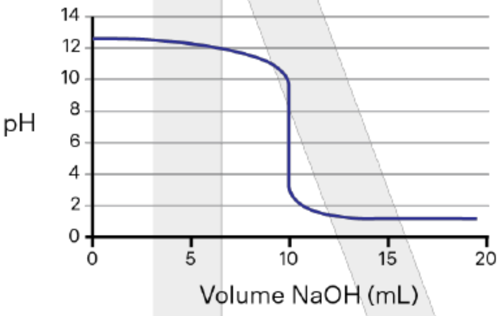
(a)



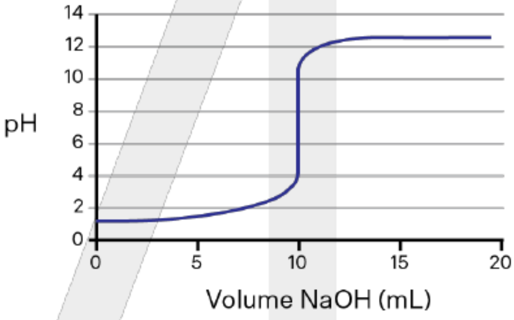
(b)



(c)



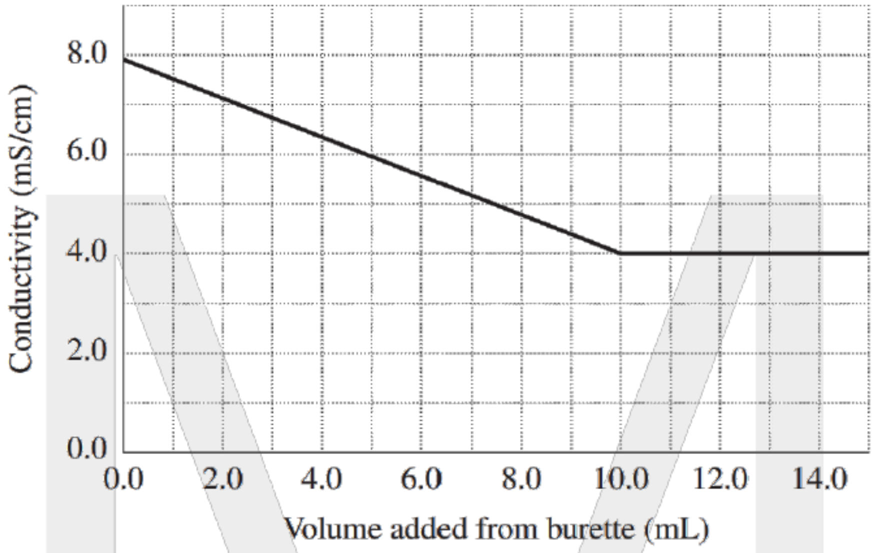
(d)



pH increases as NaOH (base) is added. HCl and NaOH are both strong, so the pH at equivalence point will be neutral, which corresponds to curve (d).

Question 10 **[2022 HSC Q15]**

A 25.00 mL sample of 0.1131 mol L^{-1} $HCl_{(aq)}$ was titrated with an aqueous ammonia solution. The conductivity of the solution was measured throughout the titration and the results graphed.



What was the concentration of the ammonia solution?

(a) 0.0452 mol L^{-1}

(b) 0.189 mol L^{-1}

(c) 0.283 mol L^{-1}

(d) 0.690 mol L^{-1}

$HCl_{(aq)} + NH_{3(aq)} \longrightarrow NH_4Cl_{(aq)}$

n(HCl) = c × V = 0.1131 × 0.025 = 2.8275 × 10^{-3} mol = n(NH_3) (1:1 stoichiometry)

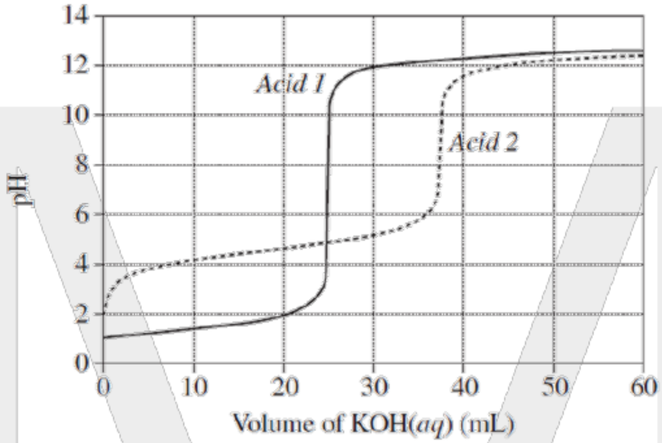
In a strong acid/weak base conductivity titration curve, the end point is indicated by a sharp turning point. This occurs at a volume of 10 mL (0.010 L).

c(NH_3) = n/V = 2.8275 × 10^{-3} mol/0.010 L = 0.283 mol L^{-1}

Part B: Extended response questions

Question 11 **[2009 HSC Q21]**

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**

Acid 1 is a strong acid, equivalence point is at pH = 7 and initial pH is at 1.

Acid 2 is a weak acid as pH at equivalence point is >7 and initial pH is >1.

Acid 2 has a higher concentration than Acid 1 as it takes more KOH to neutralise it.

(b) Name the salt produced by the reaction of an acid of the same type as Acid 2 with $KOH_{(aq)}$. **1**

Potassium acetate (potassium ethanoate), KCH_3COO

(c) Calculate the concentration of hydrogen ions when 20 mL of $KOH_{(aq)}$ has been added to Acid 1. **1**

When 20 mL of KOH has been added, pH of mixture is 2 (from graph)

Since $pH = -\log_{10}[H^+]$, $[H^+] = 10^{-2} = 0.01$ mol L^{-1}

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

Phenolphthalein changes colour at pH 8.2–10.0, which is on the vertical section of each titration curve.

Question 12

A titration is conducted between 25.0 mL of 0.405 M sodium hydroxide and acetic acid of unknown concentration.

(a) Outline how a pH probe could be used to identify the equivalence point of this titration. 3

A pH probe could be placed in the conical flask at the start of the titration. The pH should be recorded upon the addition of regular volumes of acetic acid (e.g., every 0.5 mL). A titration curve could be plotted with volume of acetic acid added vs pH. The equivalence point is located at the most vertical part of the curve.

(b) 23.15 mL of acetic acid was required to reach the equivalence point. Calculate the concentration of acetic acid in the original solution. 2

n(NaOH) = 0.405 × 0.025 = 0.010125 mol = n(CH_3COOH)

c(CH_3COOH) = 0.010125 / 0.02315 = 0.437365 = 0.437 M (3 s.f.)

(c) Calculate the concentration of sodium acetate in the conical flask at the equivalence point. Assume no additional water was added to the flask during the titration. 2

0.010125 mol in 48.15 mL = 0.21028 = 0.210 M (3 s.f.)

(d) The pK_a of acetic acid is 4.74. Calculate the pH of the equivalence point. 3

pK_b of CH_3COO^- = 14 − 4.74 = 9.26

Concentration (M)	CH_3COO^-	CH_3COOH	OH^-
Initial	0.21028	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.21028 - x$	x	x

Assume $[CH_3COO^-]_{eqm}$ = 0.21028 M, since K_b is small and dissociation does not change $[CH_3COO^-]_{initial}$ by much.

$$\frac{x^2}{0.21028} = 10^{-9.26}$$

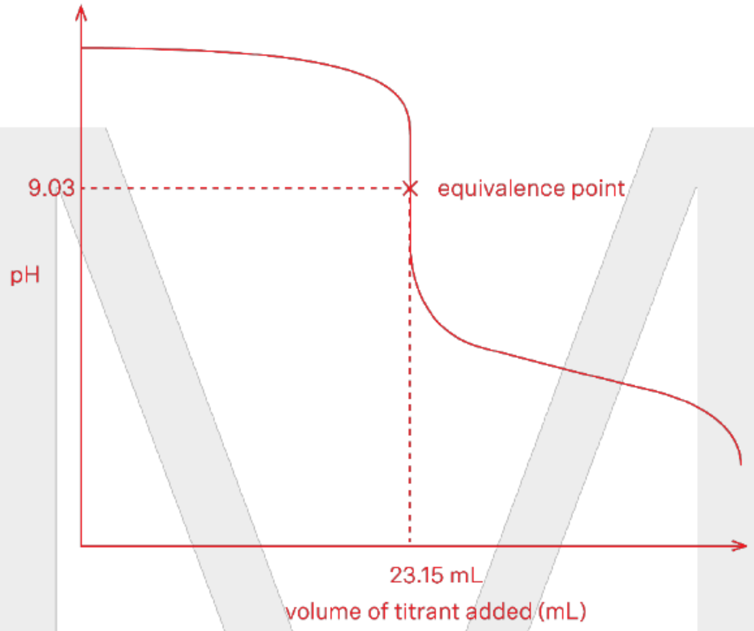
$x = 1.07498 \times 10^{-5} = [OH^-]$

pOH = 4.9686, therefore pH = 9.03 (2 s.f.)

(e) What is a suitable indicator to use for this titration? **1**

Phenolphthalein

(f) Sketch the shape of the expected titration curve when an excess of the titrant is added in the space below. **2**



Question 13 **[2020 HSC Q28]**

A chemist used the following method to determine the concentration of a dilute solution of propanoic acid (pK_a = 4.88).

The chemist weighed out 1.000 g of solid NaOH on an electronic balance and then made up the solution in a 250.0 mL volumetric flask.

The chemist then performed titrations, using bromocresol green as the indicator. This indicator is yellow below pH 3.2 and green above pH 5.2.

The results are shown in the table.

Titre	Volume of $NaOH_{(aq)}$ added (mL)
1	16.35
2	10.10
3	12.35
4	11.25

Explain why this method produces inaccurate and unreliable results. 3

NESA sample answer: NaOH is used as a primary standard, but it is not a primary standard. Solid NaOH is deliquescent, so the mass obtained by the balance is that of NaOH and water. The mass of NaOH is therefore lower than reported, and the solution made up is therefore lower in concentration than expected, resulting in a titre that is larger than it should be, giving an inaccurate result.

The indicator chosen is inappropriate. It changes colour in the flat region of the titration curve, prior to the equivalence point. Small titre volume differences will result in big variations in the observed endpoint, explaining the lack of reliability in titres 2, 3 and 4.

Question 14

Some titration curves display a buffering region. Sally performs the following titrations:

Titration	Conical flask	Burette
X	KOH	HF
Y	HBr	NaOH
Z	HNO_3	NH_3

(a) Which titration(s) would produce curves that display a buffering region? **1**

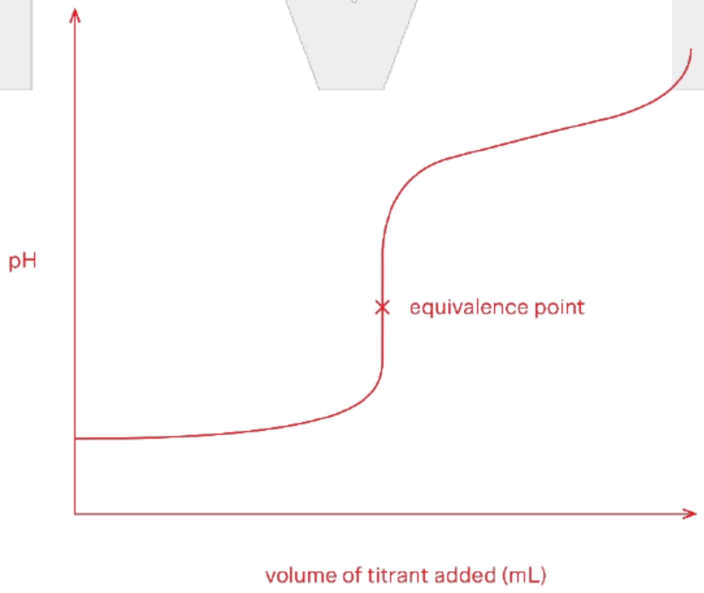
X and Z

(b) Explain why the buffering region arises in a titration you have identified above. Include a balanced chemical equation in your answer. **3**

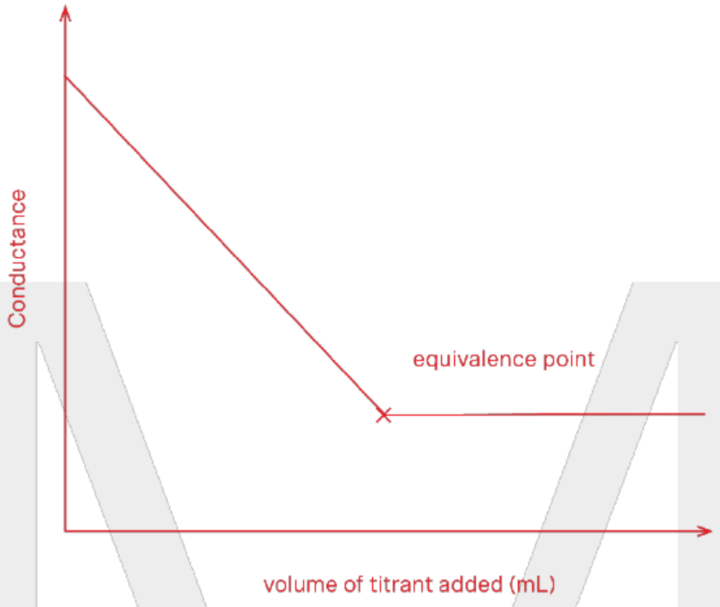
Titration Z: $HNO_{3(aq)} + NH_{3(aq)} \longrightarrow NH_4NO_{3(aq)}$

In this titration, when an excess of NH_3 is present, NH_4^+ is also present, hence there is a weak acid and its conjugate base, forming an equilibrium that can absorb excess H^+ or OH^-. This stabilises the pH, and hence forms a buffering region where the pH changes slowly.

(c) Sketch the shape of the expected titration curve for Titration Z. Label the equivalence point. **2**



(d) Sketch the shape of the expected conductivity graph for Titration Z. Label the equivalence point. 2



Question 15

Outline two advantages of using digital technologies to determine the equivalence point of a titration instead of indicators. 3

- More precise: indicators change colour over a range of pH values, therefore indicators can only show the approximate equivalence point

- Can determine the equivalence point for weak acid/weak base titrations, when the pH changes too gradually for a sharp indicator colour change to occur

- Can determine the equivalence point for intensely coloured reagents that would mask the indicator colour change

Other answers possible.

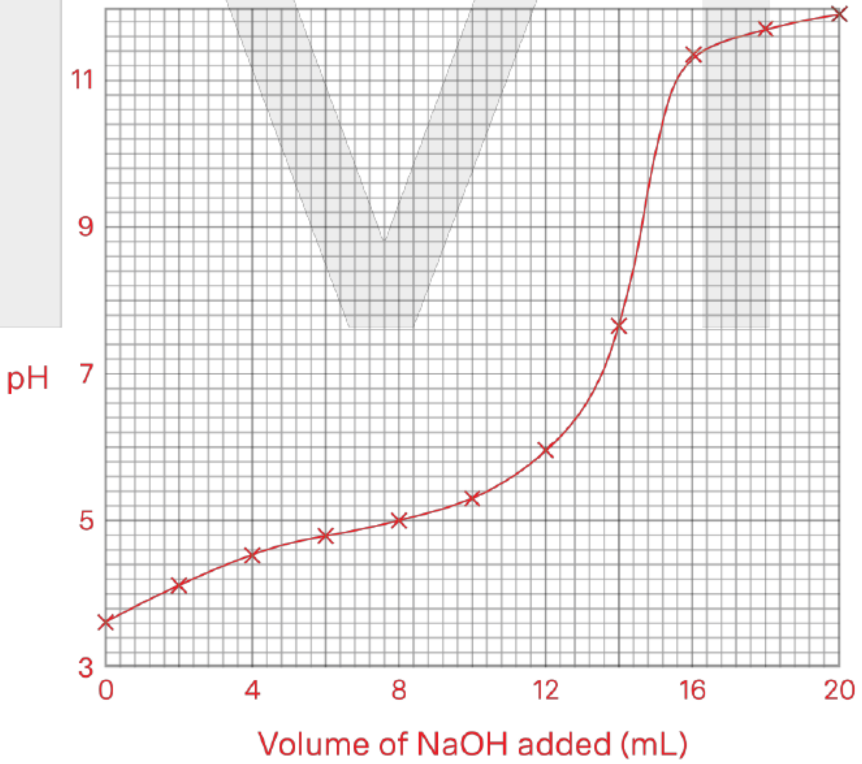
Question 16

A data logger with a pH probe attached was used in the titration of 30 mL of dilute propanoic acid with 0.010 mol L^{-1} sodium hydroxide to determine its concentration. The following results were obtained:

Volume of NaOH added (mL)	pH
0	3.6
2	4.1
4	4.5
6	4.8
8	5.0
10	5.3

Volume of NaOH added (mL)	pH
12	5.9
14	7.6
16	11.3
18	11.7
20	11.9

(a) Draw a graph of pH versus volume of NaOH added on the grid below. 3



(b) Use the graph to determine the volume of NaOH required to reach the equivalence point. 1

15 mL

(c) Propanoic acid is monoprotic. Determine the concentration of the acid from the titration results. 2

n(NaOH) = c × V = 0.01 × 0.015 = 0.00015 mol = n(propanoic acid) (1:1 ratio)

c(propanoic acid) = n / V = 0.00015 / 0.030 = 0.0050 M (2 sig. fig.)

(d) Is propanoic acid a strong or weak acid? Justify your response using two different pieces of evidence from the data and responses above. 2

It is a weak acid:

- The pH of 0.0050 M propanoic acid is 3.60, according to the data from the graph. If it were a strong acid, then it would be completely dissociated and $[H^+]$ = 0.050 M, which would result in a pH of 2.30.
- The equivalence point is at ~pH 10 from the graph. These result from weak acid/strong base titrations.

(e) Determine the K_a for propanoic acid. 2

When the concentrations of weak acid and conjugate base are equal, pH = pK_a. This occurs halfway on the titration curve to the equivalence point.

The half-equivalence point occurs at 7.5 mL. The pH at the half-equivalence point is 4.9. Thus pK_a = 4.9.

$K_a = 10^{-pka} = 10^{-4.9} = 1.3 \times 10^{-5}$

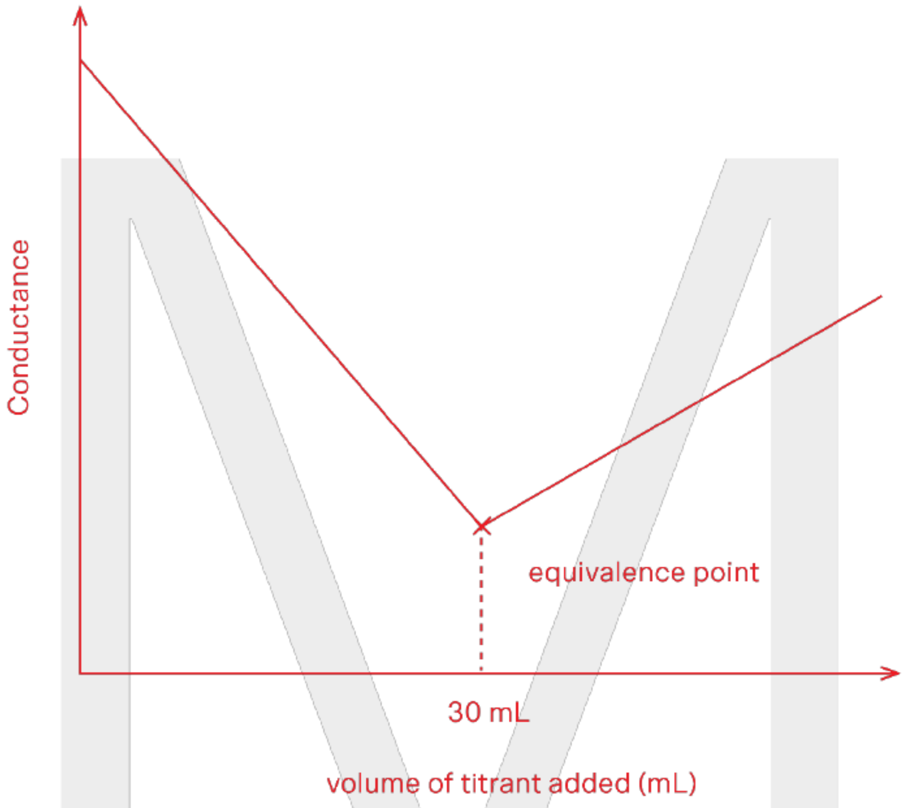
(f) Explain why the equivalence point in a titration between propanoic acid and ammonia cannot be determined using indicators. 2

It is a weak acid-weak base titration, therefore the pH changes very gradually at the equivalence point, so a sharp indicator colour change does not occur.

Question 17

30.0 mL of 0.15 M HCl is titrated against 0.15 M NaOH.

(a) Sketch the shape of the expected conductivity graph when an excess of base is added in the space below. Label the equivalence point. 2



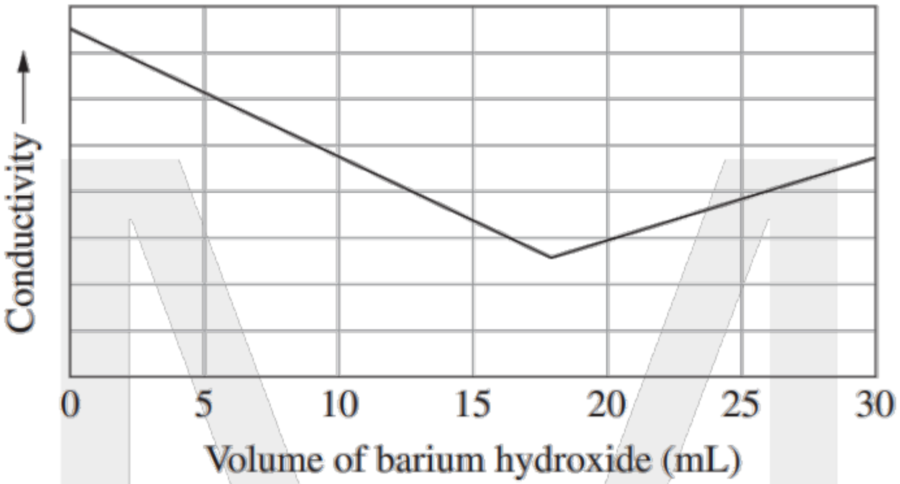
(b) Explain the changes in conductivity during the course of the experiment. 3

The conductivity initially decreases as the moles of ions in the reaction mixture stays the same, but the volume of the solution increases, so the concentration of the ions decreases resulting in a decrease in conductivity. H^+ is also replaced by Na^+, which is less conductive. The conductivity reaches a minimum at the equivalence point, when only Na^+ and Cl^- are present.

After the equivalence point, ions are added but no reaction occurs, therefore the concentration of ions increases. OH^- is also being added, which is highly conductive.

Question 18 **[2019 HSC Q24]**

A conductometric titration was undertaken to determine the concentration of a barium hydroxide solution. The solution was added to 250.0 mL of standardised 1.050×10^{-3} mol L^{-1} hydrochloric acid solution. The results of the titration are shown in the conductivity graph.



(a) Explain the shape of the titration curve. 3

The curve steadily falls as OH^- ions are added because they react and remove the highly conductive H^+ ions from the solution. At equivalence point, conductivity is lowest because the solution contains only Ba^{2+} and Cl^- ions, which are much less mobile than the H^+ or OH^- ions. The curve then rises as more OH^- ions are added.

(b) The equivalence point was reached when a volume of 17.15 mL of barium hydroxide was added. Calculate the concentration of barium hydroxide (in mol L^{-1}), and give a relevant chemical equation. 4

$2HCl_{(aq)} + Ba(OH)_{2(aq)} \longrightarrow BaCl_{2(aq)} + 2H_2O_{(l)}$

$n(HCl) = c \times V = 1.050 \times 10^{-3}$ mol $L^{-1} \times 250.0$

$n(Ba(OH)_2) = n(HCl) \times \frac{1}{2}$ (as reaction ratio is 2 :1)

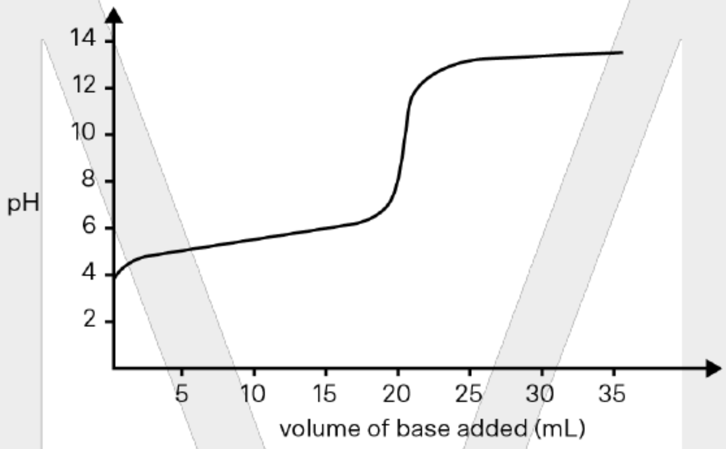
Concentration $Ba(OH)_2 = n / V = (1.050 \times 10^{-3} \times 0.2500) / 0.01715 \times 2 = 7.653 \times 10^{-3}$ mol L^{-1}

Part C: Additional practice questions

The questions in this part are intended to be used for revision purposes. However, from time to time your teacher may prescribe some of these questions for homework.

Question 19

The change in pH as a 0.10 M solution of a strong base is added to 20.0 mL of a 0.10 M solution of a weak acid is shown below.



Which indicator below would be least suitable for determining the endpoint of this titration?

(a) Phenol red (pH colour change range 6.8–8.4)

(b) Methyl orange (pH colour change range 3.1–4.4)

(c) Bromothymol blue (pH colour change range 6.0–7.6)

(d) Phenolphthalein (pH colour change range 8.2–10.0)

Equivalence point is the most vertical part of the graph; it is at pH ~9 and occurs when 20 mL of base has been added. When methyl orange is used, the end point would be observed almost immediately as base is added as the starting pH of the solution is already ~4.

Question 20

A student wishes to distinguish between 0.1 M nitric acid and 0.1 M acetic acid. Which method will be invalid?

(a) Determine the amount of NaOH required to neutralise an equal volume of each solution

(b) Test the pH of the solutions

(c) Test the electrical conductivity of the solutions

(d) All the above methods are valid

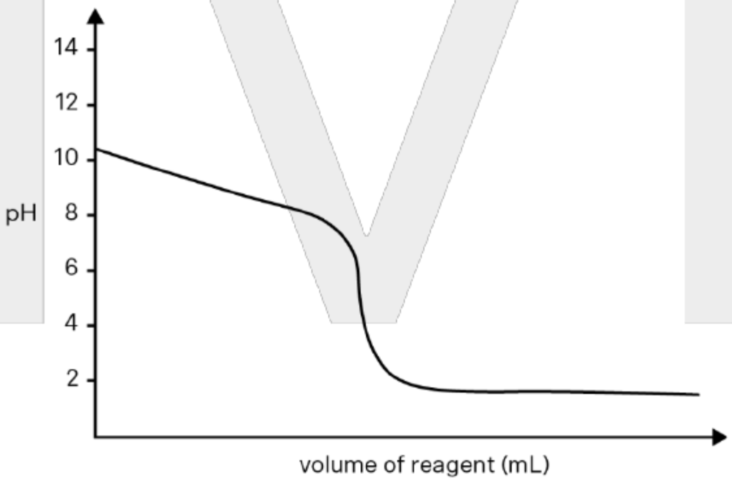
(a): Weak vs strong only reflects the ability of water to extract protons from an acid; during neutralisation, the base is capable of removing all protons from an acid, whether weak or strong. Since both are monoprotic, the same amount of base will be required

(b): The pH of the acetic acid will be higher as it is less ionised and produces less H^+

(c): As nitric acid is stronger and more ionised, it will conduct electricity better

Question 21

The diagram below represents the titration curve for an acid-base reaction.



Which reaction could this curve represent?

(a) $HCl_{(aq)} + NH_{3(aq)} \longrightarrow NH_4Cl_{(aq)}$

(b) $HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$

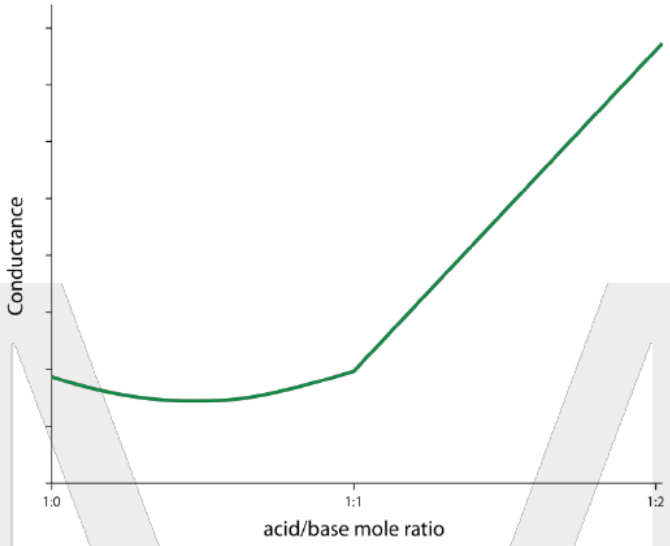
(c) $CH_3COOH_{(aq)} + NH_{3(aq)} \longrightarrow NH_4CH_3COO_{(aq)}$

(d) $CH_3COOH_{(aq)} + NaOH_{(aq)} \longrightarrow NaCH_3COO_{(aq)} + H_2O_{(l)}$

The equivalence point (the most vertical part) is acidic, which corresponds to a strong acid/weak base titration (since the weak base produces its moderately strong conjugate acid upon neutralisation).

Question 22

Which pair of reactants would produce the conductivity graph shown below?



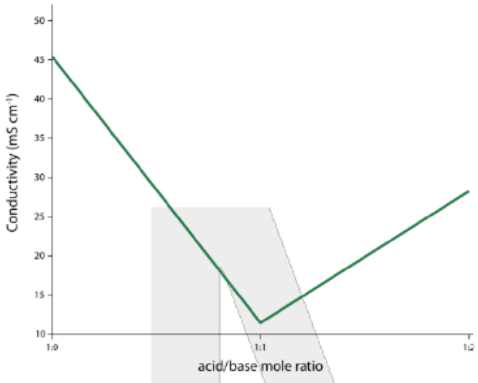
	Solution in conical flask	Solution in burette
(a)	HBr	NH_3
(b)	CH_3COOH	KOH
(c)	CH_3COOH	NH_3
(d)	KOH	HBr

This shape is typical for when a strong acid or base is added to a weak base or acid.

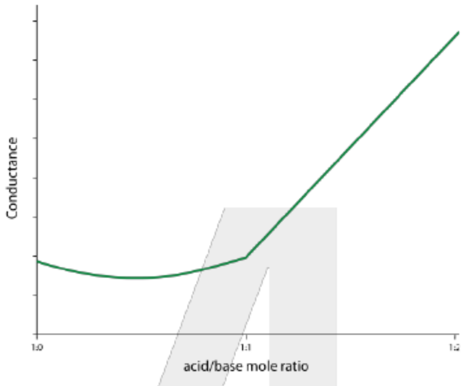
Question 23

Ammonia was added to an aliquot of nitric acid during a titration experiment. Which graph shows how the conductivity of the reaction mixture changed during the course of the titration?

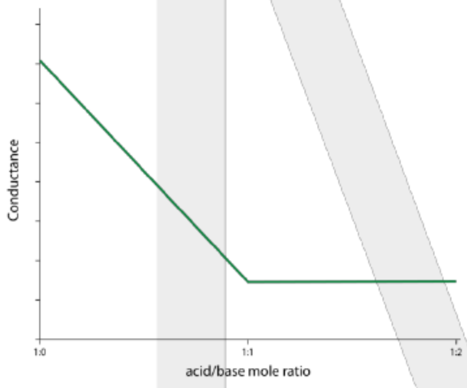
(a)



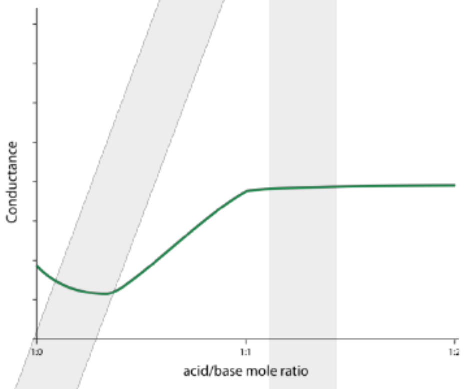
(b)



(c)



(d)



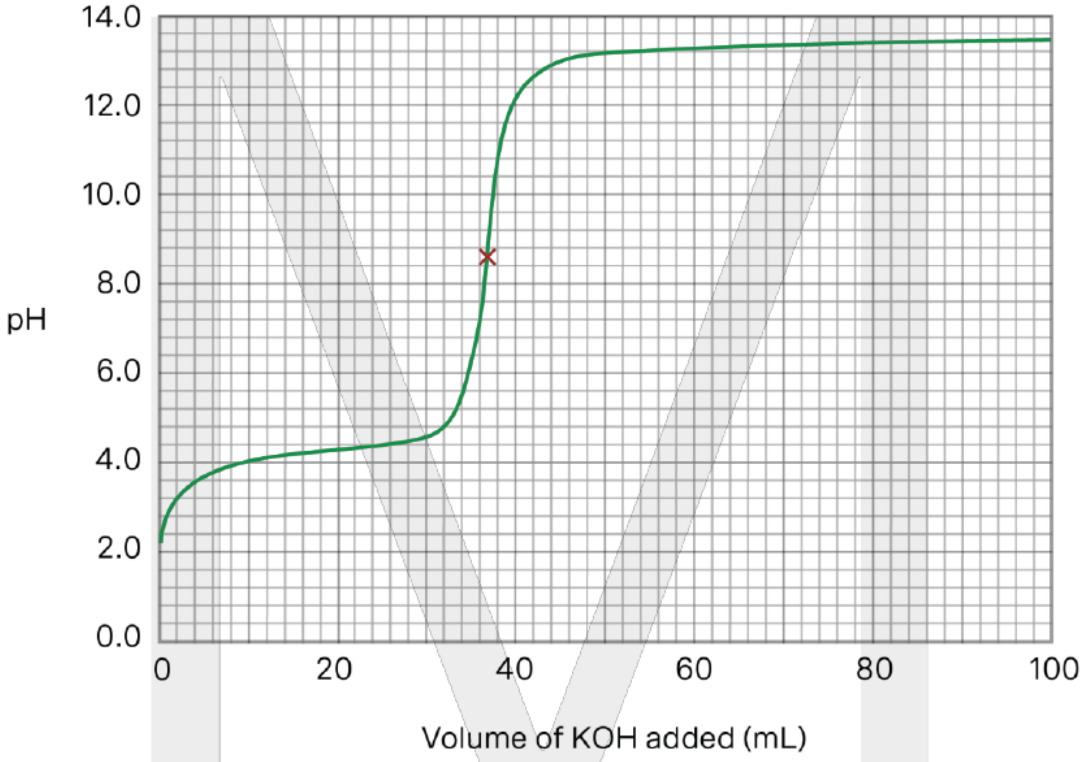
This shape is typical for when a weak base is added to a strong acid.

Question 24

Benzoic acid is a weak monoprotic acid commonly used as a food preservative. Its dissociation is shown below:

$$C_6H_5COOH_{(aq)} \rightleftharpoons C_6H_5COO^-_{(aq)} + H^+_{(aq)}$$

The titration curve from the titration of 25.0 mL of benzoic acid with 0.182 M potassium hydroxide is shown below. The equivalence point is marked.



(a) What volume of potassium hydroxide solution has been added to reach the equivalence point? **1**

37 mL

(b) Calculate the concentration of benzoic acid. **2**

n(KOH) = 0.182 × 0.037 = 6.734×10^{-3} mol

n(C_6H_5COOH) = n(KOH) = 6.734×10^{-3} mol

c(C_6H_5COOH) = 6.734×10^{-3} / 0.025 = 0.27 M (2 s.f.)

(c) Using an equation, explain why the pH at the equivalence point is not neutral. 2

$C_6H_5COOH_{(aq)} + KOH_{(aq)} \longrightarrow KC_6H_5COO_{(aq)} + H_2O_{(l)}$

The only species in the solution at equivalence point is potassium benzoate. The benzoate ion is the conjugate base of a weak acid, therefore it is basic and will accept H^+ from water to produce an excess of OH^-.

$C_6H_5COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons C_6H_5COOH_{(aq)} + OH^-_{(aq)}$

(d) Which indicator is suitable for determining the equivalence point of this titration? Explain. 2

Phenolphthalein (colour change range = pH 8.2–10.0). The indicator's colour change range needs to include the pH of the equivalence point so it can be determined accurately.

(e) The pipette used to transfer the acid was initially rinsed with water, then the wet pipette was used to take the aliquot.

(i) What effect did this error have on the measured concentration of benzoic acid compared with the actual concentration? 1

The water would dilute the acid, so this error would make the measured concentration lower than the real concentration.

(ii) What should have been done instead? 1

The pipette should be rinsed three times with benzoic acid solution before transferring the acid.

(f) A 25.0 mL aliquot of nitric acid with the same concentration as the benzoic acid solution was titrated with the 0.182 M potassium hydroxide solution.

(i) How would the pH of the equivalence point compare with the benzoic acid solution? Explain using a balanced chemical equation. 2

The pH at equivalence point would be ~7 (lower than the benzoic acid titration) as a strong acid/strong base titration results in neutral products.

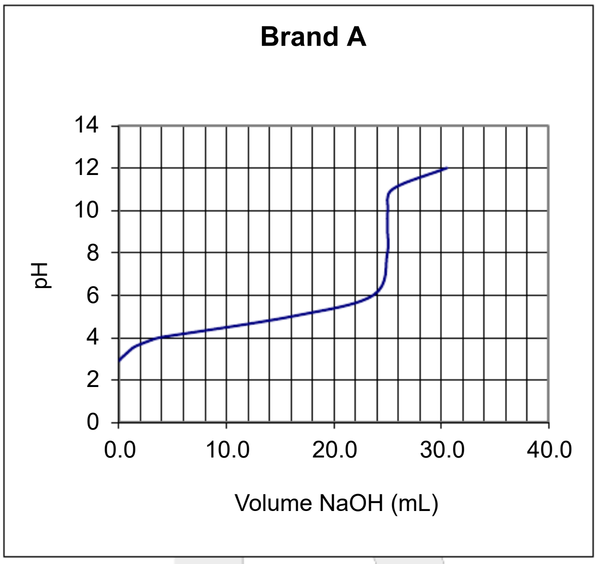
$HNO_{3(aq)} + KOH_{(aq)} \longrightarrow KNO_{3(aq)} + H_2O_{(l)}$

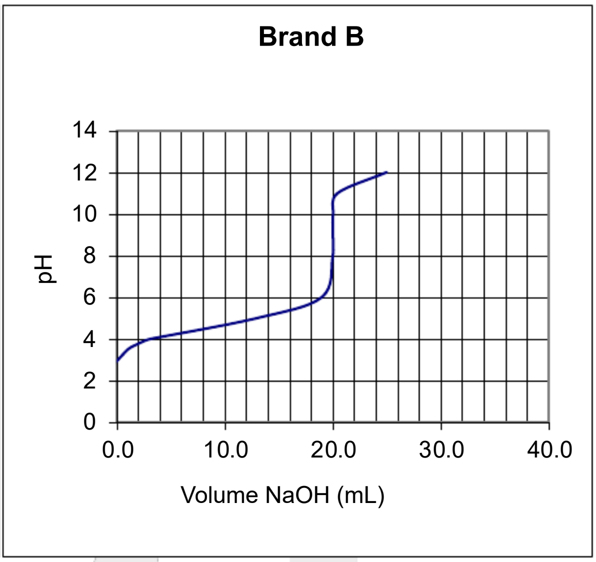
(ii) How would the volume of KOH required to reach the equivalence point compare with the benzoic acid solution? Justify your answer. 2

The volume of KOH required would be the same (37 mL). Since the concentration and volume of nitric acid is the same as that of benzoic acid, and both acids are monoprotic, the moles of ionisable protons will be the same for both acids. Thus, the same quantity of KOH would be required.

Question 25

An investigation to determine the concentration of acetic acid in two brands of vinegar was performed. Standard sodium hydroxide solution was used to titrate equal volumes of Brand A and Brand B. The results of these titrations are shown below.





(a) Describe the procedure that can be used to generate these titration curves. **3**

25 mL of Brand A vinegar can be placed in a conical flask with a digital pH probe, connected to a data logger. NaOH can be added by the millilitre and after thorough stirring, the pH can be recorded and plotted on a graph against volume of NaOH added.

(b) Identify the solution used to rinse the pipette before its final use. **1**

Acid (Brand A or Brand B respectively)

(c) Which brand contains a higher concentration of acetic acid? Explain. **1**

Brand A required a greater volume of NaOH for neutralisation (the most vertical part of the graph) – 25 mL for Brand A vs. 20 mL for Brand B.

(d) Is the salt produced by this neutralisation reaction acidic, basic or neutral? Write an equation to support your answer. **2**

The salt is basic.

$$CH_3COOH_{(aq)} + NaOH_{(aq)} \longrightarrow Na^+_{(aq)} + CH_3COO^-_{(aq)} + H_2O_{(l)}$$

$$CH_3COO^-_{(aq)} + H_2O_{(l)} \rightleftharpoons CH_3COOH_{(aq)} + OH^-_{(aq)}$$

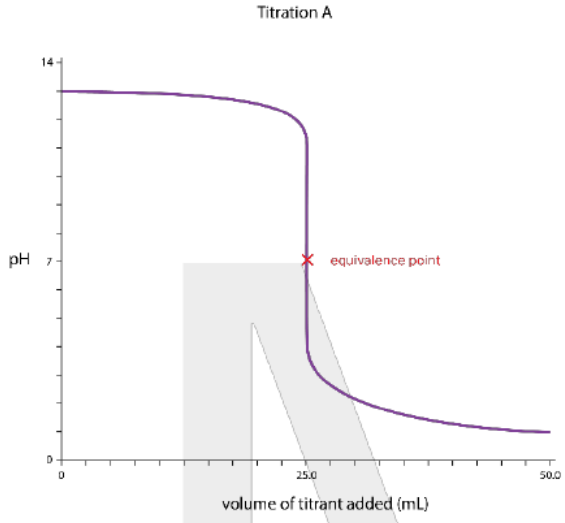
(e) Explain what is meant by the buffering region of a titration curve, and explain why it occurs.

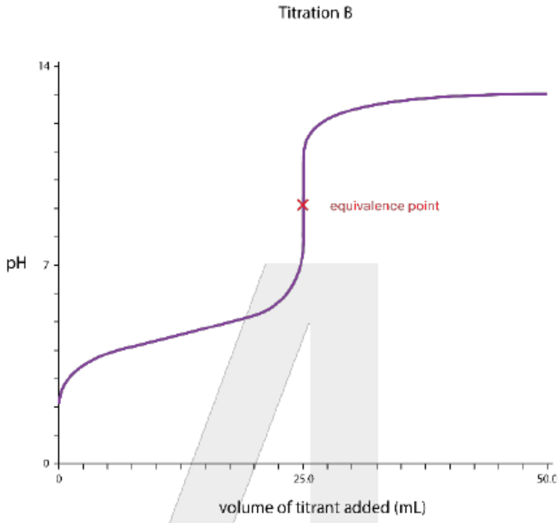
3

The buffering region occurs when the pH changes slowly. This occurs because a buffer is produced when a weak acid or weak base is partially neutralised in the course of a titration. Hence a weak acid and its conjugate base will be present, establishing an equilibrium that can shift to absorb added H^+ or OH^-, hence stabilising pH.

Question 26

The titration curves below show the change in pH for two separate titration experiments.





(a) On each of the graphs above, label the equivalence point. 2

(b) Deduce the strength of both reagents (acid and base) for each titration. Justify your answers. 3

A: Strong base + strong acid. The pH of the equivalence point is neutral, the shape has no buffering region.

B: Weak acid + strong base. The pH of the equivalence point is basic, there is a buffering region at low pH when excess acid is present

(c) For each titration, identify a suitable indicator to use to determine the equivalence point. 2

A: bromothymol blue (or methyl orange or phenolphthalein), B: phenolphthalein

(d) For titration A, identify the solutions that should be used to rinse the conical flask and burette immediately before use. 2

Conical flask: distilled water

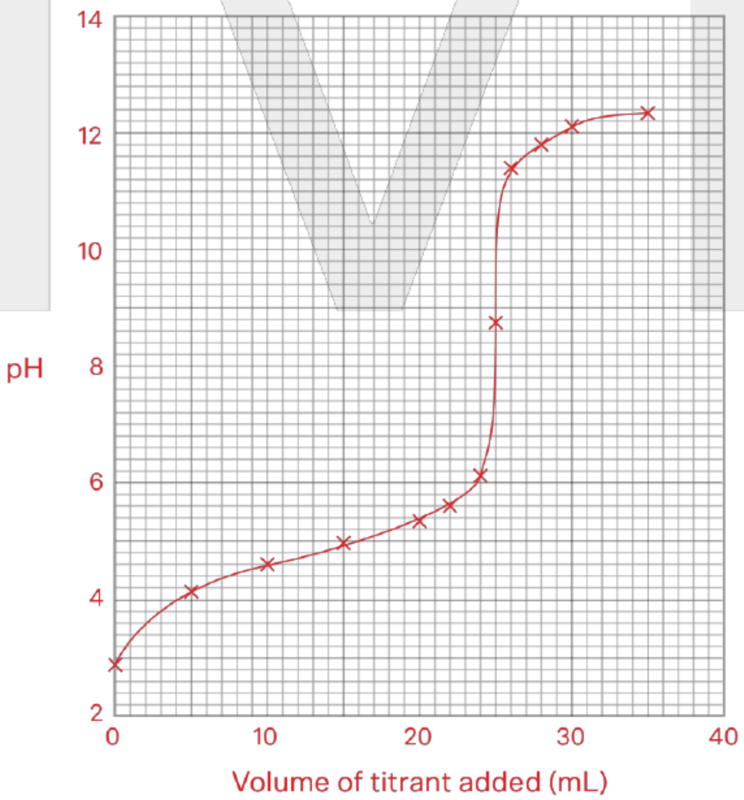
Burette: acid solution

Question 27

A group of students carried out an investigation to determine the concentration of a monoprotic acid. They used a 0.0527 mol L^{-1} standardised solution of $Ca(OH)_2$ to titrate 20 mL of the acid solution. A pH probe and data logger were used. The results are shown below.

Volume of titrant added (mL)	pH in conical flask	Volume of titrant added (mL)	pH in conical flask
0	2.88	24.0	6.13
5.0	4.15	25.0	8.77
10.0	4.59	26.0	11.40
15.0	4.96	28.0	11.80
20.0	5.36	30.0	12.10
22.0	5.60	35.0	12.34

(a) Graph the volume of titrant added against pH for this titration. 3



(b) Determine the volume of titrant required to reach equivalence point. 1

25 mL

(c) Hence calculate the concentration of acid in the solution. 2

n($Ca(OH)_2$) = 0.0527 × 0.025 = 1.3175×10^{-3} mol

n(acid) = 2.635×10^{-3} mol

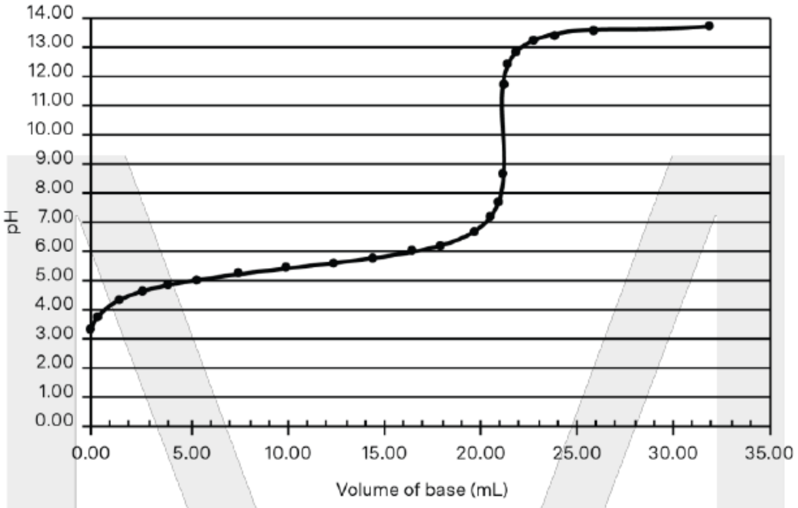
c(acid) = 0.13175 M = 0.13 M (2 s.f.)

(d) Is the acid strong or weak? Provide two reasons to support your answer. 2

Weak. The pH of the equivalence point is basic (pH = 8.77), there is a buffering region at low pH when excess acid is present

Question 28

The graph below shows the change in pH during a titration between an acid and base which react in an equimolar ratio. 25.00 mL of acid was originally transferred to the flask. It had a concentration of 0.100 M.



(a) Describe one method which may be used to measure the pH of the solution during the titration. Assess the accuracy of your chosen method. 3

Using an indicator e.g., phenolphthalein. This method has limited accuracy as indicators change colour over a range of pH values (8.2–10.0 for phenolphthalein), so determination of the equivalence point will not be as accurate as using a pH probe.

(b) Identify a possible base and acid which could have produced this titration curve. 2

Weak acid / strong base e.g., acetic acid / sodium hydroxide

(c) Using the graph, determine the concentration of the base. 2

n(acid) = c × V = 0.100 × 0.025 = 0.002500 mol = n(base) (since reaction is equimolar)

V(base) = 21.00 mL (most vertical section of graph)

c(base) = n / V = 0.0025 / 0.021 = 0.119 M (3 sig. fig.)

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2024 HIGHER SCHOOL CERTIFICATE EXAMINATION

Chemistry

FORMULAE SHEET

$n = \frac{m}{MM}$ $\quad c = \frac{n}{V}$ $\quad PV = nRT$

$q = mc\Delta T$ $\quad \Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ $\quad \text{pH} = -\log_{10}[\text{H}^+]$

$pK_a = -\log_{10}[K_a]$ $\quad A = \varepsilon lc = \log_{10}\frac{I_o}{I}$

Avogadro constant, N_A .... $6.022 \times 10^{23}\ \text{mol}^{-1}$

Volume of 1 mole ideal gas: at 100 kPa and
at 0°C (273.15 K) .... 22.71 L
at 25°C (298.15 K) .... 24.79 L

Gas constant .... $8.314\ \text{J mol}^{-1}\text{K}^{-1}$

Ionisation constant for water at 25°C (298.15 K), K_w .... 1.0×10^{-14}

Specific heat capacity of water .... $4.18 \times 10^3\ \text{J kg}^{-1}\text{K}^{-1}$

DATA SHEET

Solubility constants at 25°C

Compound	K_{sp}	*Compound*	K_{sp}
Barium carbonate	2.58×10^{-9}	Lead(II) bromide	6.60×10^{-6}
Barium hydroxide	2.55×10^{-4}	Lead(II) chloride	1.70×10^{-5}
Barium phosphate	1.3×10^{-29}	Lead(II) iodide	9.8×10^{-9}
Barium sulfate	1.08×10^{-10}	Lead(II) carbonate	7.40×10^{-14}
Calcium carbonate	3.36×10^{-9}	Lead(II) hydroxide	1.43×10^{-15}
Calcium hydroxide	5.02×10^{-6}	Lead(II) phosphate	8.0×10^{-43}
Calcium phosphate	2.07×10^{-29}	Lead(II) sulfate	2.53×10^{-8}
Calcium sulfate	4.93×10^{-5}	Magnesium carbonate	6.82×10^{-6}
Copper(II) carbonate	1.4×10^{-10}	Magnesium hydroxide	5.61×10^{-12}
Copper(II) hydroxide	2.2×10^{-20}	Magnesium phosphate	1.04×10^{-24}
Copper(II) phosphate	1.40×10^{-37}	Silver bromide	5.35×10^{-13}
Iron(II) carbonate	3.13×10^{-11}	Silver chloride	1.77×10^{-10}
Iron(II) hydroxide	4.87×10^{-17}	Silver carbonate	8.46×10^{-12}
Iron(III) hydroxide	2.79×10^{-39}	Silver hydroxide	2.0×10^{-8}
Iron(III) phosphate	9.91×10^{-16}	Silver iodide	8.52×10^{-17}
		Silver phosphate	8.89×10^{-17}
		Silver sulfate	1.20×10^{-5}

Some standard potentials

$K^+ + e^-$	$\rightleftharpoons$	$K(s)$	–2.94 V
$Ba^{2+} + 2e^-$	$\rightleftharpoons$	$Ba(s)$	–2.91 V
$Ca^{2+} + 2e^-$	$\rightleftharpoons$	$Ca(s)$	–2.87 V
$Na^+ + e^-$	$\rightleftharpoons$	$Na(s)$	–2.71 V
$Mg^{2+} + 2e^-$	$\rightleftharpoons$	$Mg(s)$	–2.36 V
$Al^{3+} + 3e^-$	$\rightleftharpoons$	$Al(s)$	–1.68 V
$Mn^{2+} + 2e^-$	$\rightleftharpoons$	$Mn(s)$	–1.18 V
$H_2O + e^-$	$\rightleftharpoons$	$\frac{1}{2}H_2(g) + OH^-$	–0.83 V
$Zn^{2+} + 2e^-$	$\rightleftharpoons$	$Zn(s)$	–0.76 V
$Fe^{2+} + 2e^-$	$\rightleftharpoons$	$Fe(s)$	–0.44 V
$Ni^{2+} + 2e^-$	$\rightleftharpoons$	$Ni(s)$	–0.24 V
$Sn^{2+} + 2e^-$	$\rightleftharpoons$	$Sn(s)$	–0.14 V
$Pb^{2+} + 2e^-$	$\rightleftharpoons$	$Pb(s)$	–0.13 V
$H^+ + e^-$	$\rightleftharpoons$	$\frac{1}{2}H_2(g)$	0.00 V
$SO_4^{2-} + 4H^+ + 2e^-$	$\rightleftharpoons$	$SO_2(aq) + 2H_2O$	0.16 V
$Cu^{2+} + 2e^-$	$\rightleftharpoons$	$Cu(s)$	0.34 V
$\frac{1}{2}O_2(g) + H_2O + 2e^-$	$\rightleftharpoons$	$2OH^-$	0.40 V
$Cu^+ + e^-$	$\rightleftharpoons$	$Cu(s)$	0.52 V
$\frac{1}{2}I_2(s) + e^-$	$\rightleftharpoons$	I^-	0.54 V
$\frac{1}{2}I_2(aq) + e^-$	$\rightleftharpoons$	I^-	0.62 V
$Fe^{3+} + e^-$	$\rightleftharpoons$	Fe^{2+}	0.77 V
$Ag^+ + e^-$	$\rightleftharpoons$	$Ag(s)$	0.80 V
$\frac{1}{2}Br_2(l) + e^-$	$\rightleftharpoons$	Br^-	1.08 V
$\frac{1}{2}Br_2(aq) + e^-$	$\rightleftharpoons$	Br^-	1.10 V
$\frac{1}{2}O_2(g) + 2H^+ + 2e^-$	$\rightleftharpoons$	H_2O	1.23 V
$\frac{1}{2}Cl_2(g) + e^-$	$\rightleftharpoons$	Cl^-	1.36 V
$\frac{1}{2}Cr_2O_7^{2-} + 7H^+ + 3e^-$	$\rightleftharpoons$	$Cr^{3+} + \frac{7}{2}H_2O$	1.36 V
$\frac{1}{2}Cl_2(aq) + e^-$	$\rightleftharpoons$	Cl^-	1.40 V
$MnO_4^- + 8H^+ + 5e^-$	$\rightleftharpoons$	$Mn^{2+} + 4H_2O$	1.51 V
$\frac{1}{2}F_2(g) + e^-$	$\rightleftharpoons$	F^-	2.89 V

Aylward and Findlay, *SI Chemical Data* (5th Edition) is the principal source of data for the standard potentials. Some data may have been modified for examination purposes.