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

| | | | | |
|---|---|---|---|---|
| 1. D | 2. C | 3. B | 4. C | 5. D |
| 6. C | 7. C | 8. D | 9. B | 10. C |
| 11. D | 12. B | 13. B | 14. C | 15. A |
| 16. All | 17. B | 18. D | 19. A | 20. C |
| 21. C | 22. B | 23. C | 24. D | 25. C |

## Section Two: Short answer (70 marks)

### Question 26 (4 marks)

(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)

(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)**

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

(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** **(8 Marks)**

The following chemical equation represents an unbalanced redox reaction.

$MnO_4^-$ (aq) + $C_2O_4^{2-}$ (aq) $\longrightarrow$ $Mn^{2+}$ (aq) + $CO_2$ (g) (4 marks)

Oxidation: **($C_2O_4^{2-}$ $\longrightarrow$ 2 $CO_2$ + 2$e^-$ ) x 5** **(1)**

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

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Overall Redox: **5 $C_2O_4^{2-}$ + 2 $MnO_4^-$ + 16 $H^+$ $\longrightarrow$ 10 $CO_2$ + 2 $Mn^{2+}$ + 8 $H_2O$** **(1)**

a) The following chemical equation represents an unbalanced redox reaction.

$IO_3^-$ (aq) + $SO_3^{2-}$ (aq) $\longrightarrow$ $I_2$ (aq) + $SO_4^{2-}$ (aq)

(4 marks)

Oxidation: (**$H_2O$ + $SO_3^{2-}$ → $SO_4^{2-}$ 2$H^+$ + 2$e^-$ ) x 5** **(1)**

Reduction: (2**$IO_3^-$ + 12 $H^+$ + 10 $e^-$ → $I_2$ + 6 $H_2O$)** **(1)**

Overall Redox: 2**$IO_3^-$ + 2 $H^+$ + 5$SO_3^{2-}$ → 5$SO_4^{2-}$ + $I_2$ + $H_2O$** **(1)**

1 MARK FOR OXIDATION AND REDUCTION IN CORRECT POSITION

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

(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)**

(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 [$Br_2$], and the solution becomes more brown. (1)**

**Question 30** **(5 Marks)**

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

| | | |
|---|---|---|
| $n(NaOH)$ | $= cV = 2.00 \times 0.025 = 0.05$ 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) |

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

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

**Question 31** **(6 marks)**

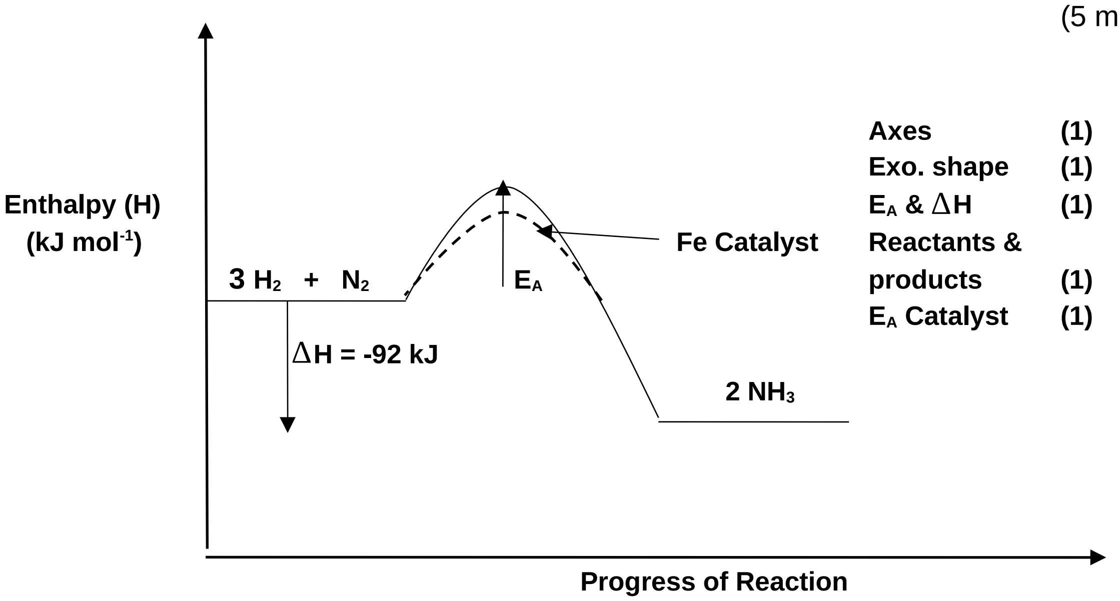
| Substance | pH (acidic, basic or neutral) | Equation |
|---|---|---|
| $Mg(CH_3COO)_2$ (aq) | **Basic (1)** | $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$ **(1)** |
| $NH_4Cl$ (aq) | **Acidic (1)** | $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ **(1)** |
| $NaHSO_4$ (aq) | **Acidic (1)**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

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

**Question 32** **(9 Marks)**

(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.

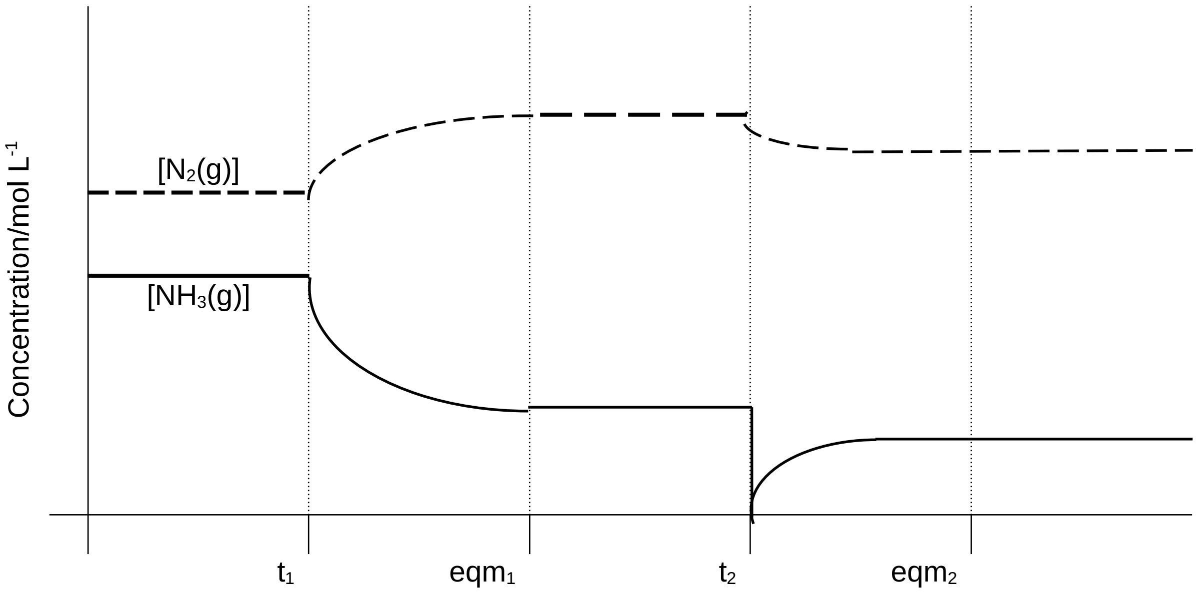
(5 marks)




| | |
|---|---|
| **Axes** | **(1)** |
| **Exo. shape** | **(1)** |
| **$E_A$ & $\Delta H$** | **(1)** |
| **Reactants & products** | **(1)** |
| **$E_A$ Catalyst** | **(1)** |

(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)**

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

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

(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 <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 34** **(9 Marks)**

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

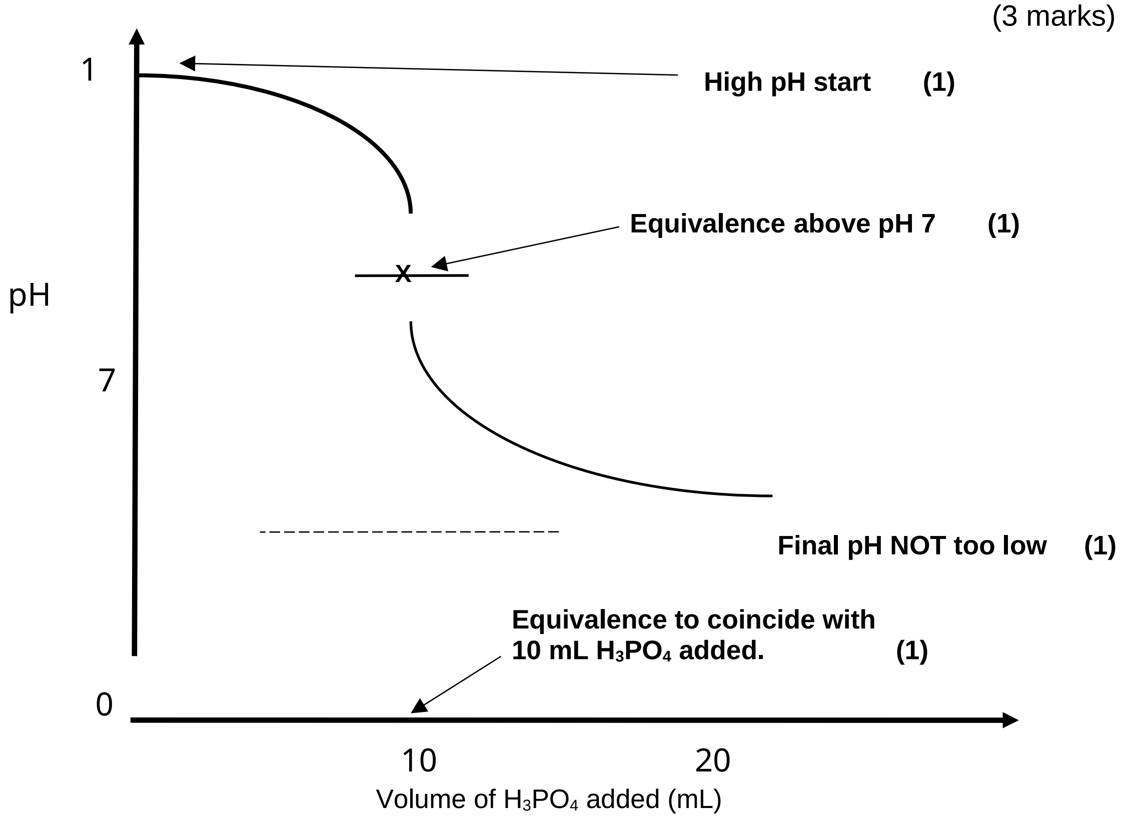
(2 marks)

**$H_3PO_4 (aq) + 3 NaOH (aq) \longrightarrow Na_3PO_4 (aq) + 3 H_2O (l)$ (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? EXPLAIN (2 marks)

**The <u>NaOH solution</u>. (1) to ensure correct concentration present, to not dilute (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)

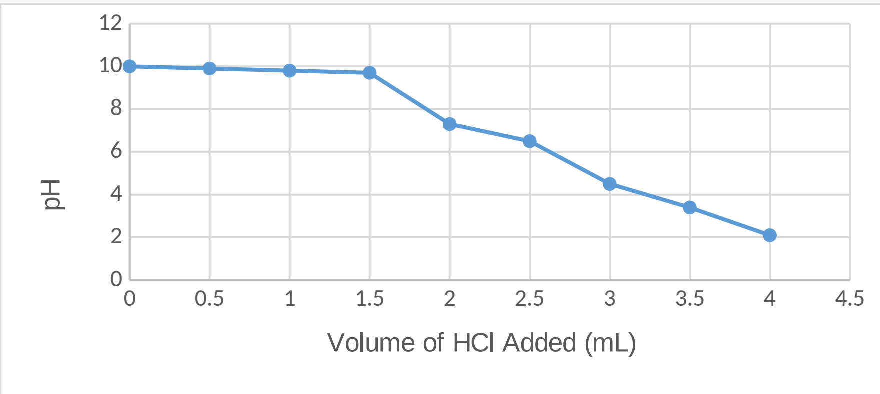
**Phenolphthalein**
(pH 8.3 – 10.0)

**Bromothymol blue**
(pH 6.0 – 7.6)

**Question 35** **(7 Marks)**

| Measurement | Volume of HCℓ(aq) (mL) | pH | Colour of solution |
|---|---|---|---|
| 1 | 0.0 | 10 | green/yellow |
| 2 | 0.5 | 9.9 | green/yellow |
| 3 | 1.0 | 9.8 | green/yellow |
| 4 | 1.5 | 9.7 | green/yellow |
| 5 | 2.0 | 7.3 | yellow |
| 6 | 2.5 | 6.5 | orange |
| 7 | 3.0 | 4.5 | orange |
| 8 | 3.5 | 3.4 | orange |
| 9 | 4.0 | 2.1 | orange |

(a) Plot a graph on the grid below showing the variation of pH against volume of hydrochloric acid added. *(a spare grid is provided at the end of the questions if required)* (4 marks)




| Description | Marks |
|---|---|
| appropriate scales | 1 |
| labelled axis (including units on x axis) | 1 |
| points plotted accurately | 1 |
| line drawn | 1 |
| **Total** | **4** |

(b) Suggest why there was no significant change in pH for the first four measurements.

**(1 mark)**

**It is an equilibrium system therefore as H+ gets used it favours the reverse reaction to produce more H+ ions therefore the pH won't change significantly for a short period of time.**

(c) Based on these results, the students concluded that potassium chromate could be used as an indicator in an acid-base titration. Evaluate this conclusion. (2 marks)

**To be an indicator you require either a weak acid and its conjugate base or a weak base and its conjugate acid. Also the acid base pairs must have two different colours. The chromate dichromate system are not acid –base conjugate pairs.**

**End of Section Two**

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

### Question 36 (11 marks)

(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. (2 marks)

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

**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)**

(b) Write a balanced ionic equation for the titration reaction. (1 mark)

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

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

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

(d) Calculate the number of moles of carbonate in the 20.0 mL aliquots. (1 mark)

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

$= 0.000880$ mol **(3SF) (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. (4 marks)

$n(CO_3^{2-})_{\text{in 250 mL}} = 250 / 20 \times 0.000879755 = 0.010997$ mol **(1)**

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

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

$\%(CaCO_3)_{\text{in shells}} = (1.10 / 2.17) \times 100 = 50.7\%$ **(3SF) (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. (2 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.**

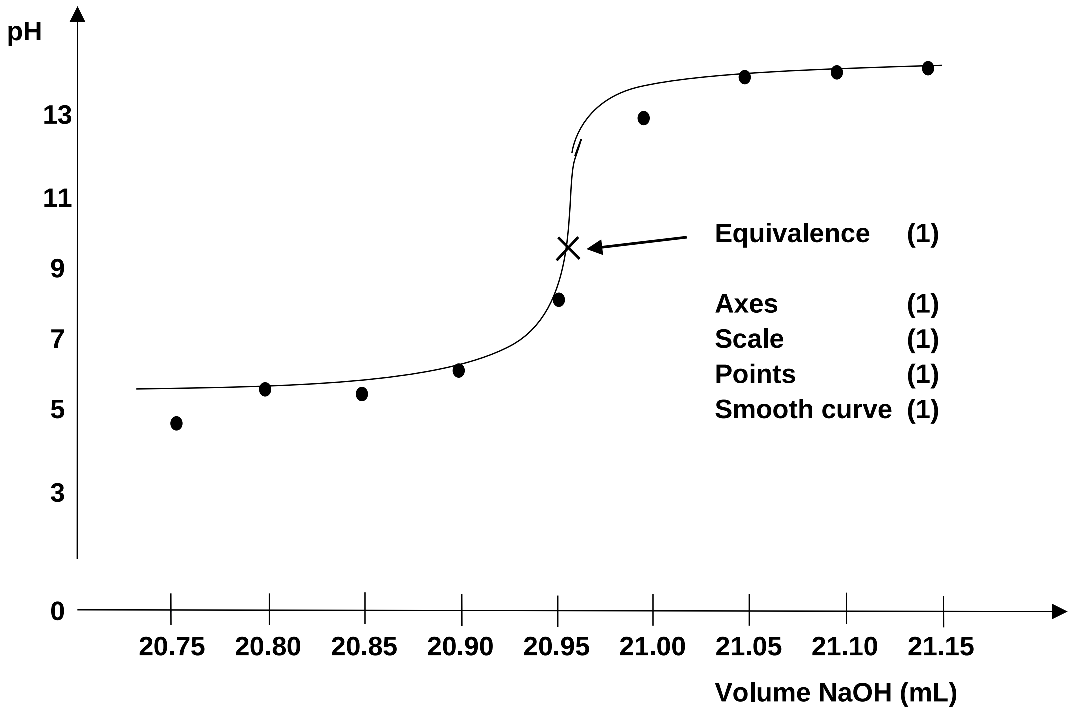
**Question 37** **(22 marks)**

| **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) 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 $CO_2$ 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, $CH_3CH_2COONa$; (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 $molL^{-1}$.

(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 / 0.0550 = 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 (3SF) (1)**

## Question 38 (14 marks)

(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)**

| | 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)

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

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

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

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

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

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

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

(e) Complete the following table (6 marks)

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

**Question 39 (18 marks)**

(a) State an environmental issue linked to the release of oxides of nitrogen into the atmosphere.
**It creates acid rain**
(1 mark)

(b) Balance the equation for this reaction shown below:

**$6NO(g) + 4NH_3(g) \rightarrow 5N_2(g) + 6H_2O(g)$** (2 marks)

(c) 500 kL of ammonia gas is added to the same volume of nitrogen oxide at 25°C and 101.5 kPa. Calculate the amount of nitrogen released into the atmosphere at this temperature and pressure.
**(Note can use gay lussacs law)**
**PV = nRT**
**$n = \dfrac{101.5 \times 500 \times 10^{-3}}{8.314 \times 298}$**
**$= 20.48 \times 10^{-3}$ mol each (2)**
**Calculation for Limiting reagent NO (2)**
**$n(N_2) = 17.06 \times 10^{-3}$ mol (1)**
**$V(N_2) = 416.6 \times 10^{-3}$ L**
**$= 4.17 \times 10^{-5}$ (1)**

(d) Calculate the volume of excess reagent not reacted.

**$n(NH_3)$ xs = 20.48 - 4/6 x 20.48**
**= 6.826 mol (2)**
**Vol $(NH_3)$ = $\frac{6.826 \times 8.314 \times 298}{101.5}$**
**= 166.6 L (1)** (3 marks)

(e) It is important to adjust the amount of ammonia mixing with the chimney gases to give the correct mole ratio of ammonia to nitrogen oxide (NO).
Explain the effect on the composition of the gases released into the atmosphere if the amount of ammonia was too low or too high.
Too low : **acidic NO released into the atmosphere**

Too high: **basic $NH_3$ released into the atmosphere** (1)

(f) The chemical reaction occurring in the ammonia generator is:

$(NH_2)_2CO(aq) + H_2O(l) \leftrightarrow 2NH_3(aq) \rightarrow CO_2(aq) \quad \Delta H = +ve$

In a particular generator a 1:1 mass ratio of urea and water is used.

i which reactant is in excess? Make mass the same ie 1000g
**n (Urea) = $\frac{1000}{60.062}$**
**= 16.649 mol**
**n(water) = $\frac{1000}{18.016}$**
**= 55.5 mol**
**Water in excess** (2 marks)

ii In this chemical reaction is the excess of one reactant an issue.

**If the excess is water it isn't an issue, however an excess of urea as aqueous is ok but vapour will cause toxic vapours to enter the environment.**

(1 mark)

Changing the temperature of the reaction mixture in the ammonia generator can control the amount of ammonia gas produced.

iii Explain the effect of increasing the temperature on the amount of ammonia formed in the generator. (2 marks)

**If the temperature is increased the system will want to decrease the temperature and favour the endothermic reaction. This is the forward reaction hence will produce a greater yield of ammonia.**

**Question 40** **(15 marks)**

$$Cl_2\ (g) + H_2S\ (aq) + H_2O\ (l) \rightarrow H_2SO_4\ (aq) + HCl\ (aq)$$

(a) Balance the equation

$$\mathbf{4\ Cl_2\ (g) + H_2S\ (aq) + 4\ H_2O\ (l) \rightarrow H_2SO_4\ (aq) + 8\ HCl\ (aq)}$$

(2 marks)

The process represented by the equation above is also a redox reaction.

(b) State which substance has been oxidised and which has been reduced. Use oxidation numbers to support your answer. (3 marks)

- **$Cl_2$ reduced**
- **(0) to (-1)**
- **$H_2S$ oxidised**
- **S from (-2) to (+6)**

The tank held 20 000 L of contaminated water. The concentration of hydrogen sulfide in the water was $7.13 \times 10^{-4}$ g $L^{-1}$.

(c) Calculate the volume of chlorine at STP that would be required to remove all the hydrogen sulfide from the water. (4 marks)

**$m(H_2S)$= cV**
**= 0.000713 x 20000**
**= 14.26 g**

**$n(H_2S)$ = m/M**
**= 14.26 / 34.086**
**= 0.4183536**

**$n(Cl_2)$ = 4 x $n(H_2S)$**
**= 0.4183536 x 4**
**= 1.673414**

**$V(Cl_2)$ = 22.71n**
**= 22.71 x 1.673414**
**= 38.00324 L**
**= 38.0 L**

(d) Calculate the final concentration of HCl (in mol $L^{-1}$) that would be present in the tank after the chlorination process was complete. (2 marks)

**n(HCl) = 8 x $n(H_2S)$**
**= 0.4183536 x 8**
**= 3.346829**

**c(HCl) = n/V**
**= 3.346829 / 20000**
**= 0.00016734 mol $L^{-1}$**
**= 0.000167 mol $L^{-1}$ <u>OR</u> $1.67 \times 10^{-4}$ mol $L^{-1}$**

Both of the products in this reaction (hydrochloric and sulfuric acid) are strong acids.

(e) Explain the difference between a strong and weak acid. (2 marks)

- **strong acid completely ionises in solution**
- **weak acid only partially ionises**

(f) Sulfuric acid is 'diprotic'. Explain what this term means, using equations to support your answer. (2 marks)

- **diprotic means 2 acidic / ionisable hydrogens**
- **$H_2SO_4 \rightarrow H^+ + HSO_4^-$ <u>OR</u> $H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$**
- **$HSO_4^- \rightleftarrows H^+ + SO_4^{2-}$ <u>OR</u> $HSO_4^- + H_2O \rightleftarrows H_3O^+ + SO_4^{2-}$**