Topic 4  Acids and Alkalis

Section A  Fill in the blanks

Unit 13  Looking at acids and alkalis

1  hydrochloric
2  sour
3  bases
4  ionize, ionization
5  hydrogen
6  mobile ions
7  basicity
8  monobasic
9  dibasic
10  Metal + dilute acid $\rightarrow$ salt + hydrogen
11  Carbonate + dilute acid $\rightarrow$ salt + water + carbon dioxide
12  Hydrogencarbonate + dilute acid $\rightarrow$ salt + water + carbon dioxide
13  Acid + metal hydroxide $\rightarrow$ salt + water
14  Acid + metal oxide $\rightarrow$ salt + water
15  salt, water
16  alkali
17  bitter
ammonia
precipitates
complex

Unit 14 The pH scale and strengths of acids and alkalis
indicator
litmus solution, methyl orange, phenolphthalein
acidity, alkalinity
acidic
alkaline
neutral
strong
weak

Unit 15 Neutralization
neutralization
sodium hydrogensulphate, sodium sulphate
normal
acid
Unit 16  Concentration of solutions and volumetric analysis

34  molarity

35  a)  electronic balance  
b)  volumetric flask  
c)  pipette, pipette filler  
d)  burette

36  indicator

Unit 17  Rate of reactions
Section B  True or False

Unit 13  Looking at acids and alkalis

1  T

2  F  An acid is a hydrogen-containing substance that gives hydrogen ions as the only positive ion when dissolved in water.

3  F  The basicity of an acid is the maximum number of hydrogen ions produced by one acid molecule. For example, ethanoic acid CH$_3$COOH is monobasic because only the hydrogen atom in the -COOH group is ionizable.

4  T

5  F  The properties of solutions of alkalis depend on the presence of mobile hydroxide ions.

6  T

7  F  Iron(II) hydroxide does not dissolve in excess ammonia solution.

8  F  Copper(II) hydroxide does not dissolve in excess dilute sodium hydroxide solution. It dissolves in excess ammonia solution to give a deep blue solution.

9  T

10  F  Concentrated nitric acid tends to decompose to nitrogen dioxide gas and oxygen gas.

\[ 4\text{HNO}_3(aq) \rightarrow 4\text{NO}_2(g) + \text{O}_2(g) + 2\text{H}_2\text{O}(l) \]
Unit 14  The pH scale and strengths of acids and alkalis

11  T

12  T

13  F  Phenolphthalein is colourless in acidic solutions.

14  F  Methyl orange is yellow in alkaline solutions.

15  T

16  F  The pH value of ammonia solution is greater than 7, but less than 14.

17  T

18  T

19  T

20  F  When ammonia gas dissolves in water, it reacts with water to give ammonium ions and hydroxide ions. However, ammonia does not react with water completely. Only very few hydroxide ions are formed.

Unit 15  Neutralization

21  T

22  F  In neutralization, salt and water are the only products. However, potassium carbonate and dilute hydrochloric acid react to give a salt, carbon dioxide gas and water. Therefore the reaction is not a neutralization reaction.

23  T
The reaction between dilute hydrochloric acid and dilute sodium hydroxide solution produces only one salt, sodium chloride. Sodium chloride is a normal salt.

Iron(II) chloride is a normal salt.

A few acid salts give alkaline solutions when dissolved in water.

Calcium sulphate, lead(II) sulphate and barium sulphate are insoluble in water.

Lithium carbonate, sodium carbonate, potassium carbonate and ammonium carbonate are soluble in water.

Farmers neutralize acidic soil by adding quicklime. Ammonium sulphate gives an acidic solution when dissolved in water.

Sodium hydroxide is highly corrosive. It is never used as an active ingredient in antacids.

**Unit 16** Concentration of solutions and volumetric analysis

To dilute 100 cm$^3$ of 1.0 M hydrochloric acid to 0.10 M, add water until the total volume of the solution is 1 000 cm$^3$.

A volumetric flask is used to prepare a solution of accurately known volume, e.g. 250.0 cm$^3$.

Potassium hydroxide absorbs moisture from the air and cannot be weighed accurately. It is
unsuitable for preparing a standard solution.

37    T

38    F    A solution of known molarity is called a standard solution.

39    T

40    F    The conical flask is to hold a specific volume of a solution (usually 25.0 cm$^3$), i.e. a specific amount of the solute. It should not be washed with the solution it is to contain before use because the additional amount of solute remaining in the flask will affect the titration results.

41    F    The last drop of the solution in the pipette should not be blown out.

42    T

Unit 17    Rate of reactions

43    T

44    F    The volume of a liquid would not affect the rate of a reaction.

45    T

46    F    The reaction rate between dilute hydrochloric acid and marble chips is slower than that between the acid and powdered marble.

47    F    Magnesium reacts with the hydrogen ions in the acid. Ethanoic acid is a weak acid and it has a lower concentration of hydrogen ions than the hydrochloric acid. Therefore, the reaction between magnesium and 2 M ethanoic acid is slower than that between magnesium and 2 M hydrochloric acid.

Section C    Multiple Choice
Unit 13  
Looking at acids and alkalis

1  A

2  C  
Option A – Caustic soda is sodium hydroxide.
Option B – Drain cleaners contain sodium hydroxide.
Option D – Slaked lime is calcium hydroxide.

3  D

4  D  
Option A – Dilute acids react with carbonates to give carbon dioxide gas.
Option B – Dilute acids have a sour taste.
Option C – Dilute acids react with reactive metals only.
Option D – Dilute acids conduct electricity due to the presence of mobile ions.

5  A  
Option B – Dilute sulphuric acid reacts with sodium hydrogencarbonate to give sodium sulphate.
\[
2\text{NaHCO}_3(\text{s or aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) + 2\text{CO}_2(\text{g})
\]
Option C – Dilute sulphuric acid reacts with calcium carbonate to give calcium sulphate.
Option D – Dilute sulphuric acid reacts with magnesium hydroxide to give magnesium sulphate and water.
\[
\text{H}_2\text{SO}_4(\text{aq}) + \text{Mg(OH)}_2(\text{s}) \rightarrow \text{MgSO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})
\]

6  C  
The reaction between dilute nitric acid and copper(II) carbonate can be represented by the following equation:
\[
\text{CuCO}_3(\text{s}) + 2\text{HNO}_3(\text{aq}) \rightarrow \text{Cu(NO}_3)_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})
\]
Copper(II) nitrate solution is blue in colour while carbon dioxide is a colourless gas.

7  A  
A car battery contains sulphuric acid. The acid reacts with marble (calcium carbonate) to give carbon dioxide gas.

8  C  
When dilute sulphuric acid reacts with calcium carbonate, insoluble calcium sulphate forms. The calcium sulphate covers the surface of calcium carbonate and prevents further reaction.
Zinc reacts with dilute nitric acid to give nitrogen monoxide, not hydrogen.

When citric acid crystals dissolve in water, the molecules ionize to give hydrogen ions $\text{H}^+(\text{aq})$. Hydrogen ions are responsible for all the acidic properties.

Phosphoric acid is a tribasic acid.

Glass cleaners usually contain ammonia.

Aluminium hydroxide redissolves in excess dilute sodium hydroxide solution due to the formation of a soluble complex salt.

Dilute sodium hydroxide solution does not give a white precipitate with potassium chloride solution. The mixture is a colourless solution.

The white precipitate is magnesium hydroxide.

<table>
<thead>
<tr>
<th>Solution</th>
<th>Colour of precipitate upon the addition of ammonia solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>Iron(II) nitrate</td>
<td>Green</td>
</tr>
<tr>
<td>Iron(III) sulphate</td>
<td>Reddish brown</td>
</tr>
<tr>
<td>Potassium nitrate</td>
<td>No precipitate</td>
</tr>
<tr>
<td>Zinc sulphate</td>
<td>White</td>
</tr>
</tbody>
</table>

Option A − $\text{CuSO}_4(\text{aq})$ gives a pale blue precipitate with dilute sodium hydroxide solution.

\[
\text{Cu}^{2+}(\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Cu(OH)}_2(\text{s})
\]

Option B − $\text{Pb(NO}_3)_2(\text{aq})$ gives a white precipitate with dilute sodium hydroxide solution.

\[
\text{Pb}^{2+}(\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Pb(OH)}_2(\text{s})
\]

Option D − $\text{Zn(NO}_3)_2(\text{aq})$ gives a white precipitate with dilute sodium hydroxide solution.
solution.

\[
\text{Zn}^{2+} \text{(aq)} + 2\text{OH}^- \text{(aq)} \rightarrow \text{Zn(OH)}_2(s)
\]

20 B FeSO₄(aq) and KOH(aq) give a green precipitate, iron(II) hydroxide, when mixed.

21 C Calcium nitrate solution and dilute sodium hydroxide solution will produce a white precipitate when mixed.

\[
\text{Ca}^{2+} \text{(aq)} + 2\text{OH}^- \text{(aq)} \rightarrow \text{Ca(OH)}_2(s)
\]

22 B

23 B Calcium reacts with dilute hydrochloric acid to give calcium ions.

\[
\text{Ca}(s) + 2\text{H}^+ \text{(aq)} \rightarrow \text{Ca}^{2+} \text{(aq)} + \text{H}_2(g)
\]

Calcium ions react with hydroxide ions in dilute sodium hydroxide solution to give a precipitate, calcium hydroxide.

\[
\text{Ca}^{2+} \text{(aq)} + 2\text{OH}^- \text{(aq)} \rightarrow \text{Ca(OH)}_2(s)
\]

Calcium hydroxide is insoluble in excess dilute sodium hydroxide solution.

24 A Ammonium carbonate gives off ammonia gas when heated with dilute sodium hydroxide solution. It gives carbon dioxide gas with dilute nitric acid.

25 B Dilute hydrochloric acid reacts with sodium carbonate to give carbon dioxide gas, but not potassium nitrate.

26 D

27 B

28 A (3) Dilute hydrochloric acid is irritant.

29 D (1) Magnesium reacts with dilute hydrochloric acid to give hydrogen gas.

(2) & (3) Sodium carbonate and potassium hydrogen carbonate react with dilute hydrochloric acid to give carbon dioxide gas.
Zinc reacts with dilute sulphuric acid to give hydrogen gas.

Magnesium carbonate and dilute sulphuric acid react to give carbon dioxide gas.

Magnesium oxide and dilute sulphuric acid react to give magnesium sulphate and water. No gas is formed.

Very dilute nitric acid acts like a typical acid. It does not react with unreactive metals, such as copper.

Calcium carbonate solid and dilute nitric acid react to give carbon dioxide gas.

There is no reaction between silver and dilute hydrochloric acid.

Dilute sodium hydroxide solution and dilute sulphuric acid react to give sodium sulphate and water.

Magnesium gives gas bubbles (hydrogen) with dilute hydrochloric acid. There is no reaction between copper and the acid.

Sodium chloride has no observable change when mixed with dilute hydrochloric acid. Sodium carbonate gives gas bubbles (carbon dioxide gas) with dilute hydrochloric acid.

Iron gives gas bubbles (hydrogen gas) with dilute hydrochloric acid. Copper(II) oxide gives a coloured solution, without gas bubbles, with dilute hydrochloric acid.

Carbon burns in air to give carbon dioxide. Carbon dioxide gives an acidic solution when dissolved in water.

Solid citric acid crystals cannot show the typical properties of an acid.
37 C (1) Ammonium chloride solution and dilute sodium hydroxide solution do not give a precipitate when mixed.

(2) Magnesium sulphate solution and dilute sodium hydroxide solution give a white precipitate when mixed.

\[
\text{Mg}^{2+}(\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Mg(OH)}_2(\text{s})
\]

(3) Iron(II) nitrate solution and dilute sodium hydroxide solution give a green precipitate when mixed.

\[
\text{Fe}^{2+}(\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Fe(OH)}_2(\text{s})
\]

38 B (1) Cu(OH)_2 dissolves in excess ammonia solution to give a deep blue solution.

(2) Pb(OH)_2 is insoluble in excess ammonia solution.

(3) Zn(OH)_2 dissolves in excess ammonia solution to give a colourless solution.

39 C (1) Ca(NO_3)_2(aq) does not give a precipitate when mixed with NH_3(aq).

(2) Pb(NO_3)_2(aq) gives a white precipitate when mixed with NaOH(aq).

\[
\text{Pb}^{2+}(\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Pb(OH)}_2(\text{s})
\]

(3) Mg(NO_3)_2(aq) gives a white precipitate when mixed with NaOH(aq).

\[
\text{Mg}^{2+}(\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Mg(OH)}_2(\text{s})
\]

40 D (1) Heating ammonium chloride solution with dilute sodium hydroxide solution liberates ammonia.

\[
\text{NH}_4^+(\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l})
\]

(2) Iron(III) sulphate solution and dilute sodium hydroxide solution give a reddish brown precipitate when mixed.

(3) Vinegar and dilute sodium hydroxide solution react to give a salt and water.

41 A (2) & (3) Both silver and copper do not react with dilute hydrochloric acid.
Copper(II) oxide reacts with dilute sulphuric acid to give copper(II) sulphate and water.

\[
\text{CuO(s)} + \text{H}_2\text{SO}_4(aq) \rightarrow \text{CuSO}_4(aq) + \text{H}_2\text{O(l)}
\]

Iron reacts with dilute hydrochloric acid to give iron(II) chloride and hydrogen.

\[
\text{Fe(s)} + 2\text{HCl(aq)} \rightarrow \text{FeCl}_2(aq) + \text{H}_2(g)
\]

Magnesium nitrate solution reacts with ammonia solution to give a white precipitate, magnesium hydroxide.

\[
\text{Mg}^{2+} (aq) + 2\text{OH}^- (aq) \rightarrow \text{Mg(OH)}_2(s)
\]

Ordinary concentrated nitric acid contains about 70% nitric acid by mass.

Concentrated sulphuric acid does not give an acid mist in air.

Ordinary concentrated sulphuric acid contains about 98% sulphuric acid by mass.

A fruit salt contains a solid acid and a solid metal carbonate or hydrogencarbonate. The solid acid gives hydrogen ions H\(^+\)(aq) when dissolved in water. The hydrogen ions react with the metal carbonate or hydrogencarbonate to give carbon dioxide gas. This gives rise to the effervescence that happens when the fruit salt is added to water.

The chemical equation for the decomposition of concentrated nitric acid:

\[
4\text{HNO}_3(aq) \rightarrow 4\text{NO}_2(g) + \text{O}_2(g) + 2\text{H}_2\text{O(l)}
\]

Many metal hydroxides are insoluble in water.

When ammonia gas dissolves in water, it reacts with water to give ammonium ions and hydroxide ions.

\[
\text{NH}_3(g) + \text{H}_2\text{O(l)} \rightarrow \text{NH}_4^+ (aq) + \text{OH}^- (aq)
\]

When dilute sulphuric acid reacts with calcium carbonate, insoluble calcium sulphate forms. The calcium sulphate covers the surface of calcium carbonate and prevents further reaction.
Unit 14  The pH scale and strengths of acids and alkalis

51  A

<table>
<thead>
<tr>
<th>Indicator</th>
<th>Colour in Acidic solution</th>
<th>Colour in Alkaline solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methyl orange</td>
<td>Red</td>
<td>Yellow</td>
</tr>
</tbody>
</table>

52  B

<table>
<thead>
<tr>
<th>Indicator</th>
<th>Colour in Acidic solution</th>
<th>Colour in Alkaline solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>Phenolphthalein</td>
<td>Colourless</td>
<td>Red</td>
</tr>
</tbody>
</table>

53  D

54  C

55  D

56  B  Slaked lime is calcium hydroxide.

57  D  Soft drinks contain carbonic acid.

58  D  Tomato juice contains vinegar (ethanoic acid).

59  B  Normally, the pH value of rainwater is about 5.6. It is slightly acidic because carbon dioxide in the air reacts with rainwater to form carbonic acid.

60  C

<table>
<thead>
<tr>
<th>Substance</th>
<th>pH value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Glass cleaner</td>
<td>~ 10</td>
</tr>
<tr>
<td>Lemon juice</td>
<td>~ 2</td>
</tr>
<tr>
<td>Milk</td>
<td>&lt; 7</td>
</tr>
</tbody>
</table>

61  D

<table>
<thead>
<tr>
<th>Substance</th>
<th>pH value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Dilute nitric acid</td>
<td>~ 1</td>
</tr>
<tr>
<td>Orange juice</td>
<td>~ 3</td>
</tr>
<tr>
<td>Window cleaner</td>
<td>~ 10</td>
</tr>
</tbody>
</table>
64  C  Option A – The pH value of calcium hydroxide solution is greater than 7.
Option B & C – Hydrochloric acid is a strong acid while ethanoic acid is a weak acid. The hydrochloric acid has a higher concentration of hydrogen ions than the ethanoic acid.
Therefore the pH value of the hydrochloric acid is lower than that of the ethanoic acid.

Option D – The pH value of sodium chloride solution is 7.

65  B  (1) & (2)  The pH values of dilute carbonic acid and dilute sulphuric acid are both less than 7. Dilute sulphuric acid is a strong acid while carbonic acid is a weak acid. The sulphuric acid has a higher concentration of hydrogen ions than the carbonic acid. Therefore the pH value of the sulphuric acid is lower than that of the carbonic acid.

(3)  The pH value of dilute sodium hydroxide solution is greater than 7.

66  D  (1)  The pH value of limewater is greater than 7.
(2)  The pH value of normal rainwater is about 5.6.

67  A

<table>
<thead>
<tr>
<th>Substance</th>
<th>pH value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ammonia solution</td>
<td>~ 10</td>
</tr>
<tr>
<td>Glass cleaner</td>
<td>~ 10</td>
</tr>
<tr>
<td>Sodium chloride solution</td>
<td>7</td>
</tr>
</tbody>
</table>

68  C  (1)  Litmus solution can be used to identify whether a solution is acidic or alkaline. It cannot be used to measure the pH value of a solution.

69  B  (2)  When ammonia gas dissolves in water, it reacts with water to give ammonium ions and hydroxide ions.
Ammonia solution reacts with an acid to give a salt and water.

Adding 1 M hydrochloric acid to the ethanoic acid makes the ethanoic acid more acidic, thus decreases the pH value of the ethanoic acid.

Solid sodium carbonate and magnesium react with the ethanoic acid, thus both increase the pH value of the ethanoic acid.

Solid citric acid and ethanoic acid react with the ammonia solution, thus both decrease the pH value of the ammonia solution.

Adding concentrated sodium hydroxide solution makes the ammonia solution more alkaline, thus increases the pH value of the ammonia solution.

The pH value is a measure of the concentration of hydrogen ions in a solution, not the concentration of an acid.

Hydrochloric acid is a strong acid while ethanoic acid is a weak acid. Therefore acid solutions of the same pH value should have the same concentration of hydrogen ions.

The solution of a strong alkali conducts electricity due to the presence of mobile ions.

Sodium hydroxide ionizes almost completely into sodium ions and hydroxide ions when dissolved in water. The solution contains mainly sodium ions Na⁺ (aq), hydroxide ions OH⁻ (aq) and water molecules.

Hydrochloric acid is a strong acid while ethanoic acid is a weak acid. Thus hydrochloric acid has a higher concentration of hydrogen ions than ethanoic acid.
the same concentration. The pH value of the hydrochloric acid is lower than that of the ethanoic acid.

(3) The hydrochloric acid has a higher concentration of ions than the ethanoic acid.

Therefore the electrical conductivity of the hydrochloric acid is higher than that of the ethanoic acid.

77 B

78 A

79 B Sulphuric acid is a strong acid because it almost completely ionizes in water. Ethanoic acid is a weak acid because it only partly ionizes in water.

80 B Concentrated hydrochloric acid is a strong acid because it almost completely ionizes in water.

Unit 15 Neutralization

81 B Slaked lime (calcium hydroxide) is an alkali.

82 A Carbon dioxide dissolves in water to give carbonic acid.

83 C Neutralization is the combination of hydrogen ions and hydroxide ions (or oxide ions) to form water molecules.

84 D Copper(II) carbonate reacts with dilute hydrochloric acid to give copper(II) chloride, carbon dioxide and water. This is not a neutralization reaction. In neutralization, salt and water are the only products.

85 B Iron(III) hydroxide reacts with dilute hydrochloric acid to give iron(III) chloride and water.

The resulting solution is yellow in colour.

\[ \text{Fe(OH)}_3(s) + 3\text{HCl(aq)} \rightarrow \text{FeCl}_3(aq) + 3\text{H}_2\text{O(l)} \]
86 D Only Zn(OH)\(_2\) dissolves in excess dilute sodium hydroxide solution due to the formation of a soluble complex salt.

87 C Options A & D − Carbonates react with dilute hydrochloric acid to give carbon dioxide gas.

Option B − Copper(II) hydroxide does not dissolve in excess dilute sodium hydroxide solution.

Option C − Lead(II) hydroxide dissolves in excess dilute sodium hydroxide solution due to the formation of a soluble complex salt.

88 B Copper(II) oxide reacts with dilute sulphuric acid to give copper(II) sulphate and water.

\[
\text{CuO(s) + H}_2\text{SO}_4(aq) \rightarrow \text{CuSO}_4(aq) + \text{H}_2\text{O(l)}
\]

Copper(II) sulphate is soluble in water while copper is insoluble. Therefore copper can be separated from the mixture by filtration.

89 B In the two reactions, the only chemical change is OH\(^-\) (aq) + H\(^+\) (aq) \rightarrow H\(_2\)O(l). Therefore the energy changes should be the same.

90 C Option A − Calcium carbonate gives gas bubbles with dilute sulphuric acid. However, it is insoluble in water.

Option B − Magnesium chloride does not give gas bubbles with dilute sulphuric acid.

Option C − Sodium carbonate is soluble in water and gives gas bubbles (carbon dioxide) with dilute sulphuric acid.

Option D − Zinc oxide does not give gas bubbles with dilute sulphuric acid.

91 D Option A − Calcium sulphate is insoluble in water.

Option B − Copper(II) oxide is not white in colour. It is also insoluble in water.

Option C − Iron(II) chloride is not white in colour.

Option D − Neutralization occurs when potassium hydroxide is mixed with dilute sulphuric
92 B Copper(II) oxide is black in colour. When copper(II) oxide is mixed with dilute sulphuric acid, a blue solution results.

\[ \text{CuO(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CuSO}_4(\text{aq}) + \text{H}_2\text{O(l)} \]

93 A

94 B

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Salt obtained</th>
</tr>
</thead>
<tbody>
<tr>
<td>Calcium carbonate + dilute sulphuric acid</td>
<td>When dilute sulphuric acid reacts with calcium carbonate, insoluble calcium sulphate forms. The calcium sulphate covers the surface of calcium carbonate and prevents further reaction.</td>
</tr>
<tr>
<td>Magnesium oxide + dilute sulphuric acid</td>
<td>Magnesium sulphate</td>
</tr>
<tr>
<td>Copper(II) hydroxide + dilute nitric acid</td>
<td>Copper(II) nitrate</td>
</tr>
<tr>
<td>Sodium hydroxide + dilute hydrochloric acid</td>
<td>Sodium chloride</td>
</tr>
</tbody>
</table>

95 A

96 A

97 C

<table>
<thead>
<tr>
<th>Option</th>
<th>Salt</th>
<th>Solubility in water</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>Calcium carbonate</td>
<td>Insoluble</td>
</tr>
<tr>
<td></td>
<td>Calcium nitrate</td>
<td>Soluble</td>
</tr>
<tr>
<td></td>
<td>Ammonium sulphate</td>
<td>Soluble</td>
</tr>
<tr>
<td>B</td>
<td>Barium sulphate</td>
<td>Insoluble</td>
</tr>
<tr>
<td></td>
<td>Potassium carbonate</td>
<td>Soluble</td>
</tr>
<tr>
<td></td>
<td>Zinc nitrate</td>
<td>Soluble</td>
</tr>
<tr>
<td>C</td>
<td>Magnesium carbonate</td>
<td>Insoluble</td>
</tr>
<tr>
<td></td>
<td>Ammonium nitrate</td>
<td>Soluble</td>
</tr>
<tr>
<td></td>
<td>Copper(II)sulphate</td>
<td>Soluble</td>
</tr>
</tbody>
</table>

98 D \[ \text{Pb}^{2+} (\text{aq}) + 2\text{Cl}^- (\text{aq}) \rightarrow \text{PbCl}_2(\text{s}) \]

white precipitate

99 D
Copper(II) sulphate is soluble in water while calcium sulphate is insoluble. Adding warm water to the mixture dissolves copper(II) sulphate but not calcium sulphate. Calcium sulphate can be separated from the copper(II) sulphate solution by filtration. Finally, evaporate the filtrate to obtain copper(II) sulphate crystals.

Magnesium chloride is white in colour. It can be prepared by the reaction between magnesium and dilute hydrochloric acid.

Na₂SO₄ can be prepared by the reaction between dilute sulphuric acid and dilute sodium hydroxide solution.

\[ \text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \]

Magnesium carbonate reacts with dilute nitric acid to give magnesium nitrate, carbon dioxide and water. This is not a neutralization reaction. In neutralization, salt and water are the only products.

Heat energy is given out in neutralization reactions.
The product is sodium nitrate, a normal salt.
Sodium nitrate is an ionic compound.

ZnCO₃ reacts with dilute hydrochloric acid to give a salt, carbon dioxide and water.

Adding excess ammonia solution will leave an alkaline solution after reacting with the acid.
Adding excess calcium carbonate will not leave an alkaline solution after reacting with the acid because calcium carbonate is insoluble in water.
Sodium chloride does not react with the acid.

Ammonium chloride is soluble in water. No precipitate is formed in the reaction.
between ammonia solution and dilute hydrochloric acid.

110  B  (2)  Sodium hydroxide is not a salt.
     (3)  Sodium hydrogencarbonate is an acid salt formed from carbonic acid.

111  C  (1)  The salt formed in the reaction is sodium chloride. It is soluble in water. Therefore no precipitate appears.
     (3)  Neutralization occurs in the beaker and heat energy is given out. Therefore the solution warms up.

112  D  Zinc sulphate can be prepared by reacting dilute sulphuric acid with either zinc, zinc oxide or zinc carbonate.

\[
\begin{align*}
\text{Zn(s)} & + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2(\text{g}) \\
\text{ZnO(s)} & + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) \\
\text{ZnCO}_3(\text{s}) & + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{ZnSO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})
\end{align*}
\]

113  D  Only the hydrogen atom in the -COOH group of the ethanoic acid is ionizable. Ethanoic acid is a monobasic acid. It cannot form acid salt.

114  C  Sodium hydroxide is corrosive and thus never used in antacids.

115  C  The student should wash the affected area with plenty of water.

116  B

117  C  Some salts are insoluble in water.

**Unit 16  Concentration of solutions and volumetric analysis**

118  C  Molarity of sodium carbonate solution =
\[
\frac{\text{Number of moles of } \text{Na}_2\text{CO}_3}{\text{Volume of solution}}
\]
\[ 0.20 \text{ mol dm}^{-3} = \frac{\text{Number of moles of } \text{Na}_2\text{CO}_3}{(250.0 / 1000)\text{dm}^3} \]

Number of moles of \( \text{Na}_2\text{CO}_3 \) = 0.20 mol dm\(^{-3} \) \( \times \) \( \frac{250.0}{1000} \) dm\(^3\)

= 0.050 mol

Molar mass of \( \text{Na}_2\text{CO}_3 \) = \( (2 \times 23.0 + 12.0 + 3 \times 16.0) \) g mol\(^{-1}\)

= 106.0 g mol\(^{-1}\)

\[ \therefore \text{ Mass of } \text{Na}_2\text{CO}_3 \text{ required} = \text{Number of moles of } \text{Na}_2\text{CO}_3 \times \text{Molar mass of } \text{Na}_2\text{CO}_3 \]

= 0.050 mol \( \times \) 106.0 g mol\(^{-1}\)

= 5.3 g

119 \hspace{1cm} B \hspace{1cm} \text{Molarity of ethanoic acid solution} = \frac{\text{Number of moles of } \text{CH}_3\text{COOH}}{\text{Volume of solution}}

\[ 0.50 \text{ mol dm}^{-3} = \frac{\text{Number of moles of } \text{CH}_3\text{COOH}}{(500.0 / 1000)\text{dm}^3} \]

Number of moles of \( \text{CH}_3\text{COOH} \) = 0.50 mol dm\(^{-3} \) \( \times \) \( \frac{500.0}{1000} \) dm\(^3\)

= 0.25 mol

Molar mass of \( \text{CH}_3\text{COOH} \) = \( (2 \times 12.0 + 4 \times 1.0 + 2 \times 16.0) \) g mol\(^{-1}\)

= 60.0 g mol\(^{-1}\)

\[ \therefore \text{ Mass of } \text{CH}_3\text{COOH} \text{ required} = \text{Number of moles of } \text{CH}_3\text{COOH} \times \text{Molar mass of } \text{CH}_3\text{COOH} \]

= 0.25 mol \( \times \) 60.0 g mol\(^{-1}\)

= 15 g

120 \hspace{1cm} D \hspace{1cm} \text{Molar mass of } (\text{COOH})_2\*\text{H}_2\text{O} = [2 \times (12.0 + 2 \times 16.0 + 1.0) + 2 \times (2 \times 1.0 + 16.0)] \text{ g mol}^{-1}
Number of moles of (COOH)$_2$$\cdot$H$_2$O present = \(\frac{\text{Mass}}{\text{Molar mass}}\)

= \(\frac{3.15\text{g}}{126.0\text{g mol}^{-1}}\)

= 0.0250 mol

Molarity of ethanedioic acid solution =

\[\frac{\text{Number of moles of (COOH)$_2$$\cdot$H$_2$O}}{\text{Volume of solution}}\]

= \(\frac{0.0250\text{mol}}{(250.0 \div 1000)\text{dm}^3}\)

= 0.100 mol dm$^{-3}$ (M)

121 B Molar mass of ZnSO$_4$$\cdot$H$_2$O = [65.0 + 32.0 + 4 x 16.0 + 7 x (2 x 1.0 + 16.0)] g mol$^{-1}$

= 287.0 g mol$^{-1}$

Number of moles of ZnSO$_4$$\cdot$H$_2$O present = \(\frac{\text{Mass}}{\text{Molar mass}}\)

= \(\frac{23\text{g}}{287.0\text{g mol}^{-1}}\)

= 0.080 mol

Molarity of zinc sulphate solution =

\[\frac{\text{Number of moles of ZnSO$_4$$\cdot$H$_2$O}}{\text{Volume of solution}}\]

= \(\frac{0.080\text{mol}}{(500.0 \div 1000)\text{dm}^3}\)

= 0.16 mol dm$^{-3}$ (M)

122 C Molar mass of KCl = (39.0 + 35.5) g mol$^{-1}$

= 74.5 g mol$^{-1}$
Number of moles of KCl present = \( \frac{\text{Mass}}{\text{Molar mass}} \)

= \( \frac{44.7 \text{ g}}{74.5 \text{ g mol}^{-1}} \)

= 0.600 mol

Molarity of potassium chloride solution = \( \frac{\text{Number of moles of KCl}}{\text{Volume of solution}} \)

\[ 2.40 \text{ mol dm}^{-3} = \frac{0.600 \text{ mol}}{\text{Volume of solution}} \]

Volume of solution = \( \frac{0.600 \text{ mol}}{2.40 \text{ mol dm}^{-3}} \)

= 0.250 dm\(^3\)

= 250 cm\(^3\)

123 A Molar mass of KOH = \((39.0 + 16.0 +1.0) \text{ g mol}^{-1}\)

= 56.0 g mol\(^{-1}\)

Number of moles of KOH present = \( \frac{\text{Mass}}{\text{Molar mass}} \)

= \( \frac{7.00 \text{ g}}{56.0 \text{ g mol}^{-1}} \)

= 0.125 mol

Molarity of potassium hydroxide solution = \( \frac{\text{Number of moles of KOH}}{\text{Volume of solution}} \)

\[ 1.00 \text{ mol dm}^{-3} = \frac{0.125 \text{ mol}}{\text{Volume of solution}} \]

Volume of solution = \( \frac{0.125 \text{ mol}}{1.00 \text{ mol dm}^{-3}} \)

= 0.125 dm\(^3\)
D  Molarity of sodium carbonate solution = 
\[
\frac{\text{Number of moles of } \text{Na}_2\text{CO}_3}{\text{Volume of solution}}
\]

\[
2.5 \text{ mol dm}^{-3} = \frac{\text{Number of moles of } \text{Na}_2\text{CO}_3}{(200.0 / 1000) \text{ dm}^3}
\]

Number of moles of Na\textsubscript{2}CO\textsubscript{3} = 2.5 \text{ mol dm}^{-3} \times \frac{200.0}{1000} \text{ dm}^3

= 0.50 \text{ mol}

### Table

<table>
<thead>
<tr>
<th>Substance</th>
<th>Number of moles of substance present</th>
<th>Number of moles of ions in one mole of substance</th>
<th>Number of moles of ions present</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl</td>
<td>0.10 mol dm\textsuperscript{-3} \times \frac{100}{1000} \text{ dm}^3 = 0.010 mol</td>
<td>2</td>
<td>2 \times 0.010 mol = 0.020 mol</td>
</tr>
<tr>
<td>Fe\textsubscript{2}(SO\textsubscript{4})\textsubscript{3}</td>
<td>0.050 mol dm\textsuperscript{-3} \times \frac{50}{1000} \text{ dm}^3 = 0.0025 mol</td>
<td>5</td>
<td>5 \times 0.0025 mol = 0.013 mol</td>
</tr>
<tr>
<td>CaCl\textsubscript{2}</td>
<td>0.080 mol dm\textsuperscript{-3} \times \frac{50}{1000} \text{ dm}^3 = 0.0040 mol</td>
<td>3</td>
<td>3 \times 0.0040 mol = 0.012 mol</td>
</tr>
<tr>
<td>MgSO\textsubscript{4}</td>
<td>0.080 mol dm\textsuperscript{-3} \times \frac{100}{1000} \text{ dm}^3 = 0.0080 mol</td>
<td>2</td>
<td>2 \times 0.0080 mol = 0.016 mol</td>
</tr>
</tbody>
</table>

100 cm\textsuperscript{3} of 0.10 M NaCl contain the greatest number of moles of ions, i.e. the greatest number of ions.

B  \( (\text{MV}) \text{ before dilution} = (\text{MV}) \text{ after dilution} \) where \( M = \text{Molarity}, \ V = \text{Volume} \)

\[
5.0 \times \frac{V}{1000} = 0.40 \times \frac{250}{1000}
\]

\[
V = 20 \text{ cm}^3
\]

C  \( (\text{MV}) \text{ before dilution} = (\text{MV}) \text{ after dilution} \) where \( M = \text{Molarity}, \ V = \text{Volume} \)
\[10 \times \frac{120}{1000} = 2 \times \frac{V}{1000}\]

\[V = 600\]

Volume of the final solution = 600 cm\(^3\)

\[\therefore \text{Volume of water added} = (600 - 120) \text{ cm}^3\]

\[= 480 \text{ cm}^3\]

128 B \[
(MV) \text{ before dilution} = (MV) \text{ after dilution where } M = \text{Molarity, } V = \text{Volume}
\]

\[0.50 \times \frac{100}{1000} = 0.10 \times \frac{V}{1000}\]

\[V = 500 \text{ cm}^3\]

Volume of the final solution = 500 cm\(^3\)

\[\therefore \text{Volume of water added} = (500 - 100) \text{ cm}^3\]

\[= 400 \text{ cm}^3\]

129 B Number of moles of \(\text{Na}_2\text{CO}_3\) in 200.0 cm\(^3\) of 2.50 M solution

= Molarity of solution \(\times\) Volume of solution

\[= 2.50 \text{ mol dm}^{-3} \times \frac{200.0}{1000} \text{ dm}^3\]

\[= 0.500 \text{ mol}\]

Number of moles of \(\text{Na}_2\text{CO}_3\) in 50.0 cm\(^3\) of 1.00 M solution

= Molarity of solution \(\times\) Volume of solution

\[= 1.00 \text{ mol dm}^{-3} \times \frac{50.0}{1000} \text{ dm}^3\]

\[= 0.0500 \text{ mol}\]

Total number of moles of \(\text{Na}_2\text{CO}_3\) in the resulting solution

= (0.500 + 0.0500) mol
\( = 0.550 \text{ mol} \)

Total volume of the resulting solution
\( = (200.0 + 50.0) \text{ cm}^3 \)
\( = 250.0 \text{ cm}^3 \)

Molarity of the resulting solution
\[
\text{Molarity} = \frac{\text{Number of moles of } \text{Na}_2\text{CO}_3}{\text{Volume of solution}}
\]
\[
= \frac{0.550 \text{ mol}}{(250.0 / 1000) \text{ dm}^3}
\]
\( = 2.20 \text{ mol dm}^{-3} \text{ (M)} \)

130  D  Carbonic acid and ethanoic acid are weak acids. They only partly ionize in water. Hydrochloric acid and sulphuric acid are strong acids. They almost completely ionize in water. Therefore 1 M hydrochloric acid and 1 M sulphuric acid have a higher concentration of hydrogen ions than the two weak acids.

Sulphuric acid is a dibasic acid while hydrochloric acid is a monobasic acid. Every sulphuric acid molecule give two hydrogen ions when dissolved in water. Every hydrogen chloride molecule gives one hydrogen ion when dissolved in water. Therefore 1 M sulphuric acid has the highest concentration of hydrogen ions among the four acids.

131  A

132  C

133  D

134  A

135  D  \( \text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)} \)
\begin{align*}
0.20 \text{ M} & \quad ? \text{ M} \\
27.5 \text{ cm}^3 & \quad 25.0 \text{ cm}^3
\end{align*}
Number of moles of HCl in 27.5 cm$^3$ solution = Molarity of solution x Volume of solution

\[
= 0.20 \text{ mol dm}^{-3} \times \frac{27.5}{1000} \text{ dm}^3
\]

= 0.0055 mol

According to the equation, 1 mole of HCl requires 1 mole of NaOH for complete neutralization.

i.e. Number of moles of NaOH = 0.0055 mol

Molarity of sodium hydroxide solution = \( \frac{\text{Number of moles of NaOH}}{\text{Volume of solution}} \)

\[
= \frac{0.0055 \text{ mol}}{(25.0 / 1000) \text{ dm}^3} = 0.22 \text{ mol dm}^{-3} \text{ (M)}
\]

136 C \( \text{H}_2\text{SO}_4(aq) + 2\text{KOH}(aq) \rightarrow \text{K}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}(l) \)

0.150 M ? M
25.0 cm$^3$ 37.5 cm$^3$

Number of moles of H$_2$SO$_4$ in 25.0 cm$^3$ solution = Molarity of solution x Volume of solution

\[
= 0.150 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}^3
\]

= 0.00375 mol

According to the equation, 1 mole of H$_2$SO$_4$ requires 2 moles of KOH for complete neutralization.

i.e. Number of moles of KOH = 2 x Number of moles of H$_2$SO$_4$

= 2 x 0.00375 mol

= 0.00750 mol
Molarity of potassium hydroxide solution = \frac{\text{Number of moles of KOH}}{\text{Volume of solution}}

= \frac{0.00750 \text{ mol}}{(37.5 / 1 \text{ 000}) \text{ dm}^3} \\
= 0.200 \text{ mol dm}^{-3} \text{ (M)}

The burette should be washed with distilled water and then the acid because any water or impurities in the apparatus will change the concentration of the acid it is to contain. This will affect the titration results.

The conical flask is to hold 25.0 cm$^3$ of the sodium hydroxide solution. It should not be washed with the solution because the additional amount of solute remaining in the flask will affect the titration results.

We can represent the dibasic acid solution by H$_2$X(aq).

\[ \text{H}_2\text{X(aq)} + 2\text{NaOH(aq)} \rightarrow \text{Na}_2\text{X(aq)} + 2\text{H}_2\text{O(l)} \]

\[ \text{? M} \quad 0.20 \text{ M} \]

\[ 12.5 \text{ cm}^3 \quad 25.0 \text{ cm}^3 \]

Number of moles of NaOH in 25.0 cm$^3$ solution = Molarity of solution \times Volume of solution

\[ = 0.20 \text{ mol dm}^{-3} \times \frac{25.0}{1 \text{ 000}} \text{ dm}^3 \]

\[ = 0.0050 \text{ mol} \]

According to the equation, 1 mole of H$_2$X requires 2 moles of NaOH for complete neutralization.

i.e. \[ \text{Number of moles of H}_2\text{X} = \frac{1}{2} \times \text{Number of moles of NaOH} \]
\[
= \frac{1}{2} \times 0.0050 \text{ mol}
\]

\[= 0.0025 \text{ mol}\]

Molarity of dibasic acid solution = \(\frac{\text{Number of moles of } H_2X}{\text{Volume of solution}}\)

\[= \frac{0.0025 \text{ mol}}{(12.5 / 1000) \text{ dm}^3}\]

\[= 0.20 \text{ mol dm}^{-3} (M)\]

141  D  We can represent the monobasic acid solution by HX(aq).

\[
\text{HX(aq)} + \text{NaOH(aq)} \rightarrow \text{NaX(aq)} + \text{H}_2\text{O(l)}
\]

0.40 M  \(?M\)

27.5 cm\(^3\)  25.0 cm\(^3\)

Number of moles of HX in 27.5 cm\(^3\) solution = Molarity of solution \(\times\) Volume of solution

\[= 0.40 \text{ mol dm}^{-3} \times \frac{27.5}{1000} \text{ dm}^3\]

\[= 0.011 \text{ mol}\]

According to the equation, 1 mole of HX requires 1 mole of NaOH for complete neutralization.

i.e. \(\text{Number of moles of NaOH} = 0.011 \text{ mol}\)

Molarity of sodium hydroxide solution = \(\frac{\text{Number of moles of NaOH}}{\text{Volume of solution}}\)

\[= \frac{0.0011 \text{ mol}}{(25.0 / 1000) \text{ dm}^3}\]

\[= 0.44 \text{ mol dm}^{-3} (M)\]
We can represent the dibasic acid solution by $\text{H}_2\text{X}(\text{aq})$.

$$\text{H}_2\text{X}(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{X}(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$$

<table>
<thead>
<tr>
<th>Molarity</th>
<th>Molarity</th>
<th>Volume</th>
<th>Volume of NaOH solution required for complete neutralization</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.0600 M</td>
<td>0.150 M</td>
<td>25.0 cm$^3$</td>
<td>20.0 cm$^3$</td>
</tr>
</tbody>
</table>

Number of moles of $\text{H}_2\text{X}$ in 25.0 cm$^3$ solution = Molarity of solution x Volume of solution

$$= 0.0600 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}^3$$

$$= 0.00150 \text{ mol}$$

According to the equation, 1 mole of $\text{H}_2\text{X}$ requires 2 moles of NaOH for complete neutralization.

i.e. Number of moles of NaOH = 2 x Number of moles of $\text{H}_2\text{X}$

$$= 2 \times 0.00150 \text{ mol}$$

$$= 0.00300 \text{ mol}$$

Volume of NaOH solution required for complete neutralization

$$= \frac{\text{Number of moles of NaOH}}{\text{Molarity of solution}}$$

$$= \frac{0.00300 \text{ mol}}{0.150 \text{ mol dm}^{-3}}$$

$$= 0.0200 \text{ dm}^3$$

$$= 20.0 \text{ cm}^3$$

$\text{H}_2\text{SO}_4(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$

According to the equation, 1 mole of $\text{H}_2\text{SO}_4$ requires 1 mole of $\text{Na}_2\text{CO}_3$ for complete reaction.

i.e. Number of mole of $\text{Na}_2\text{CO}_3 = 0.010 \text{ mol}$

$$2\text{HCl} + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$$
0.010 mol

According to the equation, 2 moles of HCl require 1 mole of Na$_2$CO$_3$ for complete reaction.

i.e. Number of moles of Na$_2$CO$_3$ = \( \frac{1}{2} \) x Number of moles of HCl

\[
= \frac{1}{2} \times 0.010 \text{ mol}
\]

\[
= 0.0050 \text{ mol}
\]

Total number of moles of Na$_2$CO$_3$ required to react with the acids
\[
= 0.010 \text{ mol} + 0.0050 \text{ mol}
\]
\[
= 0.015 \text{ mol}
\]

Total volume of Na$_2$CO$_3$ solution required to react with the acids

\[
= \frac{\text{Number of moles of Na}_2\text{CO}_3}{\text{Molarity of solution}}
\]

\[
= \frac{0.015 \text{ mol}}{0.20 \text{ mol dm}^{-3}}
\]
\[
= 0.075 \text{ dm}^3
\]
\[
= 75 \text{ cm}^3
\]

144 D

145 C The volumes of 0.10 M sodium hydroxide solution required for neutralization are: 22.0 cm$^3$, 20.0 cm$^3$, 19.9 cm$^3$ and 20.1 cm$^3$. It is reasonable to discard the first reading since it is very different from the other three and it is the result obtained in the first trial.

\[\therefore \text{ Average volume of 0.10 M sodium hydroxide solution required for neutralization} \]

\[
= \frac{20.0 + 19.9 + 20.1}{3} \text{ cm}^3
\]
\((\text{COOH})_2\text{(aq)} + 2\text{NaOH(aq)} \rightarrow (\text{COONa})_2\text{(aq)} + 2\text{H}_2\text{O(l)}\)

\[ \text{M} = 0.10 \text{ M} \]

\[
\begin{array}{c|c}
25.0 \text{ cm}^3 & 20.0 \text{ cm}^3 \\
\end{array}
\]

\[
\begin{array}{c}
250.0 \text{ cm}^3 \\
(\text{used}) 25.0 \text{ cm}^3
\end{array}
\]

Number of moles of NaOH in 20.0 \text{ cm}^3 solution

\[
\text{Molarity of solution} \times \text{Volume of solution}
\]

\[
= 0.10 \text{ mol dm}^{-3} \times \frac{20.0}{1000} \text{ dm}^3
\]

\[
= 0.0020 \text{ mol}
\]

According to the equation, 1 mole of (\text{COOH})_2 requires 2 moles of NaOH for complete neutralization.

i.e.

\[
\begin{array}{l}
\text{Number of moles of (COOH)$_2$ in 25.0 cm}^3 \text{ dilute solution} = \frac{1}{2} \times \text{Number of moles of NaOH} \\
\end{array}
\]

\[
= \frac{1}{2} \times 0.0020 \text{ mol}
\]

\[
= 0.0010 \text{ mol}
\]

Molarity of dilute acid solution

\[
\frac{\text{Number of moles of (COOH)$_2$}}{\text{Volume of solution}} = \frac{0.0010 \text{ mol}}{(25.0 / 1000) \text{ dm}^3}
\]

\[
= 0.040 \text{ mol dm}^{-3}
\]
Molarity of the original acid solution = \(0.040 \text{ mol dm}^{-3} \times \frac{250 \text{ cm}^3}{25 \text{ cm}^3}\) = 0.40 mol dm\(^{-3}\) (M)

Mass of H\(_2\)SO\(_4\) in 1 dm\(^3\) = 9.8 g

Number of moles of H\(_2\)SO\(_4\) in 1 dm\(^3\) = \(\frac{9.8 \text{ g}}{98.0 \text{ g mol}^{-1}}\)

= 0.10 mol

Molarity of the sulphuric acid = 0.10 mol dm\(^{-3}\) (M)

\[\text{H}_2\text{SO}_4(aq) + 2\text{NaOH(aq)} \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O(l)}\]

0.10 M \hspace{1cm} ? M

30.0 cm\(^3\) \hspace{1cm} 25.0 cm\(^3\)

Number of moles of H\(_2\)SO\(_4\) in 30.0 cm\(^3\) solution = Molarity of solution \times Volume of solution

= 0.10 mol dm\(^{-3}\) \times \frac{30.0}{1000} \text{ dm}^3

= 0.0030 mol

According to the equation, 1 mole of H\(_2\)SO\(_4\) requires 2 moles of NaOH for complete neutralization.

i.e.

Number of moles of NaOH = 2 \times \text{Number of moles of H}_2\text{SO}_4

= 2 \times 0.0030 \text{ mol}

= 0.0060 mol

Molarity of sodium hydroxide solution = \(\frac{\text{Number of moles of NaOH}}{\text{Volume of solution}}\)

= \(\frac{0.0060 \text{ mol}}{(25.0 / 1000) \text{ dm}^3}\)

= 0.24 mol dm\(^{-3}\) (M)
Number of moles of $\text{H}_2\text{SO}_4$ in 100 cm$^3$ solution = Molarity of solution x Volume of solution

$$= 0.9 \text{ mol dm}^{-3} \times \frac{100}{1000} \text{ dm}^3$$

$$= 0.09 \text{ mol}$$

According to the equation, 1 mole of $\text{H}_2\text{SO}_4$ requires with 2 moles of NaOH for complete neutralization, producing 1 mole of $\text{Na}_2\text{SO}_4$.

Number of moles of $\text{Na}_2\text{SO}_4$ in the resulting solution = 0.09 mol

Volume of the resulting solution = (100 + 200) cm$^3$ = 300 cm$^3$

Molarity of the resulting solution = \(\frac{\text{Number of moles of } \text{Na}_2\text{SO}_4}{\text{Volume of solution}}\)

$$= \frac{0.09 \text{ mol}}{(300/1000) \text{ dm}^3}$$

$$= 0.3 \text{ mol dm}^{-3} (\text{M})$$

Number of moles of OH$^-$ in 25.0 cm$^3$ of 1.00 M NaOH solution = Molarity of solution x Volume of solution

$$= 1.00 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}^3$$

$$= 0.0250 \text{ mol}$$
A \[
\text{NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H}_2\text{O(l)}
\]

\[
? \text{ mol} \quad 0.550 \text{ M} \\
28.0 \text{ cm}^3
\]

Number of moles of HCl in 28.0 cm\(^3\) solution = Molarity of solution x Volume of solution

\[
= 0.550 \text{ mol dm}^{-3} x \frac{28.0}{1000} \text{ dm}^3
\]

\[
= 0.0154 \text{ mol}
\]

According to the equation, 1 mole of NaOH requires 1 mole of HCl for complete neutralization.

i.e. Number of moles of OH\(^-\) in the filtrate = 0.0154 mol

C

Number of moles of OH\(^-\) used for precipitation = (0.0250 - 0.0154) mol

\[
= 0.00960 \text{ mol}
\]

\[
\text{Ni}^{2+}(aq) + 2\text{OH}^- (aq) \rightarrow \text{Ni(OH)}_2(s)
\]

0.00960 mol

According to the equation, 1 mole of Ni\(^{2+}\) reacts with 2 moles of OH\(^-\) to give 1 mole of Ni(OH)\(_2\).

\[
\therefore \quad \text{Number of moles of Ni}^{2+} = x \text{ Number of moles of OH}^- \\
= x 0.00960 \text{ mol}
\]

\[
= 0.00480 \text{ mol}
\]

Molarity of nickel(II) sulphate solution =

\[
\frac{\text{Number of moles of Ni}^{2+}}{\text{Volume of solution}}
\]

\[
= \frac{0.00480 \text{ mol}}{(25.0 / 1000) \text{ dm}^3}
\]

\[
= 0.192 \text{ mol dm}^{-3} \text{ (M)}
\]
According to the equation, 1 mole of CH₃COOH requires 1 mole of NaOH for complete reaction.

Number of moles of CH₃COOH in 20.0 cm³ solution
= Molarity of solution x Volume of solution
= 0.10 mol dm⁻³ x \( \frac{20.0}{1000} \) dm³
= 0.0020 mol

Number of moles of NaOH in 20.0 cm³ solution
= Molarity of solution x Volume of solution
= 0.10 mol dm⁻³ x \( \frac{20.0}{1000} \) dm³
= 0.0020 mol

∴ There is complete reaction.

Option B – H₂SO₄(aq) + Na₂CO₃(aq) → Na₂SO₄(aq) + CO₂(g) + H₂O(l)

According to the equation, 1 mole of H₂SO₄ requires 1 mole of Na₂CO₃ for complete reaction.

Number of moles of H₂SO₄ in 25.0 cm³ solution
= Molarity of solution x Volume of solution
= 0.10 mol dm⁻³ x \( \frac{25.0}{1000} \) dm³
= 0.0025 mol

Number of moles of Na₂CO₃ in 25.0 cm³ solution
= Molarity of solution x Volume of solution
\[ \text{Number of moles of HCl in } 20.0 \text{ cm}^3 \text{ solution} = 0.10 \text{ mol dm}^{-3} \times \frac{20.0}{1000} \text{ dm}^3 \]
\[ = 0.0020 \text{ mol} \]

\[ \text{Number of moles of Na}_2\text{CO}_3 \text{ in } 10.0 \text{ cm}^3 \text{ solution} = 0.20 \text{ mol dm}^{-3} \times \frac{10.0}{1000} \text{ dm}^3 \]
\[ = 0.0020 \text{ mol} \]

\[ \therefore \text{ Sodium carbonate solution is in excess.} \]

Option C \(-2\text{HCl(aq)} + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow 2\text{NaCl(aq)} + \text{CO}_2(\text{g}) + \text{H}_2\text{O(l)}\)

0.10 M \quad 0.20 M
20.0 cm\(^3\) \quad 10.0 cm\(^3\)

According to the equation, 2 moles of HCl require 1 mole of Na\(_2\)CO\(_3\) for complete reaction.

Option D \(-\text{HCl(aq)} + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4\text{Cl(aq)}\)

0.10 M \quad 0.20 M
25.0 cm\(^3\) \quad 25.0 cm\(^3\)

According to the equation, 1 mole of HCl requires 1 mole of NH\(_3\) for complete reaction.
\[\text{Molarity of solution} \times \text{Volume of solution} = 0.10 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}^3\]

\[= 0.0025 \text{ mol}\]

Number of moles of \(\text{NH}_3\) in 25.0 cm\(^3\) solution
\[= \text{Molarity of solution} \times \text{Volume of solution}\]
\[= 0.20 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}^3\]

\[= 0.0050 \text{ mol}\]

\[\therefore \text{ Ammonia solution is in excess.}\]

153  C  We can represent the tribasic acid solution by \(\text{H}_3\text{X(aq)}\).

\[\text{H}_3\text{X(aq)} + 3\text{KOH(aq)} \rightarrow \text{K}_3\text{X(aq)} + 3\text{H}_2\text{O(l)}\]

\[? \text{ M} \quad 1.2 \text{ M}\]
\[25.0 \text{ cm}^3 \quad 20.0 \text{ cm}^3\]

Number of moles of \(\text{KOH}\) in 20.0 cm\(^3\) solution = Molarity of solution \(
\times \) Volume of solution
\[= 1.2 \text{ mol dm}^{-3} \times \frac{20.0}{1000} \text{ dm}^3\]

\[= 0.024 \text{ mol}\]

According to the equation, 1 mole of \(\text{H}_3\text{X}\) requires 3 moles of \(\text{KOH}\) for complete neutralization.

\[\text{i.e. Number of moles of } \text{H}_3\text{X} = \frac{1}{3} \times \text{Number of moles of } \text{KOH}\]

\[= \frac{1}{3} \times 0.024 \text{ mol}\]
\[ 0.0080 \text{ mol} \]

Molarity of tribasic acid solution \[ = \frac{\text{Number of moles of } H_3X}{\text{Volume of solution}} \]

\[ = \frac{0.0080 \text{ mol}}{(25.0 / 1000) \text{ dm}^3} \]

\[ = 0.32 \text{ mol dm}^{-3} \text{ (M)} \]

154 \quad A \quad \text{Na}_2\text{CO}_3(\text{aq}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})

3.1 \text{ g} \quad 1.0 \text{ M} \quad 50.0 \text{ cm}^3

Molar Mass of \text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O} = (106.0 + 18.0n) \text{ g mol}^{-1}

\[ \text{Number of moles of } \text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O} = \frac{\text{Mass}}{\text{Molar mass}} \]

\[ = \frac{3.1 \text{ g}}{(106.0 + 18.0n) \text{ g mol}^{-1}} \]

Number of moles of HCl in 50.0 cm\(^3\) solution = Molarity of solution \times Volume of solution

\[ = 1.0 \text{ mol dm}^{-3} \times \frac{50.0}{1000} \text{ dm}^3 \]

\[ = 0.050 \text{ mol} \]

According to the equation, 1 mole of \text{Na}_2\text{CO}_3 requires 2 moles of HCl for complete reaction.

i.e. Number of moles of \text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O} = \frac{1}{2} \times \text{Number of moles of HCl}

\[ = \frac{1}{2} \times 0.050 \text{ mol} \]

\[ = 0.025 \text{ mol} \]
Number of moles of Na₂CO₃ • H₂O = \frac{3.1}{(106.0 + 18.0n)} \text{ mol} \\
= 0.025 \text{ mol} \\
\frac{n}{1} = 1

\text{Let } n \text{ be the basicity of the acid.}

We can represent the acid solution by HₙX(aq).

HₙX(aq) + nNaOH(aq) → NaₙX(aq) + nH₂O(l)

0.50 \text{ M} \quad 0.40 \text{ M} \\
10.0 \text{ cm}³ \quad 25.0 \text{ cm}³

\text{Number of moles of NaOH in 25.0 cm}³ \text{ solution} = \text{Molarity of solution} \times \text{Volume of solution}

= 0.40 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}³

= 0.010 \text{ mol}

\text{Number of moles of HₙX in 10.0 cm}³ \text{ solution} = \text{Molarity of solution} \times \text{Volume of solution} \times \text{dm}³

= 0.50 \text{ mol dm}^{-3} \times \frac{10.0}{1000}

= 0.0050 \text{ mol}

\text{According to the equation, 1 mole of HₙX requires } n \text{ moles of NaOH for complete neutralization.}

\text{Number of moles of HₙX : Number of moles of NaOH} = 1 : n

\text{i.e. } 0.0050 \text{ mol : 0.010 mol} = 1 : n

\frac{n}{2} = 1

\frac{n}{2} = 2

\frac{2}{2} = 2

\frac{2}{2} = \frac{2}{2}
We can represent the dibasic acid by $H_2X$.

\[
H_2X(aq) + 2NaOH(aq) \rightarrow Na_2X(aq) + 2H_2O(l)
\]

<table>
<thead>
<tr>
<th>Substance</th>
<th>Mass</th>
<th>Molarity</th>
<th>Volume</th>
</tr>
</thead>
<tbody>
<tr>
<td>$H_2X$</td>
<td>3.15 g</td>
<td>0.25 M</td>
<td>0.200 dm$^3$</td>
</tr>
<tr>
<td>$NaOH$</td>
<td>25.0 cm$^3$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Let $m$ g mol$^{-1}$ be the molar mass of $H_2X$.

Number of moles of $H_2X$ in 3.15 g solid

\[
= \frac{3.15}{m} \text{ mol}
\]

Number of moles of $H_2X$ in 25.0 cm$^3$ solution

\[
= \frac{1}{10} \times \frac{3.15}{m} \text{ mol}
\]

Number of moles of $NaOH$ in 20.0 cm$^3$ solution

\[
= 0.25 \text{ mol dm}^{-3} \times \frac{20.0}{1000} \text{ dm}^3
\]

\[
= 0.0050 \text{ mol}
\]

According to the equation, 1 mole of $H_2X$ requires 2 moles of $NaOH$ for complete neutralization.

\[
i.e. \text{ Number of moles of } H_2X = \frac{1}{2} \times \text{ Number of moles of } NaOH
\]

\[
= \frac{1}{2} \times 0.0050 \text{ mol}
\]

\[
= 0.0025 \text{ mol}
\]
Number of moles of $H_2X\quad = \frac{1}{10}x \frac{3.15}{m} \quad \text{mol} = 0.0025 \text{ mol}

m = 126

157 B

158 D \quad \text{Na}_2\text{CO}_3(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(l)

2.70 g \quad 1.00 \text{M}
(with \quad 25.0 \text{ cm}^3
impurity)

Number of moles of $\text{H}_2\text{SO}_4$ in 25.0 cm$^3$ solution = Molarity of solution $\times$ Volume of solution

\[= 1.00 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}^3\]

= 0.0250 mol

According to the equation, 1 mole of $\text{Na}_2\text{CO}_3$ requires 1 mole of $\text{H}_2\text{SO}_4$ for complete reaction.

i.e. Number of moles of $\text{Na}_2\text{CO}_3$ in the sample = 0.0250 mol

Molar mass of $\text{Na}_2\text{CO}_3 = (2 \times 23.0 + 12.0 + 3 \times 16.0) \text{ g mol}^{-1}

= 106.0 \text{ g mol}^{-1}

Mass of $\text{Na}_2\text{CO}_3$ in the sample = Number of moles of $\text{Na}_2\text{CO}_3 \times$ Molar mass of $\text{Na}_2\text{CO}_3$

\[= 0.0250 \text{ mol} \times 106.0 \text{ g mol}^{-1}
= 2.65 \text{ g}

\therefore \quad \text{Percentage purity of } \text{Na}_2\text{CO}_3 \text{ in the sample} \quad = \frac{2.65 \text{ g}}{2.70 \text{ g}} \times 100\%

= 98.1\%
159 C \[ \text{Fe(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{FeSO}_4(\text{aq}) + \text{H}_2(\text{g}) \]

\[
\begin{align*}
1.60 \text{ M} & \quad ? \text{ g} \\
100 \text{ cm}^3 & 
\end{align*}
\]

Number of moles of \( \text{H}_2\text{SO}_4 \) in 100 cm\(^3\) solution = Molarity of solution \times \) Volume of solution

\[
= 1.60 \text{ mol dm}^{-3} \times \frac{100}{1000} \text{ dm}^3
\]

= 0.160 mol

According to the equation, 1 mole of \( \text{H}_2\text{SO}_4 \) reacts with 1 mole of Fe to give 1 mole of \( \text{FeSO}_4 \).

i.e. Number of moles of \( \text{FeSO}_4 \) = Number of moles of \( \text{H}_2\text{SO}_4 \)

= 0.160 mol

Molar mass of \( \text{FeSO}_4 \cdot \text{H}_2\text{O} \) = \( [56.0 + 32.0 + 4 \times 16.0 + 7 \times (2 \times 1.0 + 16.0)] \text{ g mol}^{-1} \)

= 278.0 g mol\(^{-1}\)

Mass of \( \text{FeSO}_4 \cdot \text{H}_2\text{O} \) obtained = Number of moles of \( \text{FeSO}_4 \cdot \text{H}_2\text{O} \) \times \) Molar mass of \( \text{FeSO}_4 \cdot \text{H}_2\text{O} \)

= 0.160 mol \times 278.0 g mol\(^{-1}\)

= 44.5 g

160 A \[ \text{X(OH)}_2(\text{aq}) + 2\text{HNO}_3(\text{aq}) \rightarrow \text{X(NO}_3)_2(\text{aq}) + 2\text{H}_2\text{O(l)} \]

5.13 g \hspace{1cm} 0.200 M \hspace{1cm} ? \text{ cm}^3

\[
\downarrow
\]

250.0 cm\(^3\)

\[
\downarrow
\]

(used) 25.0 cm\(^3\)

Number of moles of \( \text{X(OH)}_2 \) in 250.0 cm\(^3\) solution = \( \frac{\text{Mass}}{\text{Molar mass}} \)
\[
\text{Number of moles of X(OH)}_2 \text{ in 25.0 cm}^3 \text{ solution} = \frac{1}{10} \times 0.0300 \text{ mol} = 0.00300 \text{ mol}
\]

According to the equation, 1 mole of X(OH)_2 requires 2 moles of HNO_3 for complete neutralization.

\[
\text{Number of moles of HNO}_3 = 2 \times \text{Number of moles of X(OH)}_2 = 2 \times 0.00300 \text{ mol} = 0.00600 \text{ mol}
\]

\[
\text{Volume of nitric acid required for neutralization} = \frac{\text{Number of moles of HNO}_3}{\text{Molar mass of HNO}_3} = \frac{0.00600 \text{ mol}}{0.200 \text{ mol dm}^{-3}} = 0.0300 \text{ dm}^3 = 30.0 \text{ cm}^3
\]

161 A \[\text{X}_2\text{CO}_3(aq) + 2\text{HNO}_3(aq) \rightarrow 2\text{XNO}_3(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)\]

\[
\begin{array}{c|c|c}
\text{mass} & \text{Molar} & \text{Volume} \\
\hline
2.12 \text{ g} & 0.200 \text{ M} & 250.0 \text{ cm}^3 \\
20.0 \text{ cm}^3 & \downarrow & \text{(used) 25.0 cm}^3 \\
\end{array}
\]

Let \( m \) be the relative atomic mass of X.

\[
\text{Molar mass of } \text{X}_2\text{CO}_3 = (2 \times m + 12.0 + 3 \times 16.0) \text{ g mol}^{-1} = (2m + 60.0) \text{ g mol}^{-1}
\]

\[
\text{Number of moles of } \text{X}_2\text{CO}_3 \text{ in } 2.12 \text{ g solid} = \frac{\text{Mass}}{\text{Molar mass}}
\]
Number of moles of $\text{HNO}_3$ in 20.0 cm$^3$ solution = Molarity of solution $\times$ Volume of solution

$= 0.200 \text{ mol dm}^{-3} \times \frac{20.0}{1000}$

$= 0.00400 \text{ mol}$

According to the equation, 1 mole of $X_2\text{CO}_3$ requires 2 moles of $\text{HNO}_3$ for complete reaction.

i.e. Number of moles of $X_2\text{CO}_3$ in 25.0 cm$^3$ solution = \( \frac{1}{2} \times \) Number of moles of $\text{HNO}_3$

$= \frac{1}{2} \times 0.00400 \text{ mol}$

$= 0.00200 \text{ mol}$

Number of moles of $X_2\text{CO}_3$ in 250.0 cm$^3$ solution = 0.00200 mol $\times$ \( \frac{250.0 \text{ cm}^3}{25 \text{ cm}^3} \)

$= 0.0200 \text{ mol}$

Number of moles of $X_2\text{CO}_3$ in 250.0 cm$^3$ solution = \( \frac{2.12 \text{ g}}{(2m + 60.0) \text{ g mol}^{-1}} \) = 0.0200 mol

$= 23.0$

\[ \therefore \text{ Relative atomic mass of } X \text{ is } 23.0. \]
Number of moles of HCl in 25 cm$^3$ solution = $0.8 \text{ mol dm}^{-3} \times \frac{25}{1000} \text{ dm}^3$

= 0.02 mol

According to the equation, 1 mole of HCl requires 1 mole of NaOH for complete neutralization.

\[ \text{i.e. Number of moles of NaOH} = \text{Number of moles of HCl} = 0.02 \text{ mol} \]

\[
\text{CH}_3\text{COOH(aq)} + \text{NaOH(aq)} \rightarrow \text{CH}_3\text{COONa(aq)} + \text{H}_2\text{O(l)}
\]

Number of moles of CH$_3$COOH in 40 cm$^3$ solution = Molarity of solution $\times$ Volume of solution

= $0.5 \text{ mol dm}^{-3} \times \frac{40}{1000} \text{ dm}^3$

= 0.02 mol

According to the equation, 1 mole of CH$_3$COOH requires 1 mole of NaOH for complete neutralization.

\[ \text{i.e. Number of moles of NaOH} = \text{Number of moles of CH}_3\text{COOH} = 0.02 \text{ mol} \]

Solutions X and Y require the same number of moles of NaOH for neutralization. Therefore they require the same volume of 1 M NaOH for neutralization.

(3) HCl is a strong acid while CH$_3$COOH is a weak acid. Also the concentration of
HCl is higher than that of CH₃COOH. The HCl has a higher concentration of mobile ions than the CH₃COOH. Therefore solutions X and Y have different electrical conductivity.

165  A

166  D  100 cm³ of 1 M sulphuric acid require 200 cm³ of 1 M sodium hydroxide solution for complete neutralization. Sulphuric acid is a dibasic acid. The number of hydrogen ions in 100 cm³ of 1 M sulphuric acid is twice as that of hydroxide ions in 100 cm³ of 1 M sodium hydroxide solution.

**Unit 17  Rate of reactions**

167  D

168  C  Option A − The rate of the reaction decreases as the reaction proceeds. Option B − Dilute sulphuric acid is in excess. The reaction stops when all the magnesium ribbons are used up. Option D − The reaction will be faster if magnesium powder is used.

169  C  The seaked flask was a closed system. Therefore the mass of the flask plus its contents remained the same.

170  C  Options A & B − Equal masses of marble chips and powdered marble react with excess dilute hydrochloric acid. The amount of gas produced is the same in both cases. Option D − The reaction between powdered marble and dilute hydrochloric acid is faster than that between marble chips and the acid. Therefore the curve for the reaction between powdered marble and the acid is steeper.

171  A  Option A − Zn(s) + 2HCl(aq) → ZnCl₂(aq) + H₂(g)

1 g  1.0 M, 100 cm³
Number of moles of Zn present

\[ \text{Number of moles of } \text{Zn present} = \frac{\text{mass}}{\text{Molar mass}} \]

\[ \frac{1 \text{ g}}{65.4 \text{ g mol}^{-1}} = 0.015 \text{ mol} \]

Number of moles of HCl in 100 cm\(^3\) of 1.0 M acid

\[ \text{Number of moles of } \text{HCl} = \text{Molarity of solution} \times \text{Volume of solution} \]

\[ = 1.0 \text{ mol dm}^{-3} \times \frac{100}{1000} \text{ dm}^3 \]

\[ = 0.10 \text{ mol} \]

Number of moles of HCl in 100 cm\(^3\) of 0.50 M acid

\[ = 0.50 \text{ mol dm}^{-3} \times \frac{100}{1000} \text{ dm}^3 \]

\[ = 0.050 \text{ mol} \]

According to the equation, 1 mole of Zn reacts with 2 moles of HCl. During the reaction, 0.015 mole of Zn reacts with 0.030 mole of HCl. HCl is in excess in both cases. The amount of Zn limits the mass of hydrogen produced. Therefore the amount of hydrogen produced is the same in both cases, i.e. mass lost of the beakers plus contents at the end of the reactions is the same.

Option B – The more concentrated acid takes a shorter time to complete the reaction.

Option C – The initial reaction rate is faster for the more concentrated acid.

Option D – The 1.0 M hydrochloric acid has a higher concentration of Cl\(^-\)(aq) after reacting with all the zinc.
Option A − The antacid tablets in Experiments I and II have the same surface area. Therefore they cannot be used for comparison.

Option B − The only difference between Experiments I and IV is the surface area of the antacid tablets.

Option C − The antacid tablets in Experiments II and III have the same surface area. Therefore they cannot be used for comparison.

Option D − The concentration and temperature of the acids in Experiments III and IV are different. Therefore they cannot be used for comparison.

Option A − The concentration and temperature of the acids in Experiments I and II are different. Therefore they cannot be used for comparison.

Option C − The only difference between Experiments II and III is the concentration of the acids.

Option D − The surface area of the tablets and temperature of the acids in Experiments III and IV are different. Therefore they cannot be used for comparison.

Options A & B − Sodium carbonate reacts with hydrogen ions in the acids. Sulphuric acid is a dibasic acid and nitric acid is a monobasic acid. Sulphuric acid has a higher concentration of hydrogen ions. Therefore the reaction rate between sodium carbonate and 1 M sulphuric acid is faster than that between sodium carbonate and 1 M nitric acid.

Options C & D − Rates of reactions at 40 °C are faster than that at 20 °C.

The hydrochloric acid is in excess in each case. Therefore the amount of calcium carbonate limits the amount of carbon dioxide produced, i.e. the total loss in mass of contents of the reaction flask is the same in both cases. The initial rate of the reaction increases when the
surface area of the calcium carbonate is increased.

176 B The initial rate of the reaction decreases when a less concentrated acid is used.

177 A The initial rate of the reaction increases with an increase in temperature.

178 C The volume of acid used will not affect the initial rate of the reaction.

179 A The initial rate of the reaction increases with an increase in temperature.

The concentration of ammonium nitrite solution used in Experiments I and II is the same. Therefore the volume of nitrogen produced is the same in both experiments.

180 C The initial reaction rate for Experiment I is faster than that for Experiment III because more concentrated ammonium nitrite solution is used in Experiment I.

The volume of nitrogen produced in Experiment I is greater than that in Experiment III because more concentrated ammonium nitrite solution is used in Experiment I.

181 C

182 D (1) The rate of a reaction increases when the surface area of a solid reactant is increased.

(2) In most cases, the rate of a reaction increases when the concentration of a reactant is increased.

(3) In most cases, the rate of a reaction increases when the temperature is increased.

183 C (1) The volume of the acid used does not affect the initial reaction rate.

(2) & (3) The initial reaction rate for Experiment III is faster because a more concentrated acid is used.
Magnesium reacts with the hydrogen ions in the acids. Sulphuric acid is a dibasic acid and hydrochloric acid is a monobasic acid. 2 M sulphuric acid has a higher concentration of hydrogen ions. Therefore the initial reaction rate between magnesium and 2 M sulphuric acid is faster than that between magnesium and 2 M hydrochloric acid.

(2) Hydrochloric acid is a strong acid while ethanoic acid is a weak acid. 2 M hydrochloric acid has a higher concentration of hydrogen ions. Therefore the initial reaction rate between magnesium and 2 M hydrochloric acid is faster than that between magnesium and 2 M ethanoic acid.

(3) The volume of acid used does not affect the initial reaction rate.

Sulphuric acid is a dibasic acid and hydrochloric acid is a monobasic acid. 2 M sulphuric acid has a higher concentration of hydrogen ions than 2 M hydrochloric acid. Zinc reacts with the hydrogen ions in the acids. Therefore the rate of the reaction between 1 g of zinc and 100 cm$^3$ of 2 M sulphuric acid is faster than that between 1 g of zinc and 100 cm$^3$ of 2 M hydrochloric acid.

In most cases, the rate of a reaction increases when the temperature is increased.

**Topic 4 Miscellaneous**

\[ \text{Zn}^{2+} (aq) + 2\text{OH}^{-} \rightarrow \text{Zn(OH)}_2(s) \]

<table>
<thead>
<tr>
<th>Solution</th>
<th>Reaction with dilute sodium hydroxide solution</th>
<th>Reaction with ammonia solution</th>
</tr>
</thead>
<tbody>
<tr>
<td>Iron(II) sulphate</td>
<td>A green precipitate forms.</td>
<td>A green precipitate forms.</td>
</tr>
<tr>
<td>Iron(III) sulphate</td>
<td>A reddish brown precipitate forms.</td>
<td>A reddish brown precipitate forms.</td>
</tr>
<tr>
<td>Lead(II) nitrate</td>
<td>A white precipitate forms; the precipitate dissolves in excess dilute sodium hydroxide solution to give a colourless solution.</td>
<td>A white precipitate forms.</td>
</tr>
<tr>
<td>Zinc sulphate</td>
<td>A white precipitate forms; the precipitate dissolves in excess dilute sodium hydroxide solution to give a colourless solution.</td>
<td>A white precipitate forms; the precipitate dissolves in excess ammonia solution to give a colourless solution.</td>
</tr>
</tbody>
</table>
Both lead(II) carbonate and zinc carbonate react with dilute nitric acid to give colourless carbon dioxide gas.

Lead(II) carbonate gives lead(II) nitrate solution with dilute nitric acid. Lead(II) nitrate solution reacts with sodium chloride solution to give a white precipitate, lead(II) chloride.

\[
Pb^{2+} (aq) + 2Cl^- (aq) \rightarrow PbCl_2(s)
\]

The colourless gas Y is carbon dioxide.

Option A – Dilute sodium hydroxide solution turns methyl orange yellow.

The alkaline gas evolved is ammonia.

The carbon monoxide reacts with the iron(III) oxide to give iron.

Zinc reacts with dilute hydrochloric acid to give zinc chloride solution and hydrogen. Zinc chloride solution gives a white precipitate, zinc hydroxide, with ammonia solution. Zinc hydroxide dissolves in excess ammonia solution to give a colourless solution.

\[
Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)
\]

\[
Zn^{2+}(aq) + 2OH^- (aq) \rightarrow Zn(OH)_2(s)
\]

<table>
<thead>
<tr>
<th>Substance</th>
<th>pH value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Liquid soap</td>
<td>~ 9</td>
</tr>
<tr>
<td>Rainwater</td>
<td>5.6</td>
</tr>
<tr>
<td>Tap water</td>
<td>~ 6</td>
</tr>
<tr>
<td>Vinegar</td>
<td>~ 3</td>
</tr>
</tbody>
</table>

Option A – Magnesium reacts with dilute hydrochloric acid to give magnesium chloride and hydrogen.

\[
Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)
\]
Option B – Magnesium hydroxide reacts with dilute hydrochloric acid to give magnesium chloride and water.

\[
\text{Mg(OH)}_2(s) + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2(aq) + 2\text{H}_2\text{O(l)}
\]

Option D – Copper(II) carbonate reacts with dilute hydrochloric acid to give copper(II) chloride, carbon dioxide and water.

\[
\text{CuCO}_3(s) + 2\text{HCl(aq)} \rightarrow \text{CuCl}_2(aq) + \text{CO}_2(g) + \text{H}_2\text{O(l)}
\]

198 B

199 C

Options A & B – Nitric acid and hydrochloric acid are monobasic acids.

Option C – Sulphuric acid is a dibasic acid. It contains the highest concentration of hydrogen ions. Therefore the bulb is the brightest for 1 M \(\text{H}_2\text{SO}_4\).

Option D – Ethanoic acid is a weak acid while the other three are strong acids. The 1 M \(\text{CH}_3\text{COOH}\) has the lowest concentration of hydrogen ions.

200 D

\(\text{CH}_3\text{COOH}\) is a weak acid while \(\text{H}_2\text{SO}_4\) and \(\text{HClO}_4\) are strong acids. Therefore 0.1 M \(\text{CH}_3\text{COOH}\) has the lowest concentration of hydrogen ions. The pH value of 0.1 M \(\text{CH}_3\text{COOH}\) is thus the highest among the three acids.

\(\text{H}_2\text{SO}_4\) is a dibasic acid while \(\text{HClO}_4\) is a monobasic acid. Therefore 0.1 M \(\text{H}_2\text{SO}_4\) has a higher concentration of hydrogen ions than 0.1 M \(\text{HClO}_4\). The pH value of 0.1 M \(\text{H}_2\text{SO}_4\) is thus lower than that of the 0.10 M \(\text{HClO}_4\).

The order of the pH value of the three acids is: \(\text{H}_2\text{SO}_4 < \text{HClO}_4 < \text{CH}_3\text{COOH}\).

201 D

Option A – Hydrochloric acid is a strong acid while ethanoic acid is a weak acid. The pH value of 0.1 M hydrochloric acid is lower than that of 0.1 M ethanoic acid.

Option B – 0.1 M hydrochloric acid has a higher concentration of hydrogen ions than 0.1 M ethanoic acid. The electrical conductivity of 0.1 M hydrochloric acid is
higher than that of 0.1 M ethanoic acid.

Option C −0.1 M hydrochloric acid has a higher concentration of hydrogen ions than 0.1 M ethanoic acid. The reaction rate between magnesium and 0.1 M hydrochloric acid is faster than that between magnesium and 0.1 M ethanoic acid.

Option D − Number of moles of HCl in 100 cm$^3$ of 0.1 M hydrochloric acid

\[= \text{Molarity of solution} \times \text{Volume of solution}\]

\[= 0.1 \text{ mol dm}^{-3} \times \frac{100}{1000} \text{ dm}^3\]

\[= 0.01 \text{ mol}\]

Number of moles of CH$_3$COOH in 100 cm$^3$ of 0.1 M ethanoic acid

\[= \text{Molarity of solution} \times \text{Volume of solution}\]

\[= 0.1 \text{ mol dm}^{-3} \times \frac{100}{1000} \text{ dm}^3\]

\[= 0.01 \text{ mol}\]

Both 1 mole of HCl and 1 mole of CH$_3$COOH require 1 mole of NaOH for neutralization. Therefore 100 cm$^3$ of 0.1 M hydrochloric acid and 100 cm$^3$ of 0.1 M ethanoic acid require the same number of moles of NaOH for neutralization.

Option B − The greatest amounts of H$_2$SO$_4$ and NaOH react to give the greatest amount of heat energy.

<table>
<thead>
<tr>
<th>Option</th>
<th>Volumes of 1 M H$_2$SO$_4$(aq) and 1 M NaOH (aq) reacted</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>10.0 cm$^3$ of H$_2$SO$_4$(aq) react with 20.0 cm$^3$ of NaOH (aq)</td>
</tr>
<tr>
<td>B</td>
<td>15.0 cm$^3$ of H$_2$SO$_4$(aq) react with 30.0 cm$^3$ of NaOH (aq)</td>
</tr>
<tr>
<td>C</td>
<td>12.5 cm$^3$ of H$_2$SO$_4$(aq) react with 25.0 cm$^3$ of NaOH (aq)</td>
</tr>
<tr>
<td>D</td>
<td>7.5 cm$^3$ of H$_2$SO$_4$(aq) react with 15.0 cm$^3$ of NaOH (aq)</td>
</tr>
</tbody>
</table>

Option B − The greatest amounts of H$_2$SO$_4$ and NaOH react to give the greatest amount of heat energy.

203 A Sodium carbonate is soluble in water while calcium carbonate is insoluble. Adding water to the mixture dissolves sodium carbonate but not calcium carbonate. Calcium carbonate
can be separated from the sodium carbonate solution by filtration.

<table>
<thead>
<tr>
<th>Option</th>
<th>Reaction</th>
<th>Product</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>Ba(^{2+})(aq) + SO(^{4-})(aq) → BaSO(_4)(s)</td>
<td>A white precipitate BaSO(_4) forms.</td>
</tr>
<tr>
<td>B</td>
<td>KCl (aq) + AgNO(_3) (aq) → KNO(_3)(aq) + AgCl (s)</td>
<td>A white precipitate AgCl forms.</td>
</tr>
<tr>
<td>C</td>
<td>Na(_2)SO(_4)(aq) + MgCl(_2)(aq) → 2NaCl (aq) + MgSO(_4)(aq)</td>
<td>No precipitate forms.</td>
</tr>
<tr>
<td>D</td>
<td>Cu(^{2+})(aq) + 2OH(^-)(aq) → Cu(OH)(_2)(s)</td>
<td>A pale blue precipitate Cu(OH)(_2) forms.</td>
</tr>
</tbody>
</table>

The parent acid of sodium nitrite NaNO\(_2\) is nitrous acid HNO\(_2\)(aq).

All nitrates are soluble in water.

Copper(II) sulphate is a soluble salt. To prepare it, mix a dilute acid (dilute sulphuric acid in this case) with a metal, an insoluble base or an insoluble carbonate. However copper (metal) cannot be used as it does not react with dilute sulphuric acid.

Both zinc chloride and zinc nitrate are soluble in water. Therefore it is unsuitable to prepare zinc chloride using zinc nitrate.

Option A –Calcium reacts with dilute sulphuric acid to produce insoluble calcium sulphate.
  Calcium sulphate forms a protective layer on the surface of calcium. This prevents further reaction between calcium and dilute sulphuric acid.
  Option B –Copper has no reaction with dilute hydrochloric acid.
  Option C –Iron reacts with dilute sulphuric acid to give iron(II) sulphate, not iron(III)
sulphate.

Option D – Magnesium reacts with dilute hydrochloric acid to give magnesium chloride.

211 A Molar mass of $\text{H}_2\text{SO}_4 = (2 \times 1.0 + 32.0 + 4 \times 16.0) \text{ g mol}^{-1}$

\[= 98.0 \text{ mol}^{-1}\]

Molarity of sulphuric acid = \(\frac{\text{Number of moles of } \text{H}_2\text{SO}_4}{\text{Volume of solution}}\)

\[4.00 \text{ mol dm}^{-3} = \frac{\text{Number of moles of } \text{H}_2\text{SO}_4}{2.00 \text{ dm}^3}\]

Number of moles of $\text{H}_2\text{SO}_4$ = \(4.00 \text{ mol dm}^{-3} \times 2.00 \text{ dm}^3\)

\[= 8.00 \text{ mol}\]

Mass of $\text{H}_2\text{SO}_4$ required = \(\text{Number of moles of } \text{H}_2\text{SO}_4 \times \text{Molar mass of } \text{H}_2\text{SO}_4\)

\[= 8.00 \text{ mol} \times 98.0 \text{ g mol}^{-1}\]

\[= 784 \text{ g}\]

212 D (MV) before dilution = (MV) after dilution where M = Molarity, V = Volume

\[2.0 \times \frac{100}{1000} = M \times \frac{100+400}{1000}\]

\[M = 0.40 \text{ mol dm}^{-3} (M)\]

Molarity of the diluted $\text{K}_2\text{SO}_4$ solution is 0.40 mol dm$^{-3}$ (M).

1 mole of $\text{K}_2\text{SO}_4$ contains 2 moles of K$^+$. 

\[\therefore \text{ Molarity of K}^+ \text{ in the resulting solution} = 2 \times 0.40 \text{ M}\]

\[= 0.80 \text{ M}\]

213 A (MV) before dilution = (MV) after dilution where M = Molarity, V = Volume

\[4.0 \times \frac{25.0}{1000} = 0.50 \times \frac{V}{1000}\]
Volume of the final solution = 200 cm$^3$

∴ Volume of water added = (200 - 25.0) cm$^3$
= 175 cm$^3$

Consider 1 000 cm$^3$ (i.e. 1 dm$^3$) of the sample.

Mass of 1 000 cm$^3$ of the sample = 1.64 g cm$^{-3}$ x 1 000 cm$^3$
= 1 640 g

Mass of sulphuric acid in 1 000 cm$^3$ of the sample
= Mass of 1 000 cm$^3$ of the sample x Percentage by mass of H$_2$SO$_4$ in the sample
= 1 640 g x 92.0%
= 1 509 g

Molar mass of H$_2$SO$_4$ = (2 x 1.0 + 32.0 + 4 x 16.0) g mol$^{-1}$
= 98.0 g mol$^{-1}$

Number of moles of H$_2$SO$_4$ in 1 000 cm$^3$ of the sample

= \[ \frac{\text{Mass}}{\text{Molar mass}} \]
= \[ \frac{1 509 \text{g}}{98.0 \text{ g mol}^{-1}} \]
= 15.4 mol

Molarity of sulphuric acid in the sample =

\[ \frac{\text{Number of moles of H}_2\text{SO}_4}{\text{Volume of solution}} \]
= \[ \frac{15.4 \text{ mol}}{1 \text{ dm}^3} \]
= 15.4 mol dm$^{-3}$ (M)

K$_2$CO$_3$(s) + 2HNO$_3$(aq) → 2KNO$_3$(aq) + CO$_2$(g) + H$_2$O(l)
\[
\text{KHCO}_3(s) + \text{HNO}_3(aq) \rightarrow \text{KNO}_3(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)
\]

Number of moles of HNO\textsubscript{3} required to react with 2 moles of K\textsubscript{2}CO\textsubscript{3}  
= 2 x 2 mol  
= 4 mol

Number of moles of HNO\textsubscript{3} required to react with 1 mole of KHCO\textsubscript{3}  
= 1 mol 
∴ Number of moles of HNO\textsubscript{3} required  
= (4 + 1) mol  
= 5 mol

\begin{align*}
216 & \text{ C} \\
\text{Na}_2\text{CO}_3(aq) + 2\text{HCl}(aq) & \rightarrow 2\text{NaCl}(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l) \\
0.150 \text{ M} & \quad 0.250 \text{ M} \\
25.0 \text{ cm}^3 & \quad ? \text{ cm}^3 \\
\end{align*}

Number of moles of \text{Na}_2\text{CO}_3 in 25.0 \text{ cm}^3 solution  
= Molarity of solution  
\times Volume of solution 
\begin{align*}
&\frac{25.0}{1000} \text{ dm}^3 \\
&= 0.150 \text{ mol dm}^{-3} \times \\
&= 0.00375 \text{ mol} \\
\end{align*}

According to the equation, 1 mole of \text{Na}_2\text{CO}_3 requires 2 moles of \text{HCl} for complete reaction. 

\begin{align*}
\text{i.e. Number of moles of HCl} & = 2 \times \text{Number of moles of Na}_2\text{CO}_3 \\
& = 2 \times 0.00375 \text{ mol} \\
& = 0.00750 \text{ mol} \\
\text{Volume of HCl required} & = \frac{\text{Number of moles of HCl}}{\text{Volume of solution}} \\
& = \frac{0.00750 \text{ mol}}{0.250 \text{ mol dm}^{-3}} \\
& = 0.0300 \text{ dm}^3 \\
& = 30.0 \text{ cm}^3
\end{align*}
According to the equation, 1 mole of $\text{H}_2\text{SO}_4$ requires 2 moles of $\text{NaOH}$ for complete neutralization.

$$ \text{Volume of NaOH solution required to neutralize the sulphuric acid} = 2 V \text{ cm}^3 $$

According to the equation, 1 mole of $\text{HCl}$ requires 1 mole of $\text{NaOH}$ for complete neutralization.

$$ \text{Volume of NaOH solution required to neutralize the hydrochloric acid} = 0.5 V \text{ cm}^3 $$

\[ \therefore \text{Total volume of NaOH solution required} = (2V + 0.5V) \text{ cm}^3 = 2.5V \text{ cm}^3 \]

Number of moles of $\text{H}_2\text{SO}_4$ in $25.0 \text{ cm}^3$ solution = Molarity of solution \times Volume of solution

\[ \frac{25.0}{1000} \text{ dm}^3 = 2.00 \text{ mol dm}^{-3} \times \]

\[ = 0.0500 \text{ mol} \]
According to the equation, 1 mole of $\text{H}_2\text{SO}_4$ requires 2 moles of $\text{NaOH}$ for complete neutralization, producing 1 mole of $\text{Na}_2\text{SO}_4$.

\[
\text{i.e. Number of moles of NaOH} = 2 \times \text{Number of moles of } \text{H}_2\text{SO}_4
\]

\[
= 2 \times 0.0500 \text{ mol}
= 0.100 \text{ mol}
\]

Number of moles of $\text{Na}_2\text{SO}_4$ = Number of moles of $\text{H}_2\text{SO}_4$

\[
= 0.0500 \text{ mol}
\]

Volume of $\text{NaOH}$ solution required = \[
\frac{\text{Number of moles of } \text{NaOH}}{\text{Molarity of solution}}
\]

\[
= \frac{0.100 \text{ mol}}{2.00 \text{ mol dm}^{-3}}
= 50.0 \text{ cm}^3
\]

Volume of the resulting solution = $(25.0 + 50.0) \text{ cm}^3$

\[
= 75.0 \text{ cm}^3
\]

Molarity of the resulting $\text{Na}_2\text{SO}_4$ solution = \[
\frac{\text{Number of moles of } \text{Na}_2\text{SO}_4}{\text{Volume of solution}}
\]

\[
= \frac{0.005 \text{ mol}}{(75.0 / 1000) \text{ dm}^3}
= 0.667 \text{ mol dm}^{-3} \text{ (M)}
\]

219 A Option B – A pipette should be used to measure 25.0 cm$^3$ of dilute nitric acid.
Options C & D – A burette should be used to deliver various volumes of standard sodium hydroxide solution accurately.
CH₃COOH(aq) + NaOH(aq) $\rightarrow$ CH₃COONa(aq) + H₂O(l)

$0.320 \text{ M} \times 0.00480 \text{ mol}$

According to the equation, 1 mole of CH₃COOH requires 1 mole of NaOH for complete neutralization.

i.e. Number of moles of CH₃COOH in 25.0 cm$^3$ solution = 0.00480 mol

Molarity of CH₃COOH in vinegar = \[
\frac{\text{Number of moles of CH₃COOH}}{\text{Volume of solution}} = \frac{0.00480 \text{ mol}}{(25.0 / 1000) \text{ dm}^3} = 0.192 \text{ mol dm}^{-3} (\text{M})
\]

The burette should be washed with distilled water and then the alkali it is to contain because any water or impurities in the apparatus will change the concentration of the alkali.

This will affect the titration results.
Let \( m \) g mol\(^{-1}\) be the molar mass of \( \text{H}_2\text{X} \).

Number of moles of \( \text{H}_2\text{X} \) in 250.0 cm\(^3\) solution

\[
\text{Number of moles of } \text{H}_2\text{X} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{4.60 \text{ g}}{m \text{ g mol}^{-1}}
\]

Number of moles of \( \text{H}_2\text{X} \) in 25.0 cm\(^3\) solution

\[
\text{Number of moles of } \text{H}_2\text{X} = \frac{4.60 \text{ g}}{m \text{ g mol}^{-1}} \times \frac{1}{10} \text{ mol}
\]

Number of moles of NaOH in 25.0 cm\(^3\) solution

\[
\text{Number of moles of NaOH} = \text{Molarity of solution} \times \text{Volume of solution} = 0.400 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}^3 = 0.0100 \text{ mol}
\]

According to the equation, 1 mole of \( \text{H}_2\text{X} \) requires 2 moles of NaOH for complete neutralization.

\[
i.e. \text{Number of moles of } \text{H}_2\text{X} = \frac{1}{2} \times \text{Number of moles of NaOH}
\]

\[
= \frac{1}{2} \times 0.0100 \text{ mol} = 0.00500 \text{ mol}
\]

Number of moles of \( \text{H}_2\text{X} \) in 25.0 cm\(^3\) solution

\[
\text{Number of moles of } \text{H}_2\text{X} = \frac{4.60 \text{ g}}{m \text{ g mol}^{-1}} \times \frac{1}{10} \text{ mol} = 0.00500 \text{ mol}
\]

\[
m = 92.0
\]

\[
\therefore \text{The molar mass of } \text{H}_2\text{X} \text{ is } 92.0 \text{ g mol}^{-1}.
\]
\[
\text{Na}_2\text{CO}_3(\text{aq}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})
\]

1.00 g 0.500 M
(with 15.1 cm\(^3\) impurity)

Number of moles of HCl in 15.1 cm\(^3\) solution = Molarity of solution x Volume of solution

= 0.500 mol dm\(^{-3}\) x \(\frac{15.1}{1000}\) dm\(^3\)

= 0.00755 mol

According to the equation, 1 mole of \(\text{Na}_2\text{CO}_3\) requires 2 moles of HCl for complete reaction.

i.e. Number of moles of \(\text{Na}_2\text{CO}_3\) = \(\frac{1}{2}\) x Number of moles of HCl

= \(\frac{1}{2}\) x 0.00755 mol

= 0.00378 mol

Molar mass of \(\text{Na}_2\text{CO}_3\) = (2 x 23.0 + 12.0 + 3 x 16.0) g mol\(^{-1}\)

= 106.0 g mol\(^{-1}\)

Mass of \(\text{Na}_2\text{CO}_3\) in the sample = Number of moles of \(\text{Na}_2\text{CO}_3\) x Molar mass of \(\text{Na}_2\text{CO}_3\)

= 0.00378 mol x 106.0 g mol\(^{-1}\)

= 0.401 g

\[
\text{∴ Percentage purity of } \text{Na}_2\text{CO}_3 \text{ in the sample} = \frac{0.401 \text{ g}}{1.00 \text{ g}} \times 100\%
\]

= 40.1%

Let \(n\) be the basicity of the acid, so we can represent the acid by \(\text{H}_n\text{X}\).

\[
\text{H}_n\text{X}(\text{aq}) + n\text{NaOH}(\text{aq}) \rightarrow \text{Na}_n\text{X}(\text{aq}) + n\text{H}_2\text{O}(\text{l})
\]

0.20 M 0.50 M
25.0 cm\(^3\) 30.0 cm\(^3\)
Number of moles of $H_nX$ in 25.0 cm$^3$ solution $= \text{Molarity of solution} \times \text{Volume of solution}$

$= 0.20 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}^3$

$= 0.0050 \text{ mol}$

Number of moles of NaOH in 30.0 cm$^3$ solution $= \text{Molarity of solution} \times \text{Volume of solution}$

$= 0.50 \text{ mol dm}^{-3} \times \frac{30.0}{1000} \text{ dm}^3$

$= 0.015 \text{ mol}$

\[
\frac{\text{Number of moles of } H_nX}{\text{Number of moles of NaOH}} = \frac{1}{n} = \frac{0.0050 \text{ mol}}{0.015 \text{ mol}}
\]

$\therefore n = 3$

$\therefore$ The basicity of the acid is 3.

227 D $\text{Na}_2\text{CO}_3(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$

5.72 g 1.00 M

20.0 cm$^3$

Molar mass of $\text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O} = (106.0 + 18.0n) \text{ g mol}^{-1}$

Number of moles of $\text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O}$ used $= \frac{\text{Mass}}{\text{Molar mass}}$

$= \frac{5.72 \text{ g}}{(106.0 + 18.0n) \text{ g mol}^{-1}}$

Number of moles of $\text{H}_2\text{SO}_4$ in 20.0 cm$^3$ solution $= \text{Molarity of solution} \times \text{Volume of solution}$

$= 1.00 \text{ mol dm}^{-3} \times \frac{20.0}{1000} \text{ dm}^3$

$= 0.0200 \text{ mol}$

According to the equation, 1 mole of $\text{Na}_2\text{CO}_3$ requires 1 mole of $\text{H}_2\text{SO}_4$ for complete reaction.
i.e. Number of moles of Na₂CO₃•H₂O used = 0.0200 mol

\[
\text{Number of moles of Na₂CO₃•H₂O used} = \frac{5.72}{(106.0 + 18.0n)} \text{ mol} = 0.0200 \text{ mol}
\]

\[\therefore n = 10\]

228 \(\text{C}\) \(\text{Pb(NO}_3\text{)}_2(\text{aq}) + 2\text{NaCl(aq)} \rightarrow \text{PbCl}_2(\text{s}) + 2\text{NaNO}_3(\text{aq})\)

According to the equation, 1 mole of Pb(NO₃)₂ requires 2 moles of NaCl to give 1 mole of PbCl₂.

<table>
<thead>
<tr>
<th>Option</th>
<th>Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>5 cm³ of Pb(NO₃)₂ (aq) react with 10 cm³ of NaCl (aq)</td>
</tr>
<tr>
<td>B</td>
<td>5 cm³ of Pb(NO₃)₂ (aq) react with 10 cm³ of NaCl (aq)</td>
</tr>
<tr>
<td>C</td>
<td>10 cm³ of Pb(NO₃)₂ (aq) react with 20 cm³ of NaCl (aq)</td>
</tr>
<tr>
<td>D</td>
<td>7.5 cm³ of Pb(NO₃)₂ (aq) react with 15 cm³ of NaCl (aq)</td>
</tr>
</tbody>
</table>

Option C – The greatest amounts of Pb(NO₃)₂(aq) and NaCl(aq) react to give the greatest amount of precipitate.

229 \(\text{B}\) From the curve, 45.0 cm³ of sulphuric acid are required to neutralize the sodium hydroxide solution. (The temperature of the reaction mixture is maximum at the point of complete neutralization.)

\[
\text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH(}aq\text{)} \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O(l)}
\]

\[
? \text{ M} \quad 2.0 \text{ M}
\]

\[
45.0 \text{ cm}^3 \quad 50.0 \text{ cm}^3
\]

Number of moles of NaOH in 50.0 cm³ solution = Molarity of solution x Volume of solution

\[
= 2.0 \text{ mol dm}^{-3} \times \frac{50.0}{1000} \text{ dm}^3
\]

\[
= 0.10 \text{ mol}
\]

According to the equation, 1 mole of H₂SO₄ requires 2 moles of NaOH for complete neutralization.
i.e. Number of moles of \( \text{H}_2\text{SO}_4 \) = \( \frac{1}{2} \) x Number of moles of NaOH

\[ = \frac{1}{2} \times 0.10 \text{ mol} \]

\[ = 0.050 \text{ mol} \]

Molarity of sulphuric acid = \( \frac{\text{Number of moles of } \text{H}_2\text{SO}_4}{\text{Volume of solution}} \)

\[ = \frac{0.050 \text{ mol}}{(45.0 / 1000) \text{ dm}^3} \]

\[ = 1.1 \text{ M} \]

230  B  The only difference between Experiments I and III is the pH value of the dye solution.

231  D  A reaction goes faster if:

- the surface area of a solid reactant is increased;
- the concentration of a reactant is increased; and
- the solution is hotter.

232  D  In most cases, the rate of a reaction increases when the temperature is increased.

233  A  Adding water to the acid decreases its concentration. The initial reaction rate decreases when the concentration of the acid is decreased.

234  C  The same amount of hydrogen is produced in each case as the same mass of zinc granules is used. Different temperatures affect only the reaction rates.
Option A – The concentration for both reagents in Experiments I and II are different. Therefore they cannot be used for comparison.

Option B – The temperature in Experiments I and III are different. Therefore they cannot be used for comparison.

Option C – The temperature in Experiments II and IV are different. Therefore they cannot be used for comparison.

Option D – The total volume of the samples in Experiments III and IV are the same. The only difference is the volume of the sodium thiosulphate solution used, i.e. the concentration of the sodium thiosulphate solution.

Option A – The concentration for the sodium thiosulphate solutions in Experiments I and III are different. Therefore they cannot be used for comparison.

Option C – The concentration for the acids in Experiments II and III are different. Therefore they cannot be used for comparison.

Option D – The concentration for both reagents in Experiments II and IV are different. Therefore they cannot be used for comparison.

(1) Soft drink contains carbonic acid.
(2) Grapefruits contain citric acid.
(3) Oven cleaner contains sodium hydroxide.

(1) Glass cleaner contains ammonia.
(2) Alkaline solutions are more effective than acidic solutions in creating curls of the hair. The curls are permanent. Perm solutions are thus alkaline.
(3) Limewater is a saturated solution of calcium hydroxide.

(1) Iron nails give hydrogen gas with dilute hydrochloric acid.
(2) Marble gives carbon dioxide gas with dilute hydrochloric acid.
Baking powder is sodium hydrogen carbonate. It gives carbon dioxide gas with dilute hydrochloric acid.

Copper(II) sulphate solution gives a pale blue precipitate with dilute sodium hydroxide solution.

\[ \text{Cu}^{2+}(aq) + 2\text{OH}^- (aq) \rightarrow \text{Cu}(	ext{OH})_2(s) \]

Iron(II) nitrate solution gives a green precipitate with dilute sodium hydroxide solution.

\[ \text{Fe}^{2+}(aq) + 2\text{OH}^- (aq) \rightarrow \text{Fe}(	ext{OH})_2(s) \]

Aluminium nitrate solution gives a white precipitate with ammonia solution.

\[ \text{Al}^{3+}(aq) + 3\text{OH}^- (aq) \rightarrow \text{Al}(	ext{OH})_3(s) \]

Iron(III) sulphate solution gives a reddish brown precipitate with ammonia solution.

\[ \text{Fe}^{3+}(aq) + 3\text{OH}^- (aq) \rightarrow \text{Fe}(	ext{OH})_3(s) \]

Aluminium hydroxide, lead(II) hydroxide and zinc hydroxide dissolve in excess dilute sodium hydroxide solution due to the formation of soluble complex salts.

Dilute hydrochloric acid turns phenolphthalein colourless.

2AgNO$_3$(aq) + CaCl$_2$(aq) $\rightarrow$ 2AgCl(s) + Ca(NO$_3$)$_2$(aq)

Na$_2$CO$_3$(aq) + CuSO$_4$(aq) $\rightarrow$ CuCO$_3$(s) + Na$_2$SO$_4$(aq)

Magnesium reacts with dilute hydrochloric acid while silver has no reaction.

Silver nitrate solution gives a white precipitate with dilute hydrochloric acid while potassium nitrate solution does not.

\[ \text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s) \]

Zinc carbonate gives gas bubbles (carbon dioxide gas) with dilute hydrochloric acid while zinc chloride does not.
B
(1) Heating ammonium chloride solid with dilute sodium hydroxide solution liberates ammonia gas.

\[ \text{NH}_4^+ (aq) + \text{OH}^- (aq) \rightarrow \text{NH}_3 (g) + \text{H}_2\text{O}(l) \]

(2) Lead has no reaction with water.

(3) Zinc chloride solution reacts with ammonia solution to give a white precipitate, zinc hydroxide.

\[ \text{Zn}^{2+} (aq) + 2\text{OH}^- (aq) \rightarrow \text{Zn(OH)_2(s)} \]

C
(1) Both ammonium nitrate solution and potassium chloride solution are colourless.

(2) Ammonia is liberated when ammonium nitrate solution is heated with dilute sodium hydroxide solution. The ammonia can be tested with moist red litmus paper.

\[ \text{NH}_4^+ (aq) + \text{OH}^- (aq) \rightarrow \text{NH}_3 (g) + \text{H}_2\text{O}(l) \]

(3) Potassium chloride solution gives a white precipitate with silver nitrate solution, but ammonium nitrate solution does not.

\[ \text{Ag}^+ (aq) + \text{Cl}^- (aq) \rightarrow \text{AgCl(s)} \]

D
(1) Ca(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow CaSO_4(s) + 2NaNO_3(aq)

(2) Pb(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow PbCO_3(s) + 2NaNO_3(aq)

(3) Mg^{2+}(aq) + 2\text{OH}^- (aq) \rightarrow \text{Mg(OH)_2(s)}

A
(2) Ammonia solution turns phenolphthalein red.

(3) Aluminium hydroxide is insoluble in excess ammonia solution.

A
(1) Number of moles of HCl in 25 cm^3 of 1 M hydrochloric acid.
Number of moles of CH₃COOH in 25 cm³ of 1 M ethanoic acid

= Molarity of solution x Volume of solution

= 1 mol dm⁻³ x \( \frac{25}{1000} \) dm³

= 0.025 mol

Both 1 mole of HCl and 1 mole of CH₃COOH require 1 mole of NaOH for complete neutralization.

\[ \therefore \] 0.025 mole of hydrochloric acid and 0.025 mole of ethanoic acid require the same number of mole of sodium hydroxide for complete neutralization.

(2) Hydrochloric acid is a strong acid while ethanoic acid is a weak acid. The hydrochloric acid has a higher concentration of mobile ions than the ethanoic acid. Therefore the electrical conductivity of the hydrochloric acid is higher than that of the ethanoic acid.

(3) The hydrochloric acid has a higher concentration of hydrogen ions than the ethanoic acid.

Sodium hydroxide is a strong alkali. It almost completely ionizes in water to give sodium ions and hydroxide ions. Ammonia is a weak alkali. It only partially ionizes in water, forming very few hydroxide ions.

(1) The sodium hydroxide solution has a higher concentration of hydroxide ions than
the ammonia solution. Therefore the sodium hydroxide solution has a higher pH value than the ammonia solution.

(3) The sodium hydroxide solution has a higher concentration of mobile ions than the ammonia solution. Therefore the electrical conductivity of the sodium hydroxide solution is higher than that of the ammonia solution.

255  D  (1) Heat energy is always given out in neutralization reactions.
     (3) The resulting solution conducts electricity because it contains mobile ions.

256  D  In all the reactions, the only chemical change is H⁺(aq) + OH⁻(aq) → H₂O(l). Therefore the temperature rises are the same.

257  D  (1) Mg(s) + H₂SO₄(aq) → MgSO₄(aq) + H₂(g)
     (2) MgO(s) + H₂SO₄(aq) → MgSO₄(aq) + H₂O(l)
     (3) MgCO₃(s) + H₂SO₄(aq) → MgSO₄(aq) + H₂O(l) + CO₂(g)

258  D

259  B  (1) A pipette should be used to measure and transfer the diluted vinegar.
     (3) The colour of the phenolphthalein changes from colourless to red at the end point.

260  A  (3) The volume of the acid used would not affect the rate of the reaction.

261  A  (3) Using the same mass of magnesium ribbons will decrease the initial reaction rate.

262  A

263  D  Solid citric acid does not contain hydrogen ions. Therefore solid citric acid does not show acidic properties.
Our stomach produces hydrochloric acid along with enzymes to digest food. Calcium carbonate gives carbon dioxide gas when reacting with dilute hydrochloric acid. The gas makes a person uncomfortable. Therefore calcium carbonate is seldom used in antacids nowadays.

Ethanoic acid is a weak acid. It only partially ionizes in water, forming very few hydrogen ions.

Ammonia is a weak alkali.

For example, 1 mole of sulphuric acid can neutralize two moles of sodium hydroxide.

\[ \text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH(\text{aq})} \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O(\text{l})} \]

**Section D  Short Questions**

1

<table>
<thead>
<tr>
<th>Reaction between</th>
<th>Chemical equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Magnesium and dilute hydrochloric acid</td>
<td>Mg(s) + 2HCl(aq) \rightarrow MgCl(_2)(aq) + H(_2)(g)</td>
</tr>
<tr>
<td>Magnesium and dilute sulphuric acid</td>
<td>Mg(s) + H(_2)SO(_4)(aq) \rightarrow MgSO(_4)(aq) + H(_2)(g)</td>
</tr>
<tr>
<td>Sodium carbonate solid and dilute hydrochloric acid</td>
<td>Na(_2)CO(_3)(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H(_2)O(l) + CO(_2)(g)</td>
</tr>
<tr>
<td>Sodium carbonate solution and dilute sulphuric acid</td>
<td>Na(_2)CO(_3)(aq) + H(_2)SO(_4)(aq) \rightarrow Na(_2)SO(_4)(aq) + H(_2)O(l) + CO(_2)(g)</td>
</tr>
<tr>
<td>Calcium carbonate solid and dilute nitric acid</td>
<td>CaCO(_3)(aq) + 2HNO(_3)(aq) \rightarrow Ca(NO(_3))(_2)(aq) + H(_2)O(l) + CO(_2)(g)</td>
</tr>
<tr>
<td>Sodium hydrogen carbonate solution and dilute hydrochloric acid</td>
<td>NaHCO(_3)(aq) + HCl(aq) \rightarrow NaCl(aq) + H(_2)O(l) + CO(_2)(g)</td>
</tr>
<tr>
<td>Sodium hydrogen carbonate solid and dilute sulphuric acid</td>
<td>2NaHCO(_3)(s) + H(_2)SO(_4)(aq) \rightarrow Na(_2)SO(_4)(aq) + 2H(_2)O(l) + 2CO(_2)(g)</td>
</tr>
<tr>
<td>Magnesium hydroxide solid and dilute hydrochloric acid</td>
<td>Mg(OH)(_2)(s) + 2HCl(aq) \rightarrow MgCl(_2)(aq) + 2H(_2)O(l)</td>
</tr>
</tbody>
</table>

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<table>
<thead>
<tr>
<th>Reaction between</th>
<th>Ionic equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Magnesium and sulphuric acid</td>
<td>Mg(s) + 2H(^+)(aq) \rightarrow Mg(^{2+})(aq) + H(_2)(g)</td>
</tr>
</tbody>
</table>
Sodium carbonate solution and dilute hydrochloric acid

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

Sodium hydrogencarbonate solution and dilute nitric acid

$$\text{HCO}_3^{-} (aq) + \text{H}^+ (aq) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g)$$

<table>
<thead>
<tr>
<th>Solution containing</th>
<th>Name of hydroxide formed with dilute sodium hydroxide solution</th>
<th>Colour of precipitate formed</th>
<th>Does the precipitate redissolve in excess dilute sodium hydroxide solution? (✓ or ✗)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Calcium ions</td>
<td>Calcium hydroxide</td>
<td>White</td>
<td>✗</td>
</tr>
<tr>
<td>Aluminium ions</td>
<td>Aluminium hydroxide</td>
<td>White</td>
<td>✗</td>
</tr>
<tr>
<td>Lead(II) ions</td>
<td>Lead(II) hydroxide</td>
<td>White</td>
<td>✗</td>
</tr>
<tr>
<td>Magnesium ions</td>
<td>Magnesium hydroxide</td>
<td>White</td>
<td>✗</td>
</tr>
<tr>
<td>Iron(II) ions</td>
<td>Iron(II) hydroxide</td>
<td>Green</td>
<td>✗</td>
</tr>
<tr>
<td>Iron(III) ions</td>
<td>Iron(III) hydroxide</td>
<td>Reddish brown</td>
<td>✗</td>
</tr>
<tr>
<td>Copper(II) ions</td>
<td>Copper(II) hydroxide</td>
<td>Pale blue</td>
<td>✗</td>
</tr>
<tr>
<td>Zinc ions</td>
<td>Zinc hydroxide</td>
<td>White</td>
<td>✗</td>
</tr>
</tbody>
</table>

Solution containing aluminium ions and dilute ammonia solution

Aluminium ions

Lead(II) ions

Magnesium ions

Iron(II) ions

Iron(III) ions

Copper(II) ions

Zinc ions

<table>
<thead>
<tr>
<th>Solution containing</th>
<th>Name of hydroxide formed with dilute ammonia solution</th>
<th>Colour of precipitate formed</th>
<th>Does the precipitate redissolve in excess ammonia solution? (✓ or ✗)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminium ions</td>
<td>Aluminium hydroxide</td>
<td>White</td>
<td>✗</td>
</tr>
<tr>
<td>Lead(II) ions</td>
<td>Lead(II) hydroxide</td>
<td>White</td>
<td>✗</td>
</tr>
<tr>
<td>Magnesium ions</td>
<td>Magnesium hydroxide</td>
<td>White</td>
<td>✗</td>
</tr>
<tr>
<td>Iron(II) ions</td>
<td>Iron(II) hydroxide</td>
<td>Green</td>
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<td>Iron(III) hydroxide</td>
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</tbody>
</table>

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<table>
<thead>
<tr>
<th>Reaction between</th>
<th>Ionic equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solution containing calcium ions and dilute sodium hydroxide solution</td>
<td>Ca(^{2+}) (aq) + 2OH(^-) (aq) (\rightarrow) Ca(OH)(_2) (s)</td>
</tr>
<tr>
<td>Solution containing aluminium ions and dilute sodium hydroxide solution</td>
<td>Al(^{3+}) (aq) + 2OH(^-) (aq) (\rightarrow) Al(OH)(_3) (s)</td>
</tr>
<tr>
<td>Solution containing magnesium ions and dilute sodium hydroxide solution</td>
<td>Mg(^{2+}) (aq) + 2OH(^-) (aq) (\rightarrow) Mg(OH)(_2) (s)</td>
</tr>
</tbody>
</table>
Solution containing iron(II) ions and dilute sodium hydroxide solution

\[ \text{Fe}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Fe(OH)}_2 (\text{s}) \]

Solution containing iron(III) ions and ammonia solution

\[ \text{Fe}^{3+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Fe(OH)}_3 (\text{s}) \]

Solution containing copper(II) ions and ammonia solution

\[ \text{Cu}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Cu(OH)}_2 (\text{s}) \]

Solution containing zinc ions and ammonia solution

\[ \text{Zn}^{2+} (\text{aq}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{Zn(OH)}_2 (\text{s}) \]

Solution containing ammonium compound and dilute sodium hydroxide solution (with heating)

\[ \text{NH}_4^+ (\text{aq}) + \text{OH}^- (\text{aq}) \rightarrow \text{NH}_3 (\text{g}) + \text{H}_2\text{O} (\text{l}) \]

<table>
<thead>
<tr>
<th>Reaction between</th>
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</thead>
<tbody>
<tr>
<td>Dilute hydrochloric acid and dilute sodium hydroxide solution</td>
<td>( \text{HCl} (\text{aq}) + \text{NaOH} (\text{aq}) \rightarrow \text{NaCl} (\text{aq}) + \text{H}_2\text{O} (\text{l}) )</td>
</tr>
<tr>
<td>Dilute sulphuric acid and copper(II) oxide</td>
<td>( \text{H}_2\text{SO}_4 (\text{aq}) + \text{CuO} (\text{s}) \rightarrow \text{CuSO}_4 (\text{aq}) + \text{H}_2\text{O} (\text{l}) )</td>
</tr>
<tr>
<td>Dilute nitric acid and magnesium oxide</td>
<td>( 2\text{HNO}_3 (\text{aq}) + \text{MgO(s)} \rightarrow \text{Mg(NO}_3)_2 (\text{aq}) + \text{H}_2\text{O} (\text{l}) )</td>
</tr>
<tr>
<td>Zinc oxide and dilute sulphuric acid</td>
<td>( \text{ZnO(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{ZnSO}_4 (\text{aq}) + \text{H}_2\text{O} (\text{l}) )</td>
</tr>
<tr>
<td>Zinc carbonate and dilute sulphuric acid</td>
<td>( \text{ZnCO}_3 (\text{s}) + \text{H}_2\text{SO}_4 (\text{aq}) \rightarrow \text{ZnSO}_4 (\text{aq}) + \text{H}_2\text{O} (\text{l}) + \text{CO}_2 (\text{g}) )</td>
</tr>
</tbody>
</table>

**Section E  Structured Questions**
1 a) Solid citric acid has no effect on magnesium because it does not contain hydrogen ions. When citric acid dissolves in water, the molecules ionize to give hydrogen ions. These hydrogen ions are responsible for all the acidic properties.

b) When dilute sulphuric acid reacts with calcium carbonate, insoluble calcium sulphate forms. The calcium sulphate covers the surface of calcium carbonate and prevents further reaction.

c) Neutralization occurs if a weak acid is applied to the affected area. The heat of neutralization produced would have a bad effect on the affected area. Furthermore, the chemical may interfere with the follow-up medical treatment. Therefore the affected area should be washed with water only.

d) Ammonia gas contains no hydroxide ions. However, it reacts with water to give an alkaline solution.

e) Hydrochloric acid is a strong acid. It almost completely ionizes in water to give hydrogen ions and chloride ions. Ethanoic acid is a weak acid. It only partially ionizes in water, forming very few hydrogen ions. Therefore 0.1 M hydrochloric acid has a higher concentration of mobile ions than 0.1 M ethanoic acid. The electrical conductivity of a solution is more or less proportional to the concentration of mobile ions. Thus the electrical conductivity of 0.1 M hydrochloric acid is higher than that of 0.1 M ethanoic acid.

f) The conical flask should not be washed with the solution to be contained because the additional amount of solute remaining in the flask will affect the titration results.

2 a) Gas bubbles are given off. A blue solution results.

\[
\text{CuCO}_3(s) + 2\text{HNO}_3(aq) \rightarrow \text{Cu(NO}_3)_2(aq) + \text{CO}_2(g) + \text{H}_2\text{O(l)}
\]

b) The black copper(II) oxide dissolves. A blue solution results.

\[
\text{H}_2\text{SO}_4(aq) + \text{CuO(s)} \rightarrow \text{CuSO}_4(aq) + \text{H}_2\text{O(l)}
\]

c) A green precipitate forms.
Fe²⁺(aq) + 2OH⁻(aq) → Fe(OH)₂(s)

d) A white precipitate forms first. The precipitate dissolves in excess ammonia solution. This is due to the formation of a soluble complex salt.

Zn²⁺(aq) + 2OH⁻(aq) → Zn(OH)₂(s)

3  a) Add dilute hydrochloric acid to each solution. Sodium carbonate solution reacts to give gas bubbles while sodium hydroxide solution does not give gas bubbles.

b) Add a piece of pH paper to each solution. Dilute hydrochloric acid gives a lower pH value as it is a stronger acid.

4  a) Water makes the sodium hydroxide solution more dilute. The solution is thus less alkaline and its pH value decreases.

b) Magnesium reacts with the hydrogen ions in the dilute sulphuric acid. The acid is thus less acidic and its pH value increases.

5  a) The chicken foot becomes smaller. It turns somewhat 'translucent'.

b) The chicken foot breaks up.

c) This shows the corrosive nature of concentrated sodium hydroxide solution.

d) Wash the affected area with plenty of water immediately.

6  a) A –Iron
    B –Iron(II,III) oxide
    C –Iron(II) sulphate solution
    D –Hydrogen
    E –Iron(II) hydroxide

b) i) Fe(s) + H₂SO₄(aq) → FeSO₄(aq) + H₂(g)
    ii) Fe³⁺(aq) + 2OH⁻(aq) → Fe(OH)₂(s)
7  a) A is ammonium chloride.

A is an ammonium compound because its solution gives ammonia gas (a gas that turns moist red litmus paper blue) when heated with dilute sodium hydroxide solution.

\[
\text{NH}_4^+(aq) + \text{OH}^-(aq) \rightarrow \text{NH}_3(g) + \text{H}_2\text{O}(l)
\]

A is a chloride because it gives a white precipitate with silver nitrate solution.

\[
\text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s)
\]

b) B is calcium carbonate.

B is a calcium compound because it gives a brick red flame in flame test.

B is a carbonate because it gives carbon dioxide gas (a gas that turns limewater milky) when reacted with dilute hydrochloric acid.

\[
\text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)
\]

c) C – Ammonia
D – Silver chloride
E – Carbon dioxide
F – Calcium carbonate

8   a) 2 M sodium hydroxide solution

b) i) The reagents in bottles B and C are 2 M hydrochloric acid and 2 M ethanoic acid.

Test the solutions with pH papers.

ii) The pH value of 2 M hydrochloric acid is lower than that of 2 M ethanoic acid.

Hydrochloric acid is a strong acid. It almost completely ionizes in water.

Ethanoic acid is a weak acid. It only partially ionizes in water. 2 M hydrochloric
acid has a higher concentration of hydrogen ions and thus its pH value is lower.

c)  i) The reagents in bottles D and E are 2 M sodium chloride solution and distilled water. Evaporate the two reagents to dryness.

ii) The reagent that leaves white residue upon evaporation is sodium chloride solution.

The reagent that leaves nothing behind is distilled water.

9  a)  i) To ensure all the lead(II) ions are precipitated.

ii) To remove any excess sodium chloride solution that may adhere to the precipitate.

iii) To remove any water that may adhere to the precipitate.

b) Molar mass of Pb(NO$_3$)$_2$ = [207.0 + 2 x (14.0 + 3 x 16.0)] g mol$^{-1}$

= 331.0 g mol$^{-1}$

Number of moles of Pb(NO$_3$)$_2$ present = \( \frac{\text{Mass}}{\text{Molar mass}} \)

= \( \frac{8.94 \text{ g}}{331.0 \text{ g mol}^{-1}} \)

= 0.0270 mol

Number of moles of Pb$^{2+}$ present = 0.0270 mol

Pb$^{2+}$(aq) + 2Cl$^-$(aq) \( \rightarrow \) PbCl$_2$(s)

According to the equation, 1 mole of Pb$^{2+}$ reacts with 2 moles of Cl$^-$ to give 1 mole of PbCl$_2$.

\[ \therefore \text{Number of moles of PbCl}_2 \text{ obtained} = 0.0270 \text{ mol} \]

Molar mass of PbCl$_2$ = (207.0 + 2 x 35.5) g mol$^{-1}$

= 278.0 g mol$^{-1}$
Theoretical mass of PbCl$_2$ obtained = Number of moles of PbCl$_2$ x Molar mass of PbCl$_2$

= 0.0270 mol x 278.0 g mol$^{-1}$

= 7.51 g

c) The precipitation of lead(II) chloride is incomplete.

10 a) PbCO$_3$(s) + 2HNO$_3$(aq) $\rightarrow$ Pb(NO$_3$)$_2$(aq) + CO$_2$(g) + H$_2$O(l)

b) Pb$^{2+}$(aq) + SO$_4^{2-}$(aq) $\rightarrow$ PbSO$_4$(s)

c) To remove the sodium sulphate solution and other water soluble impurities that adhere to the precipitate.

d) Use filter papers.

e) We can represent the whole process by a sequence of steps.

HNO$_3$ Na$_2$SO$_4$
PbCO$_3$ $\rightarrow$ Pb(NO$_3$)$_2$ $\rightarrow$ PbSO$_4$ (not balanced)

For the overall reaction:

PbCO$_3$ $\rightarrow$ PbSO$_4$ (not balanced)

? g 9.09 g

Molar mass of PbSO$_4$ = (207.0 + 32.0 + 4 x 16.0) g mol$^{-1}$

= 303.0 g mol$^{-1}$

Number of moles of PbSO$_4$ obtained = \( \frac{\text{Mass}}{\text{Molar mass}} \)

= \( \frac{9.09 \text{ g}}{303.0 \text{ g mol}^{-1}} \)

= 0.0300 mol

1 mole of PbCO$_3$ produces 1 mole of PbSO$_4$.

\[ \therefore \] Number of moles of PbCO$_3$ = 0.0300 mol
Molar mass of PbCO₃ = (207.0 + 12.0 + 3 x 16.0) g mol⁻¹ = 267.0 g mol⁻¹

Mass of PbCO₃ present in the sample = Number of moles of PbCO₃ x Molar mass of PbCO₃

= 0.0300 mol x 267.0 g mol⁻¹

= 8.01 g

Percentage by mass of PbCO₃ in the sample = \( \frac{8.01 \text{ g}}{8.50 \text{ g}} \times 100\% \)

= 94.2%

11 a) Action of dilute hydrochloric acid on magnesium.

b) Titration between sodium carbonate solution and dilute sulphuric acid.

c) Action of silver nitrate solution on potassium bromide solution.

d) Action of dilute nitric acid on calcium carbonate to obtain calcium nitrate solution; then action of sodium sulphate solution on calcium nitrate solution to obtain calcium sulphate.

\[
\text{HNO}_3(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow \text{Ca(NO}_3)_2(\text{aq}) + \text{CaSO}_4(\text{s})
\]

12 a) From yellow to red

b) \( \text{Mg(OH)}_2(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + 2\text{H}_2\text{O(l)} \)

0.350 g 1.00 M

? cm³
Molar mass of Mg(OH)$_2$ = [24.0 + 2 x (16.0 + 1.0)] g mol$^{-1}$
= 58.0 g mol$^{-1}$

Number of moles of Mg(OH)$_2$ in one tablet = \frac{\text{Mass}}{\text{Molar mass}}
= \frac{0.350 \text{ g}}{58.0 \text{ g mol}^{-1}}
= 0.00603 \text{ mol}

According to the equation, 1 mole of Mg(OH)$_2$ requires 2 moles of HCl for complete neutralization.

i.e. Number of moles of HCl = 2 x Number of moles of Mg(OH)$_2$
= 2 x 0.00603 \text{ mol}
= 0.0121 \text{ mol}

Volume of 1.00 M HCl required for complete neutralization = \frac{\text{Number of moles of HCl}}{\text{Molarity of HCl}}
= \frac{0.0121 \text{ mol}}{1.00 \text{ mol dm}^{-3}}
= 0.0121 \text{ dm}^3
= 12.1 \text{ cm}^3

13 a) The acid would cause changes in the pH value of water and affect water life if discharged without treatment.

b) $\text{Na}_2\text{CO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow 2\text{NaCl}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$

0.4 M
25 \text{ dm}^3 \text{ per minute}

Number of moles of HCl discharged per minute = Molarity of solution x Volume of solution
According to the equation, 2 moles of HCl require 1 mole of Na2CO3 for complete reaction.

Number of moles of Na₂CO₃ required per minute = \( \frac{1}{2} \) x Number of moles of HCl discharged per minute

\[
= \frac{1}{2} \times 10 \text{ mol}
\]

\[
= 5 \text{ mol}
\]

Molar mass of Na₂CO₃ = (2 x 23.0 + 12.0 + 3 x 16.0) g mol\(^{-1}\)
= 106.0 g mol\(^{-1}\)

Mass of Na₂CO₃ required per minute = Number of moles of Na₂CO₃ x Molar mass of Na₂CO₃

\[
= 5 \text{ mol} \times 106.0 \text{ g mol}^{-1}
\]

\[
= 530 \text{ g}
\]

14  a) Pipette

b) Wash the pipette with water and then with sodium hydroxide solution.

c) From red to colourless
e) \[ \text{H}_2\text{SO}_4(\text{aq}) + 2\text{NaOH}(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \]

\[ ? \text{ M} \quad 0.600 \text{ M} \]

\[ 15.0 \text{ cm}^3 \quad 25.0 \text{ cm}^3 \]

Number of moles of NaOH in 25.0 cm\(^3\) solution = Molarity of solution \times Volume of solution

\[ = 0.600 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}^3 \]

\[ = 0.0150 \text{ mol} \]

According to the equation, 1 mole of H\(_2\)SO\(_4\) requires 2 moles of NaOH for complete neutralization.

i.e. Number of moles of H\(_2\)SO\(_4\) = \(\frac{1}{2}\) \times Number of moles of NaOH

\[ = \frac{1}{2} \times 0.0150 \text{ mol} \]

\[ = 0.00750 \text{ mol} \]

Molarity of sulphuric acid = \(\frac{\text{Number of moles of H}_2\text{SO}_4}{\text{Volume of solution}}\)

\[ = \frac{0.00750 \text{ mol}}{(15.0 / 1000) \text{ dm}^3} \]

\[ = 0.500 \text{ mol dm}^{-3} (\text{M}) \]

f) Sodium sulphate

g) Mix 25 cm\(^3\) of the sodium hydroxide solution and 15 cm\(^3\) of the sulphuric acid. Evaporate the reaction mixture slowly to obtain the salt.

15 a) Choose a 25.0 cm\(^3\) pipette for delivering the acid sample. Use a 250.0 cm\(^3\) volumetric
flask for dilution.

- Wash a 25.0 cm³ pipette first with water and then with the original acid.
- Deliver exactly 25.0 cm³ of the original acid into a 250.0 cm³ volumetric flask using the pipette and pipette filler.

- Add distilled water to the flask until the meniscus is about 2 cm below the graduation mark.
- Add distilled water using a dropper until the meniscus reaches the graduation mark.
- Stopper the flask. Turn it upside down several times to mix the solution well.

b) From colourless to red

c) \[
\frac{30.0 + 30.1 + 29.9}{3} \text{ cm}^3 = 30.0 \text{ cm}^3
\]

d) We can represent the monobasic acid solution by HX(aq).

\[
\begin{align*}
\text{HX(aq)} + \text{NaOH(aq)} & \rightarrow \text{NaX(aq)} + \text{H}_2\text{O(l)} \\
? \text{ M} & \quad 0.20 \text{ M} \\
25.0 \text{ cm}^3 & \quad 30.0 \text{ cm}^3 \\
\downarrow & \\
250.0 \text{ cm}^3 & \\
\downarrow & \\
(\text{used}) 25.0 \text{ cm}^3 & 
\end{align*}
\]

Number of moles of NaOH in 30.0 cm³ solution = Molarity of solution \times Volume of solution

\[
= 0.20 \text{ mol dm}^{-3} \times \frac{30.0}{1000} \text{ dm}^3
\]

= 0.0060 mol

According to the equation, 1 mole of HX requires 1 mole of NaOH for complete neutralization.

Number of moles of HX in 25.0 cm³ of the diluted sample = 0.0060 mol
Number of moles of HX in 250.0 cm$^3$ of the diluted sample
= Number of moles of HX in 25.0 cm$^3$ of the acid sample before dilution

= 0.0060 mol $\times$ \(\frac{250.0 \text{ cm}^3}{25.0 \text{ cm}^3}\)
= 0.060 mol

Molarity of the monobasic acid sample = \(\frac{\text{Number of moles of HX}}{\text{Volume of solution}}\)
= \(\frac{0.060 \text{ mol}}{(25.0 / 1000) \text{ dm}^3}\)
= 2.4 mol dm$^{-3}$ (M)

16 a) From yellow to red

b) \(\text{Na}_2\text{CO}_3(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)\)

11.20 g \hspace{1cm} 0.200 M
20.9 cm$^3$
\downarrow
250.0 cm$^3$
\downarrow
(used) 25.0 cm$^3$

Molar mass of \(\text{Na}_2\text{CO}_3\cdot n\text{H}_2\text{O}\) = \([2 \times 23.0 + 12.0 + 3 \times 16.0 + n(2 \times 1.0 + 16.0)] \text{ g mol}^{-1}\)

= \((106.0 + 18.0n) \text{ g mol}^{-1}\)

Number of moles of \(\text{Na}_2\text{CO}_3\) in 11.20 g solid = \(\frac{\text{Mass}}{\text{Molar mass}}\)

= \(\frac{11.20 \text{ g}}{(106.0 + 18.0n) \text{ g mol}^{-1}}\)
Number of moles of \( \text{Na}_2\text{CO}_3 \) in 25.0 cm\(^3\) solution = \( \frac{11.20}{(106.0 + 18.0n)} \) mol

Number of moles of \( \text{H}_2\text{SO}_4 \) in 20.9 cm\(^3\) solution = Molarity of solution \( \times \) Volume of solution

\[ = 0.200 \text{ mol dm}^{-3} \times \frac{20.9}{1000} \text{ dm}^3 \]

\[ = 0.00418 \text{ mol} \]

According to the equation, 1 mole of \( \text{Na}_2\text{CO}_3 \) requires 1 mole of \( \text{H}_2\text{SO}_4 \) for complete reaction.

i.e. Number of moles of \( \text{Na}_2\text{CO}_3 \) in 25.0 cm\(^3\) solution = 0.00418 mol

\[ \therefore \frac{1}{10} \times \frac{11.20}{(106.0 + 18.0n)} \text{ mol} = 0.00418 \text{ mol} \]

\[ n = 9 \]

17 a) A weak acid is an acid that only partly ionizes in water.

b) Let \( n \) be the basicity of the tartaric acid.

We can represent the tartaric acid by \( \text{H}_n\text{X} \).

\[ \text{H}_n\text{X(aq)} + n\text{KOH(aq)} \rightarrow \text{K}_n\text{X(aq)} + n\text{H}_2\text{O(l)} \]

0.990 g \( \quad \) 0.500 M

26.4 cm\(^3\)

Number of moles of tartaric acid used = \( \frac{\text{Mas}}{\text{Molar mass}} \)

\[ = \frac{0.990 \text{ g}}{150.0 \text{ g mol}^{-1}} \]

\[ = 0.00660 \text{ mol} \]
Number of moles of KOH required for neutralization = Molarity of solution x Volume of solution

\[ = 0.500 \text{ mol dm}^{-3} \times \frac{26.4}{1000} \text{ dm}^3 \]

= 0.0132 mol

According to the equation, 1 mole of HnX requires n moles of KOH for complete neutralization.

\[ \frac{\text{Number of moles of } H_nX}{\text{Number of moles of KOH}} = \frac{1}{n} \]

\[ = \frac{0.00660 \text{ mol}}{0.0132 \text{ mol}} \]

\[ \therefore n = 2 \]

\[ \therefore \] Tartaric acid is a dibasic acid.

c) The amount of distilled water does not affect the number of moles of alkali required for neutralization.

d) The hydrochloric acid has a lower pH value. This is because hydrochloric acid is a stronger acid. It ionizes almost completely in water. It contains a higher concentration of hydrogen ions.

18 a) \( \frac{20.1 + 20.0 + 19.9}{3} \) cm³ = 20.0 cm³

b) \[ CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l) \]

CH₃COOH(aq) + NaOH(aq) \( \rightarrow \) CH₃COONa(aq) + H₂O(l)

50.0 cm³  \quad 0.200 M

\downarrow \quad 20.0 \text{ cm}^3
Number of moles of NaOH in 20.0 cm$^3$ solution = Molarity of solution x Volume of solution 

\[ \text{Number of moles of NaOH} = 0.200 \text{ mol dm}^{-3} \times \frac{20.0}{1000} \text{ dm}^3 = 0.00400 \text{ mol} \]

According to the equation, 1 mole of CH$_3$COOH requires 1 mole of NaOH for complete neutralization.

i.e. \( \text{Number of moles of CH}_3\text{COOH in 25.0 cm}^3 \text{ of the diluted vinegar} = 0.00400 \text{ mol} \)

Number of moles of CH$_3$COOH in 250.0 cm$^3$ of the diluted vinegar

= Number of moles of CH$_3$COOH in 50.0 cm$^3$ of the vinegar before dilution

= \( 0.00400 \text{ mol} \times \frac{250.0 \text{ cm}^3}{25.0 \text{ cm}^3} \)

= 0.0400 mol

Concentration of ethanoic acid in the vinegar = \( \frac{\text{Number of moles of CH}_3\text{COOH}}{\text{Volume of solution}} \)

\[ \text{Concentration of ethanoic acid} = \frac{0.0400 \text{ mol}}{50.0 / 1000} \text{ dm}^3 = 0.800 \text{ mol dm}^{-3} (M) \]

\[ \text{c) Molar mass of CH}_3\text{COOH} = (2 \times 12.0 + 4 \times 1.0 + 2 \times 16.0) \text{ g mol}^{-1} = 60.0 \text{ g mol}^{-1} \]

Concentration of ethanoic acid in the vinegar = \( 0.800 \text{ mol dm}^{-3} \times 60.0 \text{ g mol}^{-1} \)

= 48.0 g dm$^{-3}$
d) Mass of 1 dm $^3$ of the vinegar = Density x Volume
  = 1 g cm$^{-3}$ x 1 000 cm$^3$
  = 1 000 g

1 dm$^3$ of the vinegar contain 48.0 g of ethanoic acid.

∴ Percentage by mass of ethanoic acid in the vinegar = \[
\frac{48.0 \text{ g}}{1 000 \text{ g}} \times 100\%
\] = 4.80%
Number of moles of NaOH in 20.0 cm\(^3\) solution = Molarity of solution \(\times\) Volume of solution

\[
= 1.05 \text{ mol dm}^{-3} \times \frac{20.0}{1000} \text{ dm}^3
\]

\[= 0.0210 \text{ mol}\]

According to the equation, 1 mole of HCl requires 1 mole of NaOH for complete neutralization.

\[\text{i.e.} \quad \text{Number of moles of HCl in the excess acid} = 0.0210 \text{ mol}\]

\[\text{Number of moles of HCl reacted with the egg shells} = (0.0400 - 0.0210) \text{ mol} = 0.0190 \text{ mol}\]

iii) \(\text{CaCO}_3(s) + 2\text{HCl(aq)} \rightarrow \text{CaCl}_2(\text{aq}) + \text{CO}_2(g) + \text{H}_2\text{O(l)}\)

\[
1.07 \text{ g} \quad 0.0190 \text{ mol}
\]

According to the equation, 1 mole of CaCO\(_3\) requires 2 moles of HCl for complete reaction.

\[\text{i.e.} \quad \text{Number of moles of CaCO}_3 \text{ in egg shells} = \frac{1}{2} \times \text{Number of moles of HCl}\]

\[= \frac{1}{2} \times 0.0190 \text{ mol} \]

\[= 0.00950 \text{ mol}\]

\[\text{Molar mass of CaCO}_3 = (40.0 + 12.0 + 3 \times 16.0) \text{ g mol}\(^{-1}\) = 100.0 \text{ g mol}\(^{-1}\)\]

\[\text{Mass of CaCO}_3 \text{ in the egg shells} = \text{Number of moles of CaCO}_3 \times \text{Molar mass of CaCO}_3\]

\[= 0.00950 \text{ mol} \times 100.0 \text{ g mol}\(^{-1}\) \]

\[= 0.950 \text{ g}\]
Percentage by mass of CaCO₃ in the egg shells = 
\[
\frac{0.950 \text{ g}}{1.07 \text{ g}} \times 100\% = 88.8\%
\]

20 a)  
- Wash the burette first with water and then with the alkali.
- Close the stopcock. Fill the burette with alkali through a funnel.
- Clamp the burette vertically in a stand.
- Open the stopcock for a few seconds so as to fill the tip of the burette with alkali.

b) \[ \text{H}_2\text{SO}_4 (aq) + 2\text{NaOH} (aq) \rightarrow \text{Na}_2\text{SO}_4 (aq) + 2\text{H}_2\text{O} (l) \]

\[ ? \text{ M} \quad 1.20 \text{ M} \]
\[ 10.0 \text{ cm}^3 \quad 23.8 \text{ cm}^3 \]
\[ \downarrow \]
\[ 250.0 \text{ cm}^3 \]
\[ \downarrow \]
\[ (\text{used}) 25.0 \text{ cm}^3 \]

Number of moles of NaOH in 23.8 cm³ solution = Molarity of solution \times \text{Volume of solution}

\[ = 1.20 \text{ mol dm}^{-3} \times \frac{23.8}{1000} \text{ dm}^3 \]

\[ = 0.0286 \text{ mol} \]

According to the equation, 1 mole of H₂SO₄ requires 2 moles of NaOH for complete neutralization.

i.e.  
Number of moles of H₂SO₄ in 25.0 cm³ diluted cleaner

\[ = \frac{1}{2} \times \text{Number of moles of NaOH} \]

\[ = \frac{1}{2} \times 0.0286 \text{ mol} \]
Number of moles of $\text{H}_2\text{SO}_4$ in 250.0 cm$^3$ diluted cleaner = 0.0143 mol x
\[ \frac{250.0 \text{ cm}^3}{25.0 \text{ cm}^3} \]
= 0.143 mol

Number of moles of $\text{H}_2\text{SO}_4$ in 10.0 cm$^3$ of the drain cleaner = 0.143 mol

Molarity of $\text{H}_2\text{SO}_4$ in the drain cleaner =
\[ \frac{\text{Number of moles of } \text{H}_2\text{SO}_4}{\text{Volume of solution}} \]
\[ = \frac{0.143 \text{ mol}}{(10.0 / 1000) \text{ dm}^3} \]
\[ = 14.3 \text{ mol dm}^{-3} \text{ (M)} \]

c) The cleaner contains sulphuric acid which is corrosive.

d) Wear safety glasses and protective gloves.

21 a) To make the concentration of the vinegar comparable to that of the alkali.

b) From colourless to red

c) \[ \text{CH}_3\text{COOH(aq)} + \text{NaOH(aq)} \rightarrow \text{CH}_3\text{COONa(aq)} + \text{H}_2\text{O(l)} \]

\[
\begin{array}{ccc}
25.0 \text{ cm}^3 & 0.121 \text{ M} & 17.2 \text{ cm}^3 \\
\end{array}
\]

\[ 250.0 \text{ cm}^3 \]
\[ \downarrow \]
(used) 25.0 cm$^3$

Number of moles of NaOH in 17.2 cm$^3$ solution = Molarity of solution x
Volume of solution
\[ = 0.121 \text{ mol dm}^{-3} x \frac{17.2}{1000} \text{ dm}^3 \]
According to the equation, 1 mole of CH$_3$COOH requires 1 mole of NaOH for complete neutralization.

\[ \text{i.e. Number of moles of CH}_3\text{COOH in 25.0 cm}^3 \text{ diluted vinegar} = 0.00208 \text{ mol} \]

\[ \text{Number of moles of CH}_3\text{COOH in 250.0 cm}^3 \text{ diluted vinegar} = 0.00208 \text{ mol} \times \frac{250.0 \text{ cm}^3}{25.0 \text{ cm}^3} = 0.0208 \text{ mol} \]

\[
\text{Molarity of CH}_3\text{COOH in vinegar} \quad = \quad \frac{\text{Number of moles of CH}_3\text{COOH}}{\text{Volume of solution}} \\
= \quad \frac{0.0208 \text{ mol}}{(25.0 / 1000) \text{ dm}^3} \\
= \quad 0.832 \text{ mol dm}^{-3} (M) 
\]

d) Brand A

\[
\text{Mass of CH}_3\text{COOH in one bottle} \quad = \quad 60 \text{ g dm}^{-3} \times \frac{300}{1000} \text{ dm}^3 \\
= \quad 18 \text{ g} 
\]

\[
\text{Price per g of CH}_3\text{COOH} \quad = \quad \frac{10}{18 \text{ g}} \\
= \quad 0.55 \text{ g}^{-1} 
\]

Brand B

\[
\text{Molar mass of CH}_3\text{COOH} \quad = \quad (2 \times 12.0 + 4 \times 1.0 + 2 \times 16.0) \text{ g mol}^{-1} \\
= \quad 60.0 \text{ g mol}^{-1} 
\]

\[
\text{Concentration of CH}_3\text{COOH in vinegar} \quad = \quad 0.832 \text{ mol dm}^{-3} \times 60.0 \text{ g mol}^{-1} 
\]
Mass of CH₃COOH in one bottle = 49.9 g dm⁻³ x \( \frac{600}{1000} \) dm³
= 29.9 g

Price per g of CH₃COOH = \( \frac{20}{29.9} \) g
= $0.67 g⁻¹

\[ \therefore \] Brand A is the better buy.

22  a) \( \text{Cu}^+\text{aq} + 2\text{OH}^-\text{aq} \rightarrow \text{Cu(OH)}_2\text{(s)} \)

b) 

c) To remove the hydroxide ions adhere to the copper(II) hydroxide.

d) Methyl orange / phenolphthalein

e) i) \( \text{NaOH(aq)} + \text{HNO}_3\text{(aq)} \rightarrow \text{NaNO}_3\text{(aq)} + \text{H}_2\text{O(l)} \)

\[ \text{Number of moles of HNO}_3 \text{ in } 15.5 \text{ cm}^3 \text{ solution} = \text{Molarity of solution} \times \text{Volume of solution} \]
\[ \text{Volume of solution} = 0.800 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}^3 \]

\[ = 0.0200 \text{ mol} \]

According to the equation, 1 mole of NaOH requires 1 mole of HNO3 for complete neutralization.

\[ \text{i.e. Number of moles of OH}^-\text{in the filtrate} = 0.00775 \text{ mol} \]

ii) Number of moles of OH\(^-\) added in Step I = Molarity of solution \(\times\) Volume of solution

\[ = 0.800 \text{ mol dm}^{-3} \times \frac{25.0}{1000} \text{ dm}^3 \]

\[ = 0.0200 \text{ mol} \]

iii) Number of moles of OH\(^-\) used for precipitation = (0.0200 - 0.00775) mol

\[ = 0.0123 \text{ mol} \]

\[ \text{Cu}^{2+}(aq) + 2\text{OH}^-(aq) \rightarrow \]

\[ 0.0123 \]

According to the equation, 1 mole of Cu\(^{2+}\) reacts with 2 moles of OH\(^-\) to give 1 mole of Cu(OH)\(_2\).

\[ \therefore \text{Number of moles of Cu}^{2+} \text{ in } \text{CuSO}_4 \text{ solution} = \frac{1}{2} \times \text{Number of moles of OH}^- \]

\[ = \frac{1}{2} \times 0.0123 \text{ mol} \]

\[ = 0.00615 \text{ mol} \]
Molarity of CuSO₄

\[
\text{Molarity of CuSO}_4 = \frac{\text{Number of moles of CuSO}_4}{\text{Volume of solution}}
\]

\[
= \frac{0.00615 \text{ mol}}{(25.0 / 1000) \text{ dm}^3}
\]

\[
= 0.246 \text{ mol dm}^{-3} \text{ (M)}
\]

\[\text{H}_2\text{SO}_4(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}(l)\]

\[\text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}(l)\]

\[\text{solution} = \text{Molarity of solution} \times \text{Volume of solution}\]

\[
= 0.1 \text{ mol dm}^{-3} \times \frac{50}{1000} \text{ dm}^3
\]

\[
= 0.005 \text{ mol}
\]

According to the equation, 1 mole of \text{H}_2\text{SO}_4 requires 2 moles of \text{NaOH} for complete neutralization.

\[\text{i.e. Number of moles of NaOH in 50 cm}^3\text{ solution} = 2 \times \text{Number of moles of H}_2\text{SO}_4\]

\[= 2 \times 0.005 \text{ mol}\]
\[ \text{Molarity of NaOH solution} = \frac{\text{Number of moles of NaOH}}{\text{Volume of solution}} \]

\[ = \frac{0.01 \text{ mol}}{(50 / 1000) \text{ dm}^3} \]

\[ = 0.2 \text{ mol dm}^{-3} \text{ (M)} \]

c) Sodium hydrogensulphate

d) By reaction with dilute sodium hydroxide solution

\[ \rightarrow \text{ZnSO}_4(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O(\text{l})} \]

24 a) \( \text{ZnCO}_3(\text{s}) + \text{H}_2\text{SO}_4(\text{aq}) \)

b) When no more gas bubbles are given off.

c) To concentrate the filtrate.

d) To allow time for the formation of bigger crystals.

ii) The crystals formed would be very small.
f) \[ \text{ZnCO}_3(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{ZnSO}_4(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l) \]\n
\[
\begin{align*}
1.50 \text{ M} & \quad ? \text{ g} \\
50.0 \text{ cm}^3 & \\
\end{align*}
\]

Number of moles of \(\text{H}_2\text{SO}_4\) in 50.0 cm\(^3\) solution = Molarity of solution \(\times\) Volume of solution

\[
= 1.50 \text{ mol dm}^{-3} \times \frac{50.0}{1000} \text{ dm}^3
\]

\[
= 0.0750 \text{ mol}
\]

According to the equation, 1 mole of \(\text{H}_2\text{SO}_4\) reacts with 1 mole of \(\text{ZnCO}_3\) to give 1 mole of \(\text{ZnSO}_4\).

i.e. Number of moles of \(\text{ZnSO}_4\) = 0.0750 mol

Molar mass of \(\text{ZnSO}_4\cdot7\text{H}_2\text{O}\) = \([65.4 + 32.0 + 4 \times 16.0 + 7 \times (2 \times 1.0 + 16.0)] \text{ g mol}^{-1}\)

\[
= 287.4 \text{ g mol}^{-1}
\]

Mass of \(\text{ZnSO}_4\cdot7\text{H}_2\text{O}\) = Number of moles of \(\text{ZnSO}_4\cdot7\text{H}_2\text{O}\) \(\times\) Molar mass of \(\text{ZnSO}_4\cdot7\text{H}_2\text{O}\)

\[
= 0.0750 \text{ mol} \times 287.4 \text{ g mol}^{-1}
\]

\[
= 21.6 \text{ g}
\]

25 a) An acid in solid form

b) The reaction is quite fast when \(t = 10\) s as the slope of the curve is quite steep. The reaction is slower when \(t = 50\) s as the slope of the curve is less steep. The reaction gets slower as the reaction proceeds. The reaction is over when \(t = 160\) s as no more gas is produced and the volume of gas collected no longer changes.

c) Increase the temperature
26  a) The fizzy drink tablet contains a solid acid and sodium hydrogencarbonate. The solid acid gives hydrogen ions $H^+$ when dissolved in water. The hydrogen ions react with sodium hydrogencarbonate to give carbon dioxide gas. This gives rise to the effervescence that occurs.

\[
\text{HCO}_3^-(\text{aq}) + H^+(\text{aq}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})
\]

b) Since carbon dioxide can escape through the cotton wool, the flask plus its contents get lighter as the reaction proceeds. To determine the rate of the reaction, we can measure the loss in the mass of the flask plus its contents at regular intervals.

c) ![](image)

d) Number of moles of NaHCO$_3$ in one tablet = \( \frac{\text{Mass of NaHCO}_3}{\text{Molar mass of NaHCO}_3} \)

\[
= \frac{1.25 \text{ g}}{84.0 \text{ g mol}^{-1}}
\]

\[
= 0.0149 \text{ mol}
\]

According to the equation, 1 mole of NaHCO$_3$ gives 1 mole of CO$_2$.

i.e. Number of moles of CO$_2$ given off from one tablet = 0.0149 mol

\[\therefore\] Mass of CO$_2$ given off = Number of moles of CO$_2$ x Molar mass of CO$_2$

\[
= 0.0149 \text{ mol} \times 44.0 \text{ g mol}^{-1}
\]

\[
= 0.656 \text{ g}
\]
e) The acid will react with sodium hydrogencarbonate in the presence of moisture.

27 a) \[ \text{Mg(s) + 2HCl(aq) \rightarrow MgCl}_2(aq) + \text{H}_2(g) \]

b) 

![Image of gas syringe and dilute hydrochloric acid]

c) i) ii) 

d) 

28 a) \[ \text{CaCO}_3(s) + 2\text{HCl(aq) \rightarrow CaCl}_2(aq) + \text{CO}_2(g) + \text{H}_2\text{O(l)} \]

Number of moles of \( \text{CaCO}_3 \) present = \( \frac{\text{Mass of } \text{CaCO}_3}{\text{Molar mass of } \text{CaCO}_3} \)

\[
= \frac{1.0 \text{ g}}{100.0 \text{ g mol}^{-1}}
= 0.010 \text{ mol}
\]

Number of moles of HCl in 30 cm³ solution = Molarity of solution x Volume of solution

\[
= 2.0 \text{ mol dm}^{-3} \times \frac{30}{1000} \text{ dm}^3
\]
According to the equation, 1 mole of CaCO$_3$ reacts with 2 moles of HCl. During the reaction, 0.010 mole of CaCO$_3$ reacts with 0.020 mole of HCl. Therefore HCl is in excess.

b) Use powdered marble instead of marble lumps.
   • Use more concentrated hydrochloric acid.

c) This is because when the sulphuric acid reacts with the marble lumps, insoluble calcium sulphate forms. The calcium sulphate covers the surface of the marble lumps and prevents further reaction.

29

a)

b) The reaction rate is faster at X than at Y. This is because the reaction gets slower as it proceeds.

c) The loss in mass of the contents of the reaction flask for sample A is greater than that for sample B, i.e. sample A produces more carbon dioxide gas. Therefore sample A has a higher purity of calcium carbonate than sample B.
   • The initial rate for sample B is faster than that for sample A. It is because sample B has a smaller particle size.
d) Mass of CO\(_2\) given off = 3.7 g

\[
\text{CaCO}_3(s) + 2\text{HCl(aq)} \rightarrow \text{CaCl}_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O(l)}
\]

Number of moles of CO\(_2\) given off = \(\frac{\text{Mass of CO}_2}{\text{Molar mass of CO}_2}\)

\[
= \frac{3.7 \text{ g}}{44.0 \text{ g mol}^{-1}}
\]

= 0.0841 mol

According to the equation, 1 mole of CaCO\(_3\) reacts with 2 moles of HCl to give 1 mole of CO\(_2\).

i.e. Number of moles of CaCO\(_3\) in sample A = 0.0841 mol

Mass of CaCO\(_3\) in sample A = Number of moles of CaCO\(_3\) \times \text{Molar mass of CaCO}_3

= 0.0841 mol \times 100.0 \text{ g mol}^{-1}

= 8.41 g

∴ Percentage purity of CaCO\(_3\) in sample A = \(\frac{8.41 \text{ g}}{10.0 \text{ g}} \times 100\%\)

= 84.1%

30 a) Mg(s) + 2HCl(aq) \rightarrow MgCl\(_2\)(aq) + H\(_2\)(g)

1 g 1 M
100 cm\(^2\)

Number of moles of Mg present = \(\frac{\text{Mass of Mg}}{\text{Molar mass of Mg}}\)

= \(\frac{1 \text{ g}}{24.0 \text{ g mol}^{-1}}\)

= 0.04 mol
Number of moles of HCl in 100 cm$^3$ solution = Molarity of solution x Volume of solution

= 1 mol dm$^{-3}$ x $\frac{100}{1000}$ dm$^3$

= 0.1 mol

According to the equation, 1 mole of Mg reacts with 2 moles of HCl.

During the reaction,

0.04 mole of Mg reacts with 0.08 mole of HCl. Therefore HCl is in excess.

b) All the magnesium dissolved.

c) The rate of the reactions would increase.

d) Magnesium reacts with the hydrogen ions in the acids. Sulphuric acid is a dibasic acid while hydrochloric acid is a monobasic acid. Therefore 1 M sulphuric acid has a higher concentration of hydrogen ions than 1 M hydrochloric acid. The time required for the reaction between magnesium and 1 M sulphuric acid to complete is less than 200 s.

Hydrochloric acid is a strong acid while ethanoic acid is a weak acid. Therefore 1 M hydrochloric acid has a higher concentration of hydrogen ions than 1 M ethanoic acid. The time required for the reaction between magnesium and 1 M ethanoic acid to complete is longer than 200 s.

31 a) The time required to reach a certain level of turbidity can be taken to be the reaction time.

b) i) To keep the total volume of each sample constant.

ii) To ensure that the only variable is the change in the concentration of sodium thiosulphate solution.

c)
First add sodium carbonate to the reagents. Effervescence occurs for both the 2 M hydrochloric acid and 2 M sulphuric acid. 2 M sulphuric acid gives a white precipitate when mixed with barium chloride solution. However, 2 M hydrochloric acid does not give a precipitate with barium chloride solution.

Add lead(II) nitrate solution to the remaining reagents. Both 2 M ammonia solution and 2 M sodium hydroxide solution give a white precipitate. The precipitate will dissolve in excess sodium hydroxide solution. However, the precipitate does not dissolve in excess ammonia solution.

The remaining reagent is distilled water.

Neutralization has many useful applications, including soil treatment, treatment of factory waste and treatment of indigestion.

Most plants grow best when the pH value of the soil is close to 7. Soil may become acidic due to various reasons, such as acid rain. If the soil is too acidic, most plants will not grow well in it. Farmers neutralize acidic soil by adding quicklime (calcium oxide) to it. If the soil becomes too alkaline, farmers may add ammonium sulphate to it.

Liquid waste from factories often contains acid. The waste will affect water life if it is discharged into rivers and lakes without treatment. We can add slaked lime (calcium hydroxide) to the waste to neutralize the acid. On the other hand, we can control the pH value of alkaline waste by adding sulphuric acid to it.
Our stomach is an organ for digesting food. It produces hydrochloric acid along with enzymes to digest the food. However, if the stomach produces too much hydrochloric acid, we will suffer from stomach ache and even indigestion.

Antacids are drugs that can cure the pain caused by excess acid in the stomach. Antacids contain bases such as magnesium hydroxide and aluminium hydroxide. The bases neutralize the excess acid in the stomach and so the pain can be relieved.

34 Weigh the sample using an electronic balance. Record the mass of the sample (X g). Put the sample into a beaker. Then add dilute sulphuric acid to the sample. Magnesium carbonate reacts with the acid and effervescence occurs. Silver remains undissolved because there is no reaction between silver and dilute sulphuric acid. Filter off the solid and wash it with distilled water. Then dry the solid in an oven. Finally weigh the solid using an electronic balance. Record the mass of the solid (Y g).

\[
\text{Percentage by mass of silver in the sample} = \frac{Y \text{ g}}{X \text{ g}} \times 100\%
\]

35 Copper(II) sulphate can be prepared by the reaction between dilute sulphuric acid and copper(II) oxide.

Place 20 cm³ of dilute sulphuric acid in a beaker. Add a spatula measure of copper(II) oxide. Warm and stir with a glass rod. Add more copper(II) oxide until some of it remain in the beaker. Remove the excess copper(II) oxide by filtration.

Heat the copper(II) sulphate solution to evaporate about half of the water. Set the concentrated solution aside to cool and crystallize. Filter the crystals from the remaining solution. Wash the crystals with a little distilled water. Dry the crystals using filter paper.

36 Calculate the mass of sodium carbonate solid required. Weigh an empty dry weighing bottle.
accurately. Add the solid to the bottle until we have approximately the mass calculated. Weigh the weighing bottle and the solid accurately. Tip out the solid into a beaker as much as possible. Reweigh the weighing bottle and any remaining solid accurately.

Add 100 cm$^3$ of distilled water to the beaker. Stir with a glass rod until all the solid has dissolved. After dissolving, transfer the solution into a volumetric flask with the aid of a filter funnel. Wash the beaker, the glass rod and the filter funnel with a little distilled water several times. Pour all the washings into the flask. Add distilled water to the flask until the meniscus is about 2 cm below the graduation mark. Add distilled water using a dropper until the meniscus reaches the graduation mark. Stopper the flask. Turn it upside down several times to mix the solution well.

37 First connect the pressure sensor to a data-logger interface linked to a computer. Then execute the data-logger software on the computer. Open a graph display with a plot of pressure against time.

Put a piece of clean magnesium ribbon into the suction flask. Pour 20 cm$^3$ of 2 M hydrochloric acid in a plastic bottle. Carefully put the plastic bottle into the suction flask. Seal the suction flask with a rubber stopper and a little vacuum sealant. Tilt the suction flask to mix the acid with the magnesium ribbon. Immediately start recording the change in pressure inside the suction flask with time for about 3 minutes.

Repeat the experiment with the same mass of magnesium and 20 cm$^3$ of dilute hydrochloric acid of different concentrations. Compare the plots of pressure against time for acid of different concentrations to find out how the concentration of hydrochloric acid affects the rate of its reaction with magnesium.