Course 3-pH Titres Indicators & pH

Question 1

(*b*) Use the data below to sketch (on graph paper) the pH curve for a titration between 20 cm^3 of ethanoic acid and a sodium hydroxide solution added from a burette. (18)

Volume of NaOH added (cm^3)	0	5	10	15	17	19	21	23	25	30	35	40
pH	3.1	4.2	4.8	5.2	5.6	6.0	11.2	11.6	12.1	12.4	12.5	12.6

What indicator would you use for this titration? Use the graph to explain your choice. (7)

Question 2

(*b*) (*i*) Explain how an acid-base indicator, which is itself a weak acid, and may be represented by **HX**, functions. (9)

(*ii*) Draw a clearly labelled diagram of the titration curve you would expect to obtain when 50 cm^3 of a 0.1 M sodium hydroxide (**NaOH**) solution is added slowly to 25 cm^3 of a 0.1 M ethanoic acid (**CH_3COOH**) solution. (9)

(*iii*) Explain with reference to your diagram why phenolphthalein is a suitable indicator for a titration of sodium hydroxide with ethanoic acid. (6)

Question 3

8. (*a*) Define (*i*) acid, (*ii*) conjugate acid, according to the Brønsted-Lowry theory. (8)

In acting as an acid-base indicator methyl orange behaves like a weak acid. Letting **HX** represent methyl orange, it dissociates as follows:

$$\mathbf{HX} \rightleftharpoons \mathbf{H^+} + \mathbf{X^-}$$

In aqueous solution, the undissociated form (**HX**) is red and the dissociated form (**X^-**) is yellow.

Distinguish between a strong acid and a weak acid. (6)

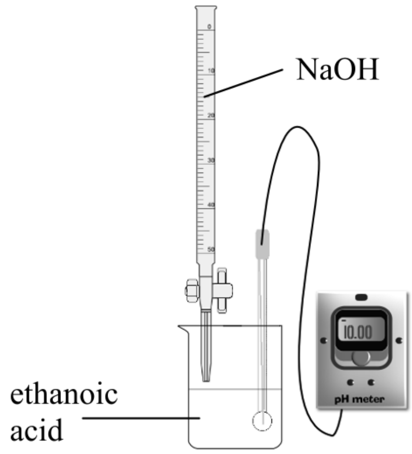
What is the conjugate base of **HX**? (3)

(*b*) State and explain the colour observed when a few drops of the methyl orange solution is added to (*i*) a 0.1 M solution of **HCl**, (*ii*) a 0.1 M solution of **NaOH**. (12)

(*c*) Calculate the pH of (*i*) a 0.1M solution of **NaOH**, (*ii*) a 0.004 M solution of methyl orange, if methyl orange has a K_a value of 3.5×10^{-4}. (9)

Draw a clearly labelled diagram of the pH curve you would expect to obtain when 50 cm^3 of 0.1 M **NaOH** solution is added slowly to 25 cm^3 of a 0.1 M **HCl** solution. Explain by referring to the curve why almost any acid-base indicator can be used in this titration. (12)

Question 4



9. (*a*) Define *acid* according to the theory of

(*i*) Arrhenius,

(*ii*) Brønsted-Lowry. (6)

(*b*) Define pH.

State one limitation of the pH scale. (9)

A sodium hydroxide solution was titrated with a solution of ethanoic acid using the apparatus shown on the right. A number of pH readings, together with the corresponding volumes of sodium hydroxide solution added, are shown in the table.

Volume NaOH (cm^3)	**pH**
0.0	3.3
2.0	4.1
5.0	4.5
10.0	4.8
15.0	5.1
17.5	5.6
19.0	6.2
19.5	6.6
19.8	7.0
20.0	8.9
20.2	10.7
20.5	11.1
22.0	11.6
25.5	12.1
30.0	12.3
40.0	12.4

(*c*) Graph the pH curve for the titration. (15)

(*d*) (*i*) Calculate the initial concentration of ethanoic acid ($\boldsymbol{K}_a = 1.8 \times 10^{-5}$) in the beaker.

(*ii*) Make use of your graph to deduce the volume of sodium hydroxide solution required for neutralisation. (12)

(*e*) What indicator could be used to detect the end point in this titration?

Refer to your pH curve to justify your answer. (8)

Question 5

9. (*a*) Define (*i*) a base, (*ii*) a conjugate acid-base pair, according to Brønsted-Lowry theory. What is the conjugate acid of $\mathbf{H_2O}$? (9)

(*b*) Define pH. (3)

The pH values of 0.10 M solutions of sulfuric acid, hydrochloric acid and methanoic acid are 0.70, 1.00 and 2.37 respectively.

In terms of hydrogen ions, account for the difference in the pH values

(*i*) of 0.10 M hydrochloric acid and 0.10 M sulfuric acid,

(*ii*) of 0.10 M hydrochloric acid and 0.10 M methanoic acid. (12)

Use the pH value given for 0.10 M methanoic acid to calculate

(*iii*) the value of the acid dissociation constant (K_a) for methanoic acid,

(*iv*) the pH of a 0.05 M solution of methanoic acid. (9)

(*c*) Draw a clear, labelled diagram of the pH curve you would expect to obtain when 50 cm^3 of a 0.10 M **NaOH** solution are added gradually to 25 cm^3 of a 0.10 M methanoic acid solution.

Name a suitable indicator for a titration between these two solutions.

Explain your selection with reference to your pH curve. (17)

Question 6

(*b*) What is the Arrhenius definition of a base?

Why is **NH_3** considered to be a base according to Brønsted-Lowry theory? (6)

Calculate the pH value of

(*i*) a 0.50 M solution of hydrochloric acid,

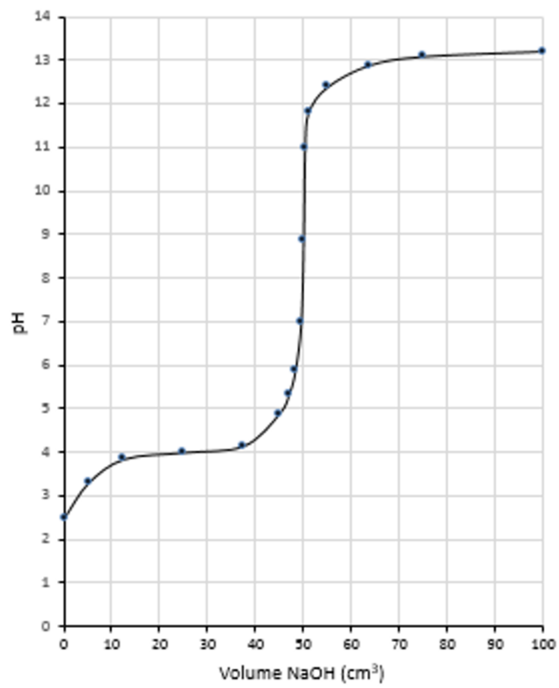
(*ii*) a 0.50 M solution of ethanoic acid. The acid dissociation constant (***K***$_a$) for ethanoic acid is 1.8×10^{-5}. (9)

The pH curve shown was obtained when 100 cm^3 of 0.50 M sodium hydroxide solution were gradually added to 50 cm^3 of one of the two acid solutions mentioned above.

Which of the two acid solutions was neutralised?

State two pieces of evidence from the pH curve in support of your answer.

What is the essential property of an indicator used to detect the end point in a titration between this acid and **NaOH**? (10)



Question 7

9. (*a*) Define pH.
Write an expression for $\boldsymbol{K}_w$, the ionic product (the dissociation constant) of water. (8)

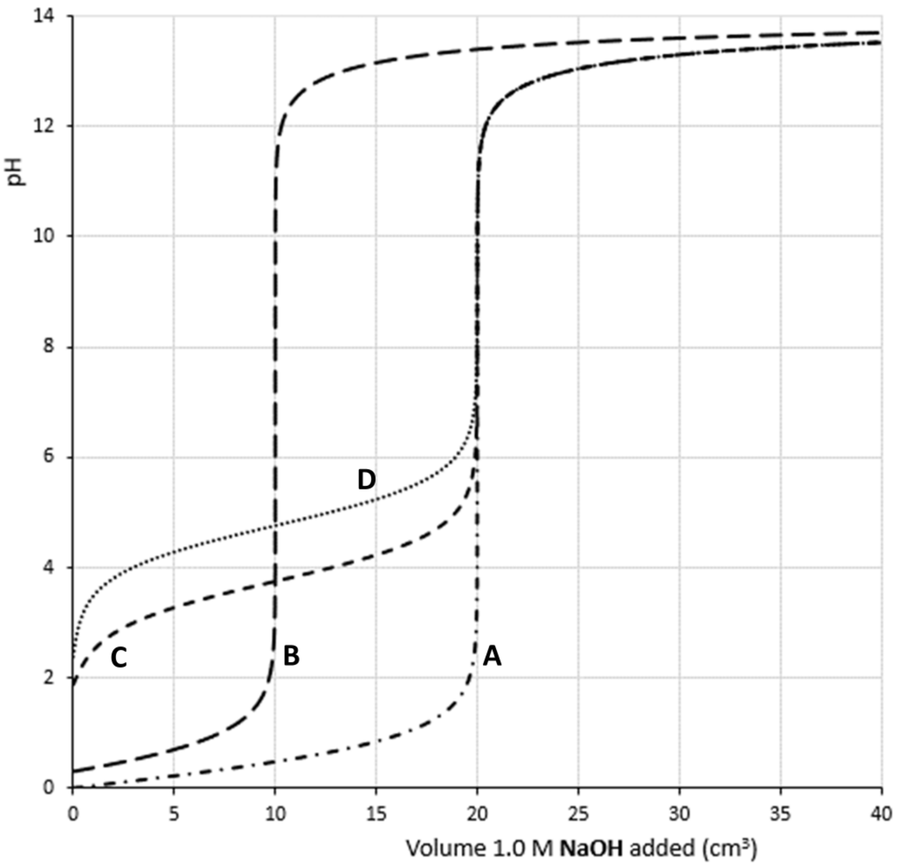
(*b*) The pH scale is most effective from 0 to 14.
State another limitation of the pH scale. (6)

(*c*) The pH of a certain mouthwash solution containing a weak acid is 5.65.
Calculate

(*i*) the molar concentration of a monobasic strong acid (**HX**) with the same pH,

(*ii*) the molar concentration of $\mathbf{OH^-}$ in the mouthwash. (12)

(*d*) In the graph below four curves are shown for the changes in pH values of four different 20 cm^3 solutions **A**, **B**, **C** and **D**, each containing one monobasic acidic substance, as 40 cm^3 of 1.0 M **NaOH** solution were slowly added.



(*i*) What volume of the **NaOH** solution neutralised solution **A**?

(*ii*) Deduce the molar concentrations of solutions **A**, **B**, **C** and **D**.

(*iii*) Which of the weak acids **C** and **D** has the greater acid dissociation constant ($\boldsymbol{K}_a$) value? Justify your answer.

(*iv*) **D** is $\mathbf{CH_3COOH}$. What is the conjugate base of **D**? (24)

Question 8

7. (*a*) Sulfuric acid is a strong dibasic acid; chloric(I) acid (**HOCl**) is a weak monobasic acid.

(*i*) Define an acid according to the Brønsted-Lowry theory.

(*ii*) What is a strong acid?

(*iii*) Identify the conjugate base of **HOCl**.

(*iv*) Identify the conjugate acid of $\mathbf{HSO_4^-}$. (12)

(*b*) Calculate, correct to one decimal place, the pH of

(*i*) a 0.04 M solution of sulfuric acid,

(*ii*) a 0.04 M solution of chloric(I) acid ($K_a = 3.0 \times 10^{-8}$). (12)

Question 9

7. (*a*) Define an acid according to the theory of (*i*) Arrhenius, (*ii*) Brønsted and Lowry.

(*iii*) State one limitation of Arrhenius' acid-base theory.

(*iv*) What is a conjugate acid-base pair in Bronsted-Lowry theory? (17)

(*b*) Distinguish between a strong acid and a weak acid. (6)

(*c*) Sulfuric acid is a strong dibasic acid. Write a balanced equation to show that the conjugate base of $\mathbf{H_2SO_4}$ acts as a Brønsted-Lowry acid in water. (6)

(*d*) (*i*) Write a balanced equation to show the dissociation into ions in water of a weak monobasic acid represented by **HA**.

If **HA** is 1.5% dissociated in a 0.10 M solution,

(*ii*) find the concentrations of $\mathbf{H_3O^+}$ ion and $\mathbf{A^-}$ ion in moles per litre in the solution,

(*iii*) calculate the pH of the 0.10 M **HA** solution,

(*iv*) hence or otherwise, calculate the value of the acid dissociation constant K_a for **HA**, assuming ionisation does not alter the concentration of the non-ionised form, i.e. take [**HA**] = 0.10 M after dissociation. (21)

Question 10

7. (*a*) According to Brønsted-Lowry theory, what is an acid?
How does this theory distinguish between a strong and a weak acid? (8)

(*b*) What is (*i*) the conjugate acid, (*ii*) the conjugate base, of $\mathbf{HSO_4^-}$? (6)

When carbon dioxide is pumped into still mineral water it becomes sparkling water. The information shown in the table was given on the label of a bottle of sparkling mineral water.

Typical Analysis	
$\mathbf{Ca^{2+}}$	164 mg/l
$\mathbf{Mg^{2+}}$	50 mg/l
$\mathbf{HCO_3^-}$	243 mg/l
$\mathbf{NO_3^-}$	3 mg/l
$\mathbf{SO_4^{2-}}$	402 mg/l
$\mathbf{Cl^-}$	49 mg/l
pH	5.6
pH at source	7.2

(*c*) Describe how you could test a sample of this mineral water for the presence of the sulfate ion. (6)

(*d*) The pH of the sparkling water is 5.6 while the pH of the source water was 7.2. Calculate

(*i*) the $\mathbf{H_3O^+}$ ion concentration of the *sparkling* water,

(*ii*) the $\mathbf{OH^-}$ ion concentration of the *source* water. (12)

An acid-base indicator, which is a weak acid, may be represented by **HA**. The indicator is red in its undissociated form and yellow in the form $\mathbf{A^-}$.

(*e*) Calculate the pH of an aqueous 1.5×10^{-3} M solution of the indicator which has an acid dissociation constant (K_a) of 4.0×10^{-4}. (6)

(*f*) (*i*) Write an equilibrium equation for the dissociation of the indicator in water.

(*ii*) What is observed when a few drops of the indicator are added to a solution of a strong base? Explain your answer. (12)

Question 11

7. Nitrous acid ($\mathbf{HNO_2}$) is a weak acid that is readily oxidised to the strong acid, nitric acid ($\mathbf{HNO_3}$).

(*a*) Distinguish between a *strong acid* and a *weak acid* according to the Brønsted-Lowry theory. (8)

(*b*) Nitrous acid dissociates in water as follows:

$$\mathbf{HNO_2 \;+\; H_2O \;\leftrightharpoons\; H_3O^+ \;+\; NO_2^-}$$

Identify the two substances acting as bases in this equilibrium. (6)

(*c*) Define pH.
The pH of a 0.2 M solution of nitrous acid is 2.0 at a temperature of 25 °C.
Calculate the concentration of $\mathbf{H_3O^+}$ ion in this solution in moles per litre.
Explain clearly how this $\mathbf{H_3O^+}$ ion concentration confirms that nitrous acid is a *weak* acid.
What concentration of nitric acid would have the same $\mathbf{H_3O^+}$ ion concentration?
Calculate the $\mathbf{OH^-}$ ion concentration in both of these acidic solutions. (18)

(*e*) Explain how high nitrate levels can result in a reduction in the dissolved oxygen content of lakes and rivers. (6)

Question 12

(*b*) Define *an acid* in terms of the Brønsted-Lowry theory.

What is a *conjugate pair*? (7)

A certain water soluble acid-base indicator represented by **HIn** is a weak acid which dissociates as follows in water.

$$\underset{\text{green}}{\mathbf{HIn}} \leftrightharpoons \mathbf{H^+} + \underset{\text{purple}}{\mathbf{In^-}}$$

State and explain the colour observed when a few drops of a solution of the indicator are added to a 0.5 M **NaOH** solution. (6)

Calculate the pH of (*i*) the 0.5 M **NaOH** solution, (*ii*) a 0.1 M solution of the indicator, given that its $\boldsymbol{K}_a$ value is 2.0×10^{-5}. (12)

Question 13

7. (*a*) Define (*i*) *acid*, (*ii*) *conjugate pair*, according to the Brønsted-Lowry theory. (8)

Identify the two conjugate pairs in the following dissociation of nitrous acid ($\mathbf{HNO_2}$):

$$\mathbf{HNO_2} + \mathbf{H_2O} \rightleftharpoons \mathbf{NO_2^-} + \mathbf{H_3O^+} \quad (6)$$

Distinguish between a strong acid and a weak acid. (6)

(*b*) Calculate the pH of 0.1 M nitrous acid ($\mathbf{HNO_2}$); the value of the acid dissociation constant ($\boldsymbol{K}_\mathbf{a}$) for nitrous acid is 5.0×10^{-4}.

What is the pH of a nitric acid ($\mathbf{HNO_3}$) solution of the same concentration? (15)

Question 14

8. (*a*) Define (i) *acid*, (ii) *base*, according to the Brønsted-Lowry theory. (8)

(*b*) Identify **one** species acting as an acid, and also identify its conjugate base, in the following system.

$$\mathbf{H_2F^+} + \mathbf{Cl^-} \rightleftharpoons \mathbf{HCl} + \mathbf{HF} \quad (6)$$

(*c*) Calculate the pH of a 0.002 M solution of methanoic acid (**HCOOH**). The value of K_a for methanoic acid is 1.8×10^{-4}. (12)

Question 15

8. (*a*) Define (*i*) *an acid*, (*ii*) *a base* according to the Brønsted-Lowry theory. (8)

Identify the *acid*, and *conjugate acid* in the following system. (6)

$$\mathbf{H_2S} \quad + \quad \mathbf{O^{2-}} \quad \rightleftharpoons \quad \mathbf{OH^-} \quad + \quad \mathbf{SH^-}$$

(*b*) Define pH. (6)

A bottle of vinegar is labelled 6% (w/v) acetic acid (ethanoic acid). The dissociation constant, K_a, for ethanoic acid is 1.8×10^{-5}. Calculate the approximate pH of the vinegar solution. (12)

Question 16

7. Sulfuric acid is a strong dibasic acid. The formula **HA** represents a weak monobasic acid.

(*a*) How do strong acids differ from weak acids in their behaviour in water according to (*i*) the Arrhenius theory, (*ii*) the Brønsted-Lowry theory? (12)

(*b*) What is the conjugate base of (*i*) sulfuric acid, (*ii*) the weak acid **HA**? Which of these conjugate bases is the stronger? Explain. (12)

(*c*) Explain, by giving a balanced equation for its dissociation in water, that the conjugate base of sulfuric acid is itself an acid. (6)

(*d*) Define pH. (6)

Calculate the pH of a 0.2 M solution of a weak acid, **HA**, the value of whose acid dissociation constant, $\boldsymbol{K_a}$, is 6.3×10^{-5} mol l^{-1}.

What is the concentration of a sulfuric acid solution that has the same pH? (14)

Question 17

(*b*) Define a base according to (*i*) the Arrhenius theory, (*ii*) the Brønsted-Lowry theory. (7)

Give (*i*) the conjugate acid, (*ii*) the conjugate base, of $\mathbf{HPO_4^{2-}}$. (6)

Ammonium hydroxide ($\mathbf{NH_4OH}$) is produced by dissolving gaseous ammonia in water.

Calculate the pH of an ammonium hydroxide solution that contains 7.0 g $\mathbf{NH_4OH}$ per litre. The value of the base dissociation constant ($\boldsymbol{K_b}$) for ammonium hydroxide is 1.8×10^{-5}. (12)

Question 18

8. (*a*) (*i*) Write an expression for the self-ionisation of water. (5)

(*ii*) Define $\boldsymbol{K}_{\mathbf{w}}$, the ionic product of water.

The value of $\boldsymbol{K}_{\mathbf{w}}$ at 25 ºC is 1.0×10^{-14}. Show that the pH of pure water is 7.0 at 25 ºC. (12)

(*iii*) Calculate the pH of a 0.5 M solution of a strong monobasic (monoprotic) acid.

Calculate the pH of a 0.5 M solution of a weak monobasic acid with a $\boldsymbol{K}_{\mathbf{a}}$ value of 1.8×10^{-5}. (12)

Question 19

10. Answer any **two** of the parts (*a*), (*b*) and (*c*). (2 × 25)

(*a*) (*i*) Distinguish between a strong acid and a weak acid. (6)

(*ii*) Calculate the molarity of a sulfuric acid solution that has a pH of 2.0.

The acid dissociation constant ($\boldsymbol{K}_a$) for a weak monobasic acid is 1.8×10^{-4}. Calculate the molarity of a solution of this acid that also has a pH of 2.0. (12)

(*iii*) Define $\boldsymbol{K}_w$ (the ionic product of water).

The value for $\boldsymbol{K}_w$ at 59 °C is 9.0×10^{-14}. What is the hydrogen ion concentration in pure water at this temperature? (7)

Question 20

7. (*a*) Use equations to show that, when dissolved in water,

(*i*) **HCl** acts as a Brønsted-Lowry acid,

(*ii*) $\mathbf{NH_3}$ acts as a Brønsted-Lowry base.

Explain why

(*iii*) **HCl** has a weak conjugate base,

(*iv*) $\mathbf{NH_3}$ has a strong conjugate acid. (12)

Temperature (°C)	K_w
0	0.1×10^{-14}
10	0.3×10^{-14}
20	0.7×10^{-14}
30	1.4×10^{-14}
40	2.9×10^{-14}
50	5.3×10^{-14}
60	9.3×10^{-14}

(*b*) Pure water self-ionises as follows:

$$2H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

The ionic product of water (K_w) measures the extent of this self-ionisation. A table of K_w values for pure water at different temperatures is given.

(*i*) Write the self-ionisation constant (K_w) expression for water.

(*ii*) Are the data in the table consistent with the self-ionisation of water being an exothermic or an endothermic process? Justify your answer. (12)

(*iii*) Plot a graph of K_w *versus* temperature (°C).

(*iv*) Use your graph to predict the value of K_w at body temperature, 37 °C.

(*v*) Use this value to calculate the $\mathbf{H_3O^+}$ ion concentration of pure water at 37 °C.

(*vi*) The pH of pure water is close to 6.77 at one of the temperatures in the table. Find, by calculation, this temperature. (26)