

CHEMISTRY

UNIT 3

2017

MARKING GUIDE

Section One: Multiple-choice

(50 marks)

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18	<input checked="" type="checkbox"/>	a <input type="checkbox"/> b <input type="checkbox"/> c <input checked="" type="checkbox"/> d
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(2 marks per question)

Section Two: Short answer**(70 marks)****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)

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)

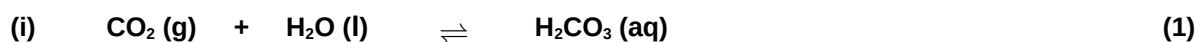
A blue solid dissolves in a colourless solution to produce a blue solution. (2)

(*Must have two observations for both marks).

Question 27**(6 Marks)**

The uptake of carbon dioxide from the atmosphere by the oceans is leading to gradual acidification of the oceans (i.e. the oceans are becoming more acidic). When carbon dioxide dissolves, it reacts with water to form carbonic acid, which in turn forms hydrogen carbonate and then carbonate ions.

- (a) Write balanced chemical equations showing carbon dioxide reacting with water to form carbonic acid, and then the two successive ionisation reactions that carbonic acid undergo in water. (3 marks)



One of the most significant consequences of ocean acidification is the effect that it has on shellfish and other marine life that produce calcium carbonate and relies on it as a major component of the exoskeleton or other supporting structure. If the water is sufficiently acidic, the carbonate structures may not form completely. Ocean acidification is thought to lead to a reduction in the availability of carbonate ions. Further reaction of the dissolved carbon dioxide occurs as shown below.



- (b) Identify a conjugate acid-base pair in this reaction, and explain why it is classified as a Brønsted – Lowry acid-base reaction. (3 marks)

Conjugate A/B pair = $\text{CO}_3^{2-} / \text{HCO}_3^-$ (1) *Also accept $\text{HCO}_3^- / \text{CO}_3^{2-}$

This equation is classified as a Brønsted – Lowry acid-base reaction because in the forward reaction, H_2O donates a proton, thus acting as a B-L acid, (1)
while CO_3^{2-} accepts a proton, thus acting as a B-L base. (1)

Question 28

(6 Marks)

The Bronsted-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
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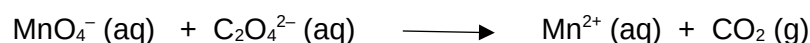
$\text{Mg}(\text{CH}_3\text{COO})_2 (\text{aq})$	Basic (1)	$\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^- \quad (1)$
$\text{NH}_4\text{Cl} (\text{aq})$	Acidic (1)	$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+ \quad (1)$
$\text{NaHSO}_4 (\text{aq})$	Acidic (1)	$\text{HSO}_4^- + \text{H}_2\text{O} \rightleftharpoons \text{SO}_4^{2-} + \text{H}_3\text{O}^+ \quad (1)$

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

Question 29

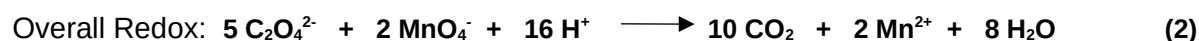
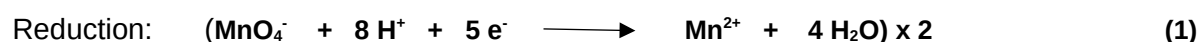
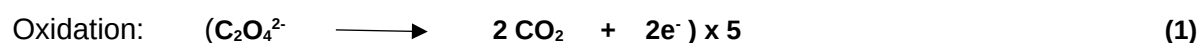
(4 Marks)

The following chemical equation represents an unbalanced redox reaction.



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

(4 marks)



Question 30

(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:



In aqueous solution, bromine, $\text{Br}_2 (\text{aq})$ is brown. Hypobromous acid, $\text{HBrO} (\text{aq})$, and bromide ions, $\text{Br}^- (\text{aq})$ 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 (H₂O). (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 [Br₂] decreases causing the brown colour to fade. (1)**

(b) 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 [Br₂], and the solution becomes more brown. (1)**

Question 31

(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)



$$n(\text{NaOH}) = cV = 2.00 \times 0.025 = 0.05 \text{ mol} \quad (1)$$

$$n(\text{HCl}) = cV = 1.00 \times 0.052 = 0.052 \text{ mol} \quad (1)$$

$$n(\text{HCl})_{\text{excess}} = (0.052 - 0.05) = 0.002 \text{ mol} \quad (1)$$

$$[\text{HCl}] = \frac{n(\text{H}^+)}{V_{\text{Tot}}} = \frac{0.002}{0.077} = 0.025974 \text{ mol L}^{-1} \quad (1)$$

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

Question 32

(9 Marks)

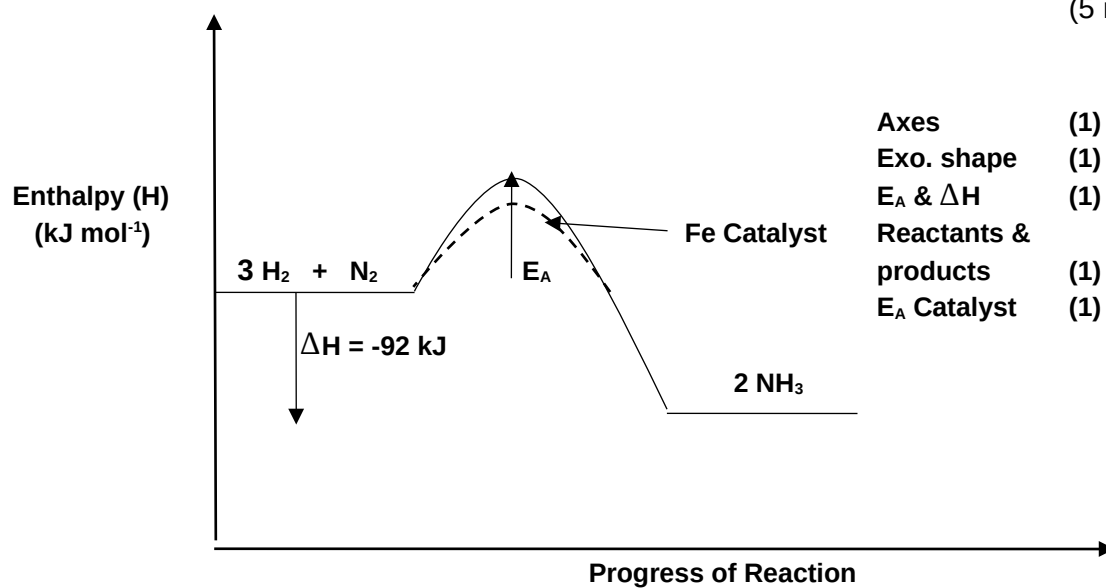
The manufacture of ammonia on an industrial scale is carried out using the Haber process, which relies on the reversible reaction of nitrogen and hydrogen in the presence of an iron catalyst, as shown in the following equation:



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)

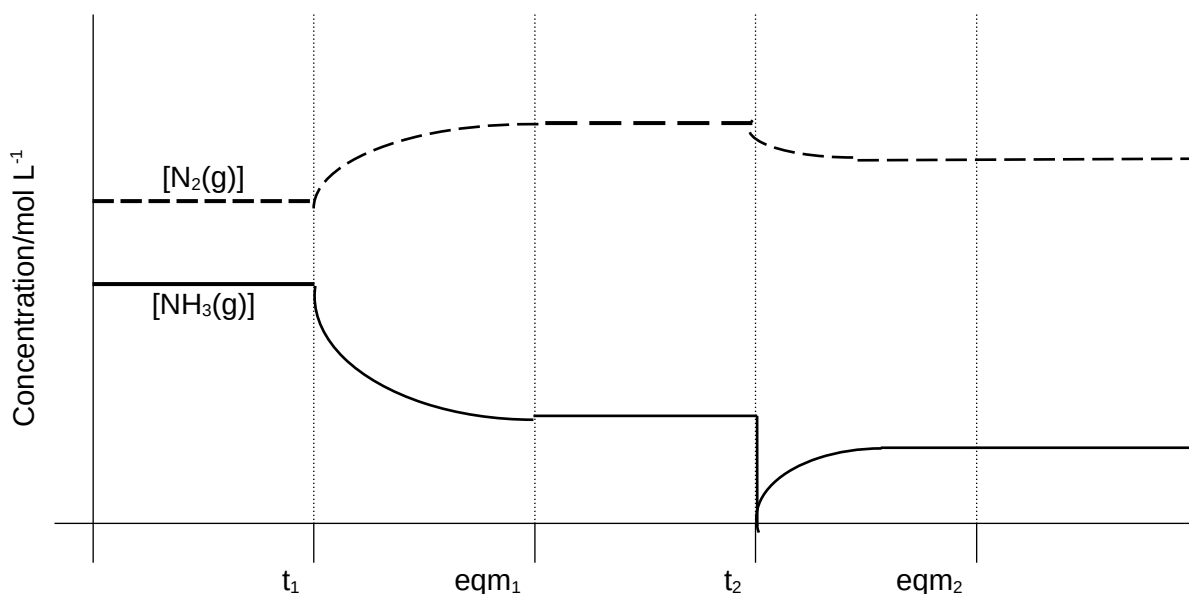


A sealed vessel containing an equilibrium mixture of nitrogen, hydrogen and ammonia was subjected to the following changes in conditions:

- 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 33

(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:



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

$$K = \frac{[(\text{Al}(\text{OH})(\text{H}_2\text{O})_5)^{2+}][\text{H}_3\text{O}^+]}{[(\text{Al}(\text{H}_2\text{O})_6)^{3+}]} \quad (1)$$

(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 $[(\text{Al}(\text{H}_2\text{O})_6)^{3+}]$. (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 $[\text{H}_3\text{O}^+]$ 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 $[\text{H}_3\text{O}^+]$ 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 forward 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 $[\text{H}^+]$, 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)

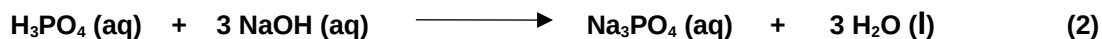
Question 34

(8 Marks)

Phosphoric acid is a weak, **triprotic** acid. In an experiment, a solution of approximately 0.2 mol L^{-1} phosphoric acid (H_3PO_4) 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 phosphoric acid added from the burette.

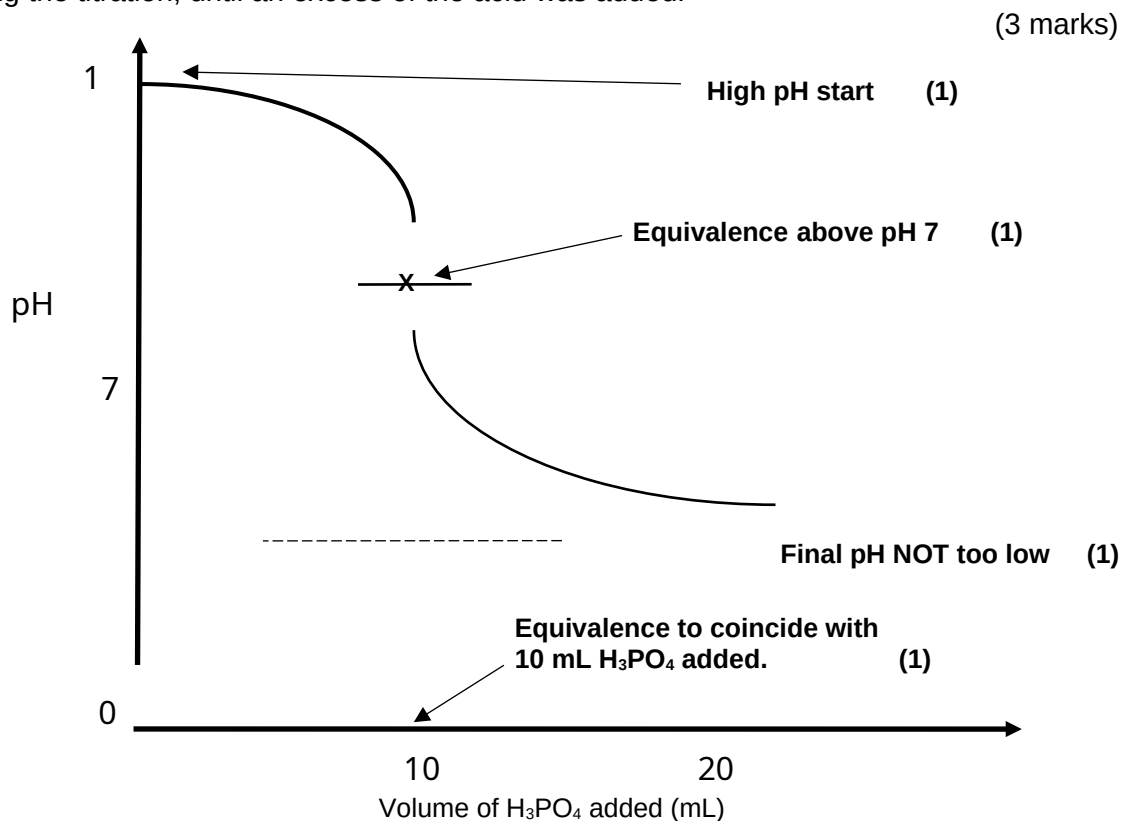
- (a) Write a balanced molecular equation, including state symbols, for the reaction occurring.

(2 marks)



**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)



- (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)

Methyl orange
(pH 3.1 – 4.4)

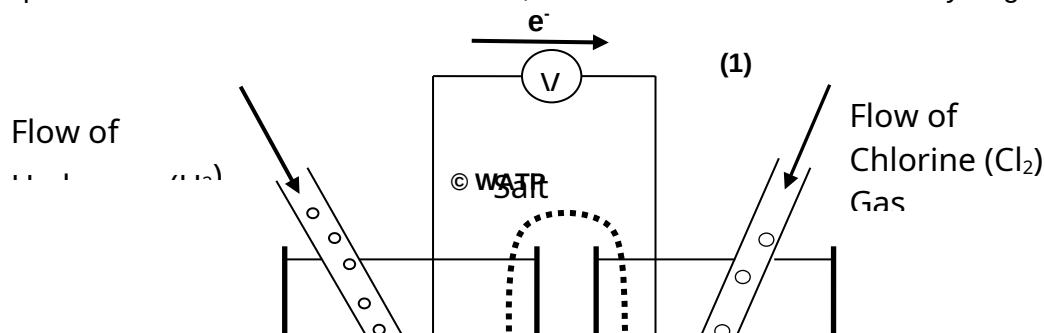
Phenolphthalein
(pH 8.3 – 10.0)

Bromothymol blue
(pH 6.0 – 7.6)

Question 35

(6 Marks)

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



- (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. (3 marks)

Cathode half-equation:	$\text{Cl}_2 + 2\text{e}^- \longrightarrow 2\text{Cl}^-$	(1)	$E^\circ = +1.36\text{ V}$
Anode half-equation:	$\text{H}_2 \longrightarrow 2\text{H}^+ + 2\text{e}^-$	(1)	$E^\circ = 0.00\text{ V}$
Overall equation:	$\text{Cl}_2 + \text{H}_2 \longrightarrow 2\text{Cl}^- + 2\text{H}^+$	(1)	

- (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)

$$\text{E.m.f.} = (+1.36) + (0.00) = +1.36\text{ V} \quad (1)$$

- (c) Show the direction of the flow of electrons in the external circuit by means of an **arrow** “(→)” in the diagram above.

***See on Diagram above.** (1 mark)

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

Platinum is INERT so it will not take part in the reaction. (1)
***Can also accept, “will allow for electron transfer”.**

Question 36

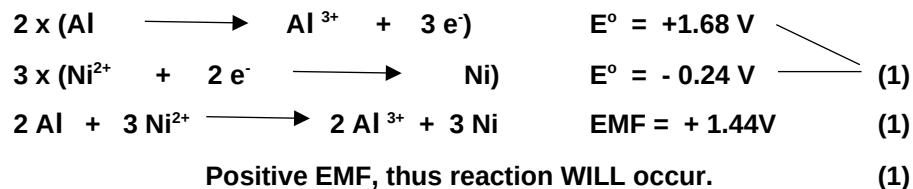
(6 Marks)

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° 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.

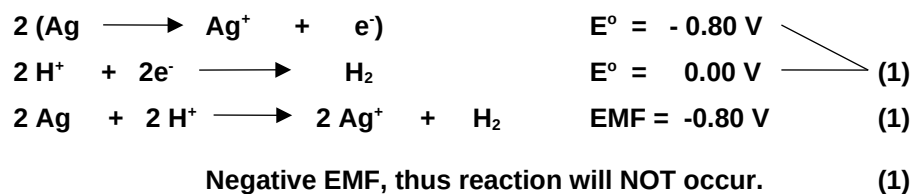
- (a) A piece of aluminium metal is placed in a 1.00 mol L⁻¹ nickel nitrate solution.

(3 marks)



- (b) Silver metal is added to a 1.00 mol L⁻¹ sulfuric acid solution.

(3 marks)



***Note:** Overall redox equation NOT necessary, as reaction will not occur.

End of Section Two

Turn to next page

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 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: 70 minutes.

Question 37

(16 marks)

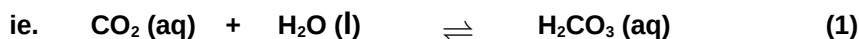
Rising carbon dioxide levels in the atmosphere are believed to play an important role in the life of organisms known as calcifiers, a group that includes many forms of coral and crustaceans. These organisms use a precipitation reaction between calcium ions and carbonate ions present in sea-water to form 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 $p(\text{CO}_2(\text{g}))$ will lead to an increase in $[\text{CO}_2(\text{aq})]$ in the oceans. (1)



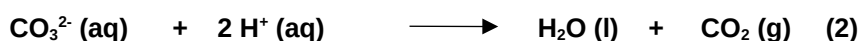
Thus an increase in $[\text{CO}_2(\text{aq})]$ will lead to an increased rate of collision of reactants, thus favouring the Fwd reaction rate, leading to more $\text{H}_2\text{CO}_3(\text{aq})$, hence a higher $[\text{H}^+(\text{aq})]$ and a lower pH. (1)

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 $0.0502 \text{ mol L}^{-1}$.
- 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)



(c) Calculate the number of moles of nitric acid titrated from the burette. (1 mark)

$$n(\text{HNO}_3) = cV = 0.0502 \times 0.03505 = 0.00176 \text{ mol} \quad (3\text{SF}) \quad (1)$$

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

$$n(\text{CO}_3^{2-})_{\text{in 20 mL}} = \frac{1}{2} n(\text{HNO}_3) \quad (1)$$

$$= 0.000880 \text{ mol} \quad (3\text{SF}) \quad (1)$$

- (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. (5 marks)

$$n(\text{CO}_3^{2-}) \text{ in 250 mL} = 250 / 20 \times 0.000879755 = 0.010997 \text{ mol} \quad (2)$$

$$n(\text{CaCO}_3) = n(\text{CO}_3^{2-}) = 0.010997 \text{ mol} \quad (1)$$

$$m(\text{CaCO}_3) = nM = 0.010997 \times 100.09 = 1.10 \text{ g} \quad (1)$$

$$\%(\text{CaCO}_3) \text{ in shells} = (1.10 / 2.17) \times 100 = 50.7\% \text{ (3SF)} \quad (1)$$

- (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)

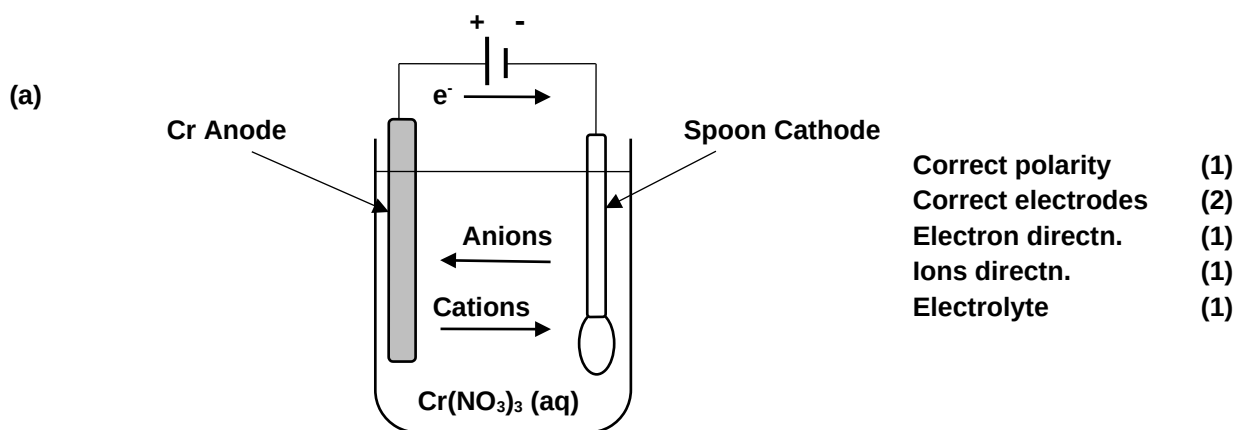
Question 38

(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:

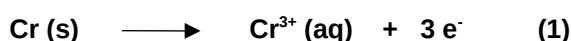
- (a) How the cell can be constructed. (A diagram with clear labels for the anode, cathode, electrolyte, direction of flow of electrons and ions). (6 marks)
- (b) Describe the processes occurring at each electrode. (Including half-equations). (4 marks)
- (c) Observations made at each electrode. (2 marks)
- (d) The role of the electrolyte. (1 mark)
- (e) An example for the industrial importance or application of the process. (1 marks)



- (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)



At the anode, the positive terminal of the cell ensures the oxidation of the chromium electrode to produce chromium ions in solution. (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 39

(22 marks)

Propanoic acid, $\text{CH}_3\text{CH}_2\text{COOH}$, 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
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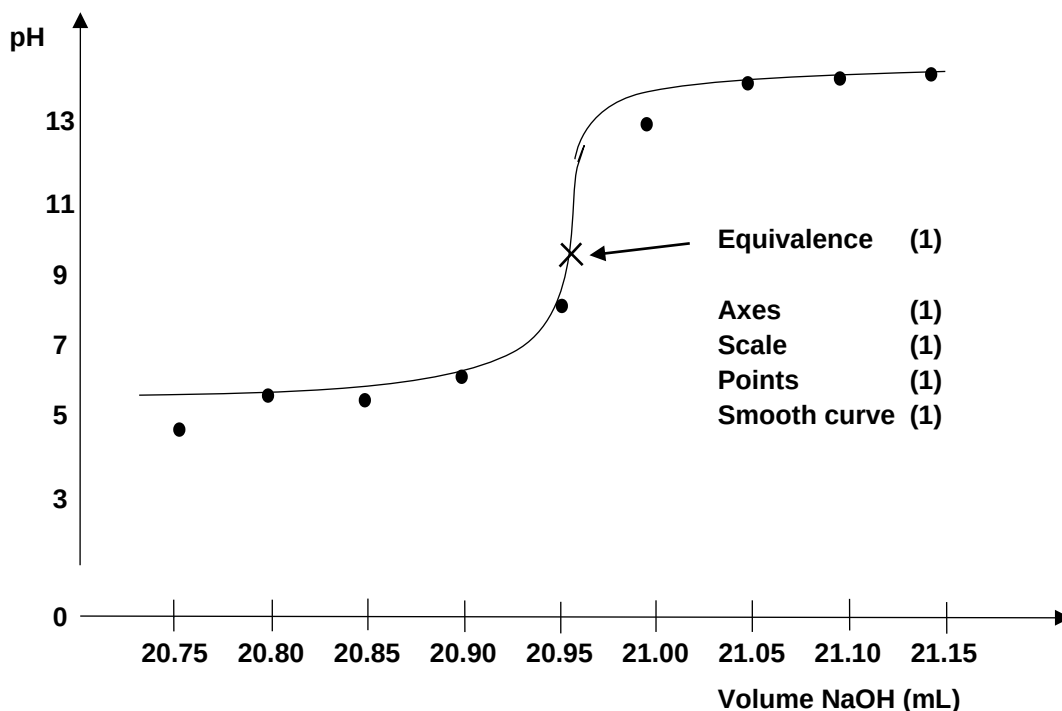
pH of solution	4.7	5.3	5.2	5.6	7.9	12.7	13.0	13.2	13.3
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- (a) Explain why a failure to standardise the sodium hydroxide solution would have led to a systematic error, and what effect it would have on the calculated value for the concentration of the 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 CO₂ in the air, and/or relatively low molar mass may lead to a significant increase in % weighing of error.

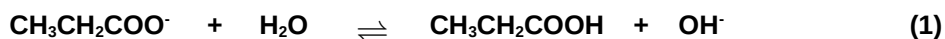
- (b) Plot the results from the experiment on the graph paper provided below, and use your graph to estimate the pH at the equivalence point. Include clearly labelled axes and an appropriate scale. (5 marks)



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

- (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, CH₃CH₂COONa; (1)
and since propanoic acid is a weak acid, the propanoate ion will hydrolyse as follows:



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 $\text{CH}_3\text{CH}_2\text{COOH}$ and $\text{CH}_3\text{CH}_2\text{COONa}$ in the flask, (ie. the weak acid and its salt – a buffer solution). (1)



As NaOH was added, $\text{OH}^- + \text{H}^+ \rightleftharpoons \text{H}_2\text{O}$. 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 $[\text{H}^+]$ 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)

$$n(\text{NaOH}) = cV = 0.976 \times 0.030 = 0.02928 \text{ mol} \quad (1)$$

$$n(\text{CH}_3\text{CH}_2\text{COOH}) = cV = 0.815 \times 0.025 = 0.020375 \text{ mol} \quad (1)$$

$$n(\text{OH})_{\text{excess}} = 0.0293 - 0.0204 = 0.008905 \text{ mol} \quad (1)$$

$$[\text{OH}^-] = 0.0089 / 0.550 = 0.1619 \text{ mol L}^{-1} \quad (1)$$

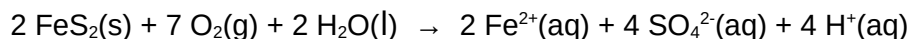
$$\text{Thus } [\text{H}^+] = 10^{-14} / 0.162 = 6.1766 \times 10^{-14} \text{ mol L}^{-1} \quad (1)$$

$$\text{Hence pH} = -\log [\text{H}^+] = -\log (6.18 \times 10^{-14}) = 13.2 \quad (3\text{SF}) \quad (1)$$

Question 40

(14 marks)

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



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.

As the reaction proceeds, H^+ are produced, thus increasing $[\text{H}^+]$, 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 mol L^{-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)

$$\text{Av Titre} = \frac{3.95 + 3.90 + 4.00}{3} = 3.95 \text{ mL} \quad (1)$$

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

$$n = cV = 0.0050 \times 0.00395 = 1.975 \times 10^{-5} \text{ mol} \quad (3 \text{ x SF}) \quad (1)$$

(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)

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

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

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

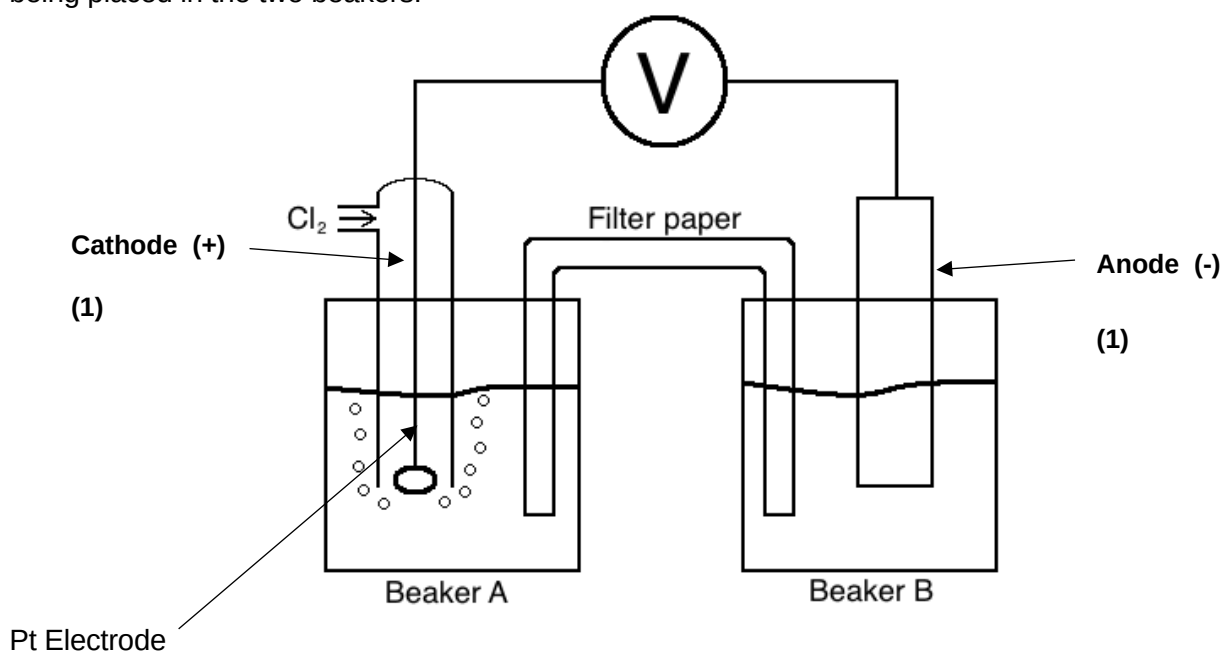
(e) Complete the following table (6 marks)

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 41

(14 marks)

The cell, $\text{Cu(s)} / \text{Cu}^{2+}(\text{aq})$ and $\text{Cl}_2(\text{g}) / \text{Cl}^-(\text{aq})$ 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.



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

Suitable electrolyte = Copper (II) nitrate or $\text{Cu}(\text{NO}_3)_2$ 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)

$$\text{EMF} = (+1.36) + (-0.34) = +1.02 \text{ V} \quad (2)$$

(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 $\text{Cl}_2 (\text{g})$ may not be at STP, or reaction not carried out at 25°C . *Accept any one valid reason. (1)

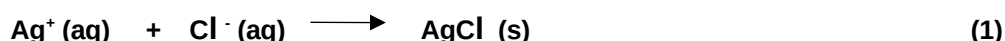
(f) Describe the changes that would be observed in **Beaker B** during the operation of the cell? (2 marks)

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:



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

