Section 2 Written questions **36 marks**

Question 1 **(4 marks)**

Sodium hydroxide and ammonia are both bases. When added to water, ammonia is considered to be a Bronsted – Lowry base but sodium hydroxide is not. Explain why, using appropriate equations to illustrate your answer.

✓ NH_3 is a proton acceptor ∴ Bronsted - Lowry

✓ $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

✓ NaOH dissociates as it dissolves ∴ Not Bronsted-Lowry

✓ $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$

Question 2 **(6 marks)**

Increasing carbon dioxide levels in the atmosphere are responsible for increasing ocean acidification. The carbon dioxide dissolves in the water and reacts to form carbonic acid. The carbonic acid then undergoes a series of ionisations and eventually affects the level of carbonate ions in the ocean.

(a) Write equations to represent the dissolution of carbon dioxide and the series of reactions described above

✓ $CO_2(g) \rightleftharpoons CO_2(aq)$

✓ $CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$

(or H^+) ✓ $H_2CO_3(aq) + H_2O \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$

✓ $HCO_3^-(aq) + H_2O \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq)$

(4 marks)

(b) The increased level of acidity in the ocean is having an effect on organisms that produce calcium carbonate (shellfish and crustaceans). Write an ionic equation for the reaction between the acid and calcium carbonate.

$CaCO_3 + 2H_3O^+ \rightleftharpoons Ca^{2+} + 3H_2O + CO_2$

(or H^+)

(2 marks)

Question 3 **(9 marks)**

Hydrogen fluoride is a weak acid.

(a) Write an equation to show what happens when hydrofluoric acid is added to water and identify the conjugate acid – base pairs.

$HF + H_2O \rightleftharpoons H_3O^+ + F^-$

(1 mark)

Pair 1	**Pair 2**
Acid HF	Acid H_3O^+
Base F^-	Base H_2O

(2 marks)

(b) Write an expression for K_a for hydrofluoric acid in the space below.

$$K_a = \frac{[H_3O^+][F^-]}{[HF]}$$

(1 mark)

(c) Write an equation for the reaction between solutions of hydrofluoric acid and barium hydroxide.

$HF + OH^- \rightarrow F^- + H_2O$

(2 marks)

(d) Considering that the concentration of hydronium ions in a hydrogen fluoride solution is quite low, explain why 20mL of a solution of hydrofluoric acid would be neutralized by 10mL of a solution of barium hydroxide of equal concentration.

✓ $n(OH^-) = n(\text{acid})$ $H_3O^+ + OH^- \rightleftharpoons 2H_2O$

- As H_3O^+ is consumed the rate of the

✓ reverse reaction in (a) is decreased

- HF continues to hydrolyse until

✓ it is completely consumed.

(3 marks)

Question 4 **(10 marks)**

(a) Hydrogen cyanide (HCN) is a weak acid. One method of preparing a buffer using hydrogen cyanide is to dissolve a salt of the cyanide ion (e.g. NaCN) in the acid solution so that the number of moles of the acid and conjugate base are equal.

Other methods can be used to prepare buffer solutions.
How would you prepare 150 mL of a HCN/CN^- buffer using 1.0 $molL^{-1}$ hydrogen cyanide and 1.0 $molL^{-1}$ sodium hydroxide solutions. .Show your reasoning.

✓ $HCN + OH^- \rightleftharpoons CN^- + H_2O$

In 100 mL HCN $n(HCN) = .1 \times 1 = 0.1$ mol

✓ In 50 mL NaOH $n(OH^-) = .05 \times 1 = .05$ mol

✓ $n(HCN)$ remaining $= .1 - .05 = .05$ mol

✓ $n(CN^-)$ produced $= .05$ mole

(4 marks)

(b) If the initial concentration of the hydrogen cyanide solution was identical in both of the above methods, which one of the buffers will have the greatest buffer capacity? Show your reasoning.

✓ Buffer 1 (HCN + NaCN). The concentration of the acid / conjugate base is

✓ higher than Buffer 2 (HCN + NaOH)

(2 marks)

(c) 100 mL of a 1.00 molL^{-1} hydrogen cyanide/ cyanide ion buffer solution has 10 drops of 1.00 molL^{-1} sodium hydroxide added to it. Use appropriate equations and collision theory to explain why there is a minimal change in the pH of the solution.

✓ $HCN + H_2O \rightleftharpoons CN^- + H_3O^+$

H_3O^+ reacts with the additional OH^-

$H_3O^+ + OH^- \rightleftharpoons 2H_2O$

✓ The rate of the reverse reaction

✓ decreases due to the consumption of H_3O^+.

Once equilibrium is re-established the H_3O^+

has been almost completely replaced.

✓ Minimal change in $[H_3O^+]$ means minimal change in pH.

(4 marks) ✓

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Question 5 **(7 marks)**

A 25.0 mL solution of nitric acid at 25°C contains 8.5 x 10^{-3} moles of acid.

(a) Calculate the pH of the nitric acid.

$n(H^+) = 8.5 \times 10^{-3}$ mol.

$[H^+] = 8.5 \times 10^{-3} / .025$

✓ $= 0.34$ mol L^{-1}

$pH = -\log_{10}[H^+]$

✓ $= 0.47$

(2 marks)

(b) Calculate the final pH of the solution after 30.0 mL of potassium hydroxide with a pH of 12 is combined with the original 25.0 mL of nitric acid.

✓ $[OH^-] = 10^{-14} / 10^{-12}$

$= 10^{-2}$ mol L^{-1}

✓ $n(OH^-) = c \times V = .01 \times .03$

$= .0003$ mol

$H^+ + OH^- \rightarrow H_2O$

.0085 .0003

✓ $n(H^+)$ remaining $= .0085 - .0003$

$= .0082$ mol

✓ $[H^+] = .0082 / .055$

$= 0.1491$ mol L^{-1}

✓ $pH = 0.82$

(5 marks)

END OF TEST