Hale School
Year 12 Semester 1 Examination, 2017

Write your name below:

Solutions

Yr12 ATAR CHEMISTRY
UNIT 3

Circle your teacher's initials:

JWZ JJF KF AD

For Examiners only	
Section 1	
Section 2	
Section 3	
Total	

TIME ALLOWED FOR THIS PAPER

Reading time before commencing: Ten minutes
Working time for paper: Three hours

MATERIAL REQUIRED/RECOMMENDED FOR THIS PAPER

TO BE PROVIDED BY THE SUPERVISOR

This Question/Answer booklet for Sections 1 & 2.
A separate Question/Answer booklet for Section 3.
A separate Multiple Choice Answer sheet for Section 1.
A Chemistry Data Sheet.

TO BE PROVIDED BY THE CANDIDATE

Standard Items: Pens, pencils, eraser, ruler

Special Items: A calculator satisfying the conditions set by the Curriculum Council, and a '2B' pencil for the separate Multiple Choice Answer sheet.

IMPORTANT NOTE TO CANDIDATES

No other items may be taken into the examination room.

It is your responsibility to ensure that you do not have any unauthorised notes or other items of a non-personal nature in the examination room. Please check carefully, and if you have any unauthorised material with you, hand it to the supervisor **BEFORE** reading any further.

INSTRUCTIONS TO CANDIDATES

This paper consists of **THREE SECTIONS** as follows:

SECTION 1 contains **25 questions worth 2 marks each.** It is the multiple choice section.

Answer **ALL** questions in Section 1 on the Separate Multiple Choice Answer Sheet. Use a **'2B' PENCIL. DO NOT USE A BALL POINT OR INK PEN.** If you consider that two or more of the alternative answers are correct then select the BEST alternative. Marks will **NOT** be deducted for incorrect answers. This part is worth 50 marks and should take about 50 minutes.

Do not use pencil for Sections 2 & 3.

SECTION 2 contains **10 short answer questions**. You should answer **ALL** the questions. The answers are to be written in the spaces provided in this Examination booklet. This part is worth 70 marks and should take about 60 minutes.

SECTION 3 contains **5 extended response and calculation questions**. You should answer **ALL** the questions in detail in the **separate question/answer booklet provided.** This part is worth 80 marks and should take about 70 minutes.

At the end of the examination make sure that your **name** is on this Examination paper, the separate Question/Answer Booklet for Section 3 and your Multiple Choice Answer Sheet.

SPECIAL INSTRUCTIONS

Chemical Equations

For full marks, chemical equations should refer only to those species consumed in the reaction and any new species produced. These species may be **ions** [for example $Ag^{+}(aq)$], **molecules** [for example $NH_3(g)$, $NH_3(aq)$, $CH_3COOH(l)$, $CH_3COOH(aq)$] or **solids** [for example $BaSO_4(s)$, $Cu(s)$, $Na_2CO_3(s)$].

Structure of this paper

Section	Number of questions available	Number of questions to be answered	Suggested working time (minutes)	Marks available	Percentage of exam
Section One: Multiple-choice	25	25	50	50	25
Section Two: Short answer	10	10	60	70	35
Section Three: Extended answer	5	5	70	80	40
				Total	100

Section One: Multiple-choice 25% (50 marks)

This section has **25** questions. Answer **all** questions on the separate Multiple-choice Answer Sheet provided. Use a '2B' PENCIL. DO NOT USE A BALL POINT OR INK PEN. If you consider that two or more of the alternative answers are correct then select the BEST alternative. Marks will NOT be deducted for incorrect answers. This part is worth 50 marks and should take about 50 minutes.

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1. The pH of a solution was measured with a pH meter during a titration, and was observed to decrease from 4.0 to 2.0. Which of the following statements about the hydrogen ion concentration in the solution is correct?

(a) It doubled.
(b) It decreased by half.
(c) It increased by a factor of 100.
(d) It decreased by a factor of 100.

2. The following statements refer to the chemical reaction between magnesium carbonate granules, ($MgCO_3$) and a dilute hydrochloric acid solution, ($HC\ell$). Which one of the following statements about this reaction is FALSE?

(a) The rate of the reaction decreases with increasing time.
(b) The rate of reaction increases with increasing initial temperature.
(c) The rate of reaction increases with increasing initial concentration of $HC\ell$ (aq).
(d) The initial rate of reaction is independent of the state of sub-division of $MgCO_3$ (s).

3. Which one of the following statements about the following reversible reaction is TRUE?

$$2SO_2\,(g) + O_2\,(g) \rightleftharpoons 2SO_3\,(g)$$

(a) $K = \dfrac{[SO_2]^2\,[O_2]}{[SO_3]^2}$
(b) K is constant under all reaction conditions.
(c) Sulfur trioxide is being formed when the reaction is at equilibrium.
(d) A catalyst increases the yield of sulfur trioxide by increasing ΔH.

4. In which of the following reactions at equilibrium and at constant temperature is there a shift to the "left" if the pressure of the closed system is increased?

(a) $2NO_2\,(g) \rightleftharpoons N_2O_4\,(g)$
(b) $N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g)$
(c) $H_2O\,(g) + C\,(s) \rightleftharpoons H_2\,(g) + CO\,(g)$
(d) $H_2\,(g) + F_2\,(g) \rightleftharpoons 2HF\,(g)$

5. Which choice correctly describes the properties of aqueous solutions of the following salts?

	Sodium ethanoate ($NaCH_3COO$)	Potassium nitrate (KNO_3)	Ammonium chloride ($NH_4C\ell$)
(a)	neutral	acidic	basic
(b)	basic	neutral	acidic
(c)	acidic	neutral	basic
(d)	basic	acidic	neutral

6. Consider the buffer solution represented by the chemical reaction below:

$$H_2PO_4^-(aq) + H_2O(\ell) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$$

Which of the following would be **true** after the addition of a small volume of 2.0 mol L^{-1} sodium hydroxide solution to the buffer solution?

(a) The forward reaction rate would be unaffected.
(b) The concentration of $H_2PO_4^-$ (aq) present in the system would increase.
(c) The pH of the system would decrease.
(d) The equilibrium would shift to the right.

7. In which one of the following reactions is water **not** acting as a Brønsted-Lowry acid or base?

(a) $HCO_3^-(aq) + H_2O(\ell) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$
(b) $CO_2(s) + H_2O(\ell) \rightleftharpoons H_2CO_3(aq)$
(c) $H_2CO_3(aq) + H_2O(\ell) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$
(d) $CO_3^{2-}(aq) + H_2O(\ell) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$

8. Which one of the following combinations will **not** form a buffer solution?

(a) $HNO_3(aq)$ / $NO_3^-(aq)$
(b) $H_2PO_4^-(aq)$ / $HPO_4^{2-}(aq)$
(c) $NH_3(aq)$ / $NH_4C\ell$ (aq)
(d) $H_3PO_4(aq)$ / $H_2PO_4^-(aq)$

Questions 9, 10 and 11 relate the following information:

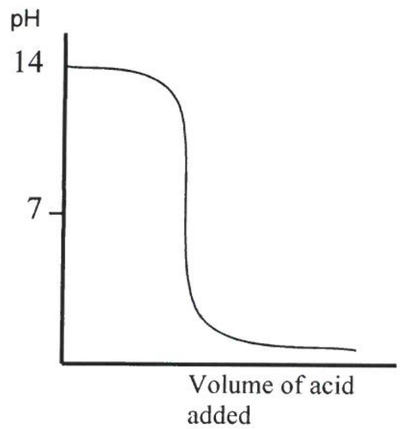
A student was asked to determine the concentration of a solution of ethanoic acid that had a concentration of approximately 0.400 mol L^{-1}. He pipetted 20.0 mL of a 0.500 mol L^{-1} solution of sodium hydroxide into a conical flask, and titrated the ethanoic acid against the standardised sodium hydroxide solution, using phenolphthalein as the indicator.

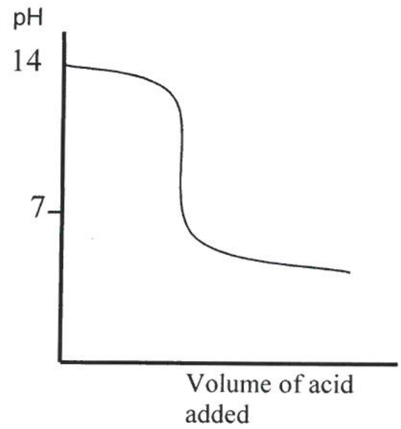
9. What is the pH of the sodium hydroxide solution at the start of the titration?

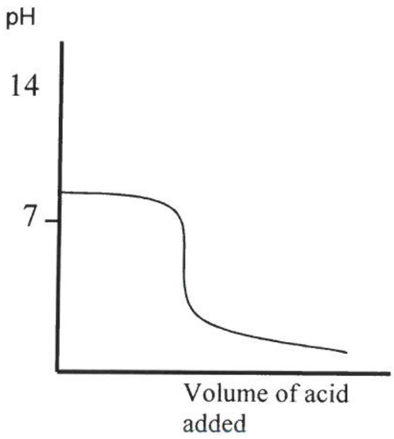
(a) 13.7
(b) 7.00
(c) 14.0
(d) 12.7

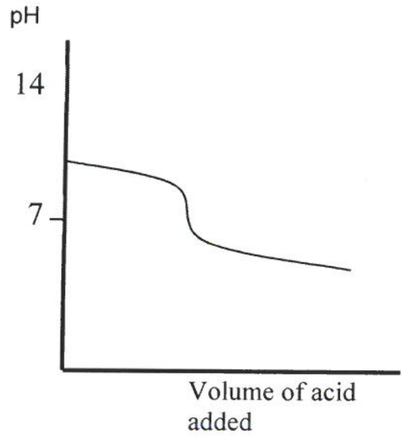
$[OH^-] = 0.500$
$[H^+] = 2 \times 10^{-14}$
$pH = -\log[\downarrow] =$

10. If the ethanoic acid was added until it was slightly in excess, which of the following pH graphs would show the variation of pH during the titration?

(a)


(c)


(b)


(d)


11. What approximate volume of ethanoic acid would the student expect to have added at the end point of the titration?

(a) 20 mL
(b) 30 mL
(c) 25 mL
(d) 35 mL

$n = cV$
$= 0.5 \times 0.02$
$= 0.01$
$V = \frac{n}{c}$
$= 0.025$

6

12. Which one of the following correctly arranges 1.0 mol L^{-1} solutions of the substances in order of increasing pH?

(a)	NaOH	$NaCH_3COO$	NH_4CH_3COO	H_2CO_3	HCℓ
(b)	H_2CO_3	HCℓ	NH_4CH_3COO	$NaCH_3COO$	NaOH
(c)	HCℓ	H_2CO_3	$NaCH_3COO$	NH_4CH_3COO	NaOH
(d)	HCℓ	H_2CO_3	NH_4CH_3COO	$NaCH_3COO$	NaOH

13. Consider the table below showing the colour changes of a range of indicators.

		Colour change	
Indicator	**pH range**	**Acid**	**Alkali**
methyl orange	3.1–4.4	red	yellow
bromothymol blue	6.0–7.6	yellow	blue
phenolphthalein	8.3–10.0	colourless	pink
litmus	5.5–8.8	red	blue

Which one of the following indicators should be used when titrating a solution of ammonia with hydrochloric acid?

(a) phenolphthalein
(b) methyl orange
(c) bromothymol blue
(d) litmus

14. During an acid-base titration, which one of the following would result in a random error?

(a) Inconsistent reading of the volume of solution in a burette.
(b) Using the incorrect mass when making up a standard solution.
(c) Rinsing the conical flask with distilled water before use.
(d) Rinsing the pipette with distilled water immediately before use.

15. Consider the equilibrium system below:

$$H_2O(\ell) + HSO_4^-(aq) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$$

Which one of the following statements is **false**?

(a) HSO_4^- is the conjugate acid of SO_4^{2-}.
(b) HSO_4^- is acting as an acid.
(c) Addition of base will favour the reverse reaction.
(d) The system could act as a buffer.

16. An experiment is set up to electroplate an antique brass spoon with silver. Which of the following statements describes how the experiment should be set up?

(a) The cathode is made of silver and the spoon is the anode.
(b) The spoon is the cathode and the electrolyte is a solution of silver nitrate.
(c) The spoon is the anode and the electrolyte is a solution of copper sulfate.
(d) The cathode is made of silver and the electrolyte is a solution of silver nitrate.

17. How many moles of electrons are required when the following half-equation is balanced using the smallest possible coefficients?

$$I_2\,(s) \;+\; H_2O\,(\ell) \;\rightleftharpoons\; IO_3^-\,(aq) \;+\; H^+\,(aq) \;+\; e^-$$

(a) 2
(b) 5
(c) 10
(d) 12

$I_2 + 6H_2O \rightarrow 2IO_3^- + 12H^+ + 10e^-$

18. Which one of the following statements **BEST** describes the function of an anode in an electrolytic cell?

(a) The anode is the electrode at which reduction occurs.
(b) The anode is the only electrode at which OH^- (aq) ions are produced.
(c) The anode is the electrode which attracts positive ions.
(d) The anode is the electrode that is oxidised.

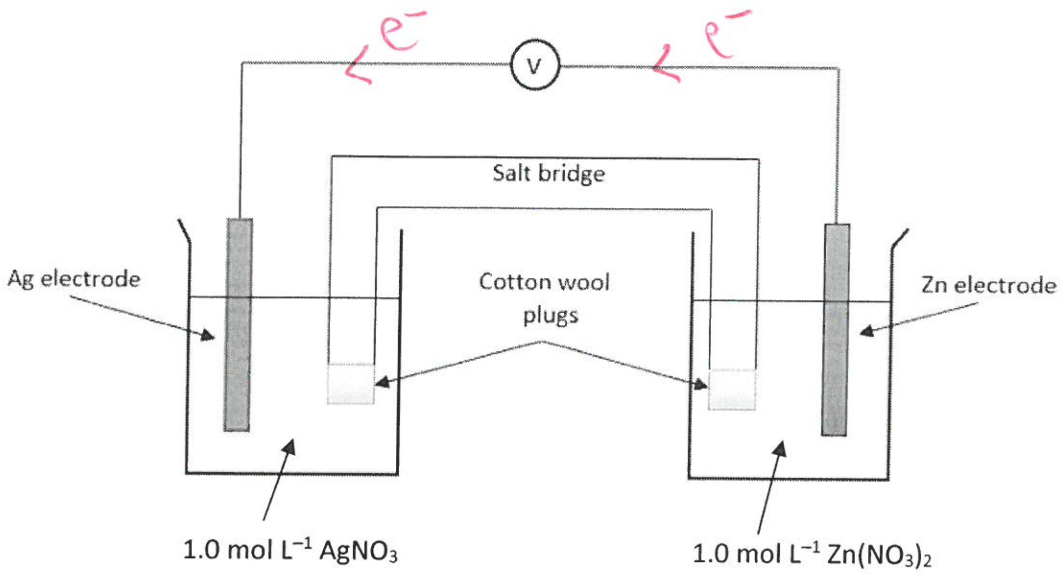
19. In which one of the following processes is the oxidation number of chlorine being increased by 2?

(a) $PC\ell_3 \;+\; C\ell_2 \;\rightarrow\; PC\ell_5$
(b) $C\ell_2 \;+\; H_2O \;\rightarrow\; C\ell^- \;+\; HC\ell O \;+\; H^+$
(c) $2\,C\ell^- \;+\; 2\,e^- \;\rightarrow\; C\ell_2$
(d) $C\ell O_3^- \;+\; H_2O_2 \;\rightarrow\; C\ell O_4^- \;+\; H_2O$
+5 +7

20. Which one of the following pairs of chemicals will undergo a spontaneous redox reaction?

(a) solid iodine and water
(b) solid lead and liquid bromine
(c) lead nitrate solution and potassium iodide solution.
(d) solid copper and a 1.0 mol L^{-1} solution of hydrochloric acid.

Questions 21, 22 and 23 relate to the following electrochemical cell at 25°C:



21. Which of the following reactions will occur during the normal operation of this cell?

(a)	$2Ag^+ (aq)$	+	$Zn (s)$	⟶	$2Ag (s)$	+	$Zn^{2+} (aq)$	$E° = 1.56$ V
(b)	$2Ag^+ (aq)$	+	$Zn (s)$	⟶	$2Ag (s)$	+	$Zn^{2+} (aq)$	$E° = 0.04$ V
(c)	$Zn^{2+} (aq)$	+	$2Ag (s)$	⟶	$Zn (s)$	+	$2Ag^+ (aq)$	$E° = 1.56$ V
(d)	$Zn^{2+} (aq)$	+	$2Ag (s)$	⟶	$Zn (s)$	+	$2Ag^+ (aq)$	$E° = 0.04$ V

22. Which of the following statements about the two electrodes is correct?

(a) The mass of the silver electrode will decrease.
(b) The zinc electrode is the cathode.
(c) The mass of the zinc electrode will decrease.
(d) The silver electrode is the anode.

23. Which of the following statements about the flow of charge is INCORRECT?

(a) Electrons will flow from the zinc electrode to the silver electrode through the external circuit.
(b) Cations will flow through the salt bridge towards the silver half-cell.
(c) Electrons will flow from the silver electrode to the zinc electrode through the salt bridge.
(d) Anions will flow through the salt bridge towards the zinc half-cell.

24. The overall redox reaction occurring in a dry cell, (Leclanché cell), is shown below.

$$Zn\,(s) + 2\,NH_4^+\,(aq) + 2\,MnO_2\,(s) \longrightarrow Zn^{2+}(aq) + Mn_2O_3\,(s) + H_2O\,(\ell) + 2NH_3\,(aq)$$

Which of the following statements regarding the dry cell are correct?

I The zinc outer casing is acting as the anode. ✓
II The oxidation state of manganese decreases from +4 to +3. ✓
III Ammonium chloride acts as an electrolyte for the cell. ✓

(a) I and III only.
(b) I and II only.
(c) II and III only.
(d) I, II and III.

25. Consider the following statements about fuel cells.

I A fuel cell converts chemical energy to electrical energy via a redox reaction. ✓
II Fuel cell technology involves the continuous supply of reactants to the cells and the continuous removal of the products. ✓
III A fuel cell can be recharged by reversing the direction of current flow through the cell. ✗
IV Fuel cells are considered a low-emission technology. ✓

Which of the above statements about fuel cells are true?

(a) I only
(b) I and II
(c) I, III and IV
(d) I, II and IV

End of section one

Section Two: Short answer **35% (70 Marks)**

This section has **ten (10)** questions. Answer **all** questions. Write your answers in the spaces provided.
When calculating numerical answers, show your working or reasoning clearly. Express numerical answers to the appropriate number of significant figures and include appropriate units where applicable.
Spare pages are included at the end of this booklet. They can be used for planning your responses and/or as additional space if required to continue an answer.

- Planning: If you use the spare pages for planning, indicate this clearly at the top of the page.
- Continuing an answer: If you need to use the space to continue an answer, indicate in the original answer space where the answer is continued, i.e. give the page number. Fill in the number of the question that you are continuing to answer at the top of the page.

Suggested working time: 60 minutes.

Question 26 **(4 marks)**

Write observations for any reactions that occur in the following procedures. In each case describe in full what you would observe, including any:

- colours
- odours
- precipitates (give the colour)
- gases evolved (give the colour or describe as colourless).

If no change is observed, you should state this.

(Note: No chemical equations necessary).

(a) Some hydrochloric acid solution is mixed with solid sodium carbonate. (2 marks)

White solid dissolves in a clear, colourless solution (1)
Bubbles of a colourless, odourless gas and clear, colourless solution. (1)

(b) Some solid copper (II) hydroxide is mixed with a dilute nitric acid solution. (2 marks)

Blue solid dissolves in a clear, colourless solution (1)
Produces a blue solution (1)

Question 27 **(4 Marks)**

The following chemical equation represents an unbalanced redox reaction.

$$MnO_4^- (aq) + C_2O_4^{2-} (aq) \rightarrow Mn^{2+} (aq) + CO_2 (g)$$

In the appropriate spaces below, write the two separate half-equations and the overall balanced redox equation. (4 marks)

Oxidation: $(C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-) \times 5$ (1)

Reduction: $(MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O) \times 2$ (1

Overall Redox: $5C_2O_4^{-} + 2MnO_4^- + 16H^+ \rightarrow 10CO_2 + 2Mn^{2+} + 8H_2O$ (2

Question 28 **(6 marks)**

Bromine water, which is a dilute aqueous solution of bromine in water, is slightly acidic because of its reaction with water, represented by the following equation:

$$Br_2 (aq) + H_2O (\ell) \rightleftharpoons HBrO (aq) + H^+ (aq) + Br^- (aq)$$

In aqueous solution, bromine, Br_2 (aq) is brown. Hypobromous acid, HBrO (aq), and bromide ions, Br^- (aq) are both colourless.

State and explain the colour changes that would be observed, if the following changes are made to the system at equilibrium.

(a) Addition of NaOH (aq). (3 marks)

Colour: Brown colour fades (1)

Explanation: addition of OH^- causes a decrease in the $[H^+]$. (1)
The rate of collision of reactants is greater than that of products.
The fwd reaction increases relative to the reverse reaction. (1)
Equilibrium shifts right and $[Br_2]$ decreases.

(b) Addition of excess HCℓ (aq). (3 marks)

Colour: Brown colour becomes more intense (1)

Explanation: addition of HCl causes an increase in the $[H^+]$ (1)
The rate of collision of products is higher than that of reactants.
The reverse reaction increases relative to the forward reaction (1)
Equilibrium shifts left and $[Br_2]$ increases.

Question 29 **(10 Marks)**

Explain how you could distinguish between the following pairs of compounds using chemical tests. For each test, write one equation for a reaction that occurred.

	Compounds	Description of Test	Observations
(a)	$NH_4Cl(aq)$	Add an aqueous solution of OH^-, PO_4^{3-}, CO_3^{2-} (1)	with $NH_4Cl(aq)$: no reaction (1)
	$MgCl_2(aq)$		with $MgCl_2(aq)$: white ppt. (1)
	Equation: $Mg^{2+}_{(aq)} + 2OH^-_{(aq)} \rightarrow Mg(OH)_{2(s)}$ (2)		
(b)	Sn(s)	Add an aqueous solution of Fe^{2+} / Ni^{2+} / Co^{2+} or etc.. (1)	with Sn(s): no reaction (1)
	Zn(s)		with Zn(s): zinc dissolves / coated with black/grey layer / colour of solution fades. (1)
	Equation: $Fe^{2+}_{(aq)} + Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + Fe_{(s)}$ (2)		

Question 30 **(13 marks)**

This question relates to the following reaction:

$$2\ NO_2(g) \rightleftharpoons N_2O_4(g) + \text{HEAT}$$

The gases are contained in a sealed vessel with a fixed volume, which is strong enough to resist changes in pressure and temperature without breaking.

(a) Complete the table by using Le Chatelier's principle or collision theory, as appropriate, to predict the effect on the position of equilibrium (shifted to the right/left/unchanged) of the following imposed changes once equilibrium is re-established. In each case explain your reasoning. One has been partially completed for you. (8 marks)

Imposed change	Effect on equilibrium position	Reasoning
Increased temperature	Shifted to the left	**The system opposes the increase in temperature by favouring the reverse endothermic reaction, which will decrease the temperature of the system.**
Additional nitrogen dioxide gas is added to the system	**Shifted to the right**	**The system opposes the increase in concentration of NO_2 by favouring the forward reaction.** **Or (better):** **The rate of the forward reaction is increased due to the increase in concentration of the NO_2 (higher rate (frequency) of collision of NO_2 molecules) and so the greater formation of N_2O_4. The rate of the reverse reaction is not initially affected.**
Neon gas is added to the system (fixed volume)	**No change**	**The addition of the neon (an inert gas) does not affect the concentrations / partial pressures of the reacting gases therefore there is no change to the position of equilibrium / rate of the forward and reverse reactions.**

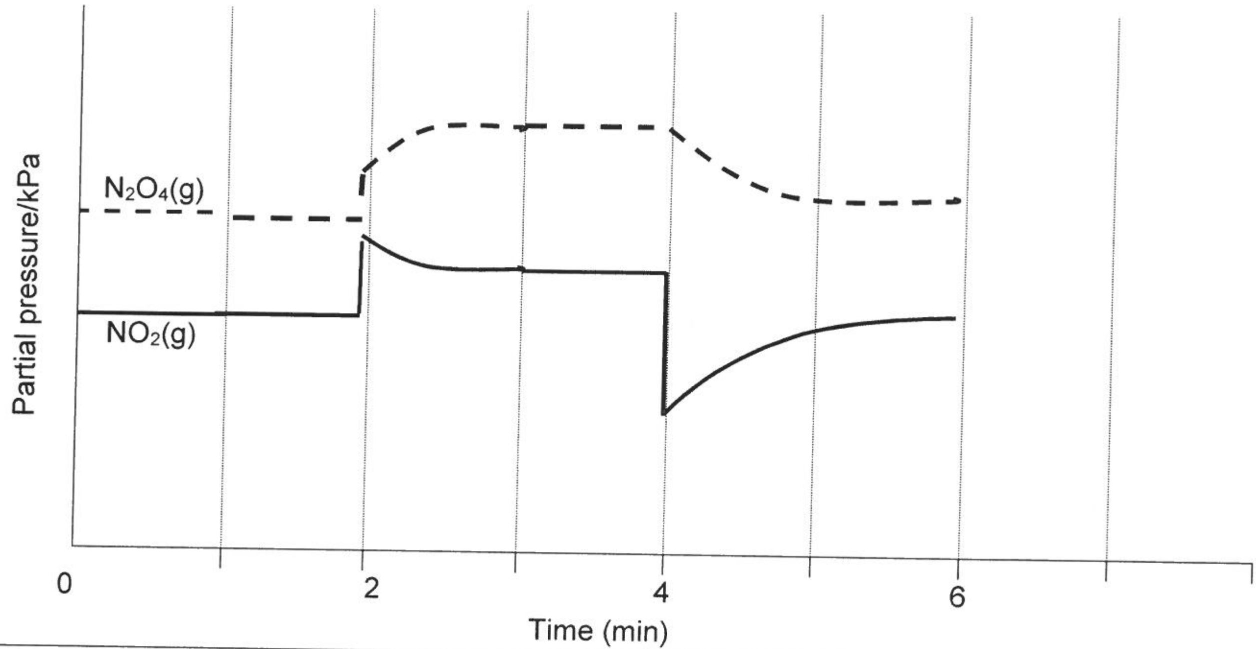
Description	Marks
Predictions	0–2
Explanations (2 marks each). **Note**: in the second scenario award only 1 mark for reference to LCP alone	0–6
Total	**8**

(b) A mixture of the gases was place into a syringe that allowed for the volume of the system to be changed. The partial pressures of the gases were measured and plotted on the graph below.

$$2\ NO_2(g) \rightleftharpoons N_2O_4(g) + \text{HEAT}$$

- After 2 minutes, the volume of the syringe was reduced. The system took one minute to re-establish equilibrium.
- 1 minute after this time, approximately half of the NO_2 was removed from the system. This time it took 2 minutes to return to a state of equilibrium.

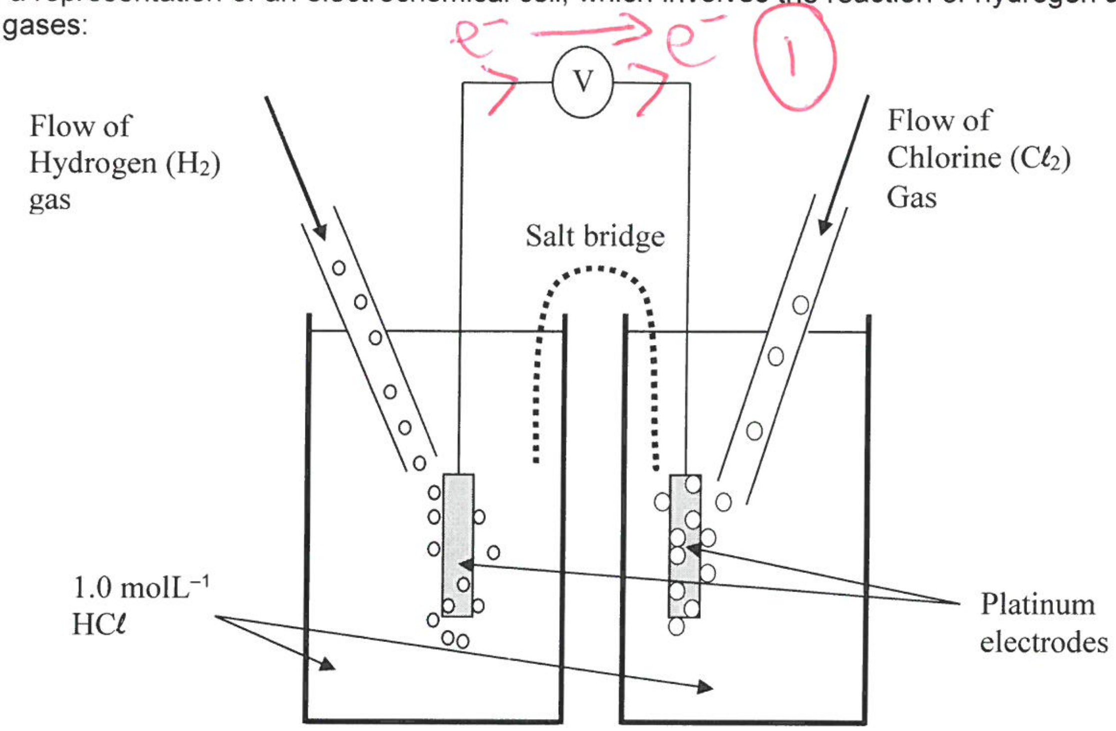
Complete the following graph showing the changes to the partial pressures of nitrogen dioxide and dinitrogen tetroxide during this time. (5 marks)



Description	Marks
Initial change at 2 minutes (both partial pressures increase)	1
NO_2 decrease as N_2O_4 increase	1
NO_2 instantaneous drop at 4 mins but no instantaneous change to N_2O_4	1
NO_2 rising between 4 and 6 minutes, N_2O_4 falling	1
NO_2 changing half as much as N_2O_4 between 2 and 3, and 4 and 6 min	1
Total	**5**

Question 31 **(6 Marks)**

Below is a representation of an electrochemical cell, which involves the reaction of hydrogen and chlorine gases:



(a) Give the half equation for the reactions occurring at the anode and at the cathode, and then write an overall balanced redox equation for the reaction occurring in the cell.

Cathode half-equation: $Cl_2 + 2e^- \rightarrow 2Cl^-$ (1) $E^\circ = +1.36V$
Anode half-equation: $H_2 \rightarrow 2H^+ + 2e^-$ (1) $E^\circ = 0.00V$
Overall equation: $Cl_2 + H_2 \rightarrow 2Cl^- + 2H^+$ (1)

(3 marks)

(b) Using the standard reduction potential values from the data sheet, calculate the maximum theoretical voltage (e.m.f.) that could be produced by this cell.

(1 mark)

+1.36V (1)

(c) Show the direction of the flow of electrons in the external circuit by means of an **arrow**, "(⟶)" in the diagram above.

(1 mark)

(d) Suggest a reason why platinum is used for the electrodes. (1 mark)

Platinum is inert so will not take part in reaction

OR Allows for electron transfer into half cell.

Question 32 **(9 Marks)**

Use the Standard Reduction Potentials from your Data Booklet to answer the following questions. In each case, write all relevant half-equations with their respective E° values. (If the reaction is likely to occur, write an overall balanced redox equation with the resultant cell voltage). Then you must state clearly if the reaction is likely or unlikely to occur as described.

(a) A piece of aluminium metal is placed in a 1.00 mol L^{-1} nickel nitrate solution. (3 marks)

$2\times(Al \rightarrow Al^{3+} + 3e^-)$ $E^\circ = +1.68V$

$3\times(Ni^{2+} + 2e^- \rightarrow Ni)$ $E^\circ = -0.24V$ ①

$2Al + 3Ni^{2+} \rightarrow 2Al^{3+} + 2Ni$ $E_{cell} = +1.44V$ ①

positive EMF / E_{cell} ∴ will occur. ①

(b) Silver metal is added to a 1.00 mol L^{-1} sulfuric acid solution. (3 marks)

$2\times(Ag \rightarrow Ag^+ + e^-)$ $E^\circ = -0.80V$

$2H^+ + 2e^- \rightarrow H_2$ $E^\circ = 0.00V$ ①

Not reqd. as won't occur → $2Ag + 2H^+ \rightarrow 2Ag^+ + H_2$ $E_{cell} = -0.80V$ ①

negative EMF / E_{cell} ∴ won't occur ①

(c) Iron(II) nitrate solution is added to a solution of hydrogen peroxide. (3 marks)

$2\times(Fe^{2+} \rightarrow Fe^{3+} + e^-)$ $E^\circ = -0.77V$

$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$ $E^\circ = +1.76V$ ①

$2Fe^{2+} + H_2O_2 + 2H^+ \rightarrow 2Fe^{3+} + 2H_2O$ $E_{cell} = 0.99V$ ①

positive EMF / E_{cell} ∴ will occur ①

Question 33 **(8 marks)**

The two initial reactions involved in the wet corrosion of iron are shown below. Both of these reactions are reversible.

Oxidation (anode) reaction: $Fe(s) \rightleftharpoons Fe^{2+}(aq) + 2e^-$

Reduction (cathode) reaction: $O_2(g) + 2H_2O(\ell) + 4e^- \rightleftharpoons 4OH^-(aq)$

(a) Combine these two half equations to write a balanced overall redox reaction for the process.

$2Fe_{(s)} + O_{2(g)} + 2H_2O_{(l)} \rightleftharpoons 2Fe^{2+}_{(aq)} + 4OH^-_{(aq)}$

species (1) balanced (1)

(2 marks)

(b) Use your knowledge of chemical equilibrium to predict whether iron will corrode more quickly in water that is slightly acidic or water that is slightly basic. Justify your reasoning.

More quickly in acidic. (1)
↑ [H⁺] will react with the OH⁻ which will increase the rate of the fwd reaction rate relative to the reverse, favouring the fwd reaction. (1)
This will in turn increase the rate of the oxidation reaction (1)

(3 marks)

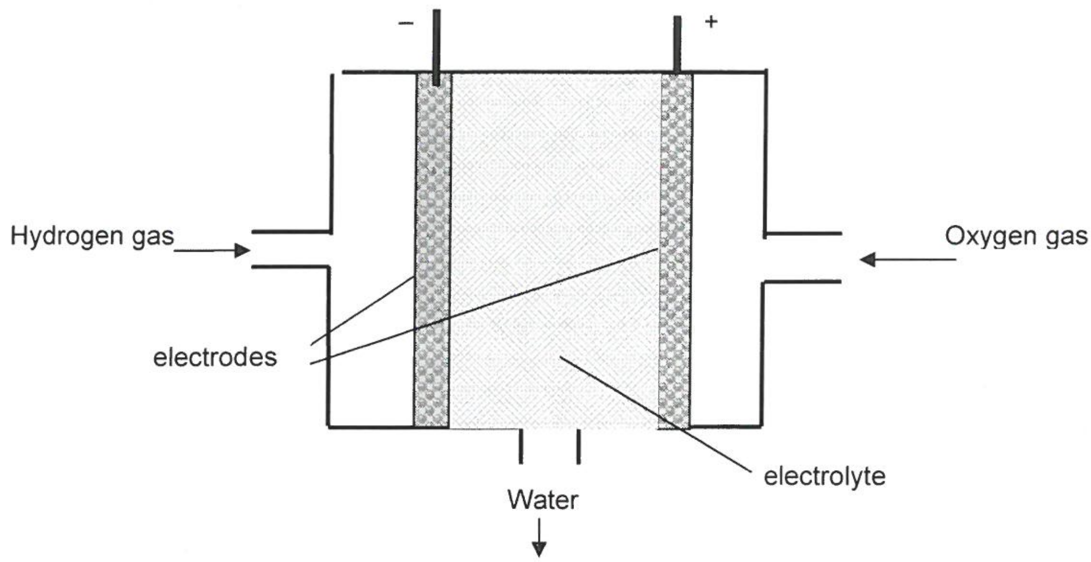
(c) Water tanks made of steel (an alloy containing a mixture of iron and carbon), can be protected from corrosion by connecting a 'sacrificial' anode made of zinc to the inside of the tank. With reference to the standard reduction potentials of iron and zinc, explain how this reduces the corrosion of iron in the tank.

The standard reduction potential for zinc (−0.76) is more negative than for iron (−0.44)
So zinc is more readily oxidised than iron (1)
So it is oxidised in preference to iron. (1)

(3 marks)

Question 34 **(5 marks)**

The following diagram represents a hydrogen fuel cell.



(a) Write the overall redox equation, including state symbols, for the reaction occurring in this cell. (2 marks)

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

① Balanced eqn ① state symbols

(b) Explain the role of the electrolyte in this fuel cell. (2 marks)

To allow charge / current / ions / hydrogen ions to move between the anode and cathode. ②

(c) Explain one environmental advantage of using fuel cells to power vehicles compared to using a fossil fuel. (1 mark)

Fuel cells only produce water as waste whereas fossil fuels release CO_2 / greenhouse gases ①

Question 35 **(5 marks)**

(a) Calculate the pH of a solution of 0.0200 mol L^{-1} sodium hydroxide.

$[OH^-] = 0.0200$ mol L^{-1}
$[H^+] = 1.00 \times 10^{-14} \div 0.0200 = 5.00 \times 10^{-13}$ mol L^{-1} (1)
$pH = -\log[H^+] = -\log(5.00 \times 10^{-13})$
$= 12.3$ (1)

(2 marks)

(b) 30.0 mL of 0.0200 mol L^{-1} sulfuric acid was added to 30.0 ml of the sodium hydroxide described in part (a). Calculate the pH of the resulting solution (assume that the sulfuric acid fully ionizes).

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$
$n(OH^-) = c.V = 0.0200 \times 0.030 = 6.00 \times 10^{-4}$ mol (½)
$n(H_2SO_4) = c.V. = 0.0200 \times 0.030 = 6.00 \times 10^{-4}$ mol (½)
$n(H^+) = 2 \times n(H_2SO_4) = 1.20 \times 10^{-3}$ mol (½)
$n(H^+)_{excess} = 1.20 \times 10^{-3} - 6.00 \times 10^{-4}$
$= 0.0006$ mol (½)
$[H^+] = n/V = 0.0006/0.060 = 0.01$ mol L^{-1} (½)
$pH = -\log[H^+] = -\log[0.01] = 2$ (½)

(3 marks)

End of section two

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Semester One Examination, 2017

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Part 3	

Section 3 Question and Answer Booklet.

Hale School
Semester One Examination, 2017

Write your name below:

____

Yr12 ATAR CHEMISTRY
UNIT 3

JWZ AD JJF KF

For Examiners only	
Part 3	

Section 3 Question and Answer Booklet.

Section Three: Extended answer **40% (80 marks)**

This section contains five **(5)** questions. You must answer **all** questions. Write your answers in the spaces provided.

Where questions require an explanation and/or description, marks are awarded for the relevant chemical content and also for coherence and clarity of expression. Lists or dot points are unlikely to gain full marks.

Final answers to calculations should be expressed to the **appropriate number** of significant figures.

Additional working space pages at the end of this Question/Answer booklet are for planning or continuing an answer. If you use these pages for planning, indicate at the original answer, the page number it is planned/continued on and write the question number being planned/continued on the additional working space page.

Suggested working time: 70 minutes.

Question 36 **(19 marks)**

Human blood contains a buffer of carbonic acid (H_2CO_3) and bicarbonate anion (HCO_3^-) in order to maintain blood pH between 7.35 and 7.45. This buffer can be represented as shown below.

$$H_3O^+(aq) + HCO_3^-(aq) \rightleftharpoons H_2CO_3(aq) + H_2O(\ell)$$

(a) Using the example above,

(i) describe the action of a buffer. (2 marks)

Description	Marks
A buffer maintains a steady pH	1
When a small amount of acid or base is added to a solution, a chemical reaction with components of the buffer occur	1
Total	**2**

(ii) explain the term "buffering capacity". (3 marks)

Description	Marks
A buffer contains a certain amount of the conjugate base and the weak acid	1
When either of these are used up the solution ceases to function as a buffer (it has reached the limit of its buffering capacity)	1
The amount of these ions present in the buffer solution determines the buffering capacity	1
Total	**3**

(b) Explain, using equations, why this system buffers the blood. (4 marks)

Description	Marks
When extra hydrogen ions are added (pH decreases), the forward reaction is favoured $H_3O^+(aq) + HCO_3^-(aq) \rightarrow H_2CO_3(aq) + H_2O(\ell)$	1
This opposes the increase in H_3O^+ ions	1
When extra hydroxide ions are added (pH increases), the reverse reaction is favoured: $H_2CO_3(aq) + OH^-(aq) \rightarrow H_2O(aq) + HCO_3^-(aq)$	1
This opposes the increase in OH^- ions	1
Total	**4**

The carbonic acid can also decompose into carbon dioxide gas and water, resulting in a second equilibrium system between carbonic acid and water:

$$H_2CO_3(aq) \rightleftharpoons H_2O(\ell) + CO_2(g)$$

(c) Use this equation, and the earlier equation that represents the buffer, to explain the changes in blood pH caused by an increase in the concentration of carbon dioxide in the blood. (4 marks)

Description	Marks
When $[CO_2]$ increases the $[H_2CO_3]$ increases	1
As the reverse reaction is favoured	1
This causes the reverse reaction here to be favoured: $H_3O^+(aq) + HCO_3^-(aq) \rightleftharpoons H_2CO_3(aq) + H_2O(\ell)$	1
This increasing the $[H_3O^+]$ and reducing the pH of the blood	1
Total	**4**

(d) In human blood, oxygen is transported by combining with haemoglobin according to the following reversible reaction to form oxyhaemoglobin that travels around the body:

HAEMOGLOBIN(aq) + $O_2(g) \rightleftharpoons$ OXYHAEMOGLOBIN(aq)

"An increase in the acidity of blood causes this equilibrium to shift towards the left, thus releasing oxygen from the oxyhaemoglobin molecules, which allows for cell respiration to provide energy". Suggest, with an explanation, how an increase in carbon dioxide levels during exercise would affect the concentration of haemoglobin in blood. (2 marks)

Description	Marks
During exercise, blood CO_2 levels increase will decrease blood pH and thereby cause an equilibrium shift to de-oxygenate oxyhaemoglobin (i.e. equilibrium shift to the left)	1
Therefore the concentration of haemoglobin in blood will increase	1
Total	**2**

(e) Buffers can be used to maintain pH in a range of situations, including for scientific research. For example, when preparing biological samples for electron microscopy a buffer of the monoprotic cacodylic acid ($C_2H_7AsO_2$) and its conjugate base is used.

Write an equation to show the equilibrium system for this buffer and describe how the system will respond to an **increase** in pH. (4 marks)

Description	Marks
$H_3O^+(aq) + C_2H_6AsO_2^-(aq) \rightleftharpoons C_2H_7AsO_2(aq) + H_2O(\ell)$	1
When the pH increases the $[OH^-]$ increases	1
The extra OH^- ions react with the cacodylic acid: $C_2H_7AsO_2(aq) + OH^-(aq) \rightarrow H_2O(aq) + C_2H_6AsO_2^-(aq)$	1
Which opposes the increase in pH	1
Total	**4**

Question 37 **(14 marks)**

The electroplating of various metals plays an extremely important role in industry. These reactions can be carried out on a small scale in the laboratory using standard laboratory equipment. A typical spoon can be chrome electroplated utilising a chromium electrode and an acidified aqueous chromium nitrate solution. Using a labelled diagram, explain the process involved in electroplating the spoon.

Your answer should pay particular attention to the following areas:

(a) How the cell can be constructed. (A diagram with clear labels for the anode, cathode, electrolyte, direction of flow of electrons and ions). (6 marks)

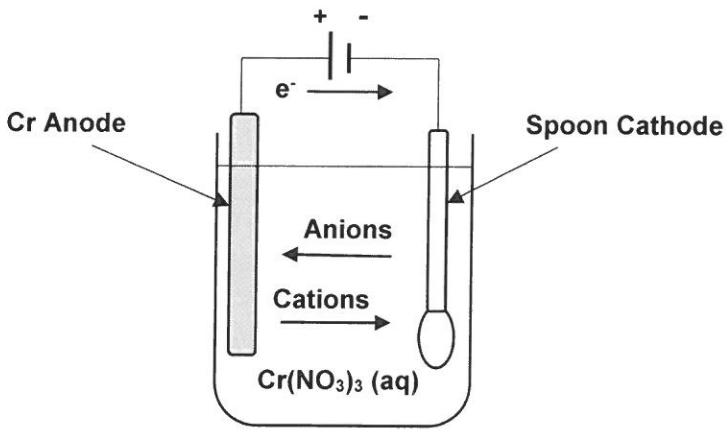
(b) Describe the processes occurring at each electrode. (Including half-equations). (4 marks)

(c) Observations made at each electrode. (2 marks)

(d) The role of the electrolyte. (1 mark)

(e) An example for the industrial importance or application of the process. (1 marks)

(a)



Correct polarity	**(1)**
Correct electrodes	**(2)**
Electron directn.	**(1)**
Ions directn.	**(1)**
Electrolyte	**(1)**

(b) **At the cathode, the negative terminal of the cell provides electron for the reduction of chromium ions to chromium metal being deposited on the spoon. (1)**

$Cr^{3+}(aq) + 3e^- \longrightarrow Cr(s)$ (1)

At the anode, the positive terminal of the cell ensures the oxidation of the chromium electrode to produce chromium ions in solution. (1)

$Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^-$ (1)

(c) **At the cathode, the mass of the spoon will increase. (1)**

At the anode, the mass of the electrode will decrease. (1)

(d) **The electrolyte allows the transfer of ions; ie. (Cr^{3+}) cations towards the cathode and (NO_3^-) nitrate ions towards the anode, in order to balance the charges during the normal operation of the cell. (1)**

(e) **Industrial importance of electrolysis includes the coating of a cheap metal with a more noble metal, ie. Jewellery, etc.**

Or for the coating of iron and other reactive metals with a corrosion resistant metal such as chromium, etc.

***Accept any "one" realistic application! (1)**

Question 38 **(10 marks)**

Compare and contrast electrolytic and galvanic cells.

In your answer include similarities and differences between the two types of electrochemical cells in relation to energy, anodes and cathodes, electrode potentials, and the spontaneity of reactions.

Illustrate your answer with examples of your choice. Include equations and clear, labelled diagrams.

Description	Marks
Suitable examples of an electrolytic and galvanic cell	0–2
Comparison of energy (galvanic cells: chemical energy to electrical energy, electrolytic cells: electrical energy to chemical energy)	0–2
Electrodes: Anode = oxidation; Cathode = Reduction Galvanic cells: Anode (–) Cathode (+) Electrolytic cells: Anode (+) Cathode (–)	0–1
Galvanic cells: Spontaneous reactions (positive cell potentials) Electrolytic cells: Allow non spontaneous reactions to occur by overcoming negative (or zero) cell potentials	0–2
Quality of diagrams	0–1
Use of equations	0–2
Total	**10**

Question 39 **(14 marks)**

When soils containing iron pyrite (FeS_2) are exposed to air, the following reaction can occur.

$$2\ FeS_2(s) + 7\ O_2(g) + 2\ H_2O(\ell) \rightarrow 2\ Fe^{2+}(aq) + 4\ SO_4^{2-}(aq) + 4\ H^+(aq)$$

These types of soils are called acid sulfate soils. The pH of groundwater in these soils will decrease. If this groundwater discharges into lakes and rivers it will also cause their pH to decrease.

(a) Explain how this reaction causes the pH of groundwater to decrease. (2 marks)

As the reaction proceeds, H^+ are produced, thus increasing $[H^+]$, **(1)**
and DECREASING pH. **(1)**

A titration was carried out on a sample of lake water, suspected of being contaminated with acid soils, to determine its pH.

A student placed a standardised solution of 0.005 $molL^{-1}$ NaOH in the burette.
The student then titrated the NaOH solution against 50.0 mL samples of the lake water and obtained the following results.

	Trial 1	Trial 2	Trial 3	Trial 4
Final burette reading (mL)	4.25	8.05	12.00	16.05
Initial burette reading (mL)	0.00	4.10	8.10	12.05
Volume of NaOH used (mL)	**4.25**	**3.95**	**3.90**	**4.00**

Calculated titres in Table **(1)**

(b) Determine the average volume of NaOH used. (2 marks)

Ave Titre $= \dfrac{3.95 + 3.90 + 4.00}{3} = 3.95$ **mL** **(1)**

(c) Calculate the average number of moles of NaOH used to neutralise the acid. (1 mark)

$n = cV = 0.0050 \times 0.00395 = 1.98 \times 10^{-5}$ **mol** **(3 x SF)** **(1)**

(d) Assuming that the lake water is the only source of H^+ ions and that complete ionisation of the acid in the lake water has occurred, determine the pH of the lake water. (3 marks)

$n(H^+) = n(NaOH) = 1.975 \times 10^{-5}$ mol **(1)**

$[H^+] = n/V = 1.975 \times 10^{-5} / 0.050 = 3.95 \times 10^{-4}$ molL^{-1} **(1)**

$pH = -\log[H^+] = -\log(3.95 \times 10^{-4}) = 3.40$ (3 x SF) **(1)**

(e) Complete the following table (6 marks)

Equipment	What is it used for in this experiment?	What should it be rinsed with before use?
Burette	**To deliver accurate volume of NaOH. (1)**	**The NaOH solution. (1)**
Pipette	**To measure 50.0 mL of lake water. (1)**	**The lake water. (1)**
Conical flask	**Where the titration reaction takes place. (1)**	**Distilled water. (1)**

Question 40 **(23 marks)**

Citric acid has a formula of $C_6H_8O_7$ and is a weak triprotic acid. A student, Steve, was investigating the amount of citric acid present in a bottle of lemon juice concentrate.

He decided to titrate 25.00 mL samples of the lemon juice against 0.998 mol L^{-1} NaOH solution.

The reaction for the equation is:

$$3\ NaOH(aq) + C_6H_8O_7(aq) \rightarrow Na_3C_6H_5O_7(aq) + 3\ H_2O(aq)$$

Steve's results are shown below.

	Titrations			
	1	2	3	4
Final Reading (mL)	15.60	30.80	46.00	15.75
Initial Reading (mL)	0.00	15.60	30.80	0.60
Titre (mL)	**15.60**	**15.20**	**15.20**	**15.15**

(a) Calculate the **concentration** of citric acid in the lemon juice

(i) in moles per litre. (6 marks)

Description	Marks
V(NaOH) = 15.20 mL (average of titre 2 and 3) = 0.01520 L	1
n(NaOH) = 0.998 x 0.01520 = 0.01517 mol	1
$n(C_6H_8O_7)$ = (1/3) x n(NaOH)	1
= 0.005057 mol	1
$c(C_6H_8O_7)$ = 0.005057/0.0250 = 0.202 mol L^{-1}	1
3 significant figures	1
Total	**6**

(ii) as a percentage by mass (assume mass of 25.00 mL sample = 25.00 g). (3 marks)

Description	Marks
$n(C_6H_8O_7)_{in\ 25\ mL}$ = 0.005057 mol	1
$m(C_6H_8O_7)_{in\ 25\ mL}$ = 0.005057 x 192.124 = 0.9715 g	2
% by mass = (0.9743/25.0) x 100 = 3.89%	1
Total	**4**

A second student, Claire, wanted to find out the citric acid concentration in a freshly squeezed lemon. As her lemon juice was less homogeneous that the bottled juice, she decided to carry out a 'back titration' to avoid having to measure out accurate aliquots of the lemon juice. Claire added an excess amount of an alkali, sodium hydroxide solution, to the lemon juice, and then titrated the excess sodium hydroxide against hydrochloric acid.

8.00 g of the lemon juice was mixed with 25.00 mL of 0.600 mol L^{-1} $NaOH_{(aq)}$ and stirred thoroughly. The resulting solution, containing the excess sodium hydroxide solution, was filtered and immediately titrated against 0.501 mol L^{-1} hydrochloric acid, with 15.50 mL of the acid being required.

(b) (i) Calculate the total number of moles of sodium hydroxide added to the 8.00 g of fresh lemon juice. (1 mark)

Description	Marks
n(NaOH) = 0.600 x 0.025 = 0.0150 mol	1
Total	**1**

(ii) Calculate the moles of excess sodium hydroxide, and hence calculate the number of moles of sodium hydroxide that reacted with the citric acid in the lemon juice. (3 marks)

Description	Marks
n(HCl) = 0.501 x 0.0155 = 0.007766 mol	1
$n(NaOH)_{excess}$ = n(HCl) = 0.007766 mol	1
$n(NaOH)_{used}$ = 0.0150 – 0.007766 = 0.00723 mol	1
Total	**3**

(iii) Calculate the number of moles of the citric acid in the 8.00 g sample, and hence determine the percentage, by mass, of the citric acid in the fresh lemon juice. (4 marks)

Description	Marks
$n(C_6H_8O_7)$ = (1/3) x $n(NaOH)_{used}$ = (1/3) x 0.00723 = 0.00241 mol	2
$m(C_6H_8O_7)_{in\ 8.00g}$ = 0.00241 x 192.124 = 0.4633 g	1
% by mass = (0.4646 /8.00) x 100 = 5.79%	1
3 significant figures (required for full marks)	
Total	**4**

(c) Explain, using equations and sketches of graphs as required, why Claire could have used a large range of possible indicators whereas Steve had to use an indicator, such as phenolphthalein, that had a basic endpoint. (5 marks)

Description	Marks
Claire was titrating a strong acid with a strong base	1
therefore, the pH change at the equivalence point was very significant (could show with graph) and a range of indicators will change colour within this large pH range	1
Steve was titrating a weak acid with a strong base	1
and the equivalence point was at a high pH due to the citrate ion being basic:	1
$H_2O(aq) + C_6H_5O_7^{3-}(aq) \rightleftharpoons C_6H_6O_7^{-}(aq) + 3\ OH^-(aq)$ or: $3\ H_2O(aq) + C_6H_5O_7^{3-}(aq) \rightleftharpoons C_6H_8O_7(aq) + 3\ OH^-(aq)$	1
Total	**5**