CHEMISTRY
UNIT 3
2017

MARKING GUIDE

Section One: Multiple-choice **(50 marks)**

1	a □ b □ c □ d ■	11	a ■ b □ c □ d □	21	a □ b □ c ■ d □
2	a □ b □ c ■ d □	12	a □ b □ c □ d ■	22	a □ b ■ c □ d □
3	a □ b □ c □ d ■	13	a □ b ■ c □ d □	23	a □ b ■ c □ d □
4	a □ b □ c □ d ■	14	a ■ b □ c □ d □	24	a □ b □ c ■ d □
5	a □ b □ c ■ d □	15	a □ b ■ c □ d □	25	a □ b □ c □ d ■

6	a □ b ■ c □ d □	16	a □ b ■ c □ d □	
7	a □ b □ c ■ d □	17	a □ b □ c ■ d □	
8	a □ b □ c ■ d	18	a ■ b □ c □ d □	(2 marks per question)

9	☐ a ☐ b ☐ c ■ d ☐	19	a ■ b ☐ c ☐ d ☐
10	a ☐ b ■ c ☐ d ☐	20	a ☐ b ☐ c ■ d ☐

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)

(i) $CO_2\ (g) + H_2O\ (l) \rightleftharpoons H_2CO_3\ (aq)$ **(1)**

(ii) $H_2CO_3\ (aq) + H_2O\ (l) \rightleftharpoons HCO_3^-\ (aq) + H_3O^+\ (aq)$ **(1)**

(iii) $HCO_3^-\ (aq) + H_2O\ (l) \rightleftharpoons CO_3^{2-}\ (aq) + H_3O^+\ (aq)$ **(1)**

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.

$$CO_2\,(g) \quad + \quad CO_3^{2-}\,(aq) \quad + \quad H_2O\,(l) \quad \rightleftharpoons \quad 2\,HCO_3^-\,(aq)$$

(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 = CO_3^{2-} / HCO_3^- (1) *Also accept HCO_3^- / 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
$Mg(CH_3COO)_2$ (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 29 **(7 marks)**

Phosphate buffered saline (PBS) is a solution which is commonly used in biological research. It was specifically designed so that the ion concentrations of the buffer solution match those found in the human body. The table below gives a standard 'recipe' for making PBS. The four salts are dissolved in water to produce the concentrations indicated.

	Final concentration when dissolved in distilled water	
Salt	**Conc. (g L^{-1})**	**Conc. (mmol L^{-1})**
NaCl	8.0	137
KCl	0.2	2.7

Na_2HPO_4	1.42	10
KH_2PO_4	0.24	1.8

(a) Which salt components would produce the buffering effect observed in PBS? Explain your answer. (2 marks)

- **Na_2HPO_4 and KH_2PO_4**
- **The HPO_4^{2-} / $H_2PO_4^-$ are a weak conjugate acid-base pair**

(b) Write an equation showing the buffering system that would form. (1 mark)

- **$H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+$ (B1)**
  OR
- **$H_2PO_4^- + OH^- \rightleftharpoons HPO_4^{2-} + H_2O$ (B2)**

(c) Explain how this buffer is able to resist a change in pH when a small amount of NaOH(aq) is added. (3 marks)

9

- **Increase OH- concentration (1)**
- **The addition of NaOH neutralises the H_3O^+ (B1) / increases the concentration of OH^- (B2)**
- **The system then favours the forward reaction to produce more H_3O^+ (B1) / reduce the amount of OH^- (B2), thereby maintaining a constant pH**

PBS is specially designed for use in molecular biology and microbiology labs, so it is made to particular specifications.

d) Describe one way you could increase the buffering capacity of PBS if you did not have to take
into account its biological uses. (1 marks)

- **the buffering capacity of PBS could be increased by combining the HPO_4^{2-} / $H_2PO_4^-$ in equimolar amounts, OR**
- **by increasing the concentration of both HPO_4^{2-} / $H_2PO_4^-$**

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:

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

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, 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)**

e) 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 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)

$$\mathbf{NaOH + HCl \longrightarrow NaCl + H_2O}$$

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

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

$pH_{\ solution} = -\log [H^+] = -\log (0.025974) = 1.59$ (1)

Question 32 **(9 Marks)**

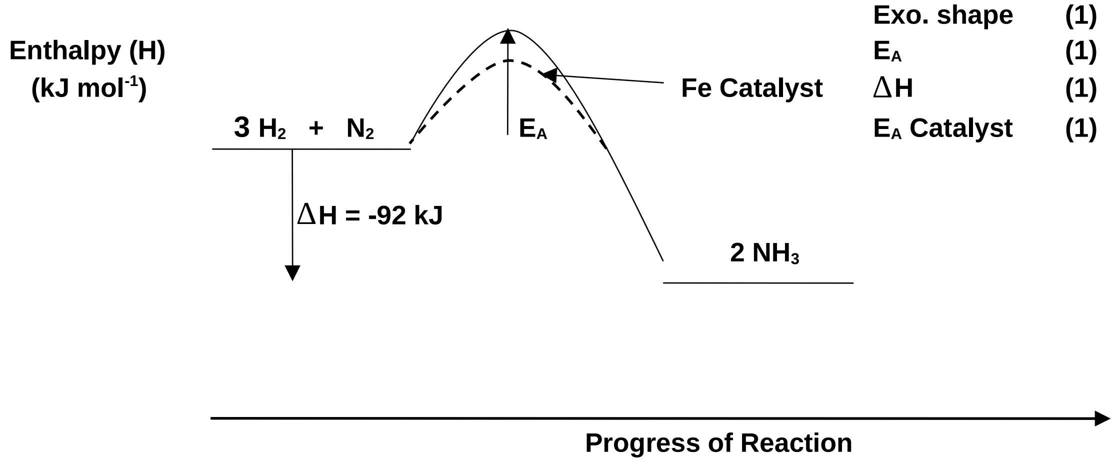
The manufacture of ammonia on an industrial scale is carried out using the Haber process, which relies on the reversible reaction of nitrogen and hydrogen in the presence of an iron catalyst, as shown in the following equation:

$$N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g) \qquad \Delta H = -\ 92\ kJ\ mol^{-1}$$

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)



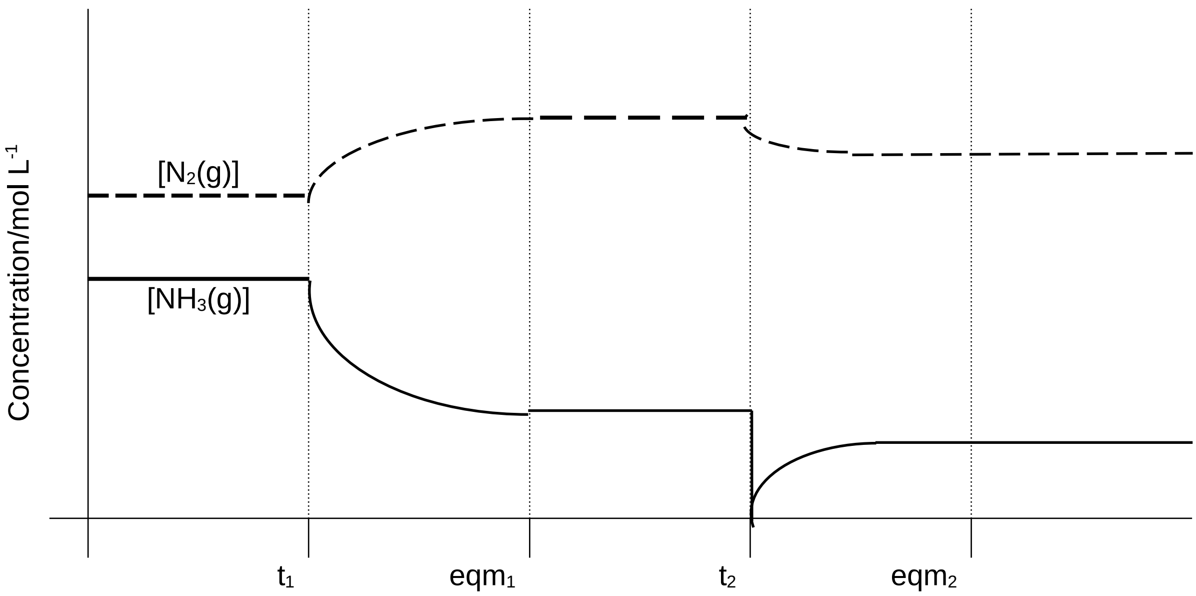
Axes	**(1)**
Exo. shape	**(1)**
E_A	**(1)**
ΔH	**(1)**
E_A Catalyst	**(1)**

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 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:

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

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

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

(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 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 $[H^+]$, 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)**

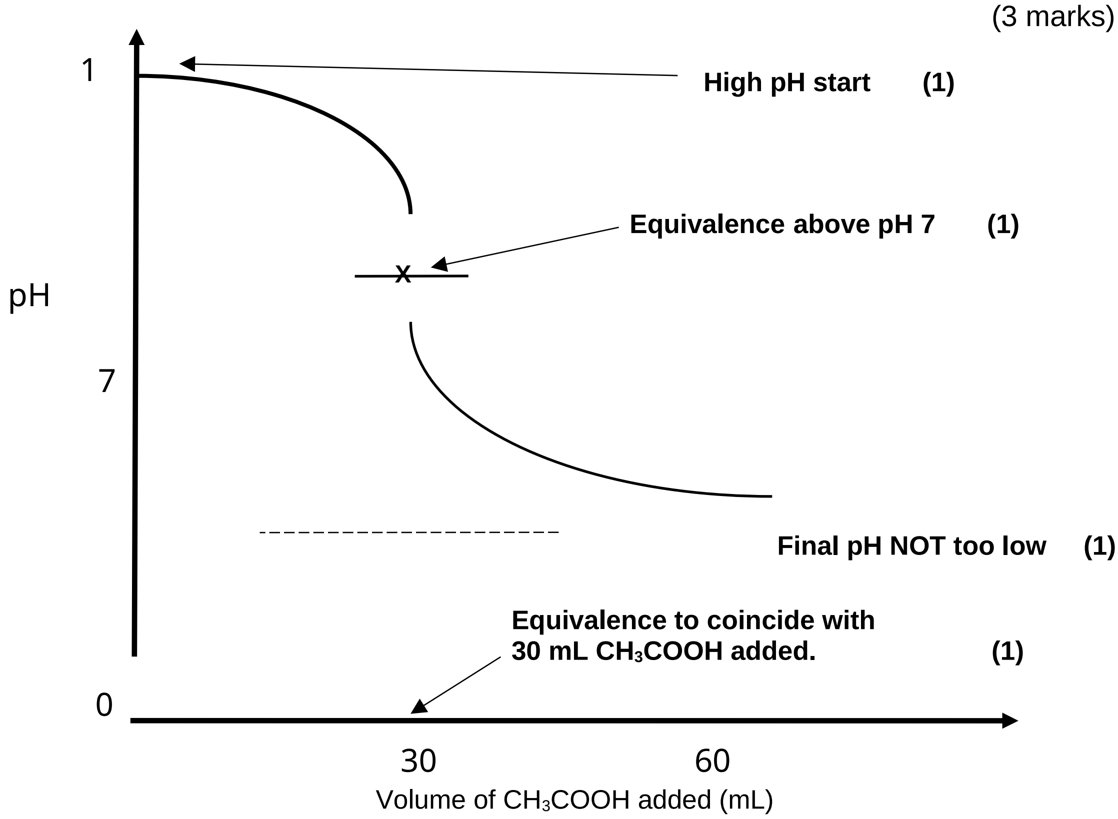
Ethanoic acid is a weak, **monoprotic** acid. In an experiment, a solution of approximately 0.2 mol L^{-1} ethanoic acid (CH_3COOH) is titrated with a standard solution of 0.200 mol L^{-1} sodium hydroxide in order to determine the accurate concentration of the acid. 30.00 mL of the sodium hydroxide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

$$CH_3COOH (aq) + NaOH (aq) \longrightarrow NaCH_3COO (aq) + H_2O (l)$$ **(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)



(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 <u>NaOH solution</u>. **(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 (pH 3.1 – 4.4)	**Phenolphthalein** (pH 8.3 – 10.0)	**Bromothymol blue** (pH 6.0 – 7.6)

Question 35 **(9 marks)**

Oxalic acid ($H_2C_2O_4$) is an organic acid, found in high levels in foods such as almonds, banana, rhubarb and spinach. It is a weak, diprotic acid which has many uses in the laboratory, such as in volumetric analysis where it can be used as a primary standard.

(a) Explain what is meant when oxalic acid is referred to as a 'weak, diprotic acid'. Use relevant chemical equations to support your answer. (4 marks)

- **'weak' indicates ionisation of oxalic acid does not go to completion**
- **'diprotic' indicates each molecule of oxalic acid contains 2 ionisable/acidic hydrogen atoms**
- $H_2C_2O_4 + H_2O \rightleftharpoons HC_2O_4^- + H_3O^+$ **OR** $H_2C_2O_4 \rightleftharpoons HC_2O_4^- + H^+$
- $HC_2O_4^- + H_2O \rightleftharpoons C_2O_4^{2-} + H_3O^+$ **OR** $HC_2O_4^- \rightleftharpoons C_2O_4^{2-} + H^+$

Some oxalic acid dihydrate crystals were used to produce a primary standard for use in a titration. 4.434 g of $H_2C_2O_4.2H_2O(s)$ was dissolved in water and made up to 250.0 mL in a volumetric flask.

(b) Calculate the concentration of the oxalic acid primary standard. (2 marks)

$n(H_2C_2O_4.2H_2O)$ = m/M
= 4.434 / 126.068
= 0.0351715 mol

$c(H_2C_2O_4)$ = n/V
= 0.0351715 / 0.2500
= 0.140686 mol L^{-1}
= 0.1407 mol L^{-1} (4SF)

The oxalic acid solution was then used to standardise some aqueous potassium hydroxide. A 20.00 mL sample of KOH(aq) required 17.85 mL of oxalic acid to reach equivalence. The relevant chemical equation for the titration is shown below.

$$2\ KOH(aq) + H_2C_2O_4(aq) \rightarrow 2\ H_2O(l) + K_2C_2O_4(aq)$$

(c) Calculate the concentration of KOH(aq). (3 marks)

$n(H_2C_2O_4)$ = cV
= 0.140686 x 0.01785
= 0.0025112 mol

n(KOH) = 2 x $n(H_2C_2O_4)$
= 0.0050225 mol

c(KOH) = n/V
= 0.0050225 / 0.02000
= 0.251124 mol L^{-1}
= 0.2511 mol L^{-1} (4SF)

End of Section Two

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

$$CO_3^{2-}(aq) + 2H^+(aq) \longrightarrow H_2O(l) + CO_2(g)$$ **(2)**

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

$n(HNO_3) = cV = 0.0502 \times 0.03505 = 0.00176$ mol (3SF) (1)

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

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

$= 0.000880$ mol (3SF) (1)

(d) 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)

$n(CO_3^{2-})_{\text{in 250 mL}}$ = 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)

$\%(CaCO_3)_{\text{in shells}}$ = (1.10 / 2.17) x 100 = 50.7% (3SF) (1)

(e) 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 37 (14 marks)

In the process of cheese making a bacterial culture is added to milk, which causes the milk to separate into the curds (solid cheese) and whey (remaining liquid). During this process the bacteria convert the lactose present in milk, into lactic acid, $CH_3CHOHCOOH$. Lactic acid is a weak, monoprotic, organic acid. Cheese makers use the concentration of lactic acid in the whey to determine when the reaction has proceeded to the extent that the cheese (curds) are ready for consumption or storage.

The concentration of lactic acid present in the whey can be determined at any time during the cheese making process via a simple titration. This usually involves taking a 10 mL sample of whey and titrating it against some standard sodium hydroxide solution, NaOH(aq), using phenolphthalein as an indicator.

The reaction that took place in the titration is shown below.

$$CH_3CHOHCOOH(aq) + NaOH(aq) \rightarrow H_2O(l) + CH_3CHOHCOONa(aq)$$

(a) Explain why phenolphthalein indicator is used. Use a chemical equation to support your answer. (3 marks)

- **the equivalence point of this titration is basic**
- **due to the production of the basic salt $CH_3CHOHCOONa$**
- **$CH_3CHOHCOO^- + H_2O \rightleftharpoons CH_3CHOHCOOH + OH^-$**

A cheese maker added 4.00 L of milk to a small sample of bacterial culture. He knew that once the concentration of lactic acid reached 1.25×10^{-2} mol L^{-1} the cheese would be ready. He took a 10.00 mL sample of the whey, added several drops of phenolphthalein, and titrated the sample against a 0.111 mol L^{-1} NaOH solution. 1.15 mL of NaOH was required for equivalence.

(b) Determine the concentration of lactic acid in the whey, and comment on whether or not the cheese maker should allow the reaction to proceed for longer before isolating the curds. (4 marks)

n(NaOH) = cV
= 0.111 x 0.00115
= 1.2765×10^{-4} mol

$n(CH_3CHOHCOOH)$ = $n(NaOH)$
= 1.2765×10^{-4} mol

$c(CH_3CHOHCOOH)$ = n/V
= $1.2765 \times 10^{-4} / 0.01000$
= 1.2765×10^{-2} mol L^{-1}
= 1.28×10^{-2} mol L^{-1} (3SF)

reaction should be stopped now as concentration of lactic acid is > 1.25×10^{-2} mol L^{-1}
<u>therefore the cheese maker should isolate the curds now</u>

(c) Determine the percent by mass of lactic acid present in the whey at this point in time, if the 10.00 mL sample was taken from a total volume of 3.10 L of whey. The density (mass/volume) of the whey is 1.040 kg L^{-1}. (5 marks)

$n(CH_3CHOHCOOH$ in 3.10 L) = cV
= $1.2765 \times 10^{-2} \times 3.1$
= 0.0395715 mol

$m(CH_3CHOHCOOH)$ = nM
= 0.0395715×90.078
= 3.56452 g

m(whey) = V
= 1.040×3.10
= 3.224 kg

= 3224 g

% lactic acid in whey = 3.56452 / 3224 x 100
= 0.11056 %
= 0.111 % (3SF)

(d) The cheese maker only took one 10.00 mL sample of whey to examine, Explain how he could improve the accuracy of his calculated lactic acid concentration. (2 marks)

- **repeat trials**
- **averaging data to calculate titre averages**
- **larger sample size**
- **eliminating outliers/ use concordant results**
- **Use a pipette to measure the 10ml accurately**

Question 38 **(10 marks)**

The fertiliser superphosphate, calcium dihydrogen phosphate ($Ca(H_2PO_4)_2$), was mined for many years on the Pacific island of Nauru. Phosphorus is an essential macronutrient to animals and plants. The fertiliser is now manufactured industrially by reacting sulfuric acid (H_2SO_4) with calcium phosphate "rock phosphate", ($Ca_3(PO_4)_2$).

$Ca_3(PO_4)_2(s)$ + $H_2SO_4(aq)$ → $Ca(H_2PO_4)_2(s)$ + $CaSO_4(s)$ [***unbalanced***]

(a) Write a balanced chemical equation for this process. (1 mark)

$Ca_3(PO_4)_2(s) + 2\ H_2SO_4(aq) \rightarrow Ca(H_2PO_4)_2(s) + 2\ CaSO_4(s)$

In a given day a reactor combines 35 000 kg of impure rock phosphate (75.0% purity, by mass) with 15 000 L of 18.0 M H_2SO_4.

(b) Determine which reactant is the limiting reagent. (5 marks)

$m(Ca_3(PO_4)_2)$ = 35000/100 x 75 = 2.62 x 10^7g 1 mark

$M(Ca_3(PO_4)_2)$ = (3 x 40.08) + (2 x 30.97) + (8 x 16)
= 122.04 + 61.94 + 128 = 310.18g/mol

$n(Ca_3(PO_4)_2)$ = 2.62 x 10^7g/312.34g/mol = 8.46 x 10^4 mol 1 mark

15 000 L of 18.0 M H_2SO_4
n(H_2SO_4) c =n /v 18 = n/15000 = 2.70 x 10^5 1 mark
Assume that H_2SO_4 is the LR
2.70 x 10^5 x ½ = 1.35 x 10^5mol $Ca_3(PO_4)_2$ 1 mark
therefore $Ca_3(PO_4)_2$ is LR 1 mark

(c) Determine the mass of excess reactant remaining after the reaction.

(2 marks)

$nCa_3(PO_4)_2(s)$ → $n(H_2SO_4)$
8.46 x 10^4 x 2 = 1.69 x 10^5
270 000 - 1.69 x 10^5= 1.01 x 10^5 1 mark

1.01 x 10^5 = m /98.076 = 9.88 x 10^6 g 1 mark

(d) What mass of superphosphate (in kg) would be produced, if the conversion process is 80.0 % efficient?

(2 marks)

$nCa_3(PO_4)_2(s)$ = 8.46 x 10^4 mol

$m(Ca(H_2PO_4)_2)$ = 234.02g n=m/M = 83882.9 mol = m/234.02g = 1.98 x 10^7

80% efficient is 19620276.26g/100 x 80 = 1.58 x 10^7 g

Question 39 **(14 marks)**

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

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

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.

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)

Av 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 x 0.00395 = 1.975 x 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 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 (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 **(16 marks)**

Ammonia is an industrially important gas produced by the Haber process, as illustrated by the reaction below:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) ; \quad \Delta H = -92 \text{ kJ mol}^{-1} \text{ (@ } 25^oC)$$

The reaction is catalysed by iron(III) oxide.
The following graph shows the partial pressures (gas concentrations) of the three species involved in the reaction.

(a) Answer the following questions about the above graph**.** (4 marks)

(i) Why does the concentration of H_2 decrease more rapidly than that of the N_2?

Due to the stoic ratio, it is consumed at a rate which is 3 times that of N_2?

(ii) Why do the concentrations of each of the three species stabilise between 20 and 30 minutes?

After 20 mins the system has reached equilibrium i.e. the rate at which products and reactants form is the same

(iii) What has occurred at the 30-minute mark to cause the changes shown in the graph?

Nitrogen has been added to the system i.e. its partial pressure or conc has been increased

(iv) By the 40-minute mark, what difference will the change imposed at the 30-minute mark have made to the rate of:

the forward reaction? Increase
the reverse reaction? Increase

(b) Using the collision theory, explain why the rate of forward reaction is affected by the imposed change at the 30-minute mark.

(2 marks)

Increase [N_2] Increase the number of collisions therefore Increase the number of collisions with the correct orientation and sufficient energy, Increase the rate of the forward reaction

(d) In the Haber process, at 200 atm pressure the yield of ammonia doubles if the temperature is dropped from 500°C to 400°C. Why do the manufacturers continue to use higher temperatures?

(2 marks)

Although the yield is doubled such a large drop in temp would dramatically reduce the rate of reaction(1) and thus take a longer time to achieve the higher yield (1)

b) Likewise if the temperature is kept the same and the pressure inside the reaction vessel is tripled, the yield doubles. Why do the manufacturers continue to use lower pressures.

(2 marks)

Safety of employees
Economic reasons
Increase pressure - increase hazards -increase cost to run at high pressure

Ammonia is used as an intermediate in a number of industrial processes. One such process is the manufacture of nitric acid. The sequence of reactions in the manufacture of nitric acid are:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$
$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$
$$2NO + O_2 \rightarrow 2NO_2$$
$$4NO_2 + 2H_2O + O_2 \rightarrow 4HNO_3$$

c) Determine the mass of nitrogen and its volume at STP, required to produce 5.0 Kg of pure nitric acid.

(3 marks)

$\mathbf{2N_2 + 6H_2 + 8O_2 \rightarrow 4\,H_2O + 4HNO_3}$
moles of HNO_3 = 5000/63.018 = 79.34
mol HNO_3 → mol N_2 → g N_2
79.36 x 2/4 x 28/1 = 1111.04g N_2
n=V/22.71   38.68 = V/22.71 = 900.9L

Question 41 **(14 marks)**

a) Write the precipitation reaction that occurred in step one

$Ba^{2+}_{(aq)}$ + $SO^{2-}_{4(aq)}$ → $BaSO_{4(s)}$

Mark	Description
2	Writes the correct ionic equation.
1	Writes a molecular equation.
0	Question incorrectly answered or not attempted.

b) Write the TWO precipitation reactions that occurred in step two.

$Pb^{2+}_{(aq)}$ **+** $2\,I^{-}_{(aq)}$ → $PbI_{2(s)}$

$Pb^{2+}_{(aq)}$ **+** $SO^{2-}_{4(aq)}$ → $PbSO_{4(s)}$

Mark	Description
4	Identifies correct products. Writes the correct ionic equation.
3 – 2	Writes two molecular equations. Writes one correct ionic equation.
1	Writes a molecular equation. Writes one correct ionic equation.
0	Question incorrectly answered or not attempted.

c) Calculate the masses of potassium iodide and potassium sulfate in the original sample.

Mark	Description
8	m(K_2SO_4) = 1.63 g and m(KI) = 4.37 g
7	Allow for one mathematical error.
	Marks assigned as per guide below.
0	Question incorrectly answered or not attempted.

In the first precipitation reaction

$Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4$

n($BaSO_4$) = $\frac{0.218}{233.39}$

= 9.34×10^{-4} mol 1 mark

= n(Ba^{2+}) in 25 mL

n($BaSO_4$) in 250mL = $9.34 \times 10^{-4} \times 10$ 1 mark

= n(K_2SO_4)

m(K_2SO_4) = $9.34 \times 10^{-3} \times 174.3$

= 1.63 g 1 mark

In the second set of precipitation reactions

(i) $Pb^{2+}_{(aq)}$ **+** $SO^{2-}_{4(aq)}$ → $PbSO_{4(s)}$

Moles of $PbSO_{4(s)}$ in 25 mL = moles of SO_4^{2-} in 25 mL in the first solution

= 9.34×10^{-4} moles 1 mark

Mass of $PbSO_{4(s)}$ = $9.34 \times 10^{-4} \times 303.26$
= 0.283g 1 mark

Mass of $PbI_{2(s)}$ = (0.607 – 0.283) g
= 0.324g 1 mark

Moles of $PbI_{2(s)}$ = $\frac{0.324}{461}$

= 7.02×10^{-4}

Moles of I^- = $2 \times 7.02 \times 10^{-4}$

= 1.404×10^{-3} in 25 mL 1 mark

Moles of I^- in 250 mL = 1.404×10^{-2}

Mass of KI in damp mixture = $1.404 \times 10^{-2} \times 166$

= 2.33g 1 mark