SHENTON COLLEGE

CHEMISTRY

UNIT 3

2017

Name: **ANSWERS**

Teacher: _____

TIME ALLOWED FOR THIS PAPER
Reading time before commencing work: ten minutes
Working time for the paper: two and a half hours

MATERIALS REQUIRED/RECOMMENDED FOR THIS PAPER

To be provided by the supervisor:
This Question/Answer Booklet
Multiple-choice Answer Sheet
Chemistry Data Book

To be provided by the candidate:
Standard items: pens (blue/black preferred), pencils (including coloured), sharpener, eraser, correction tape/fluid, ruler, highlighters

Special items: up to three non-programmable calculators approved for use in the WACE examinations

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. If you have any unauthorised material with you, hand it to the supervisor **before** reading any further.

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	20	20	40	/40	/25
Section Two: Short answer	9	9	50	/58	/35
Section Three: Extended answer	4	4	60	/66	/40
					/100

Instructions to candidates

1.   Answer the questions according to the following instructions.

Section One: Answer all questions on the separate Multiple-choice Answer Sheet provided. For each questions shade the box to indicate your answer. Use only a blue or black pen to shade the boxes. If you make a mistake, place a cross through that square then shade your new answer. Do not erase or use correction fluid/tape. Marks will not be deducted for incorrect answers. No marks will be given if more than one answer is completed for any question.

Sections Two and Three: Write your answers in this Question/Answer Booklet.

2.   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.

3.   You must be careful to confine your responses to the specific questions asked and to follow any instructions that are specific to a particular question.

4.   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(s) that you are continuing to answer at the top of the page.

5.   The Chemistry Data Book is **not** handed in with your Question/Answer Booklet.

Section One: Multiple-choice   **25% (40 marks)**

This section has **20** questions. Answer **all** questions on the separate Multiple-choice Answer Sheet provided. For each question, shade the box to indicate your answer. Use only a blue or black pen to shade the boxes. If you make a mistake, place a cross through that square then shade your new answer. Do not erase or use correction fluid/tape. Marks will not be deducted for incorrect answers. No marks will be given if more than one answer is completed for any question.
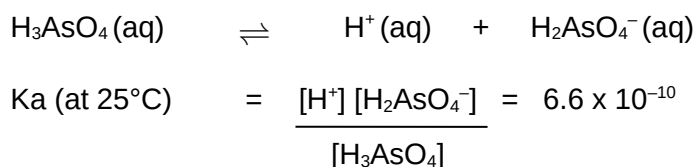
Suggested working time: 50 minutes.

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1. In a chemical reaction at constant temperature, the addition of a catalyst:

   (a) increases the concentration of the products at equilibrium.
   (b) increases the energy of the molecules so more can successfully collide.
   (c) lowers the amount of energy released in the overall reaction.
   (d) decreases the time required for equilibrium to be reached.

2. 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.

Questions 3 and 4 relate the following information:

Consider the following information for a 1.00 mol L^{-1} solution of arsenic acid, (H_3AsO_4):

$$H_3AsO_4\,(aq) \quad \rightleftharpoons \quad H^+\,(aq) \quad + \quad H_2AsO_4^-\,(aq)$$

$$Ka\,(at\ 25°C) \quad = \quad \frac{[H^+]\,[H_2AsO_4^-]}{[H_3AsO_4]} \quad = \quad 6.6 \times 10^{-10}$$

3. At equilibrium at 25°C, which of the following species will be present in the greatest concentration?

   (a) $H^+\,(aq)$
   (b) $H_2AsO_4^-\,(aq)$
   (c) $H_3AsO_4\,(aq)$
   (d) $OH^-\,(aq)$

4. Which of the following statements best describe the value of the equilibrium constant (K) for arsenic acid at 25o C?

   (a) Arsenic acid is a strong acid existing essentially as molecules.
   (b) Arsenic acid is a weak acid existing essentially as molecules.
   (c) Arsenic acid is a weak acid existing essentially as ionic species.

(d)   Arsenic acid is strong acid existing essentially as ionic species.

5.   The following statements refer to the chemical reaction between magnesium carbonate granules, ($MgCO_3$) and a dilute hydrochloric acid solution, (HCl). 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 HCl (aq).
(d)   The initial rate of reaction is independent of the state of sub-division of $MgCO_3$ (s).

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

$$2SO_2 \text{ (g)} \quad + \quad O_2 \text{ (g)} \quad \rightleftharpoons \quad 2SO_3 \text{ (g)}$$
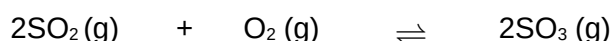
(a) $K = \dfrac{[SO_2]^2\,[O_2]}{[SO_3]^2}$

(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.

7.   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) $2NO_2 \text{ (g)} \quad \rightleftharpoons \quad N_2O_4 \text{ (g)}$
(b) $N_2 \text{ (g)} \; + \; 3H_2 \text{ (g)} \quad \rightleftharpoons \quad 2NH_3 \text{ (g)}$
(c) $H_2O \text{ (g)} \; + \; C \text{ (s)} \; \rightleftharpoons \; H_2 \text{ (g)} \; + \; CO \text{ (g)}$
(d) $H_2 \text{ (g)} \; + \; F_2 \text{ (g)} \quad \rightleftharpoons \quad 2HF \text{ (g)}$

8.   Bromophenol blue is an acid-base indicator that has a colour change from yellow to blue between pH 3.0 and 4.6. A potassium hydroxide solution (in a conical flask), containing a few drops of bromophenol blue indicator, is titrated with an acetic (ethanoic) acid solution (from a burette).

Which one of the following statements about this titration is true?

(a)   The end point and the equivalence point occur at the same time.
(b)   The end point occurs after the equivalence point.
(c)   The end point occurs before the equivalence point.
(d)   The indicator will be yellow at the equivalence point of the titration.

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

|	Sodium ethanoate ($NaCH_3COO$)	Potassium nitrate (KNO_3)	Ammonium chloride (NH_4Cl)
(a)	neutral	acidic	basic
(b)	*basic*	*neutral*	*acidic*
(c)	acidic	neutral	basic
(d)	basic	acidic	neutral

10.   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.

Questions 11 and 12 relate the following information:

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

Zn (s) + 2 NH4+ (aq) + 2 MnO2 (s) $\longrightarrow$ Zn2+(aq) + Mn2O3 (s) + H2O (l) + 2NH3 (aq)
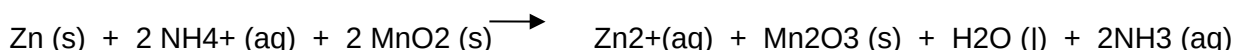
11.   Which of the following statements regarding the dry cell are correct?

        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.

12.   Which of the following will NOT increase the rate of the redox reaction?

        (a)   Increasing the concentration of ammonium ions.
        (b)   Grinding up the MnO_2 into a finer powder.
        (c)   Using a thicker zinc outer casing.
        (d)   Warming up the cell.

13.   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
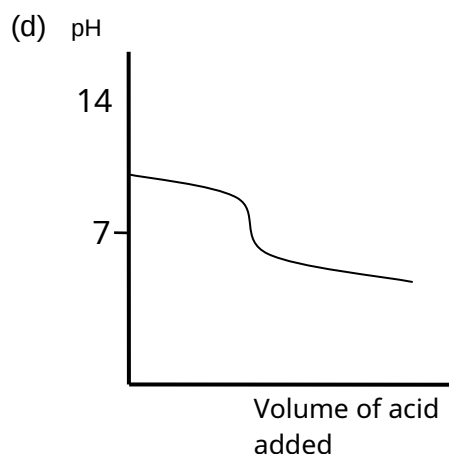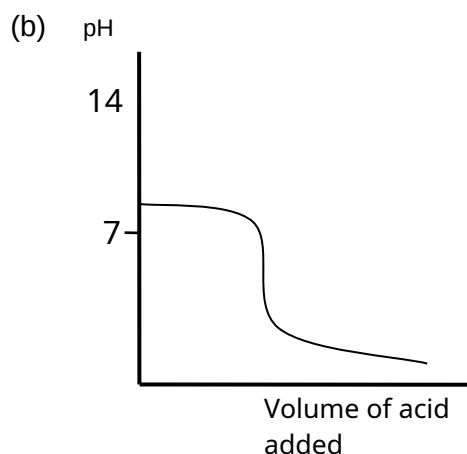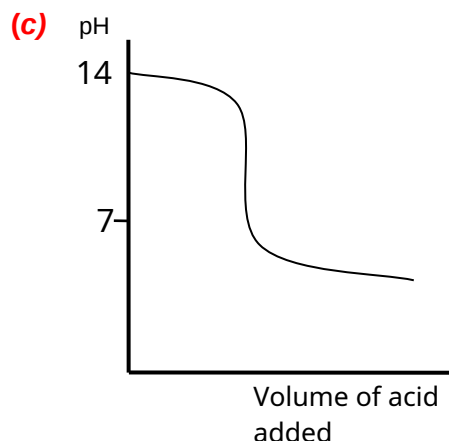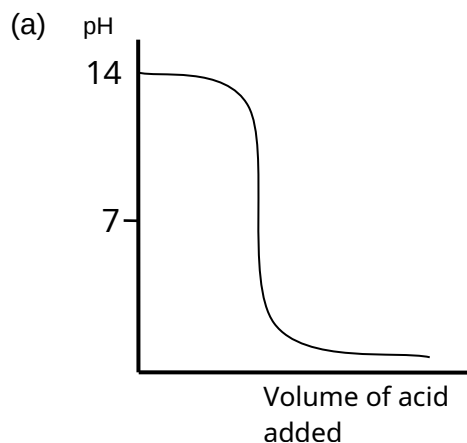   *(d)   I, II and IV*

Questions 14, 15 and 16 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.

14.   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

15.   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)   pH

14

7

Volume of acid added

(c)   pH

14

7

Volume of acid added

(b)   pH

14

7

Volume of acid added

(d)   pH

14

7

Volume of acid added

16.   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

Questions 17, 18 and 19 relate to the following electrochemical cell at 25ºC:
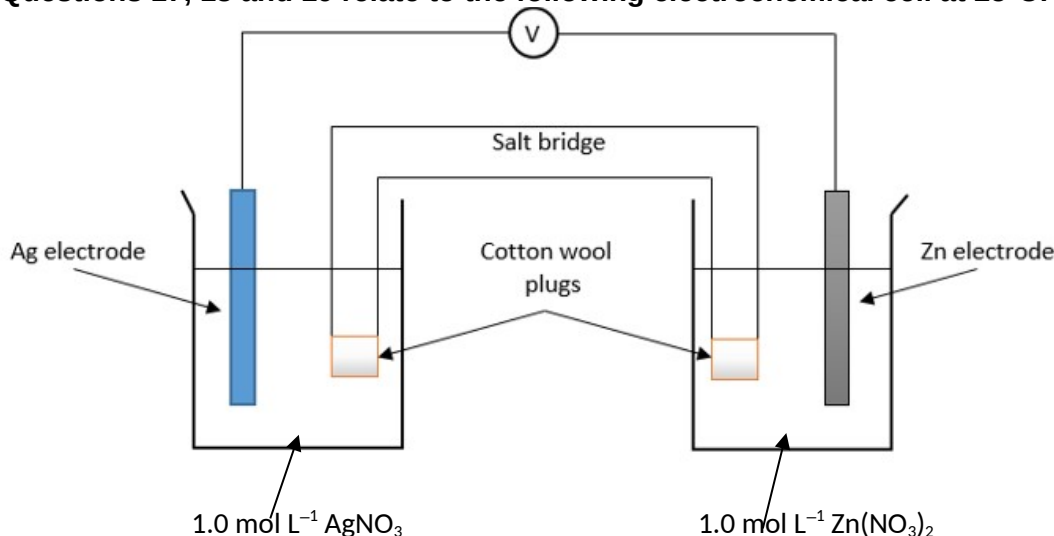
17.   Which of the following reactions will occur during the normal operation of this cell?

(a) $2Ag^+ (aq)$ + $Zn (s)$ $\longrightarrow$ $2Ag (s)$ + $Zn^{2+} (aq)$ *$E^o = 1.56$ V*
(b) $2Ag^+ (aq)$ + $Zn (s)$ $\longrightarrow$ $2Ag (s)$ + $Zn^{2+} (aq)$ $E^o = 0.04$ V
(c) $Zn^{2+} (aq)$ + $2Ag (s)$ $\longrightarrow$ $Zn (s)$ + $2Ag^+ (aq)$ $E^o = 1.56$ V
(d) $Zn^{2+} (aq)$ + $2Ag (s)$ $\longrightarrow$ $Zn (s)$ + $2Ag^+ (aq)$ $E^o = 0.04$ V

18.   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.

19.   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.

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

$$H_2PO_4^- (aq) + H_2O (l) \rightleftharpoons HPO_4^{2-} (aq) + H_3O^+ (aq)$$
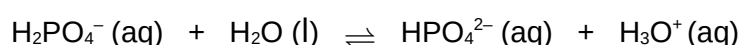
Which of the following would be **true** after the addition of a small volume of 2.0 mol L^{-1} sodium hydroxide solution to the buffer solution?

(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.

End of Section One

Section Two: Short answer   **35%**   **(58 marks)**

This section has **9** 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.

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Suggested working time: 50 minutes.

Question 21   **(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.

(<u>Note</u>: No chemical equations necessary).

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

     A white solid dissolves in a colourless solution, producing a colourless and odourless gas.(2)

     (*Must have two observations for both marks).

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

     A blue solid dissolves in a colourless solution to produce a blue solution. (2)

     (*Must have two observations for both marks).

Question 22 **(6 Marks)**

The Brønsted – Lowry theory can be used to account for the acidic and basic properties of a much wider array of substances whose properties cannot be easily explained using earlier theories.

Complete the following table by stating the pH, and give a supporting balanced chemical equation to explain the pH for each of the substances listed.

(6 marks)

Substance	pH (acidic, basic or neutral)	Equation
$Mg(CH_3COO)_2$ (aq)	**Basic (1)**	$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$ **(1)**
NH_4Cl (aq)	**Acidic (1)**	$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ **(1)**
$NaHSO_4$ (aq)	**Acidic (1)**	$HSO_4^- + H_2O \rightleftharpoons SO_4^{2-} + H_3O^+$ **(1)**

* Also accept "greater than 7" or "less than 7" respectively, for each salt.

Question 23 **(4 Marks)**

The following chemical equation represents an unbalanced redox reaction.

$$S_2O_3^{2-} (aq) + NO_3^- (aq) \longrightarrow SO_4^{2-} (aq) + NH_4^+ (aq)$$
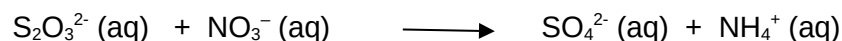
In the appropriate spaces below, write the two separate half-equations and the overall balanced redox equation.

(4 marks)

Oxidation:

$S_2O_3^{2-} + 5H_2O \rightarrow 2SO_4^{2-}$ (aq) $+ 10H^+ + 8e^-$ **(1)**

Reduction:

$NO_3^- + 10H^+ + 8e^- \rightarrow NH_4^+ + 3H_2O$ **(1)**

Overall Redox:

$S_2O_3^{2-} + NO_3^- + 2H_2O \rightarrow 2SO_4^{2-} + NH_4^+$ **(2)**

Question 24 **(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:

$$Br_2\,(aq) \; + \; H_2O\,(l) \qquad \rightleftharpoons \qquad HBrO\,(aq) \; + \; H^+\,(aq) \; + \; Br^-\,(aq)$$

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, or solution turns less brown. **(1)**

Explanation:

Addition of OH⁻ causes a decrease in the [H⁺] as the combination of the two ions produce water (H_2O). **(1)**

This will result in the rate of collision of reactants being greater than that of products, shifting the equilibrium to the right, favouring the forward reaction rate. Thus the [Br_2] decreases causing the brown colour to fade. **(1)**

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

Colour:

Brown colour becomes more intense, or solution becomes more brown. **(1)**

Explanation:
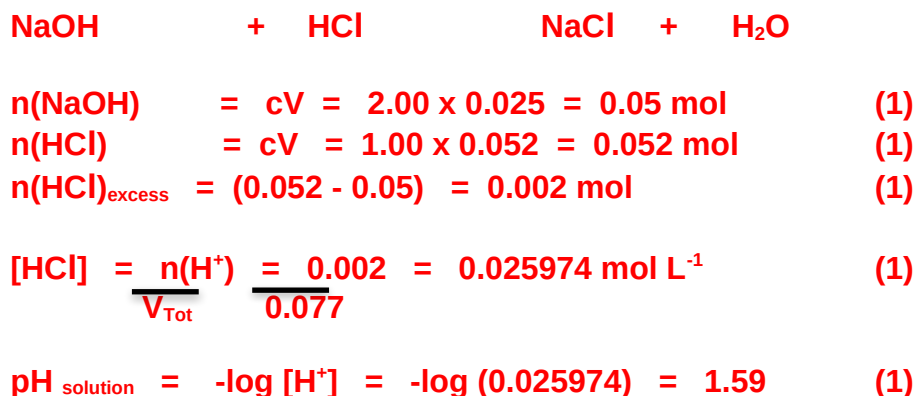
Addition of HCl causes an increase in the [H⁺] on product side, leading to a higher rate of collision of products than the reactants. **(1)**

This will shift the equilibrium to the left, favouring the reverse reaction, leading to an increase in the [Br_2], and the solution becomes more brown. **(1)**

Question 25   **(5 marks)**

Calculate the pH of the resultant solution, if 25.0 mL of 2.00 mol L^{-1} sodium hydroxide and 52.0 mL of 1.00 mol L^{-1} hydrochloric acid are mixed together.   (5 marks)

NaOH   +   HCl   NaCl   +   H₂O

n(NaOH)   = cV = 2.00 x 0.025 = 0.05 mol   (1)
n(HCl)   = cV = 1.00 x 0.052 = 0.052 mol   (1)
n(HCl)$_{excess}$ = (0.052 - 0.05) = 0.002 mol   (1)

[HCl] = n(H⁺) = 0.002 = 0.025974 mol L⁻¹   (1)
**   V$_{Tot}$ 0.077**

pH $_{solution}$ = -log [H⁺] = -log (0.025974) = 1.59   (1)

> **NOT 3 SF**
>
> **-1**

Question 26   **(9 Marks)**

The manufacture of ammonia on an industrial scale is carried out using the Haber process, which relies on the reversible reaction of nitrogen and hydrogen in the presence of an iron catalyst, as shown in the following equation:

$$N_2(g) + 3 H_2(g) \rightleftharpoons 2 NH_3(g)   \Delta H = -92 \text{ kJ mol}^{-1}$$
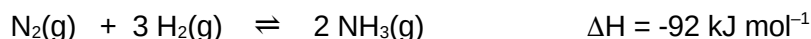
The conditions for the reaction in industry must be chosen carefully, taking into consideration not only the yield, but also the rate of the reaction. Commonly, a temperature of around 500°C is used, and the reaction operated at a pressure of around 20,000 kPa. Since ammonia has a much higher boiling point than the other gases, it can easily be removed from the equilibrium mixture by condensation.

(a)   In the space provided below, draw a fully labelled enthalpy level diagram for the Haber process, showing **ΔH**, **E$_A$**, **catalysed** and **uncatalysed** reaction pathways, and **axes with correct units** stated.

                                                                   (5 marks)

Enthalpy (H)
(kJ mol⁻¹)

3 H₂ + N₂   **E$_A$**   **Fe Catalyst**

ΔH = -92 kJ

2 NH₃

Axes   (1)
Exo. shape   (1)
E$_A$ & ΔH   (1)
Reactants &
products   (1)
E$_A$ Catalyst   (1)
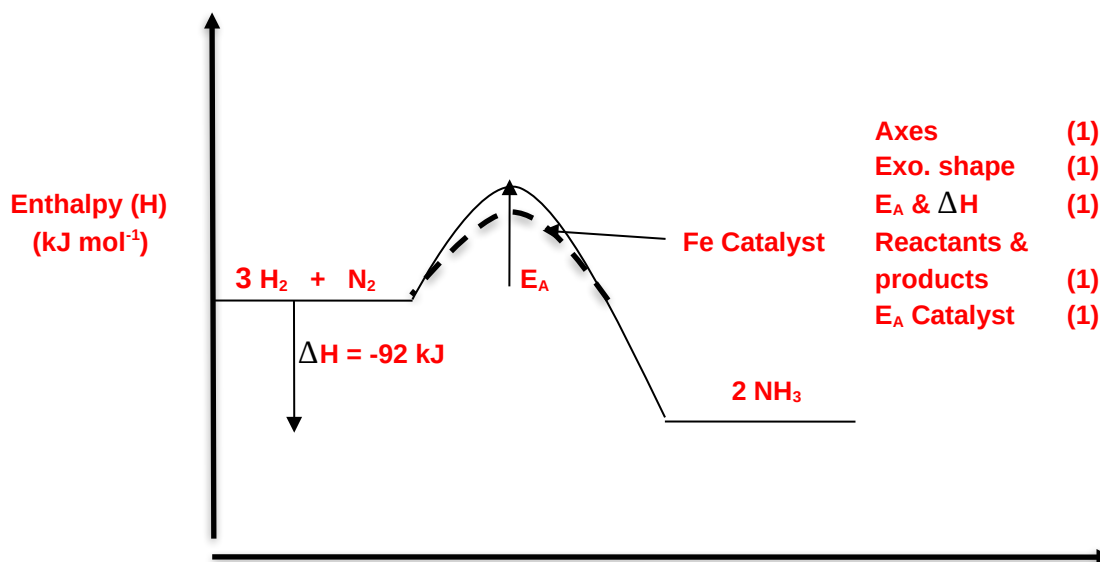
A sealed vessel containing an equilibrium mixture of nitrogen, hydrogen and ammonia was subjected to the following changes in conditions:

- At a time, t_1, the temperature of the vessel was increased
- At a time, eqm_1, the system had returned to equilibrium
- At a time, t_2, all ammonia was removed from the system
- At a time, eqm_2, the system had again returned to equilibrium

(b)   Complete the following graph, to show what happens to the concentrations of nitrogen and ammonia as the above changes are made.

(4 marks)
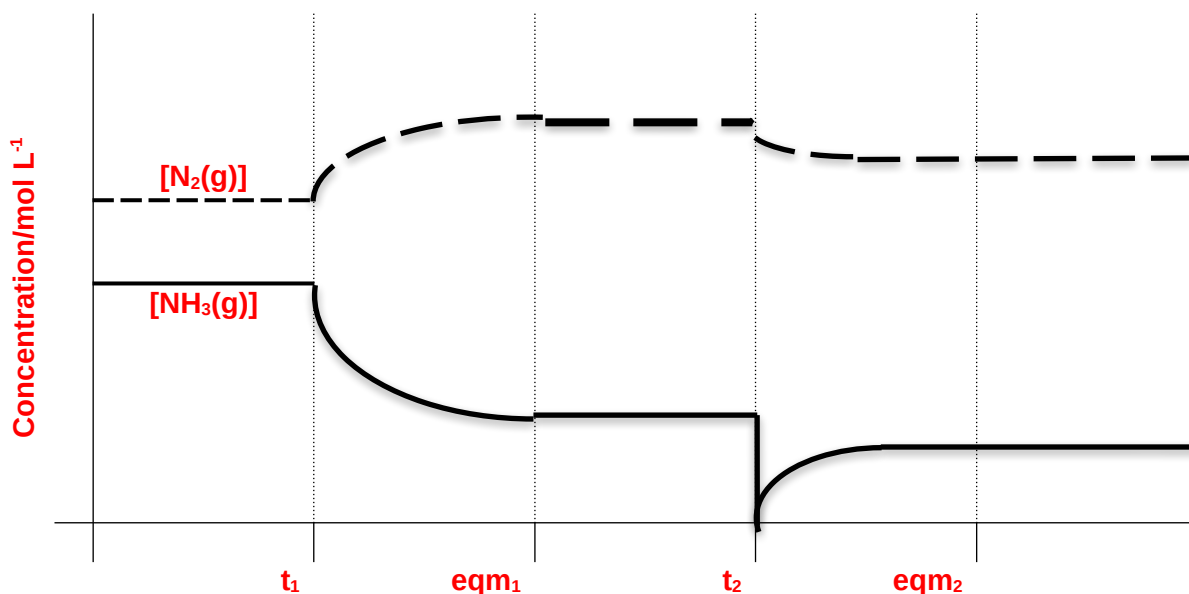
Award (2) marks for showing the correct shape and orientation for the N_2 and (2) marks for the correct shape and orientation for the NH_3 lines.

Question 27   **(10 Marks)**

Aluminium salts are acidic due to the presence of the hexaaqualuminate ion, $[Al(H_2O)_6]^{3+}$ which is formed when a soluble aluminium salt is dissolved in water. This ion undergoes hydrolysis as follows:

$$[Al(H_2O)_6]^{3+} (aq) \ + \ H_2O \ (l) \ \rightleftharpoons \ [Al(OH)(H_2O)_5]^{2+} (aq) \ + \ H_3O^+ (aq)$$

(a)   Write the equilibrium constant (K) expression for this reaction.   (1 mark)

$$K = \frac{[(Al(OH)(H_2O)_5)^{2+}]\ [H_3O^+]}{[(Al(H_2O)_6)^{3+}]} \qquad (1)$$

(b)   A solution of aluminium nitrate has a pH of 5.6.

(i)   Using the above equilibrium reaction, explain how the pH of the solution would change, if more crystals of hydrated aluminium nitrate were dissolved into the solution.

(3 marks)

The addition of a soluble Al – salt will lead to an increase in $[(Al(H_2O)_6)^{3+}]$. (1)

Thus the rate of collision of the reactants will increase, leading to an increase in the forward reaction rate.   (1)

Consequently leading to a higher $[H_3O^+]$ and a lowering in the pH.   (1)

(ii)   When a small volume of dilute sodium hydroxide was added to a sample of the original solution, the pH initially increased from 5.6 to 6.0, and then decreased back to 5.8. Explain these observations.

(3 marks)

Initially the addition of excess OH^- will cause an increase in pH to 6.0.   (1)

As the neutralisation of OH^- and H^+ takes place, the rate of collision of reactants will be higher than that of the products, thus the rate of the F'wd reaction is favoured.   (1)

This will lead to an increase in $[H_3O^+]$ and thus decrease the pH to 5.8.   (1)

(c)   It was found that when the aluminium nitrate solution was warmed, the pH of the solution decreased. From this information, deduce whether the <u>forward</u> reaction in the above equilibrium is endothermic or exothermic. Explain your reasoning.

(3 marks)

As the pH has decreased due to an increase in the $[H^+]$, caused by an increase in temp;   (1)

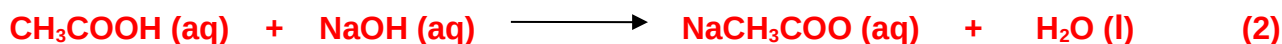**Clearly the F'wd reaction has been favoured by this imposed change, (ie. higher temp). (1)**

In order for the reaction to respond in this way, (ie. shifting the equilibrium to the right), the F'wd reaction must be ENDOTHERMIC. (1)

Question 28   **(8 Marks)**

Ethanoic acid is a weak, **monoprotic** acid. In an experiment, a solution of approximately 0.2 mol L$^-$^{1} ethanoic acid (CH_3COOH) is titrated with a standard solution of 0.200 mol L^{-1} sodium hydroxide in order to determine the accurate concentration of the acid. 30.00 mL of the sodium hydroxide solution was pipetted into a conical flask, and the ethanoic acid added from the burette.

(a)   Write a balanced molecular equation, including state symbols, for the reaction occurring.

                                                                                                            (2 marks)

CH_3COOH (aq)   +   NaOH (aq) ⟶ $NaCH_3COO$ (aq)   + H_2O (l)   (2)

        Deduct 1 x mark if missing or incorrect state symbols.

(b)   On the axis below, sketch a graph showing how the pH would be expected to change
         during the titration, until an excess of the acid was added.

                                                                                                            (3 marks)

High pH start   (1)

Equivalence above pH 7   (1)

Final pH NOT too low   (1)

**Equivalence to coincide with
30 mL CH_3COOH added.   (1)**
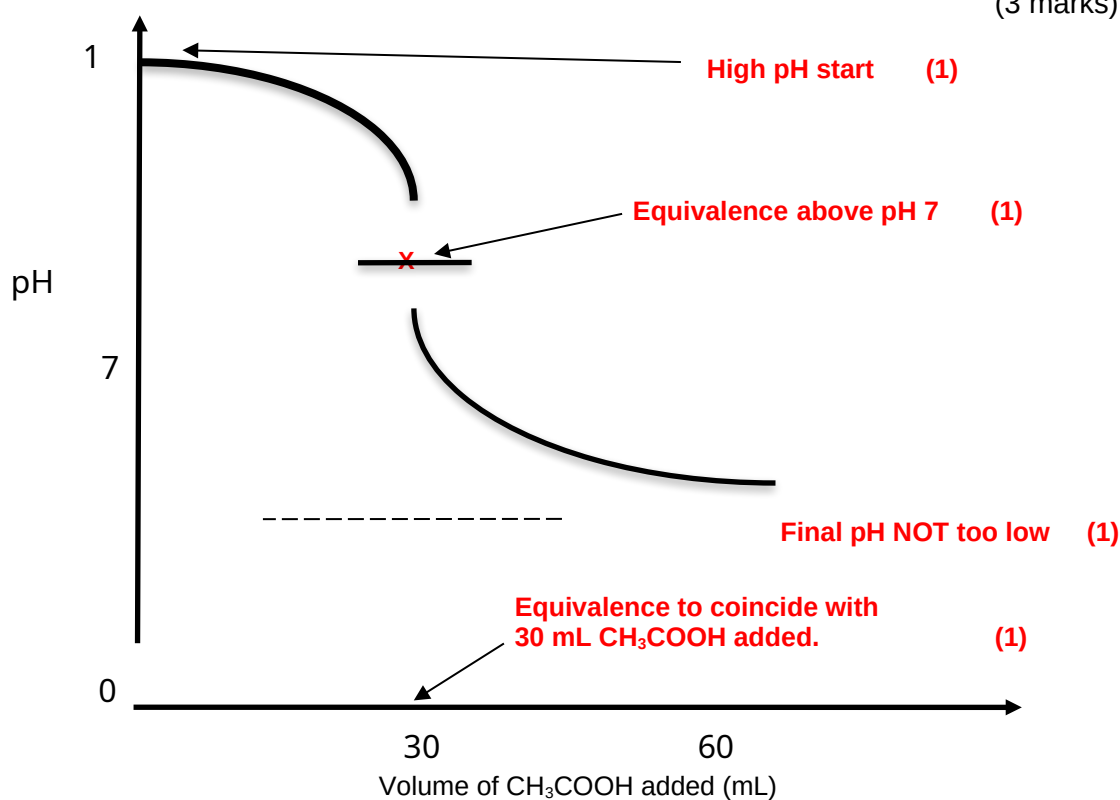
(c)   On the graph above, label the equivalence point for this reaction.   (1 mark)

(d)   What should the pipette be rinsed with, immediately prior to use?   (1 mark)

        The NaOH solution.   (1)

(e)   From the list below, circle the correct indicator, that would be suitable for use in this particular titration.   (1 mark)

(1)

Methyl orange   **Phenolphthalein**   **Bromothymol blue**
(pH 3.1 – 4.4)   (pH 8.3 – 10.0)   (pH 6.0 – 7.6)

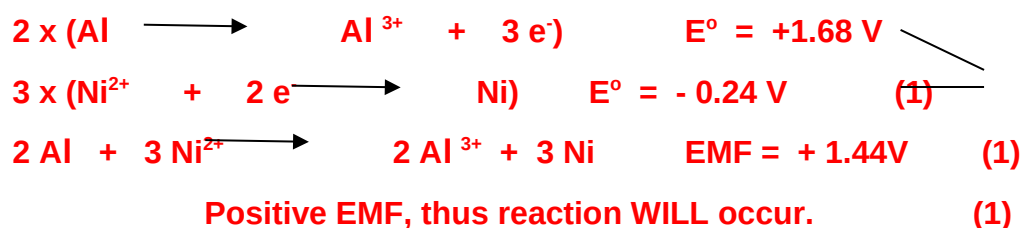**Question 29**   **(6 Marks)**

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^o 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 <u>likely or unlikely to occur</u> as described.

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

2 x (Al $\longrightarrow$ Al^{3+} + 3 e$^-$)   E^o = +1.68 V

3 x (Ni^{2+} + 2 e $\longrightarrow$ Ni)   E^o = - 0.24 V   (1)

2 Al + 3 Ni^{2+} $\longrightarrow$ 2 Al^{3+} + 3 Ni   EMF = + 1.44V   (1)

Positive EMF, thus reaction WILL occur.   (1)

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

2 (Ag $\longrightarrow$ Ag$^+$ + e$^-$)   E^o = - 0.80 V

2 H$^+$ + 2e $\longrightarrow$ H$_2$ E^o = 0.00 V   (1)

2 Ag + 2 H$^+$ $\longrightarrow$ 2 Ag$^+$ + H$_2$ EMF = -0.80 V   (1)

Negative EMF, thus reaction will NOT occur.   (1)
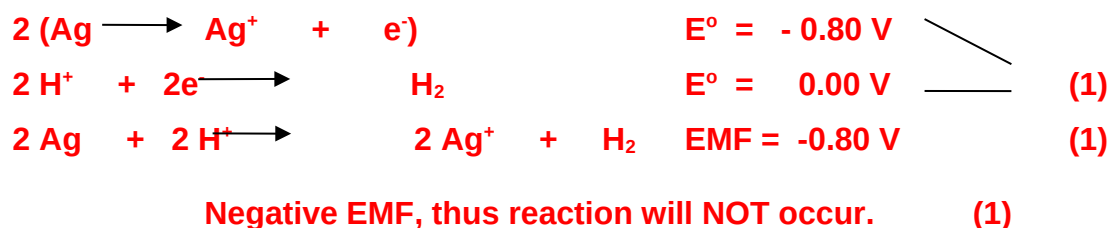
*Note: Overall redox equation NOT necessary, as reaction will not occur.

End of Section Two

Turn to next page

Section Three: Extended answer   **40%   (66 marks)**

This section contains **four (4)** questions. You must answer **all** questions. Write your answers in the spaces provided below.

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.

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(s) that you are continuing to answer at the top of the page.

Suggested working time: 60 minutes.

Question 30   **(16 marks)**

Rising carbon dioxide levels in the atmosphere are believed to play an important role in the life of organisms known as calcifiers, a group that includes many forms of coral and crustaceans. These organisms use a precipitation reaction between calcium ions and carbonate ions present in sea-water to form shells and skeletons.

Measurements have detected a fall of around 0.1 in the pH of the oceans since the beginning of the industrial revolution at the end of the 18th century. Scientists believe this acidification can be attributed to an increase in the partial pressure of carbon dioxide in the atmosphere over the same period.

(a)   Use appropriate chemical equations, to explain why a rise in the partial pressure of carbon dioxide in the atmosphere has caused a decrease in the pH of the oceans.   (3 marks)

An increase in the p(CO_2 (g)) will lead to an increase in [CO_2 (aq)] in the oceans. (1)

ie. CO_2 (aq)   + H_2O (l) $\rightleftharpoons$ H_2CO_3 (aq)   (1)

**Thus an increase in [CO_2 (aq)] will lead to an increased rate of collision of reactants, thus favouring the F'wd reaction rate, leading to more H_2CO_3 (aq), hence a higher [H^+ (aq)] and a lower pH.   (1)**

A student wished to investigate the composition of prawn shells. In order to do this, the student carried out a series of reactions to convert all the carbonate in the shells, (present as $CaCO_3$), to a soluble form, (i.e. CO_3^{2-}).
The steps that the student carried out were as follows:

- The shells of 10 prawns were ground to a fine powder using a mortar and pestle.
- 2.17 g of the powder was placed in a beaker, where it was chemically treated to convert all the carbonate into a soluble form.
- The resulting mixture was then filtered to remove any insoluble substances and the filtrate transferred to a 250 mL volumetric flask and made up to the mark with distilled water.
- 20 mL aliquots of the solution in the volumetric flask were titrated against a standard solution of nitric acid with a concentration of 0.0502 mol L^{-1}.
- All burette readings were taken from the **top of the meniscus**.
- The average titre of nitric acid used was 35.05 mL.

(b)   Write a balanced ionic equation for the titration reaction.   (2 marks)

CO_3^{2-} (aq)   +   2 H^+ (aq) $\longrightarrow$ H_2O (l)   + CO_2 (g)   (2)
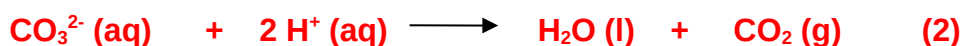
(c)   Calculate the number of moles of nitric acid titrated from the burette.   (1 mark)

n(HNO_3) = cV = 0.0502 x 0.03505 = 0.00176 mol   (1)

(d)   Calculate the number of moles of carbonate in the 20.0 mL aliquots.   (2 marks)

$n(CO_3^{2-})_{in\ 20\ mL}$ =   ½ $n(HNO_3)$ **(1)**

                              =   **0.000880 mol**   **(1)**

(e)   Calculate the number of moles of carbonate in the original 2.17 g of powdered prawn shells, and thus calculate the percentage by mass of calcium carbonate in the sample of prawn shells. (You may assume that the moles of $CaCO_3$ are equal to the moles of Na_2CO_3).

(5 marks)

$n(CO_3^{2-})_{in\ 250\ mL}$ =   **250 / 20 x 0.000879755 = 0.010997 mol**   **(2)**

$n(CaCO_3) = n(CO_3^{2-}) = 0.010997\ mol$ **(1)**

$m(CaCO_3) = nM$ = **0.010997 x 100.09   =   1.10 g**   **(1)**

$\%(CaCO_3)_{in\ shells}$ =   **(1.10 / 2.17) x 100   =   50.7%**   **(1)**

> **NOT 3 SF**
>
> **-1**

(f)   State and explain what effect the student's decision to read the burette from the top of the meniscus would have had on the calculated percentage by mass.   (3 marks)

Effect on calculated percentage (circle one)   Artificially high   (No effect)   Artificially low

                                                                                                   (1)

Explanation

As the readings were taken consistently from the top of the meniscus, and since the titre value is the difference between two readings, the systematic error would have cancelled out.   (1)

**Thus the calculated percentage would not have been affected. (1)**

Question 31   **(22 marks)**

Propanoic acid, CH_3CH_2COOH, is a weak monoprotic acid that is produced by bacteria in the skin. In an experiment to determine the concentration of an aqueous solution of propanoic acid, a student titrated 25.0 mL aliquots of the solution with a previously standardised 0.976 mol L^{-1} solution of sodium hydroxide in a conical flask, using a pH meter to monitor the change in pH. The student's results are shown in the table below.

Volume of NaOH (mL)	20.75	20.80	20.85	20.90	20.95	21.00	21.05	21.10	21.15
pH of solution	4.7	5.3	5.2	5.6	7.9	12.7	13.0	13.2	13.3

(a)   Plot the results from the experiment on the graph paper provided below, and use your graph to estimate the pH at the equivalence point. Include clearly labelled axes and an appropriate scale. (5 marks)

Equivalence   (1)

Axes   (1)
Scale   (1)
Points   (1)
Smooth curve (1)

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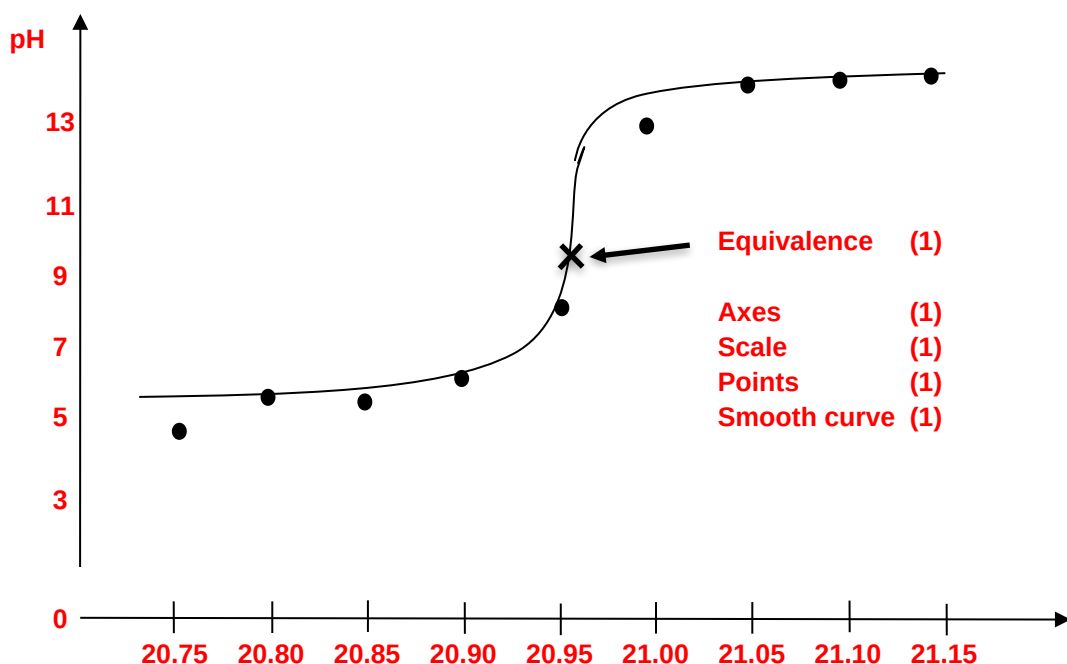
Volume NaOH (mL)

Estimated pH at equivalence point: **Accept pH   8 – 10**   **(1)**   (1 mark)

(b)   Explain why a failure to standardise the sodium hydroxide solution would have led to a
         <u>systematic error</u>, and what effect it would have on the calculated value for the concentration
         of the acid.   (3 marks)

**NaOH cannot be obtained pure and it readily absorbs moisture from the air,
so it cannot be weighed-out directly to produce a standard solution.   (1)**

**Due to impurities and high moisture content, its actual mass will always be
"less" than that weighed out.   (1)**

**This would lead to a consistently HIGHER than expected value for the
concentration of the acid being calculated. (1)**

***Can also accept other reasons like: reaction with CO_2 in the air, and/or
relatively low molar mass may lead to a significant increase in % weighing of
error.**

(c)   Use an appropriate equation, to describe and explain the pH at the equivalence point of this
         titration.
                                                                                                              (3 marks)

**Salt formed is sodium propanoate, CH_3CH_2COONa;   (1)
and since propanoic acid is a weak acid, the propanoate ion will hydrolyse as
follows:**

$CH_3CH_2COO^-$ + H_2O $\rightleftharpoons$ CH_3CH_2COOH + OH^- (1)

The resultant solution is basic, as the $[OH^-]$ is greater than the $[H^+]$.   (1)

(d)   Use an appropriate chemical equation, to describe and explain why the reaction mixture in the flask was able to act as a buffer before less than 20 mL of sodium hydroxide was added.   (4 marks)

When less than 20.0 mL NaOH were added, there was only CH_3CH_2COOH and CH_3CH_2COONa in the flask, (ie. the weak acid and its salt – a buffer solution) .   (1)

ie. CH_3CH_2COOH + H_2O $\rightleftharpoons$ $CH_3CH_2COO^-$ + H_3O^+ (1)

As NaOH was added, OH^- + H^+ $\rightleftharpoons$ H_2O.   Thus rate of collision of reactants is higher than that of products, thus F'wd reaction is favoured, producing more of the H^+ ions that were removed.   (1)

As the change in $[H^+]$ is minimised, the pH will not increase significantly. (1)

After repeating the experiment a number of times, the student found the concentration of the propanoic acid solution was 0.815 mol L^{-1}.

(e)   Using the data provided, calculate the pH of the mixture in the flask if 30.0 mL of sodium hydroxide is added to a 25.0 mL aliquot of propanoic acid.   (6 marks)

n(NaOH) = cV =   0.976 x 0.030 = 0.02928 mol   (1)
n(CH_3CH_2COOH) = cV =   0.815 x 0.025 = 0.020375 mol   (1)
n(OH) $_{excess}$ = 0.0293 – 0.0204 = 0.008905 mol   (1)
$[OH^-]$ = 0.0089 / 0.550 = 0.1619 mol L^{-1} (1)
Thus $[H^+]$ = 10^{-14} / 0.162   =   6.1766 x 10^{-14} mol L^{-1} (1)
Hence pH = -log $[H^+]$ = -log (6.18 x 10^{-14}) =   13.2   (1)

Question 32   **(14 marks)**

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

$$2\ FeS_2(s) + 7\ O_2(g) + 2\ H_2O(l) \rightarrow 2\ Fe^{2+}(aq) + 4\ SO_4^{2-}(aq) + 4\ H^+(aq)$$
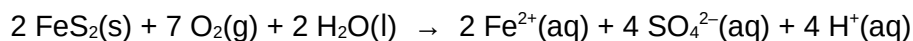
These types of soils are called acid sulfate soils. The pH of groundwater in these soils will decrease. If this groundwater discharges into lakes and rivers it will also cause their pH to decrease.

(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**

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

Av Titre = 3.95 + 3.90 + 4.00   = 3.95 mL   (1)

$$ \mathbf{3}$$

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

n = cV   =   0.0050 x 0.00395   =   1.975 x 10^{-5} mol   (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 x 10^{-5} mol   (1)

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

pH = -log[H$^+$]   = -log (3.95 x 10^{-4})   = 3.40   (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 33   **(14 marks)**

The cell, $Cu(s) / Cu^{2+}(aq)$ and $Cl_2(g) / Cl^-(aq)$ with a platinum electrode, was set up as shown in the diagram below. **Beaker A** contained a 1.00 mol L^{-1} aqueous solution of ammonium chloride, and the filter paper shown in the diagram was soaked in an aqueous solution of potassium nitrate before being placed in the two beakers.

Cathode (+)

(1)

Anode (-)

(1)

Cl$_2$

Filter paper

Beaker A   Beaker B

Pt Electrode
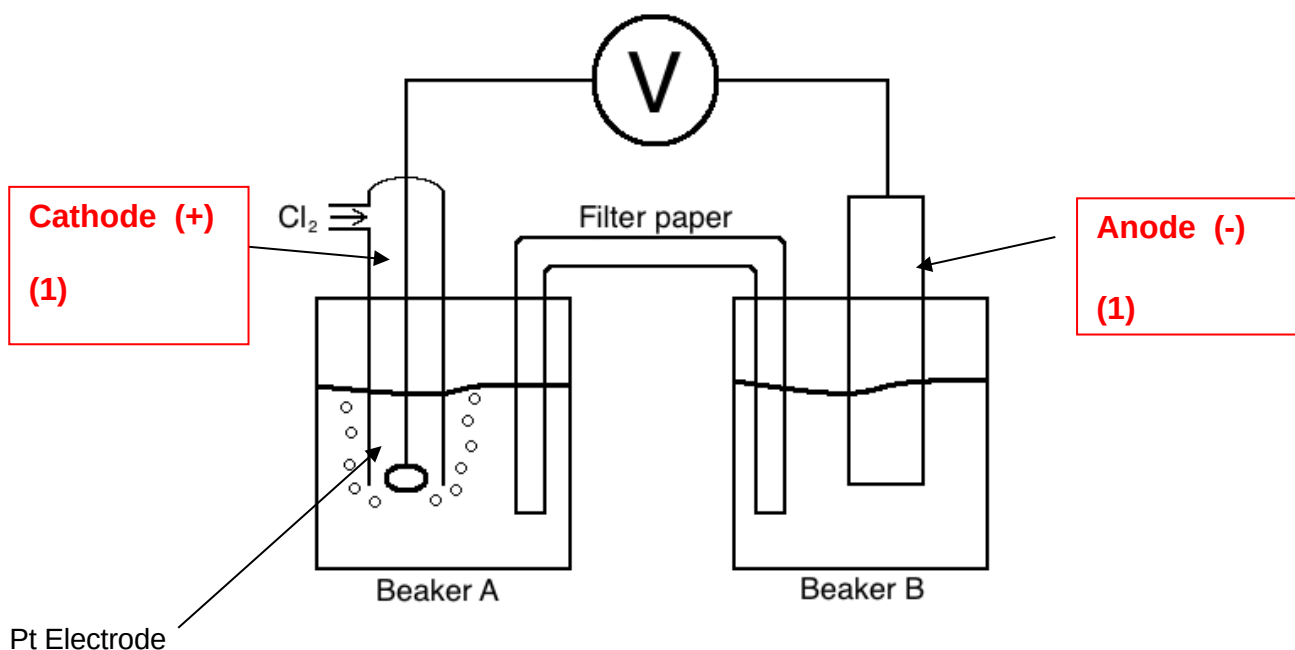
(a)   Give the name or formula of a suitable electrolyte for use in **Beaker B**.   (1 mark)

     Suitable electrolyte = Copper (II) nitrate   or   Cu(NO$_3$)$_2$ solution   (1)

(b)   Label the **anode** and **cathode** in the diagram above, including their respective **polarities**.
                                                                                  (2 marks)

(c)   Give **two** reasons why potassium nitrate was a suitable material for soaking the filter paper.
                                                                                  (2 marks)

(i) **KNO_3 is a "strong electrolyte", thus a high concentration of ions available for transfer between cells to balance the charge. (1)**

(ii) **Neither ion, (K^+) nor (NO_3^-), will form a precipitate with other ions. (1)**

(d) Calculate the maximum theoretical EMF you could measure for the cell. (2 marks)

EMF = (+1.36) + (-0.34) = + 1.02 V

(2)

NO UNIT (V) **-1**

(e) Give **one** reason why the measured cell potential might differ from the value calculated in part (d) above. (1 mark)

Concentrations may not be 1.0 mol L^{-1}, or Cl_2 (g) may not be at STP, or reaction not carried out at 25^0 C. *Accept any one valid reason. (1)

(f) Describe the changes that would be observed in **Beaker B** during the operation of the cell? (2 marks)

Blue colour of solution would intensify. (1)

Mass of salmon pink electrode would decrease. (1)

Do NOT accept "dissolve".

(g) Using relevant chemical theory and a chemical equation, state and explain how the voltmeter reading would change if a few drops of silver nitrate solution were placed in **Beaker A**. (4 marks)

The introduction of Ag^+ ions in Beaker A would cause the following reaction to occur:

Ag^+ (aq) + Cl^- (aq) $\longrightarrow$ AgCl (s) (1)

The silver ions (Ag^+) would remove chloride ions from solution, thus favouring the forward reaction and more chlorine (Cl_2) to dissolve in order to re-establish equilibrium. (1)
This would cause an INCREASE in the voltmeter reading, (cell EMF), (1) as more electrons would be required for the reduction of chlorine. (1)

*Can also accept other valid explanations; (i.e. more electrons would be required for f'wd or reduction reaction).

End of Questions