Q1. In this question, give all values of pH to 2 decimal places.

(a) The ionic product of water has the symbol K_w

(i) Write an expression for the ionic product of water.

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(1)

(ii) At 42°C, the value of K_w is 3.46×10^{-14} mol^2 dm^{-6}.

Calculate the pH of pure water at this temperature.

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(2)

(iii) At 75 °C, a 0.0470 mol dm^{-3} solution of sodium hydroxide has a pH of 11.36. Calculate a value for K_w at this temperature.

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(2)

(b) Methanoic acid (HCOOH) dissociates slightly in aqueous solution.

(i) Write an equation for this dissociation.

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(1)

(ii) Write an expression for the acid dissociation constant K_a for methanoic acid.

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(1)

(iii) The value of K_a for methanoic acid is 1.78×10^{-4} mol dm^{-3} at 25 °C.
Calculate the pH of a 0.0560 mol dm^{-3} solution of methanoic acid.

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(3)

(iv) The dissociation of methanoic acid in aqueous solution is endothermic.

Deduce whether the pH of a solution of methanoic acid will increase, decrease or stay the same if the solution is heated. Explain your answer.

Effect on pH ....

Explanation ....

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(Extra space) ....

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(3)

(c) The value of K_a for methanoic acid is 1.78×10^{-4} mol dm^{-3} at 25°C.
A buffer solution is prepared containing 2.35×10^{-2} mol of methanoic acid and 1.84×10^{-2} mol of sodium methanoate in 1.00 dm^3 of solution.

(i) Calculate the pH of this buffer solution at 25°C.

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(Extra space) ....

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(3)

(ii) A 5.00 cm^3 sample of 0.100 mol dm^{-3} hydrochloric acid is added to the buffer solution in part (c)(i).

Calculate the pH of the buffer solution after this addition.

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(Extra space) ....

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(4)
(Total 20 marks)

Q2. This question is about several Brønsted–Lowry acids and bases.

(a) Define the term *Brønsted–Lowry* acid.

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(1)

(b) Three equilibria are shown below. For each reaction, indicate whether the substance immediately **above** the box is acting as a Brønsted–Lowry acid (**A**) or a Brønsted–Lowry base (**B**) by writing **A** or **B** in each of the six boxes.

(i) $CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$

(1)

(ii) $CH_3NH_2 + H_2O \rightleftharpoons CH_3NH_3^+ + OH^-$

(1)

(iii) $HNO_3 + H_2SO_4 \rightleftharpoons H_2NO_3^+ + HSO_4^-$

(1)

(c) A 25.0 cm^3 sample of 0.0850 mol dm^{-3} hydrochloric acid was placed in a beaker. Distilled water was added until the pH of the solution was 1.25.

Calculate the total volume of the solution formed. State the units.

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(Extra space) ....

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.... **(3)**

(d) At 298 K, the value of the acid dissociation constant (K_a) for the weak acid HX in aqueous solution is 3.01×10^{-5} mol dm^{-3}.

(i) Calculate the value of pK_a for HX at this temperature.
Give your answer to 2 decimal places.

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.... **(1)**

(ii) Write an expression for the acid dissociation constant (K_a) for the weak acid HX.

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.... **(1)**

(iii) Calculate the pH of a 0.174 mol dm^{-3} solution of HX at this temperature.
Give your answer to 2 decimal places.

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(Extra space) ....

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.... **(3)**

(e) An acidic buffer solution is formed when 10.0 cm^3 of 0.125 mol dm^{-3} aqueous sodium hydroxide are added to 15.0 cm^3 of 0.174 mol dm^{-3} aqueous HX. The value of K_a for the weak acid HX is 3.01×10^{-5} mol dm^{-3}.

Calculate the pH of this buffer solution at 298 K.
Give your answer to 2 decimal places.

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(Extra space) ....

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(6)

(Total 18 marks)

Q3. Buffer solutions are important in biological systems and in industry to maintain almost constant pH values.

(a) In the human body, one important buffer system in blood involves the hydrogencarbonate ion, HCO_3^-, and carbonic acid, H_2CO_3, which is formed when carbon dioxide dissolves in water.

(i) Use the following equation to explain how this buffer maintains a constant pH of 7.41 even if a small amount of acid enters the bloodstream.

$$H_2CO_3(aq) \rightleftharpoons H^+(aq) + HCO_3^-(aq)$$

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(ii) In a sample of blood with a pH of 7.41, the concentration of HCO_3^-(aq) ions is 2.50×10^{-2} mol dm^{-3} and the concentration of H_2CO_3(aq) is 1.25×10^{-3} mol dm^{-3}. Calculate a value for the acid dissociation constant, K_a, for carbonic acid at this temperature.

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(5)

(b) In industry, the pH of a solution used to dye cloth must be controlled or else the colour varies.

A solution of dye in a beaker is buffered by the presence of ethanoic acid and sodium ethanoate. In the solution, the concentration of ethanoic acid is 0.15 mol dm^{-3} and the concentration of sodium ethanoate is 0.10 mol dm^{-3}. The value of K_a for ethanoic acid is 1.74×10^{-5} mol dm^{-3} at 298 K.

(i) A 10.0 cm^3 portion of 1.00 mol dm^{-3} hydrochloric acid is added to 1000 cm^3 of this buffered solution.

Calculate the number of moles of hydrochloric acid added.

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(ii) Calculate the number of moles of ethanoic acid and the number of moles of sodium ethanoate in the solution after addition of the hydrochloric acid.

Mol of ethanoic acid after addition ....

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Mol of sodium ethanoate after addition ....

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(iii) Hence calculate the pH of this new solution. Give your answer to 2 decimal places.

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(6)
(Total 11 marks)

Q4. Ammonia and ethylamine are examples of weak Brønsted–Lowry bases.

(a) State the meaning of the term *Brønsted–Lowry base*.

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(1)

(b) (i) Write an equation for the reaction of ethylamine ($CH_3CH_2NH_2$) with water to form a weakly alkaline solution.

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(1)

(ii) In terms of this reaction, state why the solution formed is **weakly** alkaline.

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(1)

(c) State which is the stronger base, ammonia or ethylamine. Explain your answer.

Stronger base ....

Explanation ....

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(Extra space) ....

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(3)

(d) Give the formula of an organic compound that forms an alkaline buffer solution when added to a solution of ethylamine.

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(1)

(e) Explain qualitatively how the buffer solution in part (d) maintains an almost constant pH when a small amount of hydrochloric acid is added to it.

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(Extra space) ....

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(2)
(Total 9 marks)