Please keep your worksheet duration under the maximum of 180 minutes

# CHEMISTRY
**Stage 6**

## Module 6: Acid/Base Reactions
### Quantitative Analysis

**Teacher:** Samantha Wong
**Exam Equivalent Time:** 259.5 minutes *(based on allocation of 1.5 minutes per mark)*

## Questions

### 1. CHEMISTRY, M6 2015 HSC 2 MC

Which type of glassware is used in a titration to deliver an accurate volume of a solution to a known volume of another solution?

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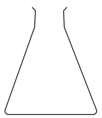
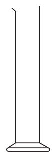
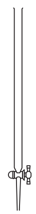


(B)


(C)


(D)


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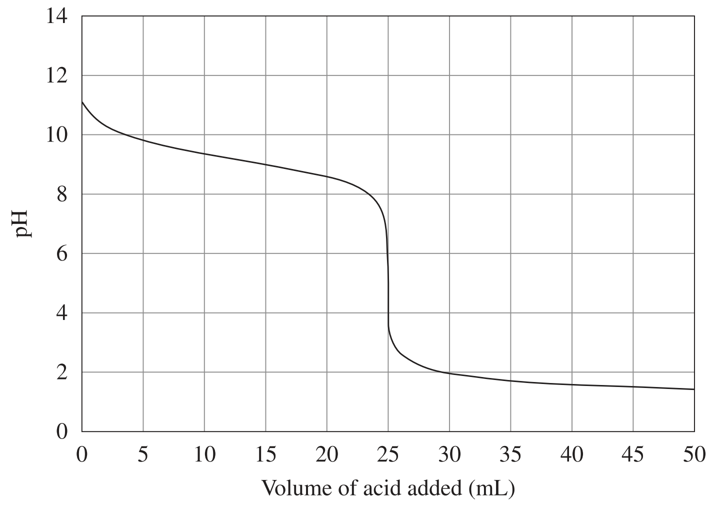
### 2. CHEMISTRY, M6 2019 HSC 5 MC

The diagram represents the titration curve for a reaction between a particular acid and a particular base.




Which indicator would be best for this titration?

| | *Indicator* | *Colour change range* (pH) |
|---|---|---|
| A. | Martius yellow | 2.0 – 3.2 |
| B. | Magdala red | 3.0 – 4.0 |
| C. | Isopicramic acid | 4.0 – 5.6 |
| D. | Cresol red | 7.2 – 8.8 |

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### 3. CHEMISTRY, M2 2010 HSC 7 MC

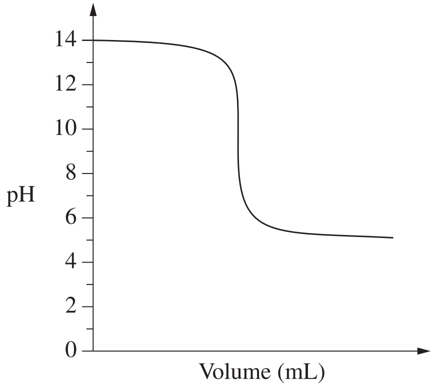
Equal volumes of four 0.1 mol $L^{-1}$ acids were titrated with the same sodium hydroxide solution.

Which one requires the greatest volume of base to change the colour of the indicator?

**A.** Citric acid
**B.** Acetic acid
**C.** Sulfuric acid
**D.** Hydrochloric acid

## 4. CHEMISTRY, M6 2015 HSC 14 MC

The graph shows the changes in pH during a titration.




Which pH range should an indicator have to be used in this titration?

**A.** $3.1 - 4.4$

**B.** $5.0 - 8.0$

**C.** $6.0 - 7.6$

**D.** $8.3 - 10.0$

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## 5. CHEMISTRY, M6 2017 HSC 1 MC

In an experiment, 30 mL of water is to be transferred into a conical flask.

Which piece of equipment would deliver the volume with the greatest accuracy?

**A.** Burette

**B.** Beaker

**C.** Test tube

**D.** Measuring cylinder

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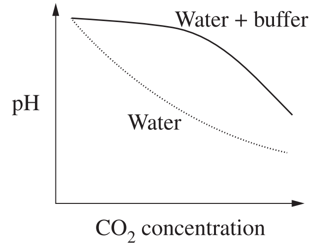
## 6. CHEMISTRY, M6 2018 HSC 17 MC

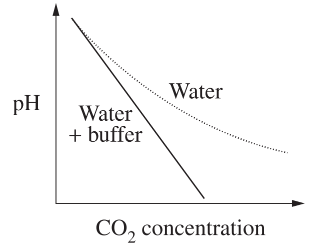
Increasing amounts of carbon dioxide were dissolved in two beakers, one containing water and one a mixture of water and a buffer. The pH in each beaker was measured and the results graphed.

Which graph best represents the results?

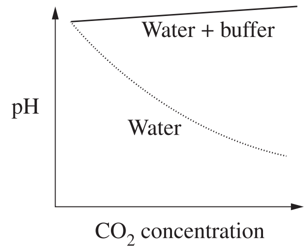
A.



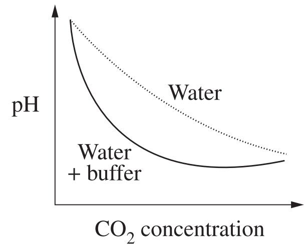

B.




C.




D.

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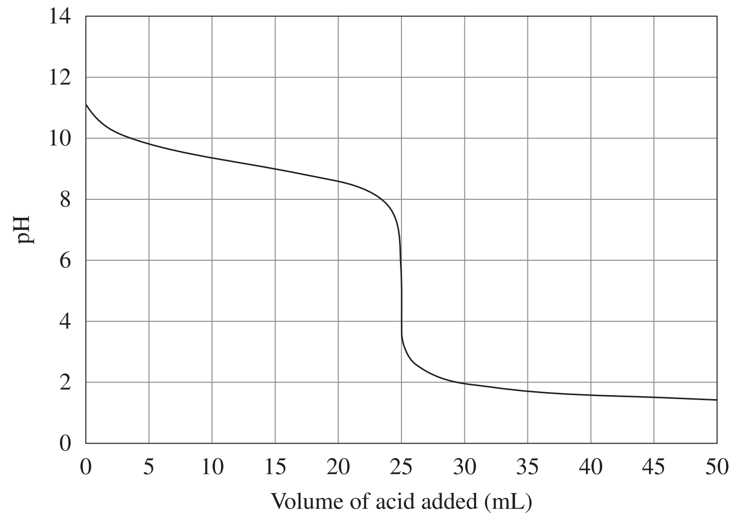
## 7. CHEMISTRY, M6 2019 HSC 6 MC

The diagram represents the titration curve for a reaction between a particular acid a particular base.




Which of the following equations best represents the reaction described by the titration curve?

**A.** $NH_3\,(aq) + HCl\,(aq)$ to $NH_4\,Cl\,(aq)$

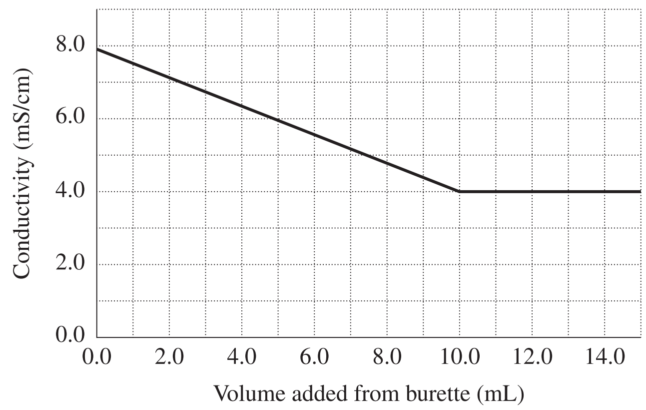
**B.** $NaOH\,(aq) + HCl\,(aq)$ to $NaCl\,(aq) + H_2O\,(l)$

**C.** $NH_3\,(aq) + CH_3\,COOH\,(aq)$ to $CH_3COONH_4\,(aq)$

**D.** $NaOH\,(aq) + CH_3\,COOH\,(aq)$ to $CH_3COONa(aq)^{+}H_2O\,(l)$

## 8. CHEMISTRY, M6 2022 HSC 15 MC

A 25.00 mL sample of 0.1131 $\text{mol L}^{-1}$ $HCl(aq)$ was titrated with an aqueous ammonia solution. The conductivity of the solution was measured throughout the titration and the results graphed.




What was the concentration of the ammonia solution?

**A.** 0.0452 $\text{mol L}^{-1}$

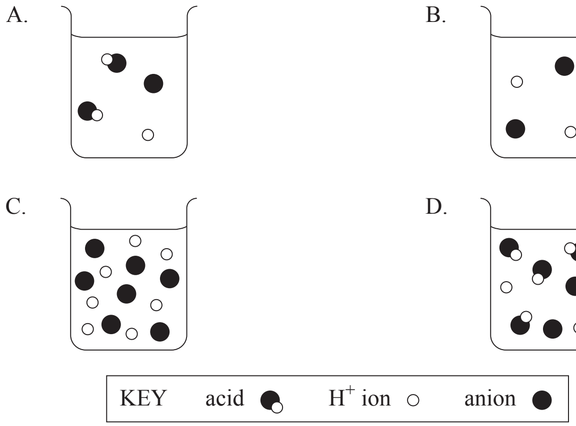
**B.** 0.189 $\text{mol L}^{-1}$

**C.** 0.283 $\text{mol L}^{-1}$

**D.** 0.690 $\text{mol L}^{-1}$

## 9. CHEMISTRY, M6 2023 HSC 5 MC

Which diagram represents the most concentrated weak acid?

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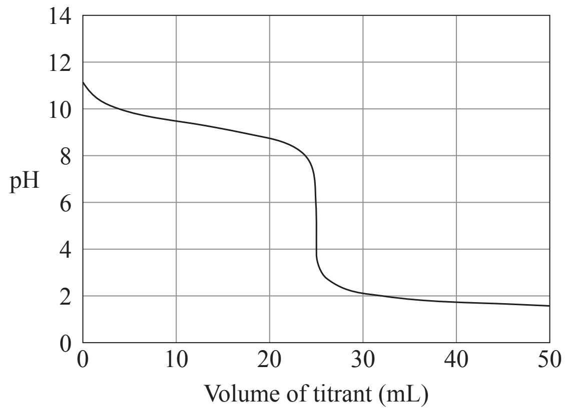
## 10. CHEMISTRY, M6 2023 HSC 9 MC

A titration was performed using two solutions of equal concentration, producing the following titration curve.




Which combination of solutions does the titration curve represent?

**A.** Addition of a weak base to a weak acid

**B.** Addition of a weak base to a strong acid

**C.** Addition of a strong acid to a weak base

**D.** Addition of a strong acid to a strong base

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## 11. CHEMISTRY, M6 2024 HSC 16 MC

Which of the following is the overall reaction that takes place when a strong acid is added to a buffer containing equal amounts of acetic acid and acetate ions?

**A.** $HCOO^- + H_3O^+ \rightarrow HCOOH + H_2O$

**B.** $CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$

**C.** $CH_3COO^- + H_3O^+ \rightarrow CH_3COOH + H_2O$

**D.** $CH_3COOH + H_3O^+ \rightarrow CH_3C(OH)_2{}^+ + H_2O$

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### 12. CHEMISTRY, M6 EQ-Bank 13 MC

The pKa of trichloroacetic acid is 0.70 and the pKa of acetic acid is 4.8.

Which of the following identifies the acid with the higher pH and explain why?

**A.** Acetic acid as it is less likely to lose a hydrogen ion

**B.** Acetic acid as it is more likely to lose a hydrogen ion

**C.** Trichloroacetic acid as it is less likely to lose a hydrogen ion

**D.** Trichloroacetic acid as it is more likely to lose a hydrogen ion

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### 13. CHEMISTRY, M6 2009 HSC 14 MC

Citric acid, the predominant acid in lemon juice, is a triprotic acid. A student titrated 25.0 mL samples of lemon juice with 0.550 mol L$^{-1}$ $\mathrm{NaOH}$. The mean titration volume was 29.50 mL. The molar mass of citric acid is 192.12 g mol$^{-1}$.

What was the concentration of citric acid in the lemon juice?

**A.** 1.04 g L$^{-1}$

**B.** 41.6 g L$^{-1}$

**C.** 125 g L$^{-1}$

**D.** 374 g L$^{-1}$

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### 14. CHEMISTRY, M6 2017 HSC 14 MC

One litre of an aqueous solution is formed from mixing equal volumes of 0.2 mol L$^{-1}$ hydrochloric acid $(\mathrm{HCl})$ and 0.2 mol L$^{-1}$ sodium chloride $(\mathrm{NaCl})$.

How effective as a buffer is the aqueous solution formed?

**A.** Ineffective, because $\mathrm{HCl}$ is a strong acid

**B.** Effective, because $\mathrm{Cl}^-$ is the conjugate base of $\mathrm{HCl}$

**C.** Ineffective, because $\mathrm{NaCl}$ forms a neutral salt solution

**D.** Effective, because the pH would change when a solution of $\mathrm{NaOH}$ is added

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### 15. CHEMISTRY, M6 2018 HSC 15 MC

A solution containing potassium dihydrogen phosphate and potassium hydrogen phosphate is a common laboratory buffer with a $\mathrm{pH}$ close to 7.

Which row of the table correctly identifies the chemistry of this buffer?

| | *Buffer equation* | *Equilibrium shift* | |
|---|---|---|---|
| | | *Acid is added to the solution* | *Alkali is added to the solution* |
| A. | $HPO_4^{2-} + H_2O \rightleftharpoons PO_4^{3-} + H_3O^+$ | Right | Left |
| B. | $HPO_4^{2-} + H_2O \rightleftharpoons PO_4^{3-} + H_3O^+$ | Left | Right |
| C. | $H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+$ | Right | Left |
| D. | $H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+$ | Left | Right |

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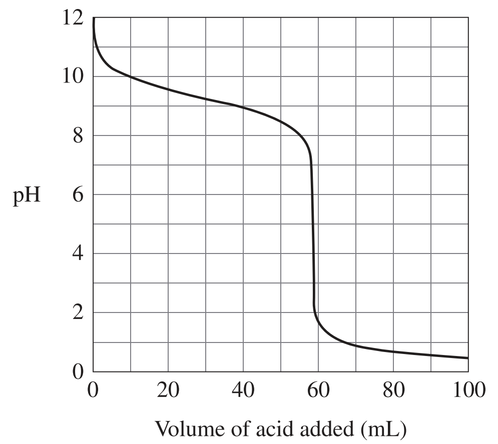
### 16. CHEMISTRY, M6 2020 HSC 8 MC

A weak base is titrated with 1.0 mol L$^{-1}$ aqueous $\mathrm{HCl}$. The $\mathrm{pH}$ curve is shown.




At which pH value would the solution be most effective as a buffer?

**A.** 5

**B.** 7

**C.** 8

**D.** 9

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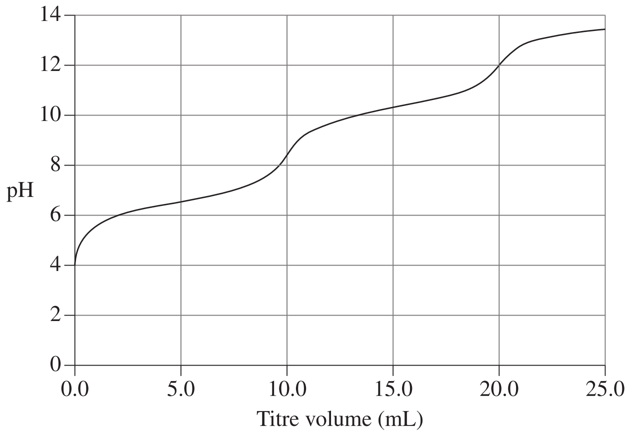
## 17. CHEMISTRY, M6 2021 HSC 16 MC

This titration curve is produced when an acid is titrated with a sodium hydroxide solution of the same concentration.




How many acidic protons does this acid possess?

**A.** 1

**B.** 2

**C.** 3

**D.** 4

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## 18. CHEMISTRY, M6 2021 HSC 5 MC

A student used the following method to titrate an acetic acid solution of unknown concentration with a standardised solution of dilute sodium hydroxide.

- Rinse burette with deionised water.
- Fill burette with sodium hydroxide solution.
- Rinse pipette and conical flask with acetic acid solution.
- Pipette 25.00 mL of acetic acid solution into conical flask.
- Add appropriate indicator to the conical flask.
- Titrate to endpoint and record volume of sodium hydroxide solution used.

Compared to the actual concentration of the acetic acid, the calculated concentration will be

**A.** lower.

**B.** higher.

**C.** the same.

**D.** different, but higher or lower cannot be predicted.

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## 19. CHEMISTRY, M6 2022 HSC 2 MC

When a solution of a primary standard is prepared for titration, which of the following is required?

**A.** A burette

**B.** A balance

**C.** An indicator

**D.** A condenser

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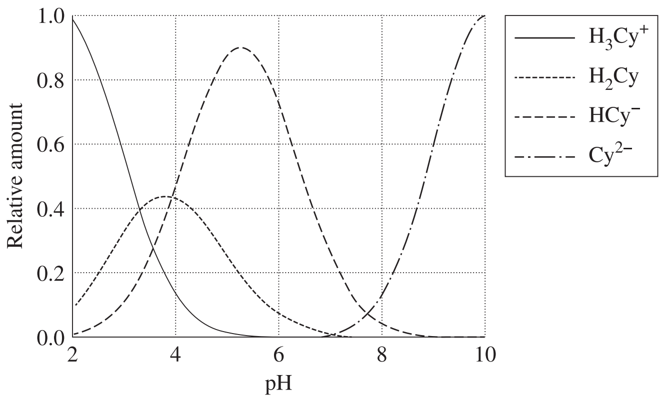
## 20. CHEMISTRY, M6 2022 HSC 20 MC

Cyanidin is a plant pigment that may be used as a pH indicator. It has four levels of protonation, each with a different colour, represented by these equilibria:

$$\underset{\text{Red}}{H_3Cy^+} \underset{+H^+}{\overset{-H^+}{\rightleftharpoons}} \underset{\text{Colourless}}{H_2Cy} \underset{+H^+}{\overset{-H^+}{\rightleftharpoons}} \underset{\text{Purple}}{HCy^-} \underset{+H^+}{\overset{-H^+}{\rightleftharpoons}} \underset{\text{Blue}}{Cy^{2-}}$$

The following graph shows the relative amount of each species present at different pH values.




What colour would the indicator be if added to a 0.75 mol L$^{-1}$ solution of hypoiodous acid, HIO $(pK_a = 10.64)$?

**A.** Red

**B.** Colourless

**C.** Purple

**D.** Blue

## 21. CHEMISTRY, M6 2023 HSC 14 MC

What volume of 0.540 mol L$^{-1}$ hydrochloric acid will react completely with 1.34 g of sodium carbonate?

**A.** 11.7 mL

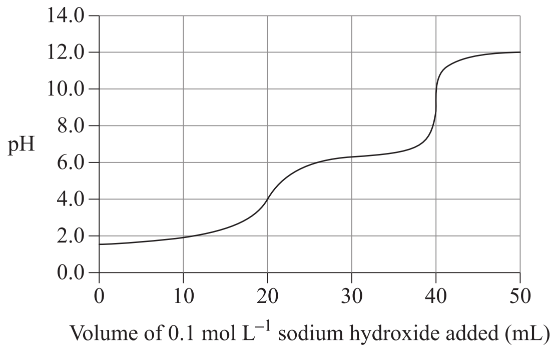
**B.** 23.4 mL

**C.** 29.9 mL

**D.** 46.8 mL

## 22. CHEMISTRY, M6 2024 HSC 17 MC

20 mL of a 0.1 mol L$^{-1}$ solution of an acid is titrated against a 0.1 mol L$^{-1}$ solution of sodium hydroxide. A graph of pH against the volume of sodium hydroxide for this experiment is shown.




Which of the following acids was used in the titration?

| | *Acid* | $pK_{a1}$ | $pK_{a2}$ |
|---|---|---|---|
| **A.** | 1 | 4.76 | – |
| **B.** | 2 | Strong | – |
| **C.** | 3 | 1.91 | 6.30 |
| **D.** | 4 | 4.11 | 9.61 |

## 23. CHEMISTRY, M6 2024 HSC 20 MC

The concentration of ascorbic acid $\left(MM = 176.124\ \text{g mol}^{-1}\right)$ in solution A was determined by titration.

- A 25.00 mL sample of solution A was titrated with potassium hydroxide solution.
- 50.00 mg of ascorbic acid was added to a second 25.00 mL sample of solution A, which was titrated in the same way.

Titration volumes for both titrations are given.

| *Solution* | *Titre* (mL) |
|---|---|
| 25.00 mL solution A | 17.50 |
| 25.00 mL solution A+ 50.00 mg of ascorbic acid | 33.10 |

What is the concentration of ascorbic acid in solution A?

**A.** $5.352 \times 10^{-3}$ mol $L^{-1}$

**B.** $6.004 \times 10^{-3}$ mol $L^{-1}$

**C.** $1.012 \times 10^{-2}$ mol $L^{-1}$

**D.** $1.274 \times 10^{-2}$ mol $L^{-1}$

## 24. CHEMISTRY, M6 EQ-Bank 14 MC

Equal volumes of four different acids are titrated with the same base at 25°.

Information about these acids is given in the table.

| *Acid* | *Concentration* (mol $L^{-1}$) | pH |
|---|---|---|
| HCl | 0.1 | 1.0 |
| $H_3PO_4$ | 0.1 | 1.6 |
| $CH_3COOH$ | 0.1 | 2.9 |
| HCN | 0.1 | 5.1 |

Which acid requires the greatest volume of base for complete reaction?

**A.** HCl

**B.** $H_3PO_4$

**C.** $CH_3COOH$

**D.** HCN

## 25. CHEMISTRY, M6 2017 HSC 13 MC

25.0 mL of a 0.100 mol $L^{-1}$ acid is to be titrated against a sodium hydroxide solution until final equivalence is reached.

Which of the following acids, if used in the titration, would require the greatest volume of sodium hydroxide?

**A.** Acetic

**B.** Citric

**C.** Hydrochloric

**D.** Sulfuric

## 26. CHEMISTRY, M6 2021 HSC 20 MC

The trimethylammonium ion, $[(CH_3)_3NH]^+$, is a weak acid. The acid dissociation equation is shown.

$$[(CH_3)_3NH]^+(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + (CH_3)_3N(aq) \quad K_a = 1.55 \times 10^{-10}$$

At 20°C, a saturated solution of trimethylammonium chloride, $[(CH_3)_3NH]Cl$, has a pH of 4.46.

What is the $K_{sp}$ of trimethylammonium chloride?

**A.** $1.26 \times 10^{-9}$

**B.** $7.76$

**C.** $60.2$

**D.** $5.01 \times 10^{10}$

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## 27. CHEMISTRY, M6 2021 HSC 6 MC

Which row of the table describes what happens when a solution of a weak acid is diluted? (Assume constant temperature.)

| | $K_a$ | *Extent of acid ionisation* |
|---|---|---|
| A. | Decreases | Increases |
| B. | Decreases | Decreases |
| C. | Remains the same | Increases |
| D. | Remains the same | Decreases |

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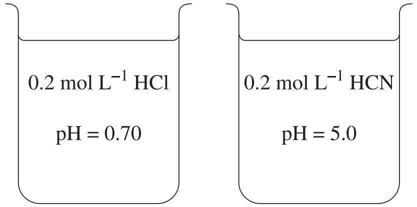
## 28. CHEMISTRY, M6 2022 HSC 25

The pH of two aqueous solutions was compared.




Explain why the $HCN(aq)$ solution has a higher pH than the $HCl(aq)$ solution. Include a relevant chemical equation for the $HCN(aq)$ solution. *(3 marks)*

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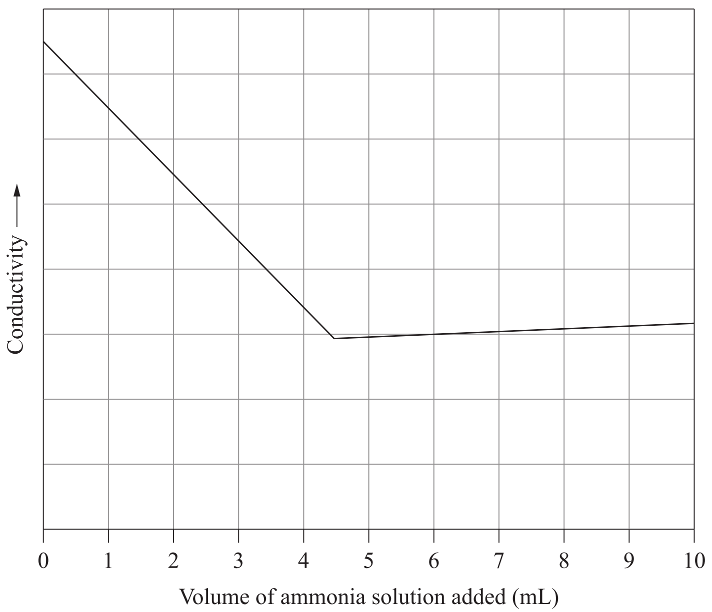
## 29. CHEMISTRY, M6 2024 HSC 34

An aqueous solution of ammonia is added to a solution containing hydrochloric acid. A plot of conductivity against volume of ammonia solution added is shown. The temperature of the solution is kept constant throughout and the conductivity of the solution is corrected for dilution.




The relative conductivities of some relevant ions are shown in the table.

| *Ion* | *Relative conductivity* |
|---|---|
| $H^+$ | 4.76 |
| $OH^-$ | 2.70 |
| $Cl^-$ | 1.04 |
| $NH_4^+$ | 1.00 |

Explain the shape of the graph. Include TWO balanced chemical equations in your answer. *(4 marks)*

## 30. CHEMISTRY, M6 EQ-Bank 27

What determines the pH of a buffer solution? *(2 marks)*

## 31. CHEMISTRY, M6 2012 HSC 30

A chemist analysed aspirin tablets for quality control. The initial step of the analysis was the standardisation of a $\mathrm{NaOH}$ solution. Three 25.00 mL samples of a 0.1034 mol L$^{-1}$ solution of standardised $\mathrm{HCl}$ were titrated with the $\mathrm{NaOH}$ solution. The average volume required for neutralisation was 25.75 mL.

a. Calculate the molarity of the $\mathrm{NaOH}$ solution. *(2 marks)*

Three flasks were prepared each containing a mixture of 25 mL of water and 10 mL of ethanol. An aspirin tablet was dissolved in each flask. The aspirin in each solution was titrated with the standardised $\mathrm{NaOH}$ solution according to the following equation:

$$\mathrm{C_9H_8O_4\,(aq) + NaOH\,(aq) \rightarrow C_9H_7O_4Na\,(aq) + H_2O\,(l)}$$

The following titration results were obtained.

| *Tablet* | *Volume* (mL) |
|---|---|
| 1 | 16.60 |
| 2 | 16.50 |
| 3 | 16.55 |

b. Calculate the average mass (mg) of aspirin per tablet. *(3 marks)*

## 32. CHEMISTRY, M6 2014 HSC 30

A batch of dry ice (solid $CO_2$) was contaminated during manufacture. To determine its purity, the following steps were carried out.

| | |
|---|---|
| Step 1: | A 0.616 gram sample of the contaminated dry ice was placed in a clean, dry flask. |
| Step 2: | 50.00 mL of 1.00 mol $L^{-1}$ sodium hydroxide was added to the flask. The sodium hydroxide was in excess. |
| Step 3: | The flask was sealed to prevent loss of carbon dioxide gas and the reaction allowed to reach completion, according to this equation: $$2NaOH(aq) + CO_2(s) \rightarrow Na_2CO_3(aq) + H_2O(l)$$ |
| Step 4: | The remaining sodium hydroxide was titrated against a 1.00 mol $L^{-1}$ solution of hydrochloric acid. The average volume of HCl used was 27.60 mL. |

a. Calculate the number of moles of $NaOH$ added in Step 2. ***(1 mark)***

b. Calculate the percentage purity by mass of this batch of dry ice. ***(4 marks)***

## 33. CHEMISTRY, M6 2016 HSC 29

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 mol $L^{-1}$.

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

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.

- 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 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. *(4 marks)*

## 34. CHEMISTRY, M6 2015 HSC 26

A sodium hydroxide solution was titrated against citric acid $(C_6H_8O_7)$ which is triprotic.

a. Draw the structural formula of citric acid (2-hydroxypropane-1,2,3-tricarboxylic acid). *(1 mark)*

b. How could a computer-based technology be used to identify the equivalence point of this titration? *(2 marks)*

c. The sodium hydroxide solution was titrated against 25.0 mL samples of 0.100 mol L $^{-1}$ citric acid. The average volume of sodium hydroxide used was 41.50 mL.

Calculate the concentration of the sodium hydroxide solution. *(4 marks)*

## 35. CHEMISTRY, M6 2017 HSC 24

A solution of sodium hydroxide was titrated against a standardised solution of acetic acid which had a concentration of 0.5020 mol $L^{-1}$.

a. The end point was reached when 19.30 mL of sodium hydroxide solution had been added to 25.00 mL of the acetic acid solution.

Calculate the concentration of the sodium hydroxide solution. *(3 marks)*

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b. Explain why the pH of the resulting salt solution was not 7. Include a relevant chemical equation in your answer. *(2 marks)*

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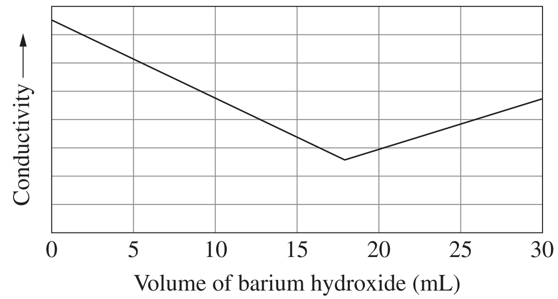
## 36. CHEMISTRY, M6 2019 HSC 24

A conductometric titration was undertaken to determine the concentration of a barium hydroxide solution. The solution was added to 250.0 mL of standardised $1.050 \times 10^{-3}$ mol $L^{-1}$ hydrochloric acid solution. The results of the titration are shown in the conductivity graph.




a. Explain the shape of the titration curve. *(3 marks)*

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b. The equivalence point was reached when a volume of 17.15 mL of barium hydroxide was added.

Calculate the concentration of barium hydroxide (in mol $L^{-1}$), and give a relevant chemical equation. *(4 marks)*

## 37. CHEMISTRY, M6 2019 HSC 27

The relationship between the acid dissociation constant, $K_a$, and the corresponding conjugate base dissociation constant, $K_b$, is given by:

$$K_a \times K_b = K_w$$

Assume that the temperature for part (a) and part (b) is 25°C.

a. The $K_a$ of hypochlorous acid $(\mathrm{HOCl})$ is $3.0 \times 10^{-8}$.

Show that the $K_b$ of the hypochlorite ion, $\mathrm{OCl}^-$, is $3.3 \times 10^{-7}$. *(1 mark)*

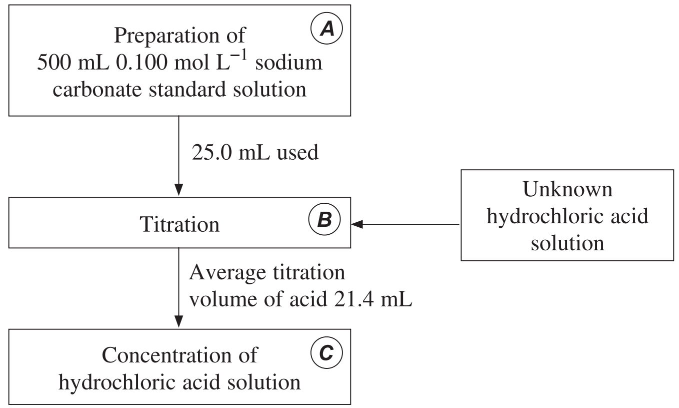
b. The conjugate base dissociation constant, $K_b$, is the equilibrium constant for the following equation:

$$\mathrm{OCl}^-(aq) + \mathrm{H_2O}(l) \rightleftharpoons \mathrm{HOCl}(aq) + \mathrm{OH}^-(aq)$$

Calculate the pH of a 0.20 mol L$^{-1}$ solution of sodium hypochlorite $(\mathrm{NaOCl})$. *(4 mark)*

## 38. CHEMISTRY, M6 EQ-Bank 28

The flowchart shown outlines the sequence of steps used to determine the concentration of an unknown hydrochloric acid solution.




Describe steps A, B and C including correct techniques, equipment and appropriate calculations. Determine the concentration of the hydrochloric acid. *(8 marks)*

## 39. CHEMISTRY, M7 2024 HSC 35

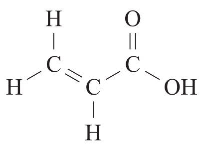
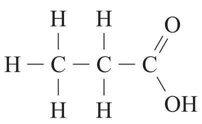
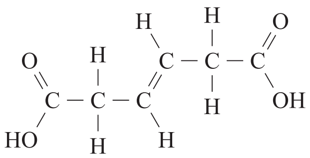
Unknown samples of three carboxylic acids, labelled X, Y and Z, are analysed to determine their identities.

- Both Y and Z react rapidly with bromine in the absence of UV light, but X does not. A 0.100 g sample of Y reacts with the same amount of bromine as a 0.200 g sample of Z.
- Separate 0.100 g samples of X, Y and Z are titrated with 0.0617 mol $L^{-1}$ sodium hydroxide solution. The titre volumes are shown.

| *Acid* | *X* | *Y* | *Z* |
|---|---|---|---|
| Volume of NaOH (mL) | 21.88 | 22.49 | 22.49 |

- Both Y and Z can undergo hydration reactions in the presence of a suitable catalyst. Two products are possible for the hydration of Y, but only one product is possible with Z.

Identify which structures 1, 2 and 3 in the table are acids X, Y and Z. Justify your answer with reference to the information provided. *(7 marks)*

| | *Structure 1* | *Structure 2* | *Structure 3* |
|---|---|---|---|
| |  |  |  |
| Molar mass $(g\ mol^{-1})$ | 72.062 g $mol^{-1}$ | 74.078 g $mol^{-1}$ | 144.124 g $mol^{-1}$ |
| Acid (X, Y or Z) | | | |

## 40. CHEMISTRY, M6 2022 HSC 34

Sodium hypochlorite $\mathrm{NaOCl}$ is the active ingredient in pool chlorine. It completely dissolves in water to produce the hypochlorite ion $\left(\mathrm{OCl}^-\right)$, which undergoes hydrolysis according to the following equilibrium.

$$\mathrm{OCl^-\,(aq) + H_2O\,(l) \rightleftharpoons HOCl\,(aq) + OH^-\,(aq)}$$

The equilibrium constant for this reaction at 25°C is $3.33 \times 10^{-7}$.

For pool chlorine to be effective the pH is maintained by a different buffer at 7.5 and the hypochlorous acid $(\mathrm{HOCl})$ concentration should be $1.3 \times 10^{-4}$ mol L$^{-1}$.

Calculate the volume of 2.0 mol L$^{-1}$ sodium hypochlorite solution that needs to be added to a $1.00 \times 10^4$ L pool to meet the required conditions. ***(4 marks)***

## 41. CHEMISTRY, M6 2019 HSC 22

A buffer was prepared with acetic acid and sodium acetate. A few drops of universal indicator were then added. When small amounts of either 0.1 mol L $^{-1}$ $\mathrm{HCl\,(aq)}$ or 0.1 mol L $^{-1}$ $\mathrm{NaOH\,(aq)}$ were added, no change in the colour of the solution was observed.

Explain these observations. Support your answer with at least ONE chemical equation. *(4 marks)*

## 42. CHEMISTRY, M6 2020 HSC 28

A chemist used the following method to determine the concentration of a dilute solution of propanoic acid $(pK_a = 4.88)$.

The chemist weighed out 1.000 g of solid $\mathrm{NaOH}$ on an electronic balance and then made up the solution in a 250.0 mL volumetric flask.

The chemist then performed titrations, using bromocresol green as the indicator. This indicator is yellow below pH 3.2 and green above pH 5.2.

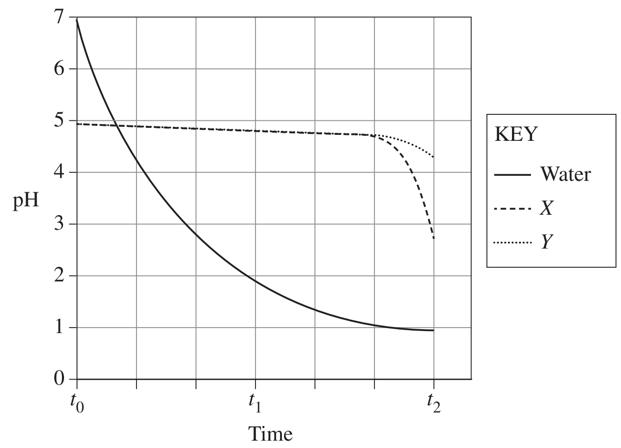
The results are shown in the table.

| *Titre* | *Volume of NaOH(aq) added* (mL) |
|---|---|
| 1 | 16.35 |
| 2 | 10.10 |
| 3 | 12.35 |
| 4 | 11.25 |

Explain why this method produces inaccurate and unreliable results. *(3 marks)*

## 43. CHEMISTRY, M6 2021 HSC 34

Gaseous $HCl$ was bubbled into water and two solutions, $X$ and $Y$. Solutions $X$ and $Y$ contain the same type of ions. The pH of each was monitored over time and recorded in the graph shown.




Explain the observed pH of the water and each of the solutions at $t_0$, $t_1$ and $t_2$. Include a relevant balanced chemical equation in your answer. ***(5 marks)***

## 44. CHEMISTRY, M6 2021 HSC 36

The $pK_a$ of sulfurous acid in the following reaction is 1.82.

$$H_2SO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HSO_3^-(aq)$$

The $pK_a$ of hydrogen sulfite in the following reaction is 7.17.

$$HSO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_3^{2-}(aq)$$

Calculate the equilibrium constant for the following reaction: *(5 marks)*

$$H_2SO_3(aq) + 2\,H_2O(l) \rightleftharpoons 2\,H_3O^+(aq) + SO_3^{2-}(aq)$$

## 45. CHEMISTRY, M6 2023 HSC 32

The ammonium ion content of mixtures can be determined by boiling the mixture with a known excess of sodium hydroxide. This converts the ammonium ions into gaseous ammonia, which is removed from the system.

$$NH_4{}^+\,(aq) + OH^-\,(aq) \rightarrow NH_3\,(g) + H_2O\,(l)$$

The excess sodium hydroxide can then be titrated with an acid solution of known concentration.

A fertiliser containing ammonium ions was analysed as follows.

- A sample of fertiliser was treated with 50.00 mL of 1.124 mol $L^{-1}$ sodium hydroxide solution and the solution boiled.
- After all of the ammonia was removed, the resulting solution was transferred to a 250.0 mL volumetric flask and made up to the mark with deionised water.
- 20.00 mL aliquots of this solution were titrated with 0.1102 mol $L^{-1}$ hydrochloric acid, giving the following results.

| *Titration* | *Volume* HCl (mL) |
|---|---|
| 1 | 22.65 |
| 2 | 22.05 |
| 3 | 22.00 |
| 4 | 21.95 |

Calculate the mass of ammonium ions in the sample of fertiliser. ***(5 marks)***

## 46. CHEMISTRY, M6 2023 HSC 35

a. A 0.2000 mol L$^{-1}$ solution of dichloroacetic acid $(CHCl_2COOH)$ has a pH of 1.107. Dichloroacetic acid is monoprotic.

Calculate the $K_a$ for dichloroacetic acid. *(3 marks)*

b. The following data apply to the ionisation of acetic acid $(CH_3COOH)$ and trichloroacetic acid $(CCl_3COOH)$.

| | $CH_3$ COOH | $CCl_3COOH$ |
|---|---|---|
| $pK_a$ | 4.76 | 0.51 |
| $\Delta H^\circ$ (kJ mol$^{-1}$) | $-0.1$ | $+1.2$ |
| $\Delta S^\circ$ (J K$^{-1}$ mol$^{-1}$) | $-91.6$ | $-5.8$ |
| $-T\Delta S^\circ$ (kJ mol$^{-1}$) | $+27.3$ | $+1.7$ |
| $\Delta G^\circ$ (kJ mol$^{-1}$) | $+27.2$ | $+2.9$ |

Explain the relative strength of these acids with reference to the data. *(3 marks)*

## 47. CHEMISTRY, M6 EQ-Bank 24

The pH of a 0.30 M aqueous propanoic acid solution was measured to be 2.7. The dissociation of propanoic acid is represented below.

$$CH_3CH_2COOH(aq) + H_2O(l) \rightleftharpoons CH_3CH_2COO^-(aq) + H_3O^+(aq)$$

Calculate the $K_a$ of the solution. *(3 marks)*

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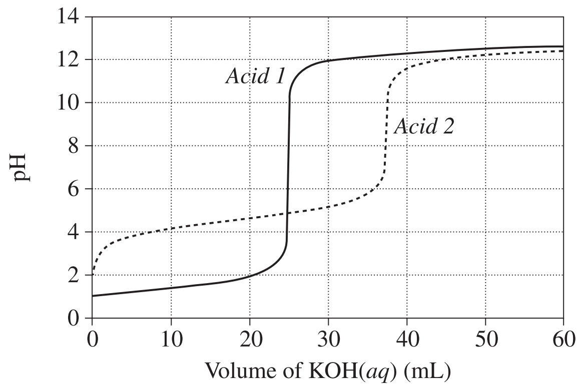
## 48. CHEMISTRY, M6 EQ-Bank 25

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




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

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## 49. CHEMISTRY, M6 EQ-Bank 29

Explain why a mixture of acetic acid (1 M) and sodium acetate (1 M) can act as a buffer while a mixture of hydrochloric acid (1 M) and sodium chloride (1 M) cannot. *(3 marks)*

## 50. CHEMISTRY, M6 2018 HSC 29

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

a. Explain why sodium carbonate is a suitable compound for preparation of a standard solution. *(2 marks)*

b. 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.

| *Initial burette reading* (mL) | *Final burette reading* (mL) | *Titre* (mL) |
|---|---|---|
| 0.00 | 22.00 | 22.00 |
| 22.00 | 43.65 | 21.65 |
| 0.00 | 21.70 | 21.70 |
| 21.70 | 43.30 | 21.60 |

Using the data from the table, calculate the concentration of the hydrochloric acid. *(3 marks)*

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

## 51. CHEMISTRY, M5 2024 HSC 39

Water and octan-1-ol do not mix. When an aqueous solution of bromoacetic acid $(\mathrm{BrCH_2COOH})$ is shaken with octan-1-ol, an equilibrium system is established between bromoacetic acid dissolved in the octan-1-ol and in the water.

$$\mathrm{BrCH_2COOH\,(aq)} \rightleftharpoons \mathrm{BrCH_2COOH}\,\textit{(octan-l-ol)}$$

The equilibrium constant expression for this system is

$$K_{eq} = \frac{[\mathrm{BrCH_2COOH}\,\textit{(octan-l-ol)}]}{[\mathrm{BrCH_2COOH}\,(aq)]}.$$

An aqueous solution of bromoacetic acid with an initial concentration of 0.1000 mol L $^{-1}$ is shaken with an equal volume of octan-1-ol. Bromoacetic acid does not dissociate in octan-1-ol but does dissociate in water, with $K_a = 1.29 \times 10^{-3}$. When the system has reached equilibrium, the $[\mathrm{H^+}]$ is $9.18 \times 10^{-3}\ \mathrm{mol\ L^{-1}}$.

Calculate the equilibrium concentration of aqueous bromoacetic acid and hence, or otherwise, calculate the $K_{eq}$ for the octan-1-ol and water system. *(4 marks)*

## 52. CHEMISTRY, M6 2015 HSC 24

a. Explain why the salt, sodium acetate, forms a basic solution when dissolved in water. Include an equation in your answer. *(2 marks)*

b. A solution is prepared by using equal volumes and concentrations of acetic acid and sodium acetate.

Explain how the pH of this solution would be affected by the addition of a small amount of sodium hydroxide solution. Include an equation in your answer. *(3 marks)*

## 53. CHEMISTRY, M6 2016 HSC 29

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 mol $L^{-1}$.

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

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.

- 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 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. *(4 marks)*

## 54. CHEMISTRY, M6 2021 HSC 32

The molar enthalpies of neutralisation of three reactions are given.

Reaction 1:

$$\mathrm{HCl}(aq) + \mathrm{KOH}(aq) \longrightarrow \mathrm{KCl}(aq) + \mathrm{H_2O}(l) \qquad \Delta H = -57.6\ \mathrm{kJ\,mol^{-1}}$$

Reaction 2:

$$\mathrm{HNO_3}(aq) + \mathrm{KOH}(aq) \longrightarrow \mathrm{KNO_3}(aq) + \mathrm{H_2O}(l) \qquad \Delta H = -57.6\ \mathrm{kJ\,mol^{-1}}$$

Reaction 3:

$$\mathrm{HCN}(aq) + \mathrm{KOH}(aq) \longrightarrow \mathrm{KCN}(aq) + \mathrm{H_2O}(l) \qquad \Delta H = -12.0\ \mathrm{kJ\,mol^{-1}}$$

Explain why the first two reactions have the same enthalpy value but the third reaction has a different value. *(4 marks)*

## 55. CHEMISTRY, M6 2022 HSC 32

The concentration of citric acid, a triprotic acid, in a carbonated soft drink was to be determined.

**Step 1:** A solution of $\mathrm{NaOH}\,(\mathrm{aq})$ was standardised by titrating it against 25.00 mL aliquots of a solution of the monoprotic acid potassium hydrogen phthalate $(\mathrm{KHP})$. The $(\mathrm{KHP})$ solution was produced by dissolving 4.989 g in enough water to make 100.0 mL of solution. The molar mass of $(\mathrm{KHP})$ is 204.22 g mol $^{-1}$.

The results of the standardisation titration are given in the table.

| *Titration* | *Volume NaOH* (mL) |
|---|---|
| 1 | 28.60 |
| 2 | 27.40 |
| 3 | 27.20 |
| 4 | 27.60 |

**Step 2:** A 75.00 mL bottle of the drink was opened and the contents quantitatively transferred to a beaker. The soft drink was gently heated to remove $\mathrm{CO_2}$.

**Step 3:** The cooled drink was quantitatively transferred to a 250.0 mL volumetric flask and distilled water was added up to the mark.

**Step 4:** 25.00 mL samples of the solution were titrated with the $\mathrm{NaOH}\,(\mathrm{aq})$ solution. The average volume of $\mathrm{NaOH}\,(\mathrm{aq})$ used was 13.10 mL.

a. Calculate the concentration of the triprotic citric acid in the soft drink. *(6 marks)*

b. Explain how your answer to part (a) would be different if the carbon dioxide was not removed from the soft drink. *(2 marks)*

## 56. CHEMISTRY, M6 2021 HSC 35

A manufacturer requires that its product contains at least 85% v/v ethanol.

The concentration of ethanol in water can be determined by a back titration. Ethanol is first oxidised to ethanoic acid using an excess of acidified potassium dichromate solution.

$$3\,C_2H_5OH(aq) + 2\,Cr_2O_7{}^{2-}(aq) + 16\,H^+(aq) \longrightarrow 3\,CH_3COOH(aq) + 4\,Cr^{3+}(aq) + 11\,H_2O$$

The remaining dichromate ions are reacted with excess iodide ions to produce iodine $(I_2)$

$$Cr_2O_7{}^{2-}(aq) + 14\,H^+(aq) + 6\,1^-(aq) \longrightarrow 2\,Cr^{3+}(aq) + 7\,H_2O(l) + 3\,I_2(aq)$$

The iodine produced is then titrated with sodium thiosulfate $(Na_2S_2O_3)$.

$$I_2(aq) + 2\,S_2O_3{}^{2-}(aq) \longrightarrow 2\,I^-(aq) + S_4O_6{}^{2-}(aq)$$

A 25.0 mL sample of the manufacturer's product was diluted with distilled water to 1.00 L. A 25.0 mL aliquot of the diluted solution was added to 20.0 mL of 0.500 mol $L^{-1}$ acidified potassium dichromate solution in a conical flask. Potassium iodide (5.0 g) was added and the solution titrated with 0.900 mol $L^{-1}$ sodium thiosulfate. This was repeated three times.

The following results were obtained.

| *Time* | *Volume of* $Na_2S_2O_3(aq)$ *added* (mL) |
|---|---|
| 1 | 29.9 |
| 2 | 28.7 |
| 3 | 28.4 |
| 4 | 28.6 |

The density of ethanol is 0.789 g $mL^{-1}$.

Does the sample meet the manufacturer's requirements? Support your answer with calculations.

***(7 marks)***

## Worked Solutions

### 1. CHEMISTRY, M6 2015 HSC 2 MC

A burette is used in a titration.

$\Rightarrow D$

---

### 2. CHEMISTRY, M6 2019 HSC 5 MC

$\rightarrow$ The pH range at which isopicramic acid exhibits a colour change includes the point at which the acid and base react in equal amounts (equivalence point), which is at approximately pH 5.

$\rightarrow$ The colour change can be used to identify when the equivalence point has been reached in a titration.

$\Rightarrow C$

---

### 3. CHEMISTRY, M2 2010 HSC 7 MC

$\rightarrow$ Citric acid is triprotic.

$\rightarrow$ Sulfuric acid is diprotic.

$\rightarrow$ Acetic acid and Hydrochloric acid are monoprotic.

$\rightarrow$ The greatest volume of $\mathrm{NaOH}$ is required for the triprotic acid as it has a 1:3, acid:base, stoichiometric ratio.

$\Rightarrow A$

---

### 4. CHEMISTRY, M6 2015 HSC 14 MC

$\rightarrow$ The indicator suitable for this titration needs to completely change colour at the centre of the vertical section of the graph (equivalence point).

$\rightarrow$The range that will achieve this is $8.3-10.0$.

$\Rightarrow D$

---

### 5. CHEMISTRY, M6 2017 HSC 1 MC

$\rightarrow$ A burette can measure and deliver to 0.05 mL accuracy, making it easily the most accurate measuring device of the given options.

$\Rightarrow A$

---

### 6. CHEMISTRY, M6 2018 HSC 17 MC

By Elimination:

$\rightarrow$ A buffer will resist changes in pH caused by the addition of an acidic or basic solution.

$\rightarrow$ If the water and buffer solution is decreasing in pH, the water solution will be declining in pH quicker and its graph will lie below (eliminate B and D).

$\rightarrow$ If the water and buffer solution is increasing in pH, the water solution should be increasing at a greater rate (eliminate C).

$\Rightarrow A$

---

### 7. CHEMISTRY, M6 2019 HSC 6 MC

$\rightarrow$The titration curve shows the progress of a titration in which a weak base $\mathrm{NH_3}$ is reacting with a strong acid $\mathrm{CHl}$ .

$\rightarrow$As the acid is added to the base, the pH of the solution moves lower.

$\Rightarrow A$

---

## 8. CHEMISTRY, M6 2022 HSC 15 MC

The equivalence point occurs at 10.0 mL of ammonia.

$\mathrm{HCl}(aq) + \mathrm{NH_3}(aq) \rightarrow \mathrm{NH_4Cl}(aq)$

$$\begin{aligned}\mathrm{n(HCl)} &= \mathrm{c} \times \mathrm{V}\\ &= 0.1131 \times 0.02500\\ &= 2.828 \times 10^{-3}\ \mathrm{mol}\end{aligned}$$

$\mathrm{n(NH_3)} = \mathrm{n(HCl)} = 2.828 \times 10^{-3}\ \mathrm{mol}$

$$\begin{aligned}\therefore\ [\mathrm{NH_3}] &= \frac{\mathrm{n}}{\mathrm{V}}\\ &= \frac{2.828 \times 10^{-3}}{0.0100}\\ &= 0.283\ \mathrm{mol\ L^{-1}}\end{aligned}$$

$\Rightarrow C$

---

## 9. CHEMISTRY, M6 2023 HSC 5 MC

→ Weak acids will only partially dissociate in solution and concentrated acids contain a large number of ions in solution.

$\Rightarrow D$

---

## 10. CHEMISTRY, M6 2023 HSC 9 MC

→ Weak base acts as a buffer, resisting an immediate decrease in pH upon addition of strong acid

$\Rightarrow C$

---

## 11. CHEMISTRY, M6 2024 HSC 16 MC

→ A buffer maintains the pH of a solution when a small amount of acid or base is added to it by shifting its equilibrium position to minimise change in the concentration of $\mathrm{H_3O^+}$ or $\mathrm{OH^-}$ ions.

→ When a strong acid is added to the solution, the $[\mathrm{H_3O^+}]$ is increased, hence the system will react to reduce the increase in $[\mathrm{H_3O^+}]$.

→ Hence, the hydronium ions will react with the weak base of the buffer solution to minimise the change in pH of the solution.

$$\mathrm{CH_3COO^-} + \mathrm{H_3O^+} \rightarrow \mathrm{CH_3COOH} + \mathrm{H_2O}$$

$\Rightarrow C$

---

## 12. CHEMISTRY, M6 EQ-Bank 13 MC

→ Higher pH corresponds to a weaker acid.

→ Higher pKa corresponds to a weaker acid (i.e. a higher pH).

→ Acetic acid the weaker acid (higher pKa), meaning it is less likely to dissociate in solution and lose a hydrogen ion.

$\Rightarrow A$

---

## 13. CHEMISTRY, M6 2009 HSC 14 MC

$\mathrm{H_3A\,(aq)} + 3\,\mathrm{NaOH\,(aq)} \longrightarrow \mathrm{Na_3A\,(aq)} + 3\,\mathrm{H_2O\,(l)}$

$\mathrm{n(NaOH)} = \mathrm{c} \times \mathrm{v} = 0.550 \times 0.0295 = 0.016225\ \mathrm{mol}$

$\mathrm{n(citric\ acid)} = \dfrac{0.016225}{3} = 0.00541\ \mathrm{mol}$

$[\mathrm{citric\ acid}] = \dfrac{0.00541}{0.025} = 0.216\ \mathrm{molL^{-1}}$

Multiply by molar mass:.

$\mathrm{Concentration\ (citric\ acid)} = 0.216 \times 192.12 = 41.5\ \mathrm{g\ L^{-1}}$

$\Rightarrow B$

---

## 14. CHEMISTRY, M6 2017 HSC 14 MC

→ $\mathrm{HCl}$ is a strong acid

→ Aqueous solution formed will be ineffective as a buffer (no equilibrium in solution will be formed).

♦ Mean mark 51%.

$\Rightarrow A$

## 15. CHEMISTRY, M6 2018 HSC 15 MC

By Elimination:

→Dihydrogen phosphate molecule is $\mathrm{H_2PO_4}^-$ and hydrogen phosphate is $\mathrm{HPO_4}^{2-}$ (eliminate A and B)

Mean mark 55%.

→When acid is added, by Le Chatelier's principle, the equation will shift to the left to reduce the $\mathrm{H_3O^+}$

$\Rightarrow D$

## 16. CHEMISTRY, M6 2020 HSC 8 MC

The solution would be most effective as a buffer at $\mathrm{pH}$ 9 because this corresponds to the half-equivalent point.

♦♦ Mean mark 37%.

$\Rightarrow D$

## 17. CHEMISTRY, M6 2021 HSC 16 MC

From the graph there are 2 equivalence points (at ~ 10 mL and 20 mL), indicating that the acid is diprotic.

♦ Mean mark 39%.

→ contains 2 acidic protons.

$\Rightarrow B$

## 18. CHEMISTRY, M6 2021 HSC 5 MC

Two points to consider in method:

♦ Mean mark 45%.

Burette rinsed with water instead of sodium hydroxide.

→ Sodium hydroxide solution diluted requiring more sodium hydroxide to neutralise the acetic acid.

This will result in greater moles of acetic acid, and thus the calculated concentration of acetic acid would be greater than the actual concentration.

Conical flask rinsed with acetic acid

→ Results in a greater number of moles of acid. More volume of sodium hydroxide would be added

→ Greater number of moles of acetic acid.

This would result in a greater number of moles, and thus the calculated concentration would be higher than the actual concentration.

In both cases, the calculated concentration would be higher than the actual concentration.

$\Rightarrow B$

## 19. CHEMISTRY, M6 2022 HSC 2 MC

→ A primary standard is a solution that is required to be accurately prepared with an accurately known concentration.

♦ Mean mark 40%.

→ It is prepared by adding an accurately measured mass of solute into a solvent. This solute is weighed out using an electronic balance.

$\Rightarrow B$

## 20. CHEMISTRY, M6 2022 HSC 20 MC

$HIO\,(aq) + H_2O\,(l) \leftrightharpoons IO^-\,(aq) + H_3O^+\,(aq)$

♦ Mean mark 46%.

| | HIO | $IO^-$ | $H_3O^+$ |
|---|---|---|---|
| Initial | 0.75 | 0 | 0 |
| Change | $-x$ | $+x$ | $+x$ |
| Equilibrium | $0.75 - x$ | $x$ | $x$ |

$$K_a = \frac{[IO^-][H_3O^+]}{[HIO]} = \frac{x^2}{(0.75 - x)}$$

$K_a$ is small $\Rightarrow 0.75 - x \approx 0.75$

$$\begin{aligned} K_a &= \frac{x^2}{0.75} \\ 10^{-10.64} &= \frac{x^2}{0.75} \\ x^2 &= 10^{-10.64} \times 0.75 \\ x &= \sqrt{10^{-10.64} \times 0.75} \\ &= 4.1 \times 10^{-6}\ \text{mol L}^{-1} \end{aligned}$$

$\therefore \left[H_3O^+\right] = 4.1 \times 10^{-6}\ \text{mol L}^{-1}$

$pH = -\log_{10}\left[4.1 \times 10^{-6}\right] = 5.38$

The major species at pH (see graph) = 5.38 is purple.

$\Rightarrow C$

---

## 21. CHEMISTRY, M6 2023 HSC 14 MC

$2\,HCl + Na_2CO_3 \rightarrow 2\,NaCl + H_2CO_3$

♦ Mean mark 45%.

$$n(Na_2CO_3) = \frac{m}{MM} = \frac{1.34}{105.99} = 0.0126\ \text{mol}$$

$n(HCl) = 0.0126 \times 2 = 0.0253\ \text{mol}$

$$\text{Vol (HCl)} = \frac{0.0253}{0.540} = 0.0468\ \text{L} = 46.8\ \text{mL}$$

$\Rightarrow D$

---

## 22. CHEMISTRY, M6 2024 HSC 17 MC

♦ Mean mark 41%.
**COMMENT:** Student's did well understanding the relationship between the $pK_a$ values and half-equivalence points.

→ There are two equivalence points on the graph, so the acid is diprotic and will have two $pK_a$ values.

→ The reaction at the first equivalence point is:

$H_2A \longrightarrow HA^- + H^+$

→ The reaction at the second equivalence point is:

$HA^- \longrightarrow A^{2-} + H^+$

→ The $pK_a$ values for each equivalence point will be equal to the pH values at the half-equivalence points.

→ The first half-equivalence point occurs at 10 mL, where the corresponding pH is 1.91 $\Rightarrow pK_{a1}$ = 1.91

→ The second half-equivalence point will occur halfway between the first two equivalence points, at 30 mL

→ The corresponding pH is 6.30 $\Rightarrow pK_{a2}$ = 6.30

$\Rightarrow C$

---

## 23. CHEMISTRY, M6 2024 HSC 20 MC

♦ Mean mark 43%.

$$\rightarrow n(\text{ascorbic acid}) = \frac{0.05}{176.124} = 0.00028389\ \text{mol}$$

→ Volume of KOH to neutralise the two samples = 15.6 mL.

→ 15.6 mL was required to neutralise the extra 0.00028389 mol of ascorbic acid.

→ Concentration of potassium hydroxide

$$= \frac{n}{V} = \frac{0.00028389}{0.0156} = 0.018198\ \text{mol L}^{-1}$$

→ $n(KOH)$ used in the first titration

$= c \times V = 0.018198 \times 0.0175 = 0.0003185\ \text{mol} = n(\text{ascorbic acid})$

→

$$[\text{ascrobic acid}] = \frac{0.0003185}{0.0250} = 0.01274 = 1.274 \times 10^{-2}\ \text{mol L}^{-1}$$

$\Rightarrow D$

## 24. CHEMISTRY, M6 EQ-Bank 14 MC

→ The volume of base required for complete reaction is independent of the strength of the acid as all neutralisation reactions go to completion.

→ Phosphoric acid ($H_3PO_4$) has three acidic hydrogens whereas the other given bases only have one. So, phosphoric acid requires three times the amount of base as any other acid given.

$\Rightarrow B$

---

## 25. CHEMISTRY, M6 2017 HSC 13 MC

→ Citric acid is triprotic (i.e. ratio moles NaOH : acid = 3 : 1). It therefore requires the greates volume of NaOH.

♦♦♦ Mean mark 29%.

→ Acetic acid and Hydrochloric acid are monoprotic (i.e. ratio moles NaOH : acid = 1 : 1)

→ Sulfuric acid is diprotic (i.e. ratio moles NaOH : acid = 2 : 1)

$\Rightarrow B$

---

## 26. CHEMISTRY, M6 2021 HSC 20 MC

$[(CH_3)_3\,NH\,]^+(aq) + H_2O\,(l) \leftrightharpoons H_3O^+(aq) + (CH_3)_3\,N\,(aq)$

$K_a = \dfrac{[(CH_3)_3N]\,[H_3O^+]}{[(CH_3)_3NH]^+}$

♦♦♦ Mean mark 19%.

To calculate $K_{sp}$ equation:

$[(CH_3)_3\,NH\,]\,Cl\,(s) \leftrightharpoons [(CH_3)_3\,NH\,]^+(aq) + Cl^-(aq)$

$K_{sp} = [(CH_3)_3NH)^+]\,[Cl^-]$

$pH = 4.46 \rightarrow [H_3O^+] = 10^{-4.46}$

Using stoichiometry;

$[(CH_3)_3N)^+] = [H_3O^+] = 10^{-4.46}$

Using $K_a$:

$1.55 \times 10^{-10} = \dfrac{(10^{-4.46} \times 10^{-4.46})}{(CH_3)_3NH^+}$

$[((CH_3)_3NH\,)^+] = \dfrac{(10^{-4.46} \times 10^{-4.46})}{1.55 \times 10^{-10}} = 7.7565\ldots mol\,L^{-1}$

$[Cl^-] = [((CH_3)_3NH)^+] = 7.7565\ldots mol\,L^{-1}$

$\therefore K_{sp} = [((CH_3)_3NH)^+] \times [Cl^-] = 7.7565\ldots \times 7.7565\ldots = 60.2$

$\Rightarrow C$

## 27. CHEMISTRY, M6 2021 HSC 6 MC

A weak acid has the following equilibrium:

$$HA(aq) + H_2O(l) \leftrightharpoons A^-(aq) + H_3O^+(aq)$$

♦♦♦ Mean mark 25%.

$$K_a = \frac{[A^-][H_3O^+]}{[HA]}$$

→ The value of $K_a$ is only affected by temperature, and thus the value of $K_a$ will remain the same.

→ When the solution is diluted, water is added. According to Le Chatelier's Principle, the equilibrium will shift to the right to counteract the change.

→ Thus, the equilibrium will shift to the right and increase the extent of ionisation.

$\Rightarrow C$

---

## 28. CHEMISTRY, M6 2022 HSC 25

→ $HCl$ is a strong acid, ie it completely ionises in water to form $H^+$ ions.

Mean mark 57%.

→ On the other hand, $HCN$ is a weak acid, ie it partially ionises in water to form $H^+$ ions.

$$HCl\,(aq) \longrightarrow H^+\,(aq) + Cl^-\,(aq)$$

$$HCN\,(aq) \rightleftharpoons H^+\,(aq) + CN^-\,(aq)$$

→ As $[H^+]$ decreases, pH increases $(pH = -\log[H^+])$

→ Therefore, at the same 0.2M, the $HCN$ solution would have a lower $[H^+]$ and thus would have a higher pH than $HCl$.

---

## 29. CHEMISTRY, M6 2024 HSC 34

→ Initially, $H^+\,(aq)$ and $Cl^-\,(aq)$ are present in the solution. Due to the significant relative conductivity of $H^+\,(aq)$, the overall conductivity of the solution is high.

→ As ammonia is added to the solution prior to the equivalence point (below 4.5 mL), hydrochloric acid is neutralized by the addition of ammonia according to the reaction:

$$H^+\,(aq) + NH_3\,(aq) \longrightarrow NH_4{}^+\,(aq)$$

→ This results in the replacement of $H^+$, which exhibit high conductivity, with $NH_4{}^+$, which have lower conductivity. Solution conductivity decreases.

→ The conductivity at the equivalence point is the lowest as only $Cl^-$ and $NH_4{}^+$ ions are present which both have low relative conductivities.

→ Beyond the equivalence point, the excess ammonia added reacts partially with water to form $NH_4{}^+$ and $OH^-$ ions according to the equation:

$$NH_3\,(aq) + H_2O\,(l) \rightleftharpoons NH_4{}^+\,(aq) + OH^-\,(aq)$$

→ Although these ions are more conductive than the reactant molecules, the ionisation of ammonia is limited due to the presence of $NH_4{}^+$ ions already in the solution (as per Le Châtelier's Principle).

→ Consequently, the conductivity increases only slightly after the equivalence point.

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## 30. CHEMISTRY, M6 EQ-Bank 27

→ The strength of the acid. The greater the pKa of (weaker) the acid, the greater the buffer pH.

→ Relative concentrations of acid and base.

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### 31. CHEMISTRY, M6 2012 HSC 30

**a.** $n(\mathrm{HCl}) = c \times V = 0.1034 \times 0.02500 = 2.585 \times 10^{-3}$ moles

$n(\mathrm{HCl}) = n(\mathrm{OH^-})$

$$[\mathrm{OH^-}] = \frac{2.585 \times 10^{-3}}{0.02575} = 0.1004\,\mathrm{mol\,L^{-1}}$$

**b.** $n(\mathrm{HCl}) = c \times V = 0.1004 \times 0.01655 = 1.661 \times 10^{-3}$ moles

$n(\mathrm{HCl}) = n(\mathrm{C_9H_8O_4}) = 1.661 \times 10^{-3}$ moles

$\mathrm{MM}(\mathrm{C_9H_8O_4}) = 9 \times 12.01 + 8 \times 1.008 + 4 \times 16.00 = 180.154\,\mathrm{g}$

Average mass of $C_9H_8O_4$ per tablet

$= n \times \mathrm{MM} = 1.661 \times 10^{-3} \times 180.154 = 0.2992\,\mathrm{g} = 299.2\,\mathrm{mg}$

---

### 32. CHEMISTRY, M6 2014 HSC 30

**a.** $n(\mathrm{NaOH}) = c \times V = 0.0500 \times 1.00 = 0.0500$ moles

**b.**

♦ Mean mark (b) 41%.

$n(\mathrm{HCl}) = c \times V = 0.0276 \times 1.00 = 0.0276\,\mathrm{mol}$ (titrate excess NaOH)

$n(\mathrm{NaOH\ to\ neutralise\ CO_2}) = 0.0500 - 0.0276 = 0.0224\,\mathrm{mol}$

Ratio $\mathrm{NaOH} : \mathrm{CO_2} = 2 : 1$ (from equation)

$n(\mathrm{CO_2}) = \frac{1}{2} \times n(\mathrm{NaOH}) = \frac{1}{2} \times 0.0224 = 0.0112\,\mathrm{mol}$

$\mathrm{MM}\ (\mathrm{CO_2}) = 12.01 + 2 \times 16 = 44.01$

$\mathrm{Mass}\ (\mathrm{CO_2}) = n \times \mathrm{MM} = 0.0112 \times 44.01 = 0.493\,\mathrm{g}$

$$\%\ \text{Dry ice (by mass)} = \frac{0.493}{0.616} \times 100\,\% = 80.0\,\%$$

---

### 33. CHEMISTRY, M6 2016 HSC 29

**a.** Substance for rinse:

→ Water should be used to rinse the conical flask as this will not change the number of moles of $Na_2CO_3$ placed in it.

**b.** $\mathrm{HCl + NaOH \rightarrow H_2O + NaCl}$

♦ Mean mark (b) 40%.

$n(\mathrm{NaOH}) = c \times V = 0.250 \times 0.0295 = 7.375 \times 10^{-3}$ moles

$n(\mathrm{HCl}) = 7.375 \times 10^{-3}$ (after reaction)

$n(\mathrm{HCl - original}) = c \times V = 0.200 \times 0.0500 = 0.0100$ moles

$n(\mathrm{HCl - used}) = 0.0100 - 7.375 \times 10^{-3} = 2.625 \times 10^{-3}$ moles

$2\,\mathrm{HCl} + \mathrm{CO_3}^{2-} \rightarrow \mathrm{H_2O} + \mathrm{CO_2} + 2\,\mathrm{Cl^-}$

$\mathrm{HCl} : \mathrm{CO_3}^{2-} = 2 : 1$

$$n(\mathrm{CO_3}^{2-}) = \frac{2.625 \times 10^{-3}}{2} = 1.3125 \times 10^{-3}\ \text{moles}$$

$m(\mathrm{CO_3}^{2-}) = 1.3125 \times 10^{-3} \times 60.01 = 0.07876\,\mathrm{g}$

$$\%\ \text{Mass}\ (\mathrm{CO_3}^{2-}) = \frac{0.07876}{0.145} \times 100\,\% = 54.3\,\%$$

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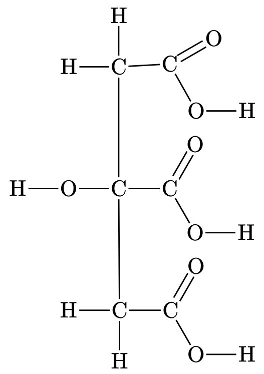
## 34. CHEMISTRY, M6 2015 HSC 26

**a.**



♦ Mean mark (a) 46%.

**b.** Technology solution

→ A digital pH probe could be placed in the flask and used to collect data that plots the pH of the solution against the volume of sodium hydroxide added.

→ The equivalence point would be identified by a steep rise in the pH on the graph.

♦ Mean mark (b) 42%.

**c.** $C_6H_8O_7 + 3\,NaOH \longrightarrow C_6H_5O_7Na_3 + 3\,H_2O$

$n(C_6H_8O_7) = c \times V = 0.100 \times 0.0250 = 0.00250\,mol$

$n(NaOH) = 3 \times 0.00250 = 0.00750\,mol$

$$[NaOH] = \frac{n}{V} = \frac{0.00750}{0.04150} = 0.181\,mol\,L^{-1}\ (3\,sig\,fig)$$

Mean mark (c) 56%.

## 35. CHEMISTRY, M6 2017 HSC 24

**a.** $CH_3COOH\,(aq) + NaOH\,(aq) \longrightarrow NaCH_3COO\,(aq) + H_2O\,(l)$

$n(CH_3COOH) = c \times V = 0.5020 \times 0.0250 = 0.01255\,mol$

$n(NaOH) = n(CH_3COOH) = 0.01255\,mol$

$$[NaOH] = \frac{n}{V} = \frac{0.01255}{0.01930} = 0.6503\,mol\,L^{-1}$$

**b.** $CH_3COO^-\,(aq) + H_2O \rightleftharpoons CH_3COOH\,(aq) + OH^-\,(aq)$

→ The acetate ion is a weak base.

→ As a result, it has accepted a proton from the water resulting in production of hydroxide ions.

→ Therefore the solution has a pH > 7.

♦♦ Mean mark (b) 34%.

## 36. CHEMISTRY, M6 2019 HSC 24

**a.** Titration curve shape:

♦ Mean mark (a) 46%.

→ The conductivity of a solution of hydrochloric acid is initially very high due to the high concentration of ions in the solution.

→ As barium hydroxide is added, the conductivity decreases because the concentration of $H^+$ decreases due to the neutralization reaction with $OH^-$ ions.

→ The conductivity reaches a minimum at the equivalence point, which is when all the $H^+$ ions have been removed and the solution only contains $Ba^{2+}$ and $Cl^-$ ions.

→ These ions are much less mobile than $H^+$ or $OH^-$ ions, so the conductivity is lower at the equivalence point.

→ After the equivalence point, the conductivity increases as more barium hydroxide is added, resulting in an increase in the concentration of $OH^-$ ions and therefore an increase in conductivity.

**b.** $2\,HCl\,(aq) + Ba(OH)_2\,(aq) \longrightarrow BaCl_2\,(aq) + 2\,H_2O\,(l)$

$n(HCl) = c \times V = 1.050 \times 10^{-3} \times 250 = 2.625 \times 10^{-4}$ mol

$n(Ba(OH)_2) = 0.5 \times n(HCl) = 1.3125 \times 10^{-4}$ mol

$$[Ba(OH)_2] = \frac{n}{V} = \frac{1.3125 \times 10^{-4}}{0.01715}$$

$$= 7.653 \times 10^{-3}\ \text{mol L}^{-1}$$

## 37. CHEMISTRY, M6 2019 HSC 27

**a.** $K_a \times K_b = K_w \quad \Rightarrow \quad K_b = \dfrac{K_w}{K_a}$

$$K_b = \frac{1.0 \times 10^{-14}}{3.0 \times 10^{-8}}$$

$$= 3.3 \times 10^{-7}$$

**b.** $OCl^-\,(aq) + H_2O\,(l) \rightleftharpoons HOCl\,(aq) + OH^-\,(aq)$

♦ Mean mark (b) 45%.

| | $OCl^-$ | $HOCl$ | $OH^-$ |
|---|---|---|---|
| Initial | 0.20 | 0 | 0 |
| Change | $-x$ | $+x$ | $+x$ |
| Equilibrium | $0.20 - x$ | $x$ | $x$ |

$$K_b = \frac{[HOCl][OH^-]}{[OCl^-]} = \frac{x^2}{(0.20 - x)}$$

Assume $0.20 - x \approx 0.20$ because $x$ is negligible:

$$3.3 \times 10^{-7} = \frac{x^2}{0.20 - x}$$

$$x = \sqrt{3.3 \times 10^{-7} \times 0.20}$$

$$= 2.5690 \times 10^{-4}\ \text{mol L}^{-1}$$

$[OH^-] = 2.5690 \times 10^{-4}$ mol $L^{-1}$

$pOH = -\log_{10}[OH^-] = -\log_{10}(2.5690 \times 10^{-4}) = 3.59$

$\therefore pH = 14 - 3.59 = 10.41$

## 38. CHEMISTRY, M6 EQ-Bank 28

Step A

→ Prepare $Na_2CO_3$ by drying. Protect solid $Na_2CO_3$ from moisture in the air by storing in a desiccator.

→ Calculate mass of dried $Na_2CO_3$ required and weigh accurately.

$$m(Na_2CO_3) = 0.1 \times 0.5 \times 105.99 = 5.30\,g$$

→ Clean and rinse a 500 mL volumetric flask with distilled water.

→ Add 5.30 grams of $Na_2CO_3$ to the volumetric flask using a funnel and wash funnel using distilled water. Add distilled water to the flask to the bottom of the meniscus.

Step B

→ Clean and rinse a 50 mL burette. Fill burette with the unknown acid and place on a retort stand.

→ Clean and rinse a 250 mL conical flask with distilled water.

→ Clean a 25 mL pipette and rinse with 0.1 M $Na_2CO_3$ solution. Fill pipette with $Na_2CO_3$ solution to bottom of meniscus.

→ Transfer all pipette solution into conical flask and add an appropriate indicator. A white background (tile) should be placed under the flask to highlight any colour changes in the solution.

→ Slowly add acid solution from the burette into the conical flask and record the volume used when the indicator changes colour.

Step C

→ The initial titration represents a test run to establish an indicative volume. Three subsequent titrations should be performed with the average titration forming the basis of $HCl$ concentration calculations.

→ Calculate the concentration of $HCl$

$$2\,H^+\,(aq) + CO_3{}^{2-}\,(aq) \rightarrow H_2CO_3\,(aq) \rightarrow H_2O\,(l) + CO_2\,(g)$$

$$M(Na_2CO_3) = 0.1 \times 0.025 = 2.5 \times 10^{-3}$$

$$M(HCl) = 2 \times M(Na_2CO_3) = 5 \times 10^{-3}$$

$$[HCl] = \frac{M(HCl)}{Vol\ HCl} = \frac{5 \times 10^{-3}}{21.4 \times 10^{-3}} = 0.234\,mol\,L^{-1}$$

---

## 39. CHEMISTRY, M7 2024 HSC 35

♦ Mean mark 55%.

→ Both sample $Y$ and $Z$ undergo an addition reaction with bromine and a hydration reaction. Therefore these samples must contain a $C = C$ bond.

→ As sample $X$, undergoes neither of these reactions, it must have no $C = C$ bond, thus sample $X$ is structure 2.

→ Both structure 1 and structure 3 contain 1 $C = C$ each $\Rightarrow$ they will react in a $1:1$ with $Br_2$. Hence the same number of moles of the carboxylic acid samples will react with the bromine.

→ Since the mass of sample $Z$ that reacts with the bromine is double the mass of sample $Y$, the molar mass of sample $Z$ must be double the molar mass of sample $Y$ following the formula $m = n \times MM$.

→ Therefore, sample $Y$ is structure 1 and sample $Z$ is structure 3.

Other information provided that could support identification includes:

→ The two products formed for the hydration of $Y$ is due to the asymmetry of structure 1 and the single product formed in the hydration of $Z$ is due to the symmetrical nature of structure 3.

The titration values are consistent with the proposed samples and their corresponding structures.

→ $n(NaOH)$ reacted with $X = 0.0617 \times 0.02188 = 1.35 \times 10^{-3}$ mol

$n_X = \dfrac{0.100}{74.078} = 0.00135$ mol. Therefore $X$ reacts in a $1:1$ molar ratio as it is a monoprotic acid.

→ $n(NaOH)$ reacted with $Y$ and $Z = 0.0617 \times 0.02249 = 1.39 \times 10^{-3}$ mol

$n_Y = \dfrac{0.100}{72.062} = 0.00139$ mol. Therefore $Y$ reacts in a $1:1$ molar ratio as it is a monoprotic acid.

$n_Z = \dfrac{0.100}{144.124} = 0.0006938$ mol. Therefore $Z$ reacts in a $2:1$ molar ratio with $NaOH$ as it is a diprotic acid.

→ The equal volumes of Y and Z used in the titration can be attributed to Z having twice the molar mass of Y and being a diprotic acid.

## 40. CHEMISTRY, M6 2022 HSC 34

$\mathrm{pOH}_{eq} = 14.00 - 7.5 = 6.5$

$[\mathrm{OH}^-]_{eq} = 10^{-\mathrm{pOH}} = 10^{-6.5}\ \mathrm{mol\ L^{-1}}$

♦ Mean mark 46%.

$$\mathrm{K}_{eq} = \frac{[\mathrm{HOCl}]_{eq}[\mathrm{OH}^-]_{eq}}{[\mathrm{OCl}^-]_{eq}}$$

$$3.33 \times 10^{-7} = \frac{(1.3 \times 10^{-4}) \times (10^{-6.5})}{[\mathrm{OCl}^-]_{eq}}$$

$$[\mathrm{OCl}^-]_{eq} = \frac{(1.3 \times 10^{-4}) \times (10^{-6.5})}{3.33 \times 10^{-7}}$$

$$= 1.246 \times 10^{-4}\ \mathrm{mol\ L^{-1}}$$

| | $\mathrm{OCl}^-$ | HOCl | $\mathrm{OH}^-$ |
|---|---|---|---|
| Initial | $x$ | 0 | – |
| Change | $-1.3 \times 10^{-4}$ | $+1.3 \times 10^{-4}$ | – |
| Equilibrium | $x - 1.3 \times 10^{-4}$ | $1.3 \times 10^{-4}$ | $10^{-6.5}$ |

$$x - 1.3 \times 10^{-4} = 1.246 \times 10^{-4}$$

$$x = 2.546 \times 10^{-4}$$

$[\mathrm{OCl}^-]_i = 2.55 \times 10^{-4}\ \mathrm{mol\ L^{-1}}$ (3 s.f.)

$[\mathrm{NaOCl}] \times \mathrm{V(NaOC)_{req}} = [\mathrm{OCl}^-](\mathrm{pool}) \times \mathrm{V(pool)}$

$$\mathrm{V(NaCl)_{req}} = \frac{2.55 \times 10^{-4} \times 10^4}{2}$$

$$= 1.3\ \mathrm{L}\ \ (2\ \mathrm{s.f.})$$

## 41. CHEMISTRY, M6 2019 HSC 22

$$\mathrm{CH_3COOH\,(aq) + H_2O\,(l) \rightleftharpoons CH_3COO^-\,(aq) + H_3O^+\,(aq)}$$

♦♦ Mean mark 38%.

→ A buffer is a solution that resists changes in pH when small amounts of acid or base are added.

→ When a small amount of acid or base is added to a buffer solution containing $\mathrm{CH_3COOH}$ and $\mathrm{CH_3COO^-}$, the equilibrium of the reaction shifts in order to minimise the disturbance and maintain a stable pH.

→ This is due Le Chatelier's Principle, which states that when a disturbance occurs in a chemical system at equilibrium, the equilibrium will shift in a way that minimises the disturbance.

→ If acid is added to the buffer solution, the equilibrium will shift to the left to use the $\mathrm{H_3O^+}$ ions, and if base is added, the equilibrium will shift to the right to increase the $\mathrm{H_3O^+}$ concentration.

→ As a result, the pH remains relatively stable and there is no change in the colour of the indicator.

## 42. CHEMISTRY, M6 2020 HSC 28

→ NaOH is hygroscopic and cannot be accurately used as a primary standard in titration experiments.

♦ Mean mark 48%.

→ When the solid NaOH is weighed, it will have absorbed water from the atmosphere, which means that the solution made from it will be more dilute than expected.

→ Furthermore, because this is a titration between a weak acid and a strong base which produces a basic salt, the pH at the equivalence point will be greater than 7.

→ Using Bromocresol green as an indicator in this case would not be suitable because it changes colour in the pH range of 3.2–5.2. This is the flat region of the titration curve, before the equivalence point.

→ The results of the titration are also unreliable because the indicator used produces a non-sharp endpoint, resulting in significantly different titres in each titration.

## 43. CHEMISTRY, M6 2021 HSC 34

When $HCl$ is added to water, hydronium ions are produced:

♦ Mean mark 48%.

$HCl(aq) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$.

At $t_0$:

→ pH (water) = 7, pH ($X$) = pH ($Y$) = 4.9

At $t_1$:

→ the pH of water significantly decreases, this is due to $HCl$ reacting with $H_2O$ to produce $H_3O^+$ ions.

→ However, the pH of $X$ and $Y$ only slightly decreases. This indicates that $X$ and $Y$ are buffer solutions, ie contain a mixture of a weak acid or base and resist changes in pH when acids or bases are added.

→ Therefore, when $HCl$ is added, $H_3O^+$ increases, and thus disturbs the equilibrium.

→ According to Le Chatelier's Principle, the system will shift to decrease the $H_3O^+$ concentration, and therefore the pH change is minimised.

$$HA(aq) + H_2O(l) \leftrightharpoons A^-(aq) + H_3O^+(aq)$$

At $t_2$:

→ the decrease in pH of water becomes more gradual because pH is calculated on a $\log_{10}$ scale, and thus requires a greater amount of change in $[H_3O^+]$ to result in a significant change in pH.

→ The pH of $X$ and $Y$ begin to significantly decrease as they have reached their buffer capacity.

→ The pH is lower for $X$ because it was initially a less concentrated buffer and thus had a lower buffer capacity.

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## 44. CHEMISTRY, M6 2021 HSC 36

♦ Mean mark 50%.

$$Ka_1 = \frac{[HSO^{3-}][H_3O^+]}{[H_2SO_3]}$$

$$Ka_2 = \frac{[SO_3^{2-}][H_3O^+]}{[HSO_3^-]}$$

$$K_{eq} = \frac{[SO_3{}^{2-}][H_3O^+]^2}{[H_2SO_3]}$$

$K_{eq}$ can be derived by multiplying $Ka_1$ with $Ka_2$

$$Ka = 10^{-pK}$$

$$Ka_1 = 10^{-1.82}$$

$$Ka_2 = 10^{-7.17}$$

Therefore $K_{eq} = 10^{-1.82} \times 10^{-7.17}$

$K_{eq} = 1.02 \times 10^{-9}$

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## 45. CHEMISTRY, M6 2023 HSC 32

Mean mark 56%.

$$\text{Average titre (HCl)} = \frac{22.05 + 22.00 + 21.95}{3} = 22.00\ \text{mL} = 0.0$$

$$n(HCl) = c \times V = 0.02200 \times 0.1102 = 2.424 \times 10^{-3}\ \text{mol}$$

$$n(\text{NaOH excess}) = 2.424 \times 10^{-3}\ \text{mol}$$

In the 250 mL flask:

$$n(\text{NaOH excess}) = \frac{250.0}{20.00} \times 2.424 \times 10^{-3} = 3.031 \times 10^{-2}\ \text{mol}$$

$$n(\text{NaOH total}) = c \times V = 0.0500 \times 1.124 = 5.620 \times 10^{-2}\ \text{mol}$$

$$n(\text{NaOH reacting with } NH_4{}^+) = 5.620 \times 10^{-2} - 3.031 \times 10^{-2} = 2.589 \times 10^{-2}\ \text{mol}$$

$$n(NH_4{}^+) = 2.589 \times 10^{-2}\ \text{mol}$$

$$MM(NH_4{}^+) = 14.01 + 4 \times 1.008 = 18.042$$

$$m(NH_4{}^+) = 2.589 \times 10^{-2} \times 18.042 = 0.4671\ \text{g}$$

## 46. CHEMISTRY, M6 2023 HSC 35

♦ Mean mark (a) 54%.

**a.** $CHCl_2COOH\,(aq) \rightleftharpoons H^+\,(aq) + CHCl_2COO^-\,(aq)$

$[H^+] = 10^{-pH} = 10^{-1.107} = 0.0782 \text{ mol L}^{-1}$

| | $CHCl_2COOH\,(aq)$ | $H^+\,(aq)$ | $CHCl_2COO^-\,(aq)$ |
|---|---|---|---|
| Initial | 0.2000 | 0 | 0 |
| Change | −0.0782 | +0.0782 | +0.0782 |
| Equilibrium | 0.1218 | 0.0782 | 0.0782 |

$$K_a = \frac{[H^+][CHCl_2COO^-]}{[CHCl_2COOH]}$$
$$= \frac{0.0782 \times 0.0782}{0.1218}$$
$$= 0.0501$$

♦ Mean mark (b) 52%.

**b.** Relative strength of acids:

→ The $pK_a$ of trichloroacetic acid is lower than the $pK_a$ of acetic acid, so trichloroacetic acid is a stronger acid than acetic acid.

→ The $\Delta S^\circ$ term for acetic acid is a significantly lower number than for the trichloroacetic acid (noting they are both negative).

→ In both cases, this value will contribute unfavourably to each acid's $\Delta G^\circ$ value, with the effect much larger for acetic acid than for trichloroacetic acid.

→ It follows from this result that the ionisation of acetic acid is less favourable than it is for trichloroacetic acid, making the latter the stronger acid.

## 47. CHEMISTRY, M6 EQ-Bank 24

$pH = -\log_{10}[H_3O^+] = 2.7$

$[H_3O^+] = 10^{-2.7} = 1.995 \times 10^{-3} \text{ mol L}^{-1}$

$[H_3O^+] = [CH_3CH_2COO^-]$

$$K_a = \frac{[H_3O^+][CH_3CH_2COO^-]}{[CH_3CH_2COOH]}$$
$$= \frac{(1.995 \times 10^{-3})(1.995 \times 10^{-3})}{(0.30 - 1.995 \times 10^{-3})}$$
$$= 1.3 \times 10^{-5}$$

## 48. CHEMISTRY, M6 EQ-Bank 25

→ Acid 1 is a strong acid. Its initial pH = 1 and its equivalence point is at pH = 7.

→ Acid 2 is a weaker acid. Its initial pH ~ 2 and its equivalence point is >7.

→ Acid 2 has a higher concentration than Acid 1 as it doesn't take that much more $KOH$ to neutralise it.

## 49. CHEMISTRY, M6 EQ-Bank 29

→ A buffer is a mixture of a weak acid and its conjugate base which is able to counteract changes in pH when small amounts of acid or base is added.

→ In this case acetic acid and sodium acetate can act as a buffer. The following equilibrium is established:

$$\mathrm{CH_3COOH\,(aq)} + \mathrm{H_2O\,(l)} \rightleftharpoons \mathrm{CH_3COO^-\,(aq)} + \mathrm{H_3O^+\,(aq)}$$

→ Any addition of acid to the mixture increases the $[\mathrm{H_3O^+}]$ and shifts the above equilibrium to the left, removing $\mathrm{H_3O^+}$ ions and minimising a change in pH.

→ Likewise, any addition of base causes $\mathrm{OH^-}$ ions to react with $\mathrm{H_3O^+}$ in solution. This shifts the above equilibrium to the right, generating more $\mathrm{H_3O^+}$ and maintaining a relatively constant pH.

→ Hydrochloric acid is a strong acid, so its conjugate base, chloride has negligible basic activity.

→ The dissociation of hydrochloric acid goes to completion, so addition of base simply lowers the pH of the solution as there is no aqueous hydrochloric acid able to generate $\mathrm{H_3O^+}$. Addition of acid will raise the pH as chloride ions are unable to react with $\mathrm{H_3O^+}$ ions.

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## 50. CHEMISTRY, M6 2018 HSC 29

**a.** → $\mathrm{Na_2CO_3}$ is a stable compound.

→ $\mathrm{Na_2CO_3}$ is a pure solid that will not readily absorb water from the atmosphere.

♦ Mean mark (a) 43%.

→ An accurate weight of $\mathrm{Na_2CO_3}$ can therefore be obtained in the experiment's measurements.

**b.** $\mathrm{Na_2CO_3\,(aq)} + 2\,\mathrm{HCl\,(aq)} \longrightarrow 2\,\mathrm{NaCl\,(aq)} + \mathrm{H_2O\,(l)} + \mathrm{CO_2\,(g)}$

$$\text{Average titre} = \frac{21.65 + 21.70 + 21.60}{3} = 21.65\,\text{mL}$$

$$\mathrm{n(Na_2CO_3)} = \mathrm{c} \times \mathrm{V} = 0.1050 \times 0.0250 = 0.002625\,\text{mol}$$

$$\mathrm{n(HCl)} = 2 \times \mathrm{n(Na_2CO_3)} = 0.005250\,\text{mol}$$

$$[\mathrm{HCl}] = \frac{\mathrm{n}}{\mathrm{V}} = \frac{0.005250}{0.02165} = 0.2425\,\text{mol L}^{-1}$$

**c.** → This is a strong acid / weak base titration.

→ Its equivalence point will occur at a pH less than seven and phenolphthalein changes colour in the pH range 10 – 8.3.

♦♦♦ Mean mark (c) 29%.

→ Phenolphthalein indicator would therefore signal the end point before equivalence (i.e. with a lower volume of acid).

→ The calculated concentration of HCl would be higher than the correct concentration.

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## 51. CHEMISTRY, M5 2024 HSC 39

→ The ionisation of bromoacetic acid in water is:

$$BrCH_2COOH\,(aq) \rightleftharpoons BrCH_2COO^-\,(aq) + H^+\,(aq)$$

→ At equilibrium $[BrCH_2COO^-\,(aq)] = [H^+\,(aq)] = 9.18 \times 10^{-3}\ \text{mol L}^{-1}$ as the are formed in a 1 : 1.

$$K_a = \frac{[H^+][BrCH_2COO^-]}{[BrCH_2COOH]_{eq}}$$

$$[BrCH_2COOH]_{eq} = \frac{[H^+][BrCH_2COO^-]}{K_a}$$

$$= \frac{(9.18 \times 10^{-3})^2}{1.29 \times 10^{-3}}$$

$$= 0.06533\ \text{mol L}^{-1}$$

$$[BrCH_2COOH]_{total} = [BrCH_2COOH\,(aq)]_{eq} + [BrCH_2COO^-\,(aq)] + [BrCH_2COOH(\text{octan-1}$$

$$[BrCH_2COOH(\text{octan-1-ol})]_{eq} = 0.1000 - 0.06533 - 9.18 \times 10^{-3} = 0.02549\ \text{mol L}^{-1}$$

→ Since the volume of the aqueous solution of bromoacetic acid and octane is the same, the concentration values between the water and octane solutions can be added/subtracted in one equation and mole calculations are not required.

$$K_{eq} = \frac{[BrCH_2COOH(\text{octan-1-ol})]_{eq}}{[BrCH_2COOH\,(aq)]_{eq}} = \frac{0.02549}{0.06533} = 0.390\ (3\ \text{sig. fig.})$$

♦♦ Mean mark 27%.
**COMMENT:** Students who identified the acid conc in the organic solvent often succeeded in this question.

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## 52. CHEMISTRY, M6 2015 HSC 24

**a.** $CH_3COO^-\,(aq) + H_2O\,(l) \rightleftharpoons CH_3COOH\,(aq) + OH^-\,(aq)$

→ Sodium acetate is a basic salt.

→ Acetate is a strong base that accepts a proton, producing hydroxide.

→ The presence of $OH^-$ ions produced by the hydrolysis of $CH_3COO^-$ increases the pH, producing a basic solution.

♦♦ Mean mark (a) 37%.

**b.** $CH_3COO^-\,(aq) + H_3O^+\,(l) \rightleftharpoons CH_3COOH\,(aq) + H_2O\,(l)$

→ The $OH^-$ ions introduced into the solution will react with the $H_3O^+$ ions, reducing their concentration in the equilibrium mixture.

→ By Le Chatelier's principle, this will subsequently move the reaction to the left to increase the $H_3O^+$ ions, thus minimising any change in pH.

♦♦♦ Mean mark (b) 25%.

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## 53. CHEMISTRY, M6 2016 HSC 29

**a.** → Distilled water.

→ This should be used to rinse the conical flask as this will not change the number of moles of $Na_2CO_3$ placed in it.

♦ Mean mark (a) 42%.

**b.** $HCl\,(aq) + NaOH\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l)$

♦♦ Mean mark (b) 40%.

$$n(NaOH) = c \times V = 0.250 \times 0.0295 = 7.375 \times 10^{-3}\text{ mol}$$

$$n(HCl)_{\text{after}} = n(NaOH) = 7.375 \times 10^{-3}\text{ mol}$$

$$n(HCl)_{\text{orig}} = c \times V = 0.200 \times 0.0500 = 0.010\text{ mol}$$

$$(HCl)_{\text{used}} = 0.0100 - 7.375 \times 10^{-3} = 2.625 \times 10^{-3}\text{ mol}$$

$$2\,HCl\,(aq) + CO_3{}^{2-}\,(aq) \longrightarrow CO_2\,(g) + 2\,Cl^-\,(aq)$$

$$n(CO_3)^{2-} = \frac{1}{2} \times 2.625 \times 10^{-3} = 1.3125 \times 10^{-3}\text{ mol}$$

$$m(CO_3{}^{2-}) = 1.3125 \times 10^{-3} \times (12.01 + 3 \times 16.00) = 0.07876\text{ g}$$

$$\therefore \%(CO_3{}^{2-}) = \frac{0.07876}{0.145} \times 100 = 54.3\,\%$$

## 54. CHEMISTRY, M6 2021 HSC 32

→ Reaction 1 and reaction 2 are both neutralisation reactions between strong acids and strong bases. These reactions completely ionise in solution when added to water.

♦ Mean mark 44%.

→ Both reactions have the same net ionic equation:

$$H^+(aq) + OH^-\,(aq) \rightarrow H_2O(l)$$

→ Therefore, the enthalpy values obtained are the same for both reactions.

→ In reaction 3, $HCN$ is a weak acid that only partially ionises in an equilibrium reaction with water.

$$HCN(aq) + H_2O(l) \leftrightharpoons CN^-\,(aq) + H_3O^+(aq).$$

→ As the reaction continues, $HCN$ will further ionise as the equilibrium shifts to the right.

→ The bond-breaking is an endothermic process and thus will consume energy to break the bonds. As a result, the overall reaction is less exothermic than reaction 1 and reaction 2.

## 55. CHEMISTRY, M6 2022 HSC 32

♦ Mean mark (a) 48%.

**a.** $\mathrm{KHP\,(aq) + NaOH\,(aq) \longrightarrow NaKP\,(aq) + H_2O\,(l)}$

$$\begin{aligned}\mathrm{n(HX)} &= \frac{\mathrm{m}}{\mathrm{MM}}\\ &= \frac{4.989}{204.22}\\ &= 0.02443\ \mathrm{mol}\end{aligned}$$

$$[\mathrm{HX}] = \frac{n}{V} = \frac{0.02443}{0.1000} = 0.2443\ \mathrm{mol\ L^{-1}}$$

$$\begin{aligned}\mathrm{n(HX)\ titrated} &= \mathrm{c \times V}\\ &= 0.2443 \times 0.02500\\ &= 0.0006107\ \mathrm{mol}\end{aligned}$$

$$\Rightarrow \mathrm{n(NaOH)} = 0.0006107\ \mathrm{mol}$$

Eliminate the first trial because it is an outlier.

$$\mathrm{V_{avg}(NaOH)} = \frac{1}{3} \times (27.40 + 27.20 + 27.60) = 27.40\ \mathrm{mL}$$

$$[\mathrm{NaOH}] = \frac{\mathrm{n}}{\mathrm{V}} = \frac{6.107 \times 10^{-}3}{0.02740} = 0.2229\ \mathrm{mol\ L^{-1}}$$

$$\mathrm{H_3X\,(aq) + 3\,NaOH\,(aq) \longrightarrow Na_3X\,(aq) + 3\,H_2O\,(l)}$$

$$\begin{aligned}\mathrm{n(NaOH)\ titrated} &= \mathrm{c \times V}\\ &= 0.2229 \times 0.01310\\ &= 2.920 \times 10^{-3}\ \mathrm{mol}\end{aligned}$$

$$\mathrm{n(H_3X)} = \frac{1}{3} \times 2.920 \times 10^{-3} = 9.733 \times 10^{-4}\ \mathrm{mol}$$

$$[\mathrm{H_3X}]\ \mathrm{diluted} = \frac{\mathrm{n}}{\mathrm{V}} = \frac{9.733 \times 10^{-4}}{0.025} = 0.03893\ \mathrm{mol\ L^{-1}}$$

$$\begin{aligned}[\mathrm{H_3X}]\ \mathrm{0riginal} &= \frac{250.0}{75.00} \times 0.03893\\ &= 0.1298\ \mathrm{mol\ L^{-1}}\end{aligned}$$

Therefore, the concentration of citric acid in the soft drink is 0.1298 mol $L^{-1}$.

♦♦ Mean mark (b) 28%.

**b.** $CO_2$ can dissolve in water to produce $H_2CO_3$:

$$\mathrm{CO_2\,(g) + H_2O\,(l) \rightleftharpoons H_2CO_3\,(aq)}$$

This would enable NaOH to react with $H_2CO_3$:

$$\mathrm{2\,NaOH\,(aq) + H_2CO_3\,(aq) \longrightarrow Na_2CO_3\,(aq) + 2\,H_2O\,(l)}$$

→ Therefore, if $CO_2$ was not removed, more NaOH would be required to reach the endpoint.

→ This would result in a higher citric acid concentration calculation.

## 56. CHEMISTRY, M6 2021 HSC 35

$$V_{avg}\ (Na_2S_2O_3) = \frac{28.7 + 28.4 + 28.6}{3}$$

$$= 28.5666\ldots \text{ mL}$$

$$= 0.0285666\ldots \text{ L}$$

♦♦♦ Mean mark 39%.

→ The first titration is an outlier and so is excluded from the average.

$$n(Na_2S_2O_3) = c \times V = 0.900 \times 0.0285666\ldots = 0.02571 \text{ mol}$$

$$n(S_2O_3{}^{2-}) = n(Na_2S_2O_3) = 0.02571 \text{ mol}$$

$I_2$ and $S_2O_3{}^{2-}$ are in a $1:2$ ratio:

$$n(I_2) = \frac{1}{2} \times n(S_2O_3^{2-}) = \frac{1}{2} \times 0.02571 = 0.012855 \text{ mol}$$

Excess $Cr_2O_7{}^{2-}$ and $I_2$ are in $1:3$ ratio:

$$n(Cr_2O_7{}^{2-}) \text{ excess} = \frac{1}{3} \times n(I_2) = \frac{1}{3} \times 0.012855 = 0.004285 \text{ mol}$$

$$n(Cr_2O_7{}^{2-}) \text{ initial} = c \times V = 0.500 \times \frac{20}{1000} = 0.01 \text{ mol}$$

$n(Cr_2O_7{}^{2-})$ reacted with ethanol

$$= n(Cr_2O_7{}^{2-}) \text{ initial} - n(Cr_2O_7{}^{2-}) \text{ excess}$$

$$= 0.01 - 0.004285$$

$$= 0.005715 \text{ mol}$$

$$n(C_2H_5OH) = \frac{3}{2} \times n(Cr_2O_7{}^{2-}) = \frac{3}{2} \times 0.005715 = 0.0085725 \text{ mol}$$

$$m(C_2H_5OH) = n \times MM = 0.0085725 \times (2 \times 12.01 + 6 \times 1.008 + 16.00) = 0.3949 \text{ g}$$

→ Thus, 0.3949 g of ethanol is in a diluted 25 mL solution.

Find the mass of ethanol in the original solution:

$$m(C_2H_5OH) \text{ original} = 0.3949\ldots \times \frac{1000}{25} = 15.796\ldots \text{ g}$$

$$D = \frac{m}{V} \quad \Rightarrow \quad V = \frac{m}{D}$$

$$V(C_2H_5OH) = \frac{15.796}{0.789} = 20.021 \text{ mL}$$

$$\%\ (C_2H_5OH) = \frac{V(\text{ethanol})}{V(\text{sample})} = \frac{20.021}{25.0} = 80.08\ldots\% \text{ v/v}$$

→ Therefore, the product doesn't meet the manufacturer's requirement as the concentration is less than 85%.