PΣΛK
PHYSICS & CHEMISTRY

**SCIENCE Years 7 – 10**
**Hieu Le**
www.peaktuition.com.au
Mobile: 0404 754 848
Office: 0497 952 888
**Level 1, 223 Canley Vale Rd, Canley Heights**

**PHYSICS & CHEMISTRY Years 11 – 12**
**Duyen Nguyen**
www.peakhsc.com.au
Mobile: 0432 637 032
Office: 0452 558 316
**Level 1, 262 Canley Vale Rd, Canley Heights**

Module 6: Acid/Base Reactions

Topic 4.4: Titrations

## Solutions

## Foundation

1. In dot-point form and using the sub-headings provided below, describe the sequence of steps required to carry out an acid-base titration where an unknown HCl solution is in the conical flask and a standardised NaOH solution is in the burette.

   (a) Preparing and filling the burette

   - Rinse the burette three times with distilled water and then the titrant (NaOH).
   - Fill the burette with the titrant using a funnel to the zero mark or below it.
   - Clamp the burette vertically.

   (b) Preparing and filling the conical flask

   - Rinse the conical flask three times with distilled water. Leave the conical flask wet.
   - Rinse a pipette three times with distilled water and then with the solution it is to deliver (HCl).
   - Using a pipette filler, fill the pipette with the solution that it is to deliver.
   - Transfer the aliquot in the pipette into the conical flask by resting the tip of the pipette against the inner glass and allowing it to drain under gravity.
   - Add three drops of an appropriate indicator into the conical flask and then place the conical flask under the burette.

   (c) Performing the acid-base titration

   - Open the tap on the burette and slowly run the titrant into the conical flask with continuous swirling.
   - Continue to add the titrant until the indicator undergoes a permanent colour change.
   - Record the volume of titrant added from the burette and then repeat the titration several times until three concordant titres are obtained.

2. In an experiment, 25.00 mL of acetic acid was titrated with a standardised 0.1200 $\text{mol L}^{-1}$ potassium hydroxide solution. 27.50 mL of potassium hydroxide was required to reach the end point.

   (a) Write a chemical equation for the reaction between acetic acid and potassium hydroxide.

   $CH_3COOH_{(aq)} + KOH_{(aq)} \rightarrow KCH_3COO_{(aq)} + H_2O_{(l)}$

   (b) Calculate the concentration of the unknown acetic acid solution.

$$\begin{aligned} n(\text{KOH}) &= 0.1200 \text{ mol L}^{-1} \times 0.02750 \text{ L} \\ &= 3.300 \times 10^{-3} \text{ mol} \\ n(\text{CH}_3\text{COOH}) &= 3.300 \times 10^{-3} \text{ mol} \\ [\text{CH}_3\text{COOH}] &= \frac{3.300 \times 10^{-3} \text{ mol}}{0.02500 \text{ L}} \\ &= 0.1320 \text{ mol L}^{-1} \end{aligned}$$

3. Write a chemical equation for the following reactions:

   (a) An unknown diprotic acid ($H_2X$) completely reacting with sodium hydroxide solution.

   $H_2X_{(aq)} + 2\,NaOH_{(aq)} \rightarrow Na_2X_{(aq)} + 2\,H_2O_{(l)}$

   (b) An unknown triprotic acid ($H_3X$) completely reacting with sodium hydroxide solution.

   $H_3X_{(aq)} + 3\,NaOH_{(aq)} \rightarrow Na_3X_{(aq)} + 3\,H_2O_{(l)}$

## Development

1. A household cleaning agent contains a weak base with the formula NaX. 1.00 g of this compound was dissolved in water to give 100.0 mL of solution. A 20.0 mL sample of the solution was titrated with 0.100 $\text{mol L}^{-1}$ hydrochloric acid, and required 24.4 mL of the acid for neutralisation.

   What is the molar mass of the weak base?

   (a) 82.0 $\text{g mol}^{-1}$

   (b) 84.0 $\text{g mol}^{-1}$

   (c) 122 $\text{g mol}^{-1}$

   (d) 410 $\text{g mol}^{-1}$

2. To determine the concentration of a hydrochloric acid solution, the acid was titrated with a sodium carbonate standard solution. In this titration, the acid was in the conical flask while the base was in the burette.

   If the student rinsed the pipette used to deliver the hydrochloric acid solution with only water before using it, what effect will this have the calculated concentration of hydrochloric acid?

   (a) The calculated concentration will be unaffected

   (b) The calculated concentration will be too high

   (c) The calculated concentration will be too low

   (d) More information is required to determine the effect on the calculated concentration

3. In a titration between an unknown strong base and a standardised strong acid, the following procedure was used:

   - A burette was rinsed with water and then filled with the standard acid solution.
   - A pipette was rinsed with some water and then with the unknown base solution.
   - A conical flask was rinsed with some unknown base solution.
   - The rinsed pipette was used to transfer 25.00 mL of the unknown base solution into the conical flask.
   - An appropriate indicator was added to the base sample and it was titrated to the end point with the acid.

   Which of the following statements is correct?

   (a) The calculated base concentration will be correct

   (b) The calculated base concentration will be too low

   (c) The calculated base concentration will be too high

   (d) No definite conclusion can be reached about the base concentration

4. A bottle contained an unknown triprotic acid with a molar mass of 192.12 $\text{g mol}^{-1}$. 25.00 mL of this acid was titrated with 0.550 $\text{molL}^{-1}$ NaOH and the average titre was found to be 14.75 mL.

What is the concentration of the unknown acid?

(a) 0.520 $\text{g L}^{-1}$

(b) $\boxed{20.8\ \text{g L}^{-1}}$

(c) 62.5 $\text{g L}^{-1}$

(d) 187 $\text{g L}^{-1}$

5. A student performed a titration to determine the concentration of a phosphoric acid solution. 10.0 mL of the acid solution was diluted to 200.0 mL in a volumetric flask. A 20.0 mL aliquot of this diluted acid was then titrated against a standardised 0.930 $\mathrm{mol\,L^{-1}}$ NaOH solution to the phenolphthalein end point. The results are shown in the table below.

| **Titration** | **Volume of NaOH used (mL)** |
|---|---|
| 1 | 22.4 |
| 2 | 21.4 |
| 3 | 21.2 |
| 4 | 21.3 |

(a) Briefly outline a suitable method that could have been used to standardise the NaOH solution. **2**

1. Prepare a primary standard solution of oxalic acid dihydrate in a volumetric flask.
2. Titrate an aliquot of the NaOH solution with the oxalic acid dihydrate primary standard. Repeat this titration until three concordant titres are obtained.
3. Calculate the concentration of the NaOH solution.

2 marks – Briefly outlines a method which can be used to standardise NaOH

(b) Calculate the concentration of the phosphoric acid solution. **3**

$$\mathrm{H_3PO_{4(aq)} + 3\,NaOH_{(aq)} \rightarrow Na_3PO_{4(aq)} + 3\,H_2O_{(l)}}$$

$$\begin{aligned} V(\mathrm{NaOH}) &= \frac{(0.0214 + 0.0212 + 0.0213)\ \mathrm{L}}{3} \\ &= 0.0213\ \mathrm{L} \\ n(\mathrm{NaOH}) &= 0.930\ \mathrm{mol\,L^{-1}} \times 0.0213\ \mathrm{L} \\ &= 0.0198\ \mathrm{mol} \\ n(\mathrm{H_3PO_4}) &= \frac{0.0198\ \mathrm{mol}}{3} \\ &= 6.60 \times 10^{-3}\ \mathrm{mol} \end{aligned}$$

These moles came from a 20.0 mL aliquot of the $\mathrm{H_3PO_4}$ solution.

$$\begin{aligned} [\mathrm{H_3PO_4}] &= \frac{6.60 \times 10^{-3}\ \mathrm{mol}}{0.0200\ \mathrm{L}} \\ &= 0.330\ \mathrm{mol\,L^{-1}} \end{aligned}$$

However, the acid solution was diluted 20-fold before the titration.

$$\begin{aligned} \therefore [\mathrm{H_3PO_4}]_{\text{undiluted}} &= 20 \times 0.330\ \mathrm{mol\,L^{-1}} \\ &= 6.60\ \mathrm{mol\,L^{-1}} \end{aligned}$$

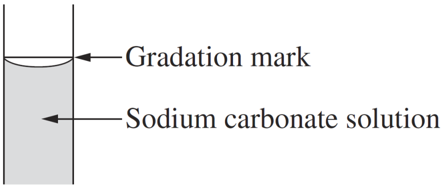
1 mark – Calculates the correct moles of NaOH

1 mark – Calculates the correct concentration of the diluted phosphoric acid

1 mark – Calculates the correct concentration of the undiluted phosphoric acid

6. A student determined the concentration of a HCl solution by using the following method.

   1. A conical flask was rinsed with distilled water.
   2. A 25.0 mL pipette was rinsed with distilled water.
   3. The student filled the pipette with a standard $Na_2CO_3$ solution to the level shown in the diagram.



   4. The standard $Na_2CO_3$ solution in the pipette was transferred to the conical flask. The student ensured that all of the $Na_2CO_3$ solution was transferred to the conical flask by blowing through the pipette.
   5. A few drops of methyl orange indicator was added to the conical flask.
   6. A burette was rinsed with distilled water and then with the unknown HCl solution. The student then carried out a titration to determine the concentration of the HCl solution.

During this experiment, the student made several mistakes.

(a) Identify the mistake the student made in Step 4 and propose a change to this step that would improve the validity of his method. **2**

The mistake in Step 4 was blowing on the pipette to ensure all the solution is transferred. To improve the validity of the method, the student should NOT blow on the pipette to remove the remaining solution since the pipette has been calibrated to take into account the remaining solution. Instead, he should just allow the solution to drain into the conical flask by holding the pipette tip against the inner surface of the conical flask.

1 mark – Identifies the mistake made by the student in Step 4

1 mark – Proposes a change to Step 4 which would improve the validity of the method

(b) Explain how the mistakes made in both Steps 2 and 3 will affect the calculation of the concentration of the HCl solution. **3**

In Step 2, the pipette should have been rinsed additionally with the $Na_2CO_3$ solution. Since the pipette was left wet with water, the $Na_2CO_3$ solution would be more dilute than expected. In Step 3, the student filled the pipette until the base of the meniscus was below the calibration line when it is supposed to be directly on it. These errors would have resulted in fewer moles of $Na_2CO_3$ being transferred to the conical flask (due to reduced concentration and reduced volume). Consequently, a smaller volume of HCl from the burette will be required to reach the end point and thus the concentration of HCl will be overestimated.

1 mark – Identifies the mistakes made in Step 3 and 4

1 mark – Relates these two mistakes to a reduction of moles of $Na_2CO_3$

1 mark – Explains that the concentration of the HCl solution will be overestimated

7. The concentration of hydrochloric acid in a solution was determined by an acid-base titration using a standard solution of sodium carbonate.

A 25.00 mL sample of 0.1050 mol L$^{-1}$ sodium carbonate solution was added to a conical flask and three drops of methyl orange indicator added. The mixture was titrated with the hydrochloric acid and the following readings were recorded.

| **Titration** | **Titre (mL)** |
|---|---|
| 1 | 22.00 |
| 2 | 21.65 |
| 3 | 21.70 |
| 4 | 21.60 |

(a) Calculate the concentration of the hydrochloric acid. **3**

$$2\,\mathrm{HCl_{(aq)}} + \mathrm{Na_2CO_{3(aq)}} \rightarrow 2\,\mathrm{NaCl_{(aq)}} + \mathrm{CO_{2(g)}} + \mathrm{H_2O_{(l)}}$$

$$\begin{aligned} n(\mathrm{Na_2CO_3}) &= 0.1050\ \mathrm{mol\,L^{-1}} \times 0.02500\ \mathrm{L} \\ &= 2.625 \times 10^{-3}\ \mathrm{mol} \\ n(\mathrm{HCl}) &= 2 \times 2.625 \times 10^{-3}\ \mathrm{mol} \\ &= 5.250 \times 10^{-3}\ \mathrm{mol} \end{aligned}$$

These moles came from an average titre of 21.65 mL (discarding the 1st titre).

$$\begin{aligned} [\mathrm{HCl}] &= \frac{5.250 \times 10^{-3}\ \mathrm{mol}}{0.02165\ \mathrm{L}} \\ &= 0.2425\ \mathrm{mol\,L^{-1}} \end{aligned}$$

1 mark – Write a correct chemical equation for the titration reaction

1 mark – Calculates the correct moles of $\mathrm{Na_2CO_3}$

1 mark – Calculates the correct concentration of the HCl

(b) Explain the effect on the calculated concentration of hydrochloric acid if phenolphthalein is used as the indicator instead of methyl orange. **2**

Phenolphthalein changes colour over the pH range 8.3 – 10.0, so a lower volume of HCl from the burette will be required to reach its end point compared to methyl orange (3.1 – 4.4). This will lead to an overestimation of the HCl concentration.

1 mark – Identifies that using phenolphthalein will cause an overestimation of the HCl concentration

1 mark – Explain why using phenolphthalein will cause an overestimation of the HCl concentration

8. A manufacturer makes lemon cordial by mixing flavouring, sugar syrup and citric acid. The concentration of the citric acid is determined by titration with NaOH.

The sodium hydroxide solution is prepared by dissolving 4.000 g of NaOH pellets in water to give 1.000 L of solution. This solution is standardised by titrating 25.00 mL with a 0.1011 $\text{mol L}^{-1}$ standardised solution of HCl. The average titration volume is found to be 24.10 mL.

To analyse the lemon cordial 50.00 mL of the cordial is diluted to 500.0 mL. Then 25.00 mL of the diluted solution is titrated with the NaOH solution to the phenolphthalein endpoint.

The following data were collected during one of the analysis runs of the lemon cordial.

| | |
|---|---|
| Titration #1 volume | 26.55 mL |
| Titration #2 volume | 27.25 mL |
| Titration #3 volume | 27.30 mL |
| Titration #4 volume | 27.20 mL |

(a) Calculate the expected concentration of the NaOH solution using the mass data. **1**

$$\begin{aligned} n(\text{NaOH}) &= \frac{4.000\text{ g}}{(22.99 + 16.00 + 1.008)\text{ g mol}^{-1}} \\ &= 0.1000\text{ mol} \\ [\text{NaOH}] &= \frac{0.1000\text{ mol}}{1.000\text{ L}} \\ &= 0.1000\text{ mol L}^{-1} \end{aligned}$$

1 mark – Calculates the correct concentration of NaOH using the mass data

(b) Calculate the concentration of the standardised NaOH solution by using the appropriate titration data and account for any discrepancies between this value and the value calculated in part (a). **2**

$$\text{HCl}_{(aq)} + \text{NaOH}_{(aq)} \rightarrow \text{NaCl}_{(aq)} + \text{H}_2\text{O}_{(l)}$$

$$\begin{aligned} n(\text{HCl}) &= 0.1011\text{ mol L}^{-1} \times 0.02410\text{ L} \\ &= 2.437 \times 10^{-3}\text{ mol} \\ n(\text{NaOH}) &= 2.437 \times 10^{-3}\text{ mol} \\ [\text{NaOH}] &= \frac{2.437 \times 10^{-3}\text{ mol}}{0.02500\text{ L}} \\ &= 0.09746\text{ mol L}^{-1} \\ &< [\text{NaOH}]\text{ from part (a)} \end{aligned}$$

The standardised concentration is lower because NaOH is hygroscopic i.e. it will absorb water as it is being weighed out, leading to a more dilute solution than anticipated.

1 mark – Calculates the correct concentration of NaOH using the titration data

1 mark – Accounts for any discrepancies between this value and the value in part (a)

(c) Calculate the concentration of citric acid in the lemon cordial. **3**

Let citric acid be denoted by $H_3X$ since it is a triprotic acid.

$$\mathrm{H_3X_{(aq)} + 3\,NaOH_{(aq)} \rightarrow Na_3X_{(aq)} + 3\,H_2O_{(l)}}$$

$$\begin{aligned} V(\mathrm{NaOH}) &= \frac{(0.02725 + 0.02730 + 0.02720)\ \mathrm{L}}{3} \\ &= 0.02725\ \mathrm{L} \\ n(\mathrm{NaOH}) &= 0.09746\ \mathrm{mol\,L^{-1}} \times 0.02725\ \mathrm{L} \\ &= 2.656 \times 10^{-3}\ \mathrm{mol} \\ n(\mathrm{H_3X}) &= \frac{2.656 \times 10^{-3}\ \mathrm{mol}}{3} \\ &= 8.853 \times 10^{-4}\ \mathrm{mol} \end{aligned}$$

These moles came from a 25.00 mL aliquot of the lemon cordial.

$$\begin{aligned} [\mathrm{H_3X}] &= \frac{8.853 \times 10^{-4}\ \mathrm{mol}}{0.02500\ \mathrm{L}} \\ &= 0.03541\ \mathrm{mol\,L^{-1}} \end{aligned}$$

However, the lemon cordial was diluted 10-fold before the titration.

$$\begin{aligned} \therefore [\mathrm{H_3X}]_{\text{undiluted}} &= 10 \times 0.03541\ \mathrm{mol\,L^{-1}} \\ &= 0.3541\ \mathrm{mol\,L^{-1}} \end{aligned}$$

1 mark – Calculates the correct moles of NaOH

1 mark – Calculates the correct concentration of citric acid in the diluted cordial

1 mark – Calculates the correct concentration of citric acid in the undiluted cordial

9. A 0.2845 g sample of impure $Na_2CO_3$ was titrated against a previously standardised solution of HCl. 24.65 mL of HCl solution was required to reach the end point. When a 0.2204 g sample of pure $Na_2CO_3$ was titrated against the same HCl solution, 20.06 mL of HCl was required to reach the end point. **4**

Calculate the percentage by mass of sodium carbonate in the impure sample.

$$2\,HCl_{(aq)} + Na_2CO_{3(aq)} \rightarrow 2\,NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

Consider the titration between the pure $Na_2CO_3$ and HCl.

$$\begin{aligned} n(Na_2CO_3) &= \frac{0.2204 \text{ g}}{(2(22.99) + 12.01 + 3(16.00)) \text{ g mol}^{-1}} \\ &= 2.079 \times 10^{-3} \text{ mol} \\ n(HCl) &= 2 \times 2.079 \times 10^{-3} \text{ mol} \\ &= 4.159 \times 10^{-3} \text{ mol} \end{aligned}$$

These moles came from a 20.06 mL solution.

$$\begin{aligned} [HCl] &= \frac{4.159 \times 10^{-3} \text{ mol}}{0.02006 \text{ L}} \\ &= 0.2073 \text{ mol L}^{-1} \end{aligned}$$

Now, consider the titration between the impure $Na_2CO_3$ and HCl.

$$\begin{aligned} n(HCl) &= 0.2073 \text{ mol L}^{-1} \times 0.02465 \text{ L} \\ &= 5.110 \times 10^{-3} \text{ mol} \\ n(Na_2CO_3) &= \frac{5.110 \times 10^{-3} \text{ mol}}{2} \\ &= 2.555 \times 10^{-3} \text{ mol} \\ m(Na_2CO_3) &= 2.555 \times 10^{-3} \text{ g} \times (2(22.99) + 12.01 + 3(16.00)) \text{ g mol}^{-1} \\ &= 0.2708 \text{ g} \\ \%Na_2CO_3 &= \frac{0.2708 \text{ g}}{0.2845 \text{ g}} \times 100\% \\ &= 95.20\% \end{aligned}$$

1 mark – Calculates the correct moles of pure $Na_2CO_3$

1 mark – Calculates the correct concentration of the HCl solution

1 mark – Calculates the correct moles of HCl titrated against the impure sample

1 mark – Calculates the correct percentage by mass of $Na_2CO_3$ in the impure sample

10. The nitrogen content of bread was determined using the following procedure:

- A sample of bread weighing 2.80 g was analysed.
- The nitrogen in the sample was converted into ammonia.
- The ammonia was collected in 50.0 mL of 0.125 $\text{mol L}^{-1}$ hydrochloric acid. All of the ammonia was neutralised, leaving an excess of hydrochloric acid.
- The excess hydrochloric acid was titrated with 23.30 mL of 0.116 $\text{mol L}^{-1}$ NaOH solution.

(a) Write chemical equations for the TWO reactions involving hydrochloric acid. **2**

$$HCl_{(aq)} + NH_{3(aq)} \rightarrow NH_4Cl_{(aq)}$$
$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

2 marks – Writes TWO correct chemical equations (1 mark each)

(b) Calculate the number of moles of excess hydrochloric acid. **1**

$$\begin{aligned} n(\text{NaOH}) &= 0.116\ \text{mol L}^{-1} \times 0.02330\ \text{L} \\ &= 2.70 \times 10^{-3}\ \text{mol} \\ n(\text{HCl})\ \text{leftover} &= 2.70 \times 10^{-3}\ \text{mol} \end{aligned}$$

1 mark – Calculates the correct moles of leftover HCl

(c) Calculate the number of moles of ammonia. **2**

$$\begin{aligned} n(\text{HCl})\ \text{added} &= 0.125\ \text{mol L}^{-1} \times 0.0500\ \text{L} \\ &= 6.25 \times 10^{-3}\ \text{mol} \\ n(\text{HCl})\ \text{reacted} &= n(\text{HCl})\ \text{added} - n(\text{HCl})\ \text{leftover} \\ &= 6.25 \times 10^{-3}\ \text{mol} - 2.70 \times 10^{-3}\ \text{mol} \\ &= 3.55 \times 10^{-3}\ \text{mol} \\ n(\text{NH}_3) &= 3.55 \times 10^{-3}\ \text{mol} \end{aligned}$$

1 mark – Calculates the correct moles of HCl added initially

1 mark – Calculates the correct moles of $NH_3$

(d) Calculate the percentage by mass of nitrogen in the bread. **2**

Each mole of ammonia contains exactly 1 mole of nitrogen.

$$\begin{aligned} n(\mathrm{N}) &= 3.55 \times 10^{-3}\ \mathrm{mol} \\ m(\mathrm{N}) &= 3.55 \times 10^{-3}\ \mathrm{mol} \times 14.01\ \mathrm{g\,mol^{-1}} \\ &= 0.0497\ \mathrm{g} \\ \%\mathrm{N} &= \frac{0.0497\ \mathrm{g}}{2.80\ \mathrm{g}} \times 100\% \\ &= 1.77\% \end{aligned}$$

1 mark – Calculates the correct mass of N

1 mark – Calculates the correct percentage by mass of N

11. A solution of hydrochloric acid was standardised by titration against a sodium carbonate solution using the following procedure.

- All glassware was rinsed correctly to remove possible contaminants.
- Hydrochloric acid was placed in the burette.
- 25.0 mL of sodium carbonate solution was pipetted into the conical flask.

The titration was performed and the hydrochloric acid was found to be 0.200 $\mathrm{mol\,L^{-1}}$.

(a) Identify the substance used to rinse the conical flask and justify your answer. **2**

Distilled water is used because it can remove impurities in the flask and it also does not affect the number of moles of $Na_2CO_3$ that will be placed into the flask.

1 mark – Identifies that the flask should be rinsed with water

1 mark – Recognises that water will not affect the moles of $Na_2CO_3$

(b) Seashells contain a mixture of carbonate compounds. The standardised hydrochloric acid was used to determine the percentage by mass of carbonate in a seashell using the following procedure. **4**

- A 0.145 g sample of the seashell was placed in a conical flask.
- 50.0 mL of the standardised hydrochloric acid was added to the conical flask.
- At the completion of the reaction, the mixture in the conical flask was titrated with $0.250\ \text{mol L}^{-1}$ sodium hydroxide.

The volume of sodium hydroxide used in the titration was 29.5 mL.

Calculate the percentage by mass of carbonate in the sample of the seashell.

$$\text{HCl}_{(aq)} + \text{NaOH}_{(aq)} \rightarrow \text{NaCl}_{(aq)} + \text{H}_2\text{O}_{(l)}$$

$$\begin{aligned}
n(\text{NaOH}) &= 0.250\ \text{mol L}^{-1} \times 0.0295\ \text{L} \\
&= 7.38 \times 10^{-3}\ \text{mol} \\
n(\text{HCl})\ \text{leftover} &= 7.38 \times 10^{-3}\ \text{mol} \\
n(\text{HCl})\ \text{added} &= 0.200\ \text{mol L}^{-1} \times 0.0500\ \text{L} \\
&= 0.0100\ \text{mol} \\
n(\text{HCl})\ \text{reacted} &= n(\text{HCl})\ \text{added} - n(\text{HCl})\ \text{leftover} \\
&= 0.0100\ \text{mol} - 7.38 \times 10^{-3}\ \text{mol} \\
&= 2.63 \times 10^{-3}\ \text{mol}
\end{aligned}$$

The initial reaction is:

$$2\,\text{HCl}_{(aq)} + \text{CO}_3^{2-}{}_{(s)} \rightarrow 2\,\text{Cl}^-{}_{(aq)} + \text{CO}_{2(g)} + \text{H}_2\text{O}_{(l)}$$

$$\begin{aligned}
n(\text{CO}_3^{2-}) &= \frac{1}{2} \times 2.63 \times 10^{-3}\ \text{mol} \\
&= 1.31 \times 10^{-3}\ \text{mol} \\
m(\text{CO}_3^{2-}) &= 1.31 \times 10^{-3}\ \text{mol} \times (12.01 + 3 \times 16.00)\ \text{g mol}^{-1} \\
&= 0.0788\ \text{g} \\
\%\text{CO}_3^{2-} &= \frac{0.0788\ \text{g}}{0.145\ \text{g}} \times 100\% \\
&= 54.3\%
\end{aligned}$$

1 mark – Calculates the correct moles of NaOH

1 mark – Calculates the correct moles of HCl reacted

1 mark – Calculates the correct moles of $\text{CO}_3^{2-}$

1 mark – Calculates the correct percentage by mass of $\text{CO}_3^{2-}$

12. A student was asked to deduce the identity of an unknown precipitate that has a chemical formula of $\mathrm{Fe(OH)_n}$ where n is a positive integer. The method used by the student used is outlined below. **4**

- 1.010 g of pure solid $\mathrm{Fe(OH)_n}$ was mixed with 20.00 mL of 2.000 $\mathrm{mol\,L^{-1}}$ HCl. All of the $\mathrm{Fe(OH)_n}$ was neutralised, leaving behind an excess of HCl.
- The excess HCl was transferred into a 200.0 mL volumetric flask and then distilled water was added until the base of the meniscus was on the engraved mark.
- A 25.00 mL aliquot of the solution from the volumetric flask was then titrated with a previously standardised 0.1000 $\mathrm{mol\,L^{-1}}$ KOH solution. The average titre was 14.56 mL.

Calculate the value of n.

$$\mathrm{HCl_{(aq)} + KOH_{(aq)} \rightarrow KCl_{(aq)} + H_2O_{(l)}}$$

$$\begin{aligned} n(\mathrm{KOH}) &= 0.1000\ \mathrm{mol\,L^{-1}} \times 0.01456\ \mathrm{L} \\ &= 1.456 \times 10^{-3}\ \mathrm{mol} \\ n(\mathrm{HCl}) &= 1.456 \times 10^{-3}\ \mathrm{mol} \end{aligned}$$

These moles came from a 25.00 mL aliquot.

$$\begin{aligned} [\mathrm{HCl}] &= \frac{1.456 \times 10^{-3}\ \mathrm{mol}}{0.02500\ \mathrm{L}} \\ &= 0.05824\ \mathrm{mol\,L^{-1}} \end{aligned}$$

The initial reaction is:

$$\mathrm{nHCl_{(aq)} + Fe(OH)_{n(s)} \rightarrow FeCl_{n(aq)} + nH_2O_{(l)}}$$

The 25.00 mL aliquot that was titrated originally came from a 200.0 mL solution.

$$\begin{aligned} \therefore n(\mathrm{HCl})\ \text{leftover} &= 0.05824\ \mathrm{mol\,L^{-1}} \times 0.2000\ \mathrm{L} \\ &= 0.01165\ \mathrm{mol} \\ n(\mathrm{HCl})\ \text{added} &= 2.000\ \mathrm{mol\,L^{-1}} \times 0.02000\ \mathrm{L} \\ &= 0.04000\ \mathrm{mol} \\ n(\mathrm{HCl})\ \text{reacted} &= n(\mathrm{HCl})\ \text{added} - n(\mathrm{HCl})\ \text{leftover} \\ &= 0.04000\ \mathrm{mol} - 0.01165\ \mathrm{mol} \\ &= 0.02835\ \mathrm{mol} \\ n(\mathrm{Fe(OH)_n}) &= \frac{0.02835\ \mathrm{mol}}{\mathrm{n}} \\ n(\mathrm{Fe(OH)_n}) &= \frac{1.010\ \mathrm{g}}{(55.85 + \mathrm{n}(16.00 + 1.008))\ \mathrm{g\,mol^{-1}}} \\ \Longrightarrow \quad \frac{0.02835\ \mathrm{mol}}{\mathrm{n}} &= \frac{1.010\ \mathrm{g}}{(55.85 + \mathrm{n}(16.00 + 1.008))\ \mathrm{g\,mol^{-1}}} \\ \Longrightarrow \quad n &= 3 \end{aligned}$$

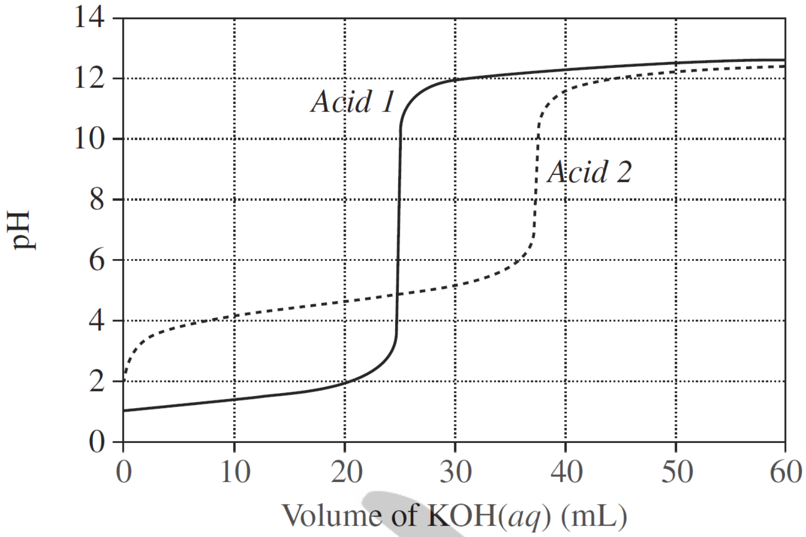
1 mark – Calculates the correct moles of KOH

1 mark – Calculates the correct moles of HCl leftover

1 mark – Calculates the correct moles of HCl reacted

1 mark – Calculates the correct value of n

13. The graph shows changes in pH for the titrations of equal volumes of solutions of two monoprotic acids, *Acid 1* and *Acid 2*.



(a) Explain the differences between *Acid 1* and *Acid 2* in terms of their relative strengths and concentrations. **3**

*Acid 1* is a strong acid, as evident by the equivalence point being at pH 7, the low initial pH and the steep pH rise at the equivalence point. *Acid 2* is a weak acid, as evident by the equivalence point being at pH $>$ 7 and the higher initial pH. *Acid 1* is more dilute relative to *Acid 2* as a lower volume of base was required to neutralise it.

2 marks – Explains that *Acid 1* is strong while *Acid 2* is weak with reference to the titration curve

1 mark – Explains that *Acid 1* is more dilute than *Acid 2* with reference to the volume of KOH used

(b) Name the salt produced by the reaction of an acid of the same type as *Acid 2* with KOH. **1**

Potassium acetate

1 mark – Names an appropriate potassium salt

(c) Calculate the concentration of hydrogen ions in the titrating mixture when 20 mL of KOH has been added to *Acid 1*. **1**

From the graph, when 20 mL of KOH is added, pH = 2.

$$\begin{aligned} \therefore [\mathrm{H}^+] &= 10^{-2}\ \mathrm{mol\,L^{-1}} \\ &= 0.01\ \mathrm{mol\,L^{-1}} \end{aligned}$$

1 mark – Calculates the correct $[\mathrm{H}^+]$

(d) Why would phenolphthalein be a suitable indicator for both titrations? **1**

Both titrations involve a reaction with a strong base (KOH), so the equivalence point is at pH $\geq 7$. Phenolphthalein changes colour in the pH range $8.3 - 10.0$ and this range coincides with the steep part of both titration curves. The phenolphthalein end-point will therefore accurately estimate the equivalence point of both titrations.

1 mark – Relates the colour change range of phenolphthalein to the equivalence point or the steep portion of the titration curve

## Extension (Optional)

1. A student is attempting to work out the identity of an unknown acid with a chemical formula of $H_nXO_3$ where n is a positive integer and X is a non-metallic element which can either be phosphorous (P), sulfur (S), carbon (C) or nitrogen (N). To determine the identity of the acid, the student performs a series of titrations. **7**

   The student first prepares a secondary standard solution of NaOH by titrating 25.00 mL of NaOH against a 0.09756 $\mathrm{mol\,L^{-1}}$ solution of potassium hydrogen phthalate. The average titre was 27.25 mL. Note that potassium hydrogen phthalate is a monoprotic acid.

   The student then dissolves 2.569 g of $H_nXO_3$ in sufficient water to make a 250.0 mL solution. This solution of $H_nXO_3$ was titrated with 25.00 mL of the standardised NaOH solution and the average titre was found to be 16.31 mL.

   Using relevant calculations, deduce the identity of the unknown acid.

Let potassium hydrogen phthalate be denoted as HA since it is a monoprotic acid.

$$\begin{aligned} \mathrm{HA_{(aq)}} + \mathrm{NaOH_{(aq)}} &\rightarrow \mathrm{NaA_{(aq)}} + \mathrm{H_2O_{(l)}} \\ n(\mathrm{HA}) &= 0.09756\ \mathrm{mol\,L^{-1}} \times 0.02725\ \mathrm{L} \\ &= 2.658 \times 10^{-3}\ \mathrm{mol} \\ n(\mathrm{NaOH}) &= 2.658 \times 10^{-3}\ \mathrm{mol} \end{aligned}$$

These moles came from a 25.00 mL aliquot.

$$\begin{aligned} [\mathrm{NaOH}] &= \frac{2.658 \times 10^{-3}\ \mathrm{mol}}{0.02500\ \mathrm{L}} \\ &= 0.1063\ \mathrm{mol\,L^{-1}} \end{aligned}$$

$$\begin{aligned} \mathrm{H_nXO_{3(aq)}} + \mathrm{nNaOH_{(aq)}} &\rightarrow \mathrm{Na_nXO_{3(aq)}} + \mathrm{nH_2O_{(l)}} \\ n(\mathrm{NaOH}) &= 0.1063\ \mathrm{mol\,L^{-1}} \times 0.02500\ \mathrm{L} \\ &= 2.658 \times 10^{-3}\ \mathrm{mol} \end{aligned}$$

Units will now be excluded from the calculations for clarity.

$$n(\mathrm{H_nXO_3}) = \frac{2.658 \times 10^{-3}}{\mathrm{n}}$$

These moles of $H_nXO_3$ came from a titre of 16.31 mL.

$$\begin{aligned} [\mathrm{H_nXO_3}] &= \frac{2.658 \times 10^{-3}}{0.01631\mathrm{n}} \\ &= \frac{0.1630}{\mathrm{n}} \end{aligned}$$

However, the acid solution was made by dissolving 2.569 g of the acid in 250.0 mL of water.

$$\begin{aligned} n(\mathrm{H_nXO_3}) &= \frac{2.569}{1.008\mathrm{n} + M(\mathrm{X}) + 3(16.00)} \\ [\mathrm{H_nXO_3}] &= \frac{2.569}{0.2500(1.008\mathrm{n} + M(\mathrm{X}) + 3(16.00))} \end{aligned}$$

$$\implies \frac{0.1630}{\mathrm{n}} = \frac{2.569}{0.2500(1.008\mathrm{n} + M(X) + 3(16.00))}$$

Rearranging and making $M(\mathrm{X})$ the subject,

$$M(\mathrm{X}) = 62.04\mathrm{n} - 48.00$$

The maximum possible value for $M(\mathrm{X})$ is the molar mass of sulfur which is $32.07\ \mathrm{g\,mol^{-1}}$. If $\mathrm{n} \geq 2$, then $M(\mathrm{X}) > 32.07\ \mathrm{g\,mol^{-1}}$ which cannot be true.

$$\begin{aligned} \therefore n &= 1 \\ M(\mathrm{X}) &= 62.04 \times 1 - 48.00 \\ &= 14.04\ \mathrm{g\,mol^{-1}} \\ &\doteqdot M(\mathrm{N}) \end{aligned}$$

Therefore, the unknown acid is $\mathrm{HNO_3}$.

7 marks – Correctly deduces that the unknown acid is $\mathrm{HNO_3}$ with relevant calculations

2. Amino acids are named so because they contain a basic amine group ($-NH_2$) and an acidic carboxyl group ($-COOH$).

Glycine is the simplest amino acid with chemical formula $NH_2CH_2COOH$. At different pH levels, glycine exists in different forms and these are summarised in the points below.

- In a solution with a low pH, both the amine and the carboxyl group of glycine are protonated so glycine exists as a cation ($^+NH_3CH_2COOH$).
- In a solution with an intermediate pH, the amine group is protonated and the carboxyl group is deprotonated so glycine exists in a zwitterionic form ($^+NH_3CH_2COO^-$).
- In a solution with a high pH, both the amine and the carboxyl group are deprotonated so glycine exists as an anion ($NH_2CH_2COO^-$).

During a university lab, a student titrated a solution of glycine with a NaOH solution, starting from a low pH.

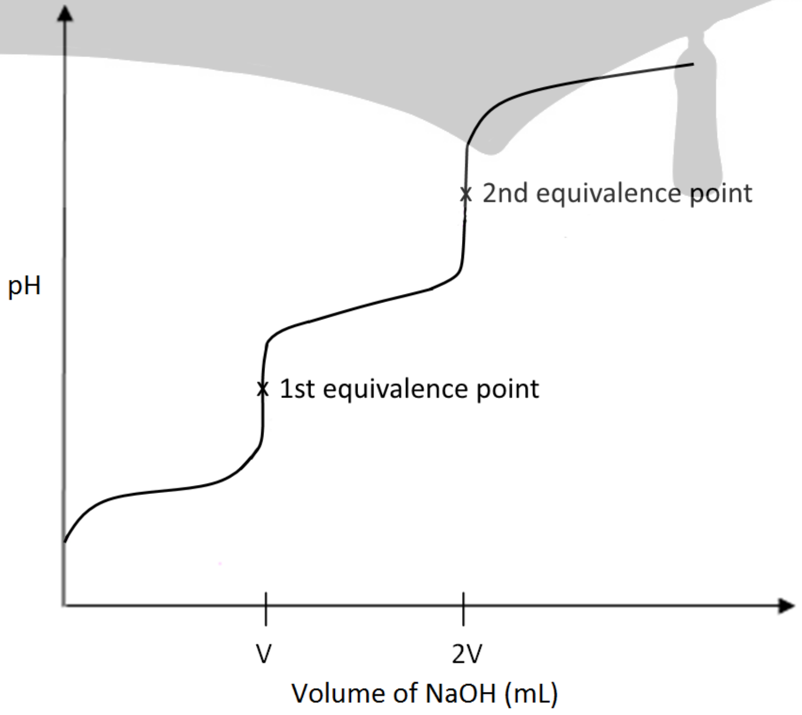
(a) Write TWO net ionic equations to represent the stepwise neutralisation reactions which occur during the titration. **2**

$$^+NH_3CH_2COOH_{(aq)} + OH^-_{(aq)} \rightarrow {}^+NH_3CH_2COO^-_{(aq)} + H_2O_{(l)}$$
$$^+NH_3CH_2COO^-_{(aq)} + OH^-_{(aq)} \rightarrow NH_2CH_2COO^-_{(aq)} + H_2O_{(l)}$$

2 marks – Writes TWO correct net ionic equations (1 mark each)

(b) Sketch a titration curve for a solution of glycine being titrated against a NaOH solution, starting from a low pH. Label the equivalence points on the titration curve. **2**

When a polyprotic acid in a conical flask is titrated with a strong base, the neutralisation steps occur stepwise so there will be several equivalence points. Glycine is a diprotic acid so there will be two neutralisation steps and two equivalence points.



| | |
|---|---|
| 2 marks – | Sketches a titration curve showing TWO equivalence points with the first occurring at exactly half the volume of the second |