Answers

Chapter 1

Checkpoint 1.1

1. (a) Carbon dioxide

   (b) Nitrogen cannot be collected by downward delivery as it is less dense than air. Hydrogen chloride will dissolve in water as it is very soluble in water. Thus, only carbon dioxide will be collected in the gas jar.

2. (a) Ethanol can be separated from water by fractional distillation. Once all the ethanol has distilled off, water remains in the distillation flask.

   (b) Chromatography can be used to separate the pigments. They are separated when the more soluble pigments move more quickly with the solvent on the chromatography paper than the less soluble pigments.

   (c) The water can be obtained as the distillate through simple distillation. A hot saturated sugar solution will remain in the distillation flask. The sugar can be obtained through crystallisation by allowing the hot saturated solution to cool.

3. (a) The students obtained the chromatograms using different solvents.

   (b) (i) Her chromatogram showed only one spot.

   (ii) The R_f value of the impurity may be very close to the R_f value of the compound in the solvent used.

   (c) He can obtain a pure sample of the compound and carry out chromatography using the same solvent and temperature. He can then determine the R_f value of the compound from the chromatogram. The spot with a different R_f value will be the spot that belongs to the impurity.

Test Station

1. **D**

   A measuring cylinder can measure variable volumes of liquids to the nearest 0.5 cm^3. A pipette can measure fixed volumes of liquids, such as 10.0 cm^3 and 25.0 cm^3. A beaker cannot be used for measuring volumes of liquids accurately. It is mainly used to contain a solution. Only a burette can measure variable volumes of liquids to the nearest 0.05 cm^3.

2. **B**

   Impurities lower the melting point of a substance and raise its boiling point.

3. **C**

   Only the spot for compound **N** has the same R_f value as the spot for compound **K** when both solvents **X** and **Y** are used.

4. (a) Silver chloride can be separated from water by filtration. [1] Silver chloride is collected as the residue while water is collected as the filtrate. [1]

   (b) Sodium chloride can be separated from water by evaporation to dryness. [1] The mixture is heated until all the water boils off. The sodium chloride that remains is collected. [1]

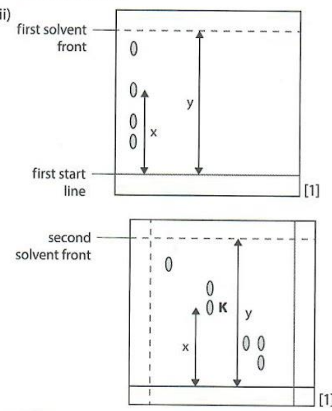
   (c) Isobutanol can be separated from water by fractional distillation. [1] Water, which has a lower boiling point than isobutanol, distils off first, while isobutanol remains in the distillation flask. [1]

   (d) Ethanal can be separated from water by simple distillation. [1] Ethanal, which has a much lower boiling point than water, distils over first while water remains in the distillation flask. [1]

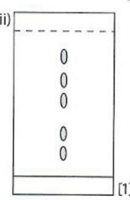
5. (a) Some substances may have very close or identical R_f values in one solvent and different R_f values in another solvent. Running the chromatogram a second time allows such substances to be separated. [1]

   (b) (i) 6 [1] (since there are 6 spots on the chromatogram that is run using the first solvent followed by the second solvent)

   (ii)



   (iii)



   (c) (i) The start line for each spot depends on its position in the second chromatogram. Hence, the distance travelled by the solvent is different for different spots. [1]

   (ii) The distance travelled by the solvent for some of the spots may be too short to allow the substances to be separated. [1]

Chapter 2

Checkpoint 2.1

1. (a) **A**: solid, solid

   B: solid, liquid

   C: liquid, gas

   D: solid, solid

   E: solid, solid

   F: gas, gas

   (b) **C**

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(c) **A**

(d) **A** and **E**. Both substances have high melting and boiling points.

2. (a) The melting point of substance **G** is 55 °C and its boiling point is 112.5 °C.

(b) (i) Solid

(ii) Liquid

(iii) Liquid and gas

(c)

(d) Substance **G** undergoes boiling and changes from the liquid state to the gaseous state between 20 min and 25 min. When substance **G** in the liquid state is heated, the particles with increased energy overcome the forces of attraction between the particles. The temperature of substance **G** remains constant until all the substance has boiled.

(e) More energy is needed to overcome the forces of attraction between the particles when substance **G** in the liquid state boils, compared to when the substance in the solid state melts. Hence, a longer time is required to boil substance **G** than to melt the same amount of the substance.

3. (a) At temperatures below the boiling point, only a small amount of particles located near the surface of a liquid have enough kinetic energy to overcome the forces of attraction between them and escape as a gas. This process is known as evaporation.

(b) Particles with greater mass move more slowly. Thus, only a small amount of particles have enough kinetic energy to overcome the forces of attraction between them.

(c) The forces of attraction between water particles are much stronger than the forces of attraction between ethanol particles. Thus, the water particles require more energy to overcome the forces of attraction between them.

(d) Boiling point

Test Station

1. **D**

The temperature –133 °C is higher than the boiling points of argon, helium, krypton and neon. Thus, these four substances would change from the liquid state to the gaseous state at –133 °C and only xenon would be in the liquid state.

2. **C**

Option **A** is not true. Substance **Z** is a liquid at 90 °C and 150 °C. This shows that its boiling point is above 150 °C.

Option **B** is not true. Substance **Z** is a solid at 30 °C and 50 °C, and a liquid at 90 °C. This shows that its melting point is between 50 °C and 90 °C.

Option **C** is true. Substance **Z** is a gas at 250 °C. Thus, its particles have more kinetic energy when it is a gas than when it is a solid.

Option **D** is not true. Substance **Z** is a solid at 50 °C and a liquid at 150 °C. Thus, its particles have more kinetic energy when it is a liquid than when it is a solid.

3. **A**

Since compound **E** is in the liquid state at 40 °C, its melting point is 40 °C or below 40 °C.

Since it is in the gaseous state at 260 °C, its boiling point is 260 °C or below 260 °C.

4. (a) The melting point of mercury is above –40 °C and equal to or lower than –38 °C. [1] The boiling point of mercury is above 356 °C and equal to or lower than 358 °C. [1]

(b) (i) When liquid mercury is heated, the mercury particles move about more quickly as they gain kinetic energy. [1] The distance between the mercury particles increases. Thus, the liquid mercury expands. [1]

(ii) When the partial vacuum in the capillary tube of the mercury thermometer is replaced with nitrogen gas, the nitrogen gas exerts a pressure on the liquid mercury and the mercury particles become closer together. [1] More energy is required for the particles to overcome the forces of attraction between them. [1] Hence, the boiling point of mercury is increased.

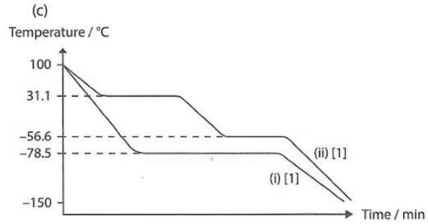
5. (a) (i) Sublimation [1]

(ii) Particles at the surface of dry ice have enough energy to overcome the forces of attraction between them. The solid thus changes into a gas. [1]

(b) (i) A liquid has a fixed volume, while a gas takes up the available volume of its container. Carbon dioxide gas occupies a larger volume than liquid carbon dioxide of the same mass. [1] Storing liquid carbon dioxide allows the size of fire extinguishers to be minimised and is thus preferred. [1]

(ii) Under high pressure, particles of carbon dioxide gas become closer together. This leads to an increase in the forces of attraction between the particles [1], until the forces of attraction become stronger than the kinetic energy of the particles. Thus, carbon dioxide gas changes into a liquid. [1]

(c)



6. (a) During evaporation, a small amount of water particles near the surface of sweat droplets on the skin have enough kinetic energy to overcome the forces of attraction between them. These particles escape as a gas and the remaining water particles have less kinetic energy. The temperature of the human body decreases and the body cools down. [1]

(b) (i) On humid days, there are more water particles in the air. With a higher concentration of water particles in the air, fewer water particles are able to evaporate from the surface of sweat droplets. [1]

(ii) Wind helps to remove water particles in the air that are near the surface of sweat droplets. [1] This lowers the concentration of water particles in the air surrounding the sweat droplets and allows water from sweat to evaporate easily. [1]

(c) Singapore is surrounded by the sea, while the Sahara Desert has very dry conditions. Thus, the air in Singapore has higher humidity than the air in the Sahara Desert. [1] As sweat evaporates more slowly where humidity is higher, one feels hotter in Singapore than in the Sahara Desert at the same temperature. [1]

(d) The tiny pores in the fabric allow water from sweat droplets to escape from the surface of the fabric and heat from the body to be transferred to the surroundings. Thus, the temperature of the body drops and the body cools down. [1]

Chapter 3

Checkpoint 3.1

1. Li: 3
   F^-: 9, 19, 9
   Ar: 18, 18, 22
   Ca^{2+}: 20, 40, 18
   I: 53, 127, 53
   Sn: 50, 50, 69
2. (a) 2, 7. Its first electron shell is fully filled with 2 electrons, while its valence electron shell is filled with 7 electrons.
   (b) E^-. **E** gains an electron to achieve a full valence electron shell and a stable electronic configuration of 2, 8.
3. (a) **X** is a neutron. **Y** is a proton. **Z** is an electron.
   (b) A hydrogen atom gains one electron to form a hydride ion.
   (c) **X**: 1
   Y: 1
   Z: 1

Test Station

1. **B**
   The particles belong to the same element as they have the same number of protons. They cannot be isotopes as isotopes have the same number of protons and electrons but different numbers of neutrons.
2. **C**
   A Mg atom loses two electrons to form a Mg^{2+} ion which has 10 electrons. A Na atom loses one electron to form a Na^+ ion which has 10 electrons. A Ne atom has 10 protons and hence 10 electrons.
3.

Element	Proton Number	Nucleon Number	Electronic Configuration
W	5	11	2, 3
X	6	12	2, 4
Y	15	31	2, 8, 5
Z	16	32	2, 8, 6

[1 mark for each row]

4. (a) **O** and **Q** [1]
   (b) **M** [1]
   (c) **N** and **O** [1]
   (d) **L** and **P** [1]
   (e) **Q** [1]
5. (a)

Sub-atomic Particle	Relative Charge	Relative Mass
neutron	0	1
proton	+1	1
electron	–1	$\frac{1}{1840}$

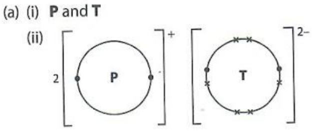
[1 mark for comparing the relative charge and mass of any two sub-atomic particles]

   (b) Lithium. [1] $^{7}_{3}Li$ [1]
   (c) The nucleus of the lithium atom has 3 protons, which are positively charged, and 4 neutrons, which are neutral. [1] The electron shells are filled with 3 electrons, which are negatively charged and move around the nucleus constantly. [1]

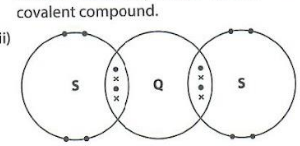
Chapter 4

Checkpoint 4.1

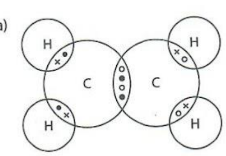
1. (a) (i) **P** and **T**
   (ii)



   (iii) Helium has a similar electronic configuration as **P**. Neon has a similar electronic configuration as **T**.
   (b) (i) $\mathbf{QS_2}$ is a covalent compound. **Q** and **S** are non-metals. Thus, they share electrons to form a covalent compound.
   (ii)



2. (a)



   (b) Each carbon atom in ethene has two electrons that are not involved in bonding. Hence, the total number of non-bonding electrons in ethene is 2 + 2 = 4.
   (c) Similarity: The total number of covalent bonds in both ethene and ethane is 6.
   Difference: There is a double covalent bond between the carbon atoms in ethene, whereas there is only a single covalent bond between the carbon atoms in ethane. / Four covalent bonds are formed between carbon and hydrogen in ethene, while six covalent bonds are formed between carbon and hydrogen in ethane. (Any one)
3. (a) **X** is an anion. **Y** is a cation.

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(b) The compound has a similar structure as sodium chloride. Its structure contains ions. This shows that it is an ionic compound. Sodium chloride is an ionic compound while hydrogen chloride is a covalent compound.

Test Station

1. **D**

Element **D** (aluminium) is a metal, while element **E** (oxygen) is a non-metal. Element **D** loses three electrons to form a $\mathbf{D}^{3+}$ ion, while element **E** gains two electrons to form a $\mathbf{E}^{2-}$ ion. Thus, ionic bonding occurs between elements **D** and **E**. Two $\mathbf{D}^{3+}$ ions and three $\mathbf{E}^{2-}$ ions are needed for the total charge of positive ions to cancel out the total charge of negative ions. Hence, the chemical formula of the compound is $\mathbf{D}_2\mathbf{E}_3$.

2. **C**

Option **C** is not true. Element **P** can share its valence electron with a non-metal to form a covalent compound. It can also lose the electron to form a positive ion, or gain an electron to form a negative ion.

Option **A** is true. A carbon atom shares its four valence electrons with four atoms of element **P** to form the covalent compound $C\mathbf{P}_4$.

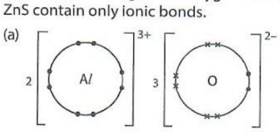
Option **B** is true. A magnesium atom loses its two valence electrons to form a Mg^{2+} ion, while an atom of element **P** gains one electron to form a $\mathbf{P}^-$ ion. Magnesium ions and $\mathbf{P}^-$ ions combine in the ratio of 1 : 2 to form the ionic compound $Mg\mathbf{P}_2$.

Option **D** is true. A potassium atom loses its valence electron to form a K^+ ion, while an atom of element **Q** gains three electrons to form a $\mathbf{Q}^{3-}$ ion. K^+ ions and $\mathbf{Q}^{3-}$ ions combine in the ratio of 3 : 1 to form the ionic compound $K_3\mathbf{Q}$.

3. **B**

In KNO_3, the NO_3^- ion contains covalent bonds formed between the nitrogen and oxygen atoms. CaO, NaCl and ZnS contain only ionic bonds.

4. (a)

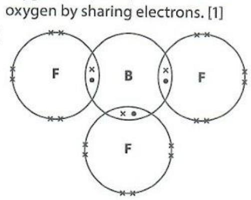


[1 mark each for Al^{3+} and O^{2-} ions drawn correctly. 1 mark for the correct number of Al^{3+} and O^{2-} ions.]

(b) 4 valence electrons. Since silicon has one more proton than aluminium, it will have one more electron. [1]

(c) No. Silicon is in Group 14 and is a non-metal like oxygen. [1] Thus, it forms chemical bonds with oxygen by sharing electrons. [1]

5. (a)



[1 mark for the correct number of electrons involved in bonding. 1 mark for the correct number of valence electrons.]

(b) Boron has only three valence electrons and hence can only share three electrons with three fluorine atoms. [1] Thus, boron would only have a maximum number of six valence electrons, which is two electrons short of a stable electronic configuration. [1]

(c) Aluminium is a metal while chlorine is a non-metal. Ionic bonds usually exist between a metal and a non-metal. [1] However, covalent bonds exist in aluminium chloride instead of ionic bonds. [1]

Chapter 5

Checkpoint 5.1

1. (a) Element **P**: 2, 8, 7

Element **Q**: 2, 8, 8, 1

(b) **QP**

(c) The compound would likely exist in the solid state at room temperature. It has a high melting point due to the large amount of energy required to overcome the strong forces of electrostatic attraction between the $\mathbf{Q}^+$ and $\mathbf{P}^-$ ions.

(d) Similarity: The element and compound are pure substances with fixed boiling and melting points.

Difference: The compound can be broken down into its constituent elements using chemical means, but the element cannot be broken down into simpler substances.

2. (a) Sodium chloride and magnesium oxide (Any two ionic compounds)

(b) At 100 °C, calcium oxide is in the solid state. It is unable to conduct electricity as its ions are held together in fixed positions and cannot function as mobile charge carriers.

At 3000 °C, calcium oxide is in the molten or liquid state. It is able to conduct electricity as its ions are free to move about and can function as mobile charge carriers.

Checkpoint 5.2

1. (a) **X**: Group 17

Z: Group 16

(b) **Y** has one more electron shell than **Z**.

(c) No. $\mathbf{YZX}_2$ does not have a lower boiling point than O_2 and Cl_2. It has a larger molecular size than O_2 and Cl_2 and hence stronger intermolecular forces of attraction. Since more energy is needed to overcome its intermolecular forces of attraction, $\mathbf{YZX}_2$ has a higher boiling point.

(d) **X**, **Y** and **Z** share electrons to form the covalent compound $\mathbf{YZX}_2$. There are no mobile electrons in the simple molecular structure of $\mathbf{YZX}_2$. Thus, the compound is not able to conduct electricity.

2. (a) The electrons in the structures of copper and zinc are mobile and able to carry charge. Thus, copper and zinc are able to conduct electricity.

(b) When a force is applied to copper, the layers of copper atoms slide over each other. This allows copper to change its shape. Thus, it is malleable.

(c) Brass is harder than copper and zinc.

Test Station

1. **C**

When an ionic compound is in the molten state, its ions are mobile and able to act as charge carriers. Thus, the

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compound is able to conduct electricity in the molten state. Since the compound is pure, it melts and boils at a fixed temperature.

A compound can only be broken down by chemical means, unlike a mixture which can be separated by physical means. A compound is made up of at least two elements.

2. **A**

Substance **V** has low melting and boiling points, does not conduct electricity and is not soluble in water. These properties show that it is a simple covalent substance.

Substance **W** conducts electricity in the liquid state but not in the solid state. Thus, it is likely to be an ionic compound. It can conduct electricity in the liquid state due to the presence of mobile ions.

Substance **Y** has high melting and boiling points and conducts electricity in the solid and liquid state. Thus, it is likely to be a metal.

Substance **X** does not conduct electricity in the solid and liquid state. This shows that it is not a metal.

3. **D**

The electrical conductivity of metals is due to the presence of mobile electrons in their structures.

4. (a) **S** and **U**. [1] They are able to conduct electricity in the solid state. [1]
   (b) **T**. [1] It has a range of melting points. [1]
   (c) (i) A compound is a substance made up of two or more elements chemically combined in a fixed ratio. [1]
   (ii) **P**. [1] It has a high melting point, does not conduct electricity in the solid state and is soluble in water. These are the properties of an ionic compound. [1]

5. (a) Ionic bonding [1]
   (b) When substance **J** is in the liquid and aqueous state, its ions are mobile [1] and hence able to carry charge [1]. Thus, substance **J** is able to conduct electricity in the liquid and aqueous state.
   (c) Substance **J** has a higher melting point than ice. Substance **J** has a giant ionic crystal lattice structure, in which strong electrostatic forces of attraction exist between its cations and anions. [1] Ice, which is water in the solid state, has a simple molecular structure, with weak intermolecular forces of attraction between its molecules. [1] More energy is required to overcome the strong electrostatic forces of attraction in substance **J** than the weak intermolecular forces of attraction between the water molecules in ice. [1]

Chapter 6

Checkpoint 6.1

1. (a) $N_2H_4(l) \longrightarrow N_2(g) + 2H_2(g)$
   (b) Hydrazine, nitrogen and hydrogen are all molecules. Thus, there are no ions involved in the reaction.

2. (a) $OH^-(aq) + NH_4^+(aq) \longrightarrow NH_3(g) + H_2O(l)$
   (b) $OH^-(aq) + H^+(aq) \longrightarrow H_2O(l)$

Test Station

1. **C**

The charge of one Ca^{2+} ion is balanced by the charge of one CO_3^{2-} ion. Thus, the chemical formula of the ionic compound formed is $CaCO_3$.

The charge of one Zn^{2+} ion is balanced by the charge of one CO_3^{2-} ion. Thus, the chemical formula of the ionic compound formed is $ZnCO_3$.

Carbon has a valency of 4 while hydrogen has a valency of 1. Thus, one carbon atom shares its four valence electrons with four hydrogen atoms to form the stable covalent molecule CH_4.

Sulfur has a valency of 2. Thus, one carbon atom shares its four valence electrons with two sulfur atoms to form a stable covalent molecule. Hence, the chemical formula of the covalent molecule formed is CS_2, not CS.

2. **C**

The nitrite ion, NO_2^-, has a valency of 1. The iron(III) ion, Fe^{3+}, has a valency of 3. The charge of one Fe^{3+} ion is balanced by the total charge of three NO_2^- ions. Thus, the chemical formula of iron(III) nitrite is $Fe(NO_2)_3$.

3. **D**

The chemical formula of carbon monoxide is CO. Only option **D** shows the correct and balanced chemical equations for the two reactions described.

4. (a)

Name of Ion	Chemical Formula of Ion
manganese(II)	Mn^{2+}
manganese(IV)	Mn^{4+}
manganese(V)	Mn^{5+}
manganese(VI)	Mn^{6+}
manganese(VII)	Mn^{7+}

[1 mark for all correct names of ions and 1 mark for all correct chemical formulae of ions]

   (b) $MnO_2(s) + 4HCl(aq) \longrightarrow MnCl_4(aq) + 2H_2O(l)$
   [1 mark for the correct chemical formulae of reactants and products. 1 mark for a balanced equation with correct state symbols.]
   (c) (i) $Mn_2(CO_3)_5 \longrightarrow 5CO_2 + MnO_2 + MnO_3$
   [1 mark for the correct chemical formulae of reactants and products. 1 mark for a balanced equation.]
   (ii) Manganese(VI) oxide, MnO_3 [1]

5. (a) Nitrogen monoxide: NO [1]
   Nitrogen dioxide: NO_2 [1]
   Dinitrogen oxide: N_2O [1]
   (b) $2N_2O \longrightarrow 2NO + N_2$ [1]
   The product that is an element is nitrogen. [1]
   (c) $2CO + 2NO \longrightarrow 2CO_2 + N_2$
   $4CO + 2NO_2 \longrightarrow 4CO_2 + N_2$
   [1 mark for balanced chemical equations and 1 mark for correct chemical formulae of reactants and products]

Chapter 7

Checkpoint 7.1

1. (a) Molar mass of CH_4
   $= 12 + (4 \times 1)$
   $= 16$ g/mol
   Mass of 0.5 mol of CH_4
   $= 0.5 \times 16$
   $= 8.00$ g

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(b) Molar mass of NH_3
$= 14 + (3 \times 1)$
$= 17$ g/mol
Mass of 0.5 mol of NH_3
$= 0.5 \times 17$
$= 8.50$ g
No. 0.5 mol of ammonia gas and 0.5 mol of methane gas do not have the same mass as they have different molar masses.

2. (a) Molar mass of **M**CO_3
$= \frac{62.5}{0.5}$
$= 125$ g/mol

(b) Let A_r of **M** be x.
M_r of **M**$CO_3 = 125$
$x + 12 + (3 \times 16) = 125$
$x = 125 - 60$
$x = 65$
Hence, **M** is zinc.

Checkpoint 7.2

1. $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$
Number of moles of CH_4
$= \frac{10}{1000} \div 24$
$= 0.000417$ mol
Number of moles of O_2
$= \frac{80}{1000} \div 24$
$= 0.00333$ mol
From the equation, 1 mol of CH_4 reacts with 2 mol of O_2.
Number of moles of O_2 required to react completely with CH_4
$= 2 \times 0.000417$
$= 0.000834$ mol
Since 0.00333 mol of O_2 is present, it is in excess. Thus, CH_4 is the limiting reactant.
From the equation, 1 mol of CH_4 reacts to form 2 mol of H_2O.
Number of moles of H_2O formed
$= 2 \times 0.000417$
$= 0.000834$ mol
Molar mass of H_2O
$= (2 \times 1) + 16$
$= 18$ g/mol
Mass of H_2O formed
$= 0.000834 \times 18$
$= 0.0150$ g

2. (a) Number of moles of NH_3
$= \frac{250}{1000} \div 24$
$= 0.0104$ mol
Concentration of NH_3 solution
$= 0.0104 \div \frac{200}{1000}$
$= 0.0521$ mol/dm^3

(b) Number of moles of NH_3 solution reacted
$= 0.0521 \times \frac{20}{1000}$
$= 0.00104$ mol
From the equation, 1 mol of NH_3 reacts with 1 mol of HNO_3.
Number of moles of HNO_3 reacted $= 0.00104$ mol
Volume of HNO_3 used in the titration
$= \frac{0.00104}{0.0400}$
$= 0.0260$ dm^3
$= 26.0$ cm^3

3. (a) $Na_2CO_3 + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O + CO_2$
Molar mass of Na_2CO_3
$= (2 \times 23) + 12 + (3 \times 16)$
$= 106$
Number of moles of Na_2CO_3 added
$= \frac{5}{106}$
$= 0.0472$ mol
Number of moles of H_2SO_4 present
$= \frac{20}{1000} \times 1.0$
$= 0.0200$ mol
From the equation, 1 mol of Na_2CO_3 reacts with 1 mol of H_2SO_4.
Number of moles of H_2SO_4 required to react with $Na_2CO_3 = 0.0472$ mol
Since only 0.0200 mol of H_2SO_4 is present, it is the limiting reactant.

(b) From the equation, 1 mol of H_2SO_4 reacts to form 1 mol of CO_2.
Number of moles of CO_2 = number of moles of H_2SO_4 $= 0.0200$ mol
Volume of CO_2 produced
$= 0.0200 \times 24$
$= 0.480$ dm^3

(c) Number of moles of H_2SO_4 reacted = number of moles of $Na_2CO_3 = 0.0200$ mol
Number of moles of excess Na_2CO_3
$= 0.0472 - 0.0200$
$= 0.0272$ mol
Volume of solution $= 20$ cm^3
$= 0.0200$ dm^3
Concentration of excess Na_2CO_3
$= \frac{0.0272}{0.0200}$
$= 1.36$ mol/dm^3

Test Station

1. **B**
Molar mass of $CaCO_3$
$= 40 + 12 + (3 \times 16)$
$= 100$ g/mol

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ber of moles of 20 g of $CaCO_3$

)

)

00 mol

Molar mass of NH_3

$= 14 + (3 \times 1)$

$= 17$ g/mol

Number of moles of 17 g of NH_3

$= \frac{17}{17}$

$= 1.00$ mol

Number of moles of 50 g of $CaCO_3$

$= \frac{50}{100}$

$= 0.500$ mol

Number of moles of 12 g of Mg

$= \frac{12}{24}$

$= 0.500$ mol

Number of moles of 54 g of Ag

$= \frac{54}{108}$

$= 0.500$ mol

Molar mass of H_2

$= 2 \times 1$

$= 2$ g/mol

Number of moles of 0.5 g of H_2

$= \frac{0.5}{2}$

$= 0.250$ mol

Molar mass of $CuSO_4$

$= 64 + 32 + (4 \times 16)$

$= 160$ g/mol

Number of moles of 320 g of $CuSO_4$

$= \frac{320}{160}$

$= 2.00$ mol

Molar mass of $FeCO_3$

$= 56 + 12 + (3 \times 16)$

$= 116$ g/mol

Number of moles of 177 g of $FeCO_3$

$= \frac{177}{116}$

$= 1.53$ mol

From the above, 50 g of $CaCO_3$ and 12 g of Mg have an equal number of moles.

2. **A**

Let the abundance of carbon-14 be x % and that of carbon-12 be $(100 - x)$ %.

Relative atomic mass of C = 12.01

$[(14 \times x) + 12(100 - x)] \div 100 = 12.01$

$14x + 1200 - 12x = 12.01 \times 100$

$2x = 1201 - 1200$

$2x = 1$

$x = 0.5$

The abundance of carbon-14 is 0.5 %.

3. **B**

Number of moles of HC*l* used in the titration

$= \frac{25}{1000} \times 0.800$

$= 0.0200$ mol

From the equation, 2 mol of HC*l* reacts with 1 mol of Na_2CO_3.

Number of moles of Na_2CO_3 reacted

$= \frac{0.0200}{2}$

$= 0.0100$ mol

Molar mass of Na_2CO_3

$= (2 \times 23) + 12 + (3 \times 16)$

$= 106$ g/mol

Mass of Na_2CO_3 used

$= 0.0100 \times 106$

$= 1.06$ g

4. (a) Molar mass of Na_2SO_4

$= (2 \times 23) + 32 + (4 \times 16)$

$= 142$ g/mol

Number of moles of Na_2SO_4

$= \frac{8.52}{142}$

$= 0.0600$ mol [1]

(b) Number of moles of $Na_2SO_4.xH_2O$

= Number of moles of Na_2SO_4

$= 0.0600$ mol

Relative molecular mass of $Na_2SO_4.xH_2O$

$= \frac{15}{0.600}$

$= 250$ [1]

(c) Relative molecular mass of Na_2SO_4

$= 142$ [1]

Relative molecular mass of xH_2O

$= 250 - 142$

$= 108$ [1]

(d) Relative molecular mass of H_2O

$= (2 \times 1) + 16$

$= 18$

$x = \frac{108}{18}$

$= 6$ [1]

5. (a) (i) Number of moles of H_2SO_4

$= \frac{50}{1000} \times 1.0$

$= 0.0500$ mol [1]

(ii) Molar mass of Na_2CO_3

$= (2 \times 23) + 12 + (3 \times 16)$

$= 106$ g/mol

Number of moles of Na_2CO_3 used

$= \frac{4.2}{106}$

$= 0.0396$ mol [1]

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(b) From the equation, 1 mol of H_2SO_4 reacts with 1 mol of Na_2CO_3.

Number of moles of H_2SO_4 that reacted completely with Na_2CO_3 = 0.0396 mol [1]

Since 0.0500 mol of H_2SO_4 is present, it is in excess. Thus, Na_2CO_3 is the limiting reactant. [1]

(c) From the equation, 1 mol of Na_2CO_3 reacts to form 1 mol of CO_2.

Number of moles of CO_2 produced = 0.0396 mol [1]

Volume of CO_2 produced

$= 0.0396 \times 24$

$= 0.950\ dm^3$ [1]

(d) Number of moles of excess H_2SO_4

$= 0.0500 - 0.0396$

$= 0.0104$ mol [1]

From the equation, 1 mol of H_2SO_4 reacts with 2 mol of KOH.

Number of moles of KOH required to react completely with H_2SO_4

$= 2 \times 0.0104$

$= 0.0208$ mol [1]

Volume of KOH required

$= \frac{0.0208}{1.0}$

$= 0.0208\ dm^3$

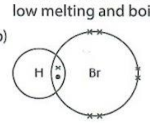
$= 20.8\ cm^3$ [1]

Chapter 8

Checkpoint 8.1

1. **A** is copper(II) carbonate. **B** is zinc. **C** is copper(II) oxide.
2. (a) Hydrogen bromide has a simple molecular structure. A small amount of energy is needed to overcome the weak forces of attraction between hydrogen bromide molecules. Hence, hydrogen bromide has low melting and boiling points.

(b)



(c) H^+, Br^-

(d) In aqueous solution, hydrogen bromide exists as H^+ and Br^- ions, which move freely and are able to conduct electricity. However, hydrogen bromide gas exists as simple molecules. There are no free electrons to conduct electricity.

3. (a) $(NH_4)_2CO_3 \longrightarrow 2NH_3 + CO_2 + H_2O$

(b) (i) Carbon dioxide

(ii) Ammonia

Checkpoint 8.2

1. (a) $4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$

(b) (i) Purple

(ii) Dilute hydrochloric acid (or other dilute acids)

2. (a) Calcium hydroxide would react with the acids present in the soil and thus increase the pH of the soil.

(b) (i) $Ca(OH)_2 + (NH_4)_2SO_4 \longrightarrow CaSO_4 + 2NH_3 + 2H_2O$

(ii) If ammonium sulfate was added immediately after calcium hydroxide was added to the soil, calcium hydroxide would react with ammonium sulfate to form calcium sulfate, water and ammonia gas. Since calcium hydroxide would not react with the acids in soil, the pH of the soil would not increase. As ammonia gas would be given off, the nitrogen from the fertiliser would be lost and not enter the soil as nutrients for the crops.

Test Station

1. **C**

Magnesium loses two valence electrons to form the Mg^{2+} ion. The chemical formula of hydrogen gas is H_2. Only option **D** shows the chemical formulae of magnesium and hydrogen correctly.

To write the chemical equation, determine the products first. The products are magnesium nitrate and hydrogen. Then, construct the equation by writing the chemical formulae and state symbols for all the reactants and products and balancing the equation if required.

$Mg(s) + 2HNO_3(aq) \longrightarrow Mg(NO_3)_2(aq) + H_2(g)$

To obtain the ionic equation, write the chemical formulae of the constituent ions of the reactant and product that are in aqueous state.

$Mg(s) + 2H^+(aq) + 2NO_3^-(aq) \longrightarrow$
$Mg^{2+}(aq) + 2NO_3^-(aq) + H_2(g)$

Then, remove the chemical formulae of the spectator ions from the equation.

$Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_2(g)$

2. **B**

Option **A** is incorrect. Solution **P** has a higher concentration of H^+ ions than solution **Q**. Thus, it reacts more vigorously with magnesium.

Option **B** is correct. The lower the pH, the higher the concentration of H^+ ions in the solution.

Option **C** is incorrect. Carbon dioxide is given off when acids react with ammonium carbonate.

Option **D** is incorrect. Only solution **P** turns Universal Indicator red as its pH is 2. Solution **Q** turns Universal Indicator orange as its pH is 3.9.

3. **D**

Oxides **X** and **Y** are gases. Thus, they are likely non-metallic oxides which are either acidic oxides or neutral oxides. Oxide **X** is a neutral oxide as it does not react with aqueous potassium hydroxide or dilute hydrochloric acid. Oxide **Y** is an acidic oxide as it reacts with aqueous potassium hydroxide but not with dilute hydrochloric acid.

4. (a)

Name of Fertiliser	Chemical Formula of Fertiliser	Ions Present
ammonium nitrate	NH_4NO_3	NH_4^+ and NO_3^-
ammonium sulfate	$(NH_4)_2SO_4$	NH_4^+ and SO_4^{2-}
potassium nitrate	KNO_3	K^+ and NO_3^-
calcium phosphate	$Ca_3(PO_4)_2$	Ca^{2+} and PO_4^{3-}

[$\frac{1}{2}$ mark for each answer]

(b) The alkali is aqueous ammonia and the acid is dilute nitric acid. [1]

5. (a) $3HCl + Al(OH)_3 \longrightarrow AlCl_3 + 3H_2O$ [1]

(b) Number of moles of $Al(OH)_3$ required per day

$= \frac{0.08}{3}$

$= 0.026667$ mol [1]

Relative molecular mass of $Al(OH)_3$

$= 27 + (3 \times 16) + (3 \times 1)$

$= 78$

Mass of $Al(OH)_3$ required

$= 0.026667 \times 78$

$= 2.08$ g [1]

Volume of antacid required

$= \frac{2.08}{2.6} \times 100$

$= 80.0$ cm^3 [1]

(c) $CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$ [1]

Chapter 9

Checkpoint 9.1

1. (a) (i) Add dilute nitric acid, followed by aqueous silver nitrate. If chloride ions are present, a white precipitate will be formed.

$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$

(ii) Add dilute nitric acid, followed by aqueous barium nitrate. If sulfate ions are present, a white precipitate will be formed.

$Ba^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow BaSO_4(s)$

(b) (i) A white precipitate is formed. It dissolves in excess aqueous sodium hydroxide to form a colourless solution.

(ii) A red-brown precipitate is formed. It is insoluble in excess aqueous ammonia.

(iii) A yellow precipitate is formed.

2. (a) Add aqueous barium nitrate. It will react with dilute sulfuric acid to form a white precipitate of barium sulfate. No precipitate will be observed when aqueous barium nitrate is added to dilute nitric acid.

(b) Add aqueous ammonia in excess. It will react with aqueous aluminium nitrate to form a white precipitate, which is insoluble in excess aqueous ammonia. Aqueous ammonia will react with aqueous zinc nitrate to form a white precipitate, which is soluble in excess aqueous ammonia.

3. (a) (i) Copper(II) nitrate

(ii) Carbon dioxide

(iii) Copper(II) hydroxide

(iv) Ammonia

(b) The light blue precipitate dissolves to form a dark blue solution.

Test Station

1. **C**

A white precipitate is formed when a few drops of each alkali is added to the solutions containing Zn^{2+} and Al^{3+} ions. For the solution containing Zn^{2+} ions, the white precipitate dissolves in excess aqueous sodium hydroxide and excess aqueous ammonia to form a colourless solution. For the solution containing Al^{3+} ions, the precipitate dissolves in excess aqueous sodium hydroxide to form a colourless solution, but does not dissolve in excess aqueous ammonia.

2. **A**

Carbon dioxide is colourless and odourless. It is an acidic gas and hence turns damp blue litmus paper red.

3. **C**

The cation in **X** reacted with aqueous sodium hydroxide to form a white precipitate, which was not soluble in excess sodium hydroxide. Hence, the cation is not aluminium. Since barium hydroxide and barium sulfate are insoluble, the cation in **X** should be barium. The anion in **X** reacted with aqueous silver nitrate to form a white precipitate. Hence, the anion is chloride, not nitrate. Thus, **X** is barium chloride.

When **Y** was heated with aqueous sodium hydroxide, ammonia gas was evolved. Hence, the cation in **Y** is ammonium, not zinc. The anion in **Y** reacted with aqueous silver nitrate to form a yellow precipitate but does not react with aqueous barium nitrate. Hence, the anion is iodide, not chloride or sulfate. Thus, **Y** is ammonium iodide.

4. (a) Fe^{2+} [1]

(b) He would observe the formation of a green precipitate. The green precipitate would not be soluble in excess aqueous ammonia. [1]

(c) Cl^- [1]

(d) Iron(II) chloride, $FeCl_2$ [1]

5. Add aqueous sodium hydroxide to a small amount of solution **R** in a test tube. Warm the mixture and test any gas evolved with damp red litmus paper. [1] If the ammonium ion is present in solution **R**, ammonia gas will be given off and turn damp red litmus paper blue. [1]

Continue heating the mixture until no more gas is given off. [1] Add a few pieces of aluminium foil to the mixture and warm the mixture gently. [1] Test any gas given off with damp red litmus paper. If the nitrate ion is present in solution **R**, ammonia gas will be given off and turn damp red litmus paper blue. [1]

Chapter 10

Checkpoint 10.1

1. Fluorine is reduced. The oxidation state of fluorine decreases from 0 in F_2 to −1 in HF. / F_2 gains electrons to form F^-. / F_2 gains hydrogen to form HF. (Any one)

2. (a) $Mg(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$

(b) Magnesium is oxidised. The oxidation state of magnesium increases from 0 in Mg to +2 in $MgSO_4$. / Mg loses electrons to form Mg^{2+}. (Any one)

3. (a) The oxidising agent is Br_2. The reducing agent is P.

(b) The oxidising agent is SiO_2. The reducing agent is C.

(c) The oxidising agent is Al^{3+}. The reducing agent is Li.

Test Station

1. **B**

In option **B**, Cl is oxidised as its oxidation state increases from −1 in KCl to 0 in Cl_2.

In options **A** and **C**, the oxidation state of Cl decreases from 0 in Cl_2 and +5 in $KClO_3$ to −1 in NaCl and KCl respectively.

In option **D**, the oxidation state of Cl remains as −1 in the reactant HCl and the product NaCl.

2. **D**

Iron(III) chloride is an oxidising agent because it causes the oxidation state of Fe to increase from 0 in Fe to +2 in $FeCl_2$. Iron(III) chloride itself is reduced because the oxidation state of Fe decreases from +3 in $FeCl_3$ to +2 in $FeCl_2$.

3. (a) Let the oxidation state of N in NO be x.

Oxidation state of O = −2

Sum of the oxidation states of all atoms in NO = 0

$x + (-2) = 0$

$x = +2$

The oxidation state of nitrogen in NO is +2. [1]

Let the oxidation state of N in NO be y.

Sum of the oxidation states of all atoms in NO_2 = 0

$y + [2 \times (-2)] = 0$

$y = +4$

The oxidation state of nitrogen in NO_2 is +4. [1]

(b) SO_2 is oxidised because it gains oxygen to become SO_3. [1] NO_2 is reduced because it loses oxygen to become NO. [1]

(c) $2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$ [1]

4. (a) Iodine was formed and appeared yellow-brown when dissolved in water. [1]

(b)

Element	Oxidation State in Reactant	Oxidation State in Product(s)
copper	+2	+1
sulfur	+6	+6
oxygen	−2	−2
potassium	+1	+1
iodine	−1	−1 and 0

[1 mark for each row]

(c) Iodine is oxidised as its oxidation state increases from −1 in KI to 0 in I_2. [1] Copper is reduced as its oxidation state decreases from +2 in $CuSO_4$ to +1 in CuI. [1]

5. (a) Iron. It is oxidised as its oxidation state increases from +2 in FeO to +2.67 in Fe_3O_4. [1] It is reduced as its oxidation state decreases from +2 in FeO to 0 in Fe. [1]

(b) Nitrogen. It is oxidised as its oxidation state increases from +4 in NO_2 to +5 in HNO_3. [1] It is reduced as its oxidation state decreases from +4 in NO_2 to +3 in HNO_2. [1]

(c) Carbon. It is oxidised as its oxidation state increases from +3 in $H_2C_2O_4$ to +4 in CO_2. [1] It is reduced as its oxidation state decreases from +3 in $H_2C_2O_4$ to +2 in CO. [1]

Chapter 11

Checkpoint 11.1

1. (a) Statement 2 is correct. Covalent chlorides formed from aluminium, silicon, phosphorus and sulfur hydrolyse, while ionic chlorides formed from sodium and magnesium dissolve in water.

Statement 1 is incorrect. Chlorides of metals, such as aluminium, and chlorides of non-metals, such as silicon, phosphorus and sulfur, hydrolyse.

(b) Argon does not form any compound as it has the electronic configuration of a noble gas.

2. (a) (i) Group 14

(ii) Group 18

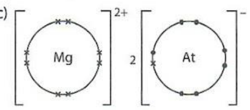
(b) (i) **ZX**

(ii) Ionic bonding

3. (a) At_2

(b) Astatide, At^-

(c)



Test Station

1. **C**

Na and Mg have fewer outer shell electrons and hence smaller group numbers than the other four elements. The proton number of the element increases from N to Ne in Period 2 and from Na to Mg in Period 3. N, O, F and Ne have 2 electron shells each, while Na and Mg have 3 electron shells each.

2. **A**

Bromine is found in Group 17 and Period 4 of the periodic table. Element **Z** is found in the period above Period 4 since it has a lower boiling point. **Z** has a smaller proton number and is more reactive than bromine. Thus, it can displace bromine from potassium bromide.

Z has a lower boiling point than bromine, which is a liquid at room temperature. Thus, **Z** is not likely to be a solid at room temperature.

Z has 7 valence electrons. Thus, it gains an electron to form the Z^- ion when it reacts with a metal.

3. **D**

A reactive halogen displaces a less reactive halogen from its salt solution. Since fluorine is more reactive than chlorine, it displaces chlorine from potassium chloride.

$F_2 + 2KCl \longrightarrow 2KF + Cl_2$

Thus, gas **X** is chlorine, which is yellow-green.

Since chlorine is more reactive than iodine, it displaces iodine from sodium iodide.

$2NaI + Cl_2 \longrightarrow 2NaCl + I_2$

Thus, gas **Y** is iodine, which is violet.

4. (a) lithium: 1; sodium: 2; potassium: 3 [1]

fluorine: 2; chlorine: 3; bromine: 4 [1]

(b) The atomic radii of the elements increase down the group [1] as the number of electron shells increases [1].

(c) (i) When aqueous chlorine is added to potassium bromide, the colourless solution turns red-brown. [1] When aqueous chlorine is added to potassium iodide, the colourless solution turns brown. [1]

(ii) $Cl_2(aq) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$

$Cl_2(aq) + 2I^-(aq) \longrightarrow 2Cl^-(aq) + I_2(aq)$

[$\frac{1}{2}$ mark for each ionic equation]

5. (a) Selenium atom has 6 valence electrons. Thus, its group number is equal to 10 plus the number of valence electrons it has. [1]

(b) It can form ionic bonds with a metal by gaining two electrons to achieve the stable noble gas electronic configuration. [1] It can also form covalent bonds with a non-metal by sharing its valence electrons. [1]

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(c) The damp blue litmus paper would turn red while the damp red litmus paper would remain red. [1] Selenium dioxide is an acidic oxide and dissolves in water to form an acidic solution. [1]

(d) Calcium [1]

Chapter 12

Checkpoint 12.1

1. (a) Lead is above hydrogen in the reactivity series, so it reacts with nitric acid. Silver is an unreactive metal and is below hydrogen in the reactivity series. Thus, it does not react with nitric acid and is removed as residue.
   (b) Sulfuric acid reacts with lead particles to form a layer of lead(II) sulfate. This layer coats the lead particles and prevents further reaction.
2. (a) **X**, zinc, iron, copper
   (b) $Fe(s) + Cu^{2+}(aq) \longrightarrow Cu(s) + Fe^{2+}(aq)$
   Iron is more reactive than copper. Thus, it displaces copper from copper(II) nitrate, and iron(II) sulfate and copper solid are formed.
   (c) Magnesium / Calcium / Aluminium (Any one)

Checkpoint 12.2

1. (a) $2PbO + C \longrightarrow 2Pb + CO_2$
   (b) Carbon. It reduces lead(II) oxide to lead while itself is oxidised to carbon dioxide.
   (c) Lead is less reactive than carbon. Thus, lead(II) oxide can be reduced by carbon to lead. Magnesium is more reactive than carbon. Thus, it is extracted from its ore by electrolysis.
2. (a) Presence of water and oxygen
   (b) Paint and plastic form a protective layer around iron, preventing it from coming into contact with water and oxygen.
   (c) Acidic gases released in industrial areas dissolve in rainwater to form acid rain. Iron rusts more quickly in the presence of acids.

Test Station

1. **A**
   Copper is unreactive and hence has no reaction with dilute hydrochloric acid as shown in test tube 2. Magnesium is more reactive than iron and hence reacts more vigorously with dilute hydrochloric acid. Thus, more bubbles are formed in test tube 3 than in test tube 1.
2. **A**
   Calcium is the most reactive metal among the three metals. It reacts with cold water while chromium and lead do not. Chromium is more reactive than lead as it is able to displace lead from lead(II) nitrate solution.
3. **C**
   Q and **R** are more reactive than **P** as they react with dilute hydrochloric acid. **Q** is more reactive than **R** as it cannot be extracted by reduction with carbon and is thus likely to be more reactive than carbon.
4. (a) Metal **Y** is more reactive than copper. Thus, it displaced copper from aqueous copper(II) sulfate to form a brown precipitate. [1] Metal **Z** reacted with the water in aqueous copper(II) sulfate to form an alkali and hydrogen gas. Thus, bubbles of hydrogen gas were observed. [1]
   (b) $\mathbf{Y} + CuSO_4 \longrightarrow Cu + \mathbf{Y}SO_4$ [1]
   $\mathbf{Z} + 2H_2O \longrightarrow \mathbf{Z}(OH)_2 + H_2$ [1]
   (c) **X**, copper, **Y**, **Z** [1]
5. (a) $Co(s) + 2HCl(aq) \longrightarrow CoCl_2(aq) + H_2(g)$ [1]
   (b) (i) Number of moles of cobalt used $= \frac{2.95}{59}$
   $= 0.05$ mol [1]
   (ii) From the equation, 1 mol of cobalt reacts to produce 1 mol of hydrogen gas.
   Number of moles of hydrogen = number of moles of cobalt = 0.05 mol [1]
   Volume of hydrogen gas produced
   $= 0.05 \times 24\ dm^3 = 12\ dm^3$ [1]
   (c) (i) **P**, cobalt, **Q** [1]
   (ii) More bubbles are observed during the reaction of metal **Q** with hydrochloric acid than during the reaction of cobalt with hydrochloric acid. [1]

Chapter 13

Checkpoint 13.1

1. Reactions I, II and III are exothermic. Reactions I and II are combustion reactions while reaction III is a neutralisation reaction. These reactions give out heat.
   Reaction IV is endothermic. Thermal decomposition of calcium carbonate requires heat.
2. (a) $N_2(g) + 2O_2(g) \longrightarrow 2NO_2(g)$
   (b) Lightning during a thunderstorm provides energy for nitrogen to react with oxygen.
   (c) The reaction is endothermic as energy is required for nitrogen and oxygen in air to react.

Test Station

1. **C**
   Heat is taken in to melt butter. Mixing of concentrated sulfuric acid with water, dissolving of anhydrous copper(II) sulfate in water and neutralisation between ammonia and hydrogen chloride give out heat.
2. **C**
   The vapour deposition of I_2 is exothermic. The thermal decomposition of $CaCO_3$ and sublimation of NH_4Cl are endothermic.
3. **D**
   Since the dissolution of ammonium nitrate in water is endothermic, the temperature of the surroundings will drop. After all the ammonium nitrate has dissolved in the water, the temperature of the surroundings will rise back to room temperature.
4. (a) The reaction between solid ammonium chloride and hydrated barium hydroxide is endothermic. [1] The reaction took in heat from the surroundings [1], causing the layer of water between the bottle and the wooden block to freeze into ice. [1]
   (b) (i) Ammonia [1]
   (ii) $2NH_4Cl + Ba(OH)_2 \longrightarrow BaCl_2 + 2NH_3 + 2H_2O$ [1]
5. (a) (i) $NaOH + HCl \longrightarrow NaCl + H_2O$ [1]
   (ii) $2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$ [1]

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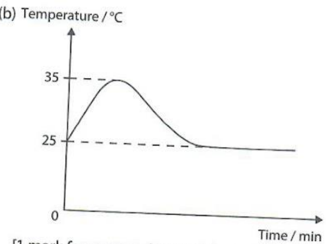
(b)



[1 mark for correct shape of the graph and 1 mark for correctly labelled graph]

Chapter 14

Checkpoint 14.1

1. (a) $ZnCO_3(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + CO_2(g) + H_2O(l)$

Molar mass of $ZnCO_3 = 65 + 12 + (3 \times 16) = 125$ g/mol

Number of moles of $ZnCO_3 = \frac{5}{125} = 0.0400$ mol

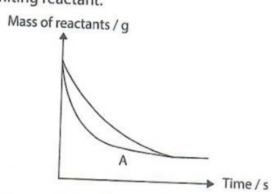
Number of moles of $HCl = \frac{40}{1000} \times 1.0 = 0.0400$ mol

From the equation, mole ratio of $ZnCO_3 : HCl = 1 : 2$

Number of moles of HCl required to react with $ZnCO_3$ $= 2 \times 0.0400 = 0.0800$ mol

Since only 0.0400 mol of HCl is present, it is the limiting reactant.

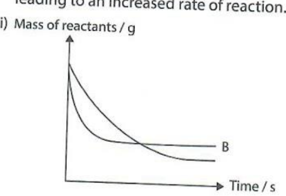
(b) (i)



(ii) The gradient of graph A is steeper as the reaction is faster. An increase in temperature results in reactant particles possessing more kinetic energy, which allows them to collide more frequently. The number of reactant particles colliding with energy greater than or equal to the activation energy increases. This causes an increase in the frequency of effective collisions, resulting in a faster reaction. Temperature does not affect the amount of products formed.

(c) (i) The rate of reaction increased. As the concentration of hydrochloric acid was increased, there were more reactant particles present per unit volume. The particles were closer to each other, resulting in more frequent collisions. The frequency of effective collisions also increased, leading to an increased rate of reaction.

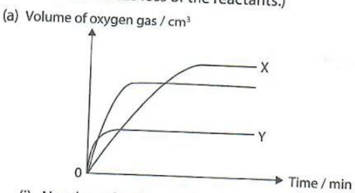
(ii)



(Number of moles of $HCl = \frac{20}{1000} \times 1.5 = 0.0300$ mol

As the number of moles of hydrochloric acid used had decreased from 0.0400 mol to 0.0300 mol, less zinc carbonate reacted with the acid, resulting in a smaller mass loss of the reactants.)

2. (a)



(i) Number of moles of 80 cm³ 1.0 mol/dm³ hydrogen peroxide $= \frac{80}{1000} \times 1.0 = 0.0800$ mol

Number of moles of 20 cm³ 0.8 mol/dm³ hydrogen peroxide $= \frac{20}{1000} \times 0.8 = 0.0160$ mol

Total number of moles of hydrogen peroxide $= 0.08 + 0.016 = 0.0960$ mol

Final concentration of 100 cm³ hydrogen peroxide $= \frac{0.096}{0.1} = 0.960$ mol/dm³

Since the final concentration of hydrogen peroxide is slightly lower than 1.0 mol/dm³, graph X would have a less steep gradient than the original graph. As the volume of hydrogen peroxide has slightly increased, the volume of oxygen produced would be slightly greater than the original volume of oxygen gas produced.

(ii) Number of moles of 80 cm³ 1.0 mol/dm³ hydrogen peroxide = 0.0800 mol

Number of moles of 20 cm³ 2.0 mol/dm³ hydrogen peroxide $= \frac{20}{1000} \times 2.0 = 0.0400$ mol

Since the number of moles of hydrogen peroxide is halved, graph Y would show half the original volume of oxygen gas produced. As the concentration of hydrogen peroxide is increased, the rate of reaction would increase. Thus, graph Y would have a steeper gradient than the original graph.

(b) If the concentration of hydrogen peroxide is decreased, there would be fewer reactant particles per unit volume. The reactant particles would be further apart from one another, resulting in a lower frequency of collisions and hence a lower frequency of effective collisions. Thus, the rate of reaction would decrease.

Test Station

1. **C**

The powdered magnesium in experiment 1 has a larger surface area in contact with nitric acid than the magnesium ribbon in experiments 2 and 3, resulting in a higher frequency of effective collisions. Thus, experiment 1 has a faster rate of reaction than experiment 2.

The volume of hydrogen gas produced in experiments 1 and 2 is the same as nitric acid of the same volume and concentration is used.

Number of moles of nitric acid in experiments 1 and 2

$= 1.0 \times \frac{20}{1000}$

$= 0.0200$ mol

Number of moles of nitric acid in experiment 3

$= 0.5 \times \frac{80}{1000}$

$= 0.0400$ mol

Experiment 3 produces twice the volume of hydrogen gas produced in experiments 1 and 2. Since the concentration of nitric acid used in experiment 3 is half the concentration of nitric acid used in experiment 2, experiment 3 takes place more slowly than experiment 2.

2. **A**

Bubbling more chlorine gas into the reaction mixture increases the concentration of aqueous chlorine, thus increasing the rate of reaction. Increasing the pressure on the reaction mixture does not affect the rate of reaction as there are no gaseous reactants present. Lowering the temperature of the reaction mixture using an ice bath decreases the rate of reaction.

3. **C**

Decreasing the particle size of a zinc ribbon increases the surface area in contact with hydrochloric acid, thus increasing the rate of reaction. Increasing the concentration and temperature of hydrochloric acid also increases the rate of reaction. Increasing the pressure on the reactants does not affect the rate of reaction as there are no gaseous reactants involved.

4. (a) Hot dilute sulfuric acid was used to increase the rate of reaction. [1] The high temperature of sulfuric acid used resulted in reactant particles possessing more kinetic energy which allowed them to collide more frequently. The number of reactant particles colliding with energy greater than or equal to the activation energy also increased. [1] Thus, the frequency of effective collisions increased, resulting in a faster reaction. [1]

   (b) (i) The same volume of hydrogen gas was produced in both experiments 1 and 2 as the same number of moles of sulfuric acid, which is the limiting reactant, was used. [1] The graph for experiment 2 has a steeper gradient than the graph for experiment 1 [1] because zinc, which was used in experiment 2, is a more reactive metal than iron, which was used in experiment 1. [1]

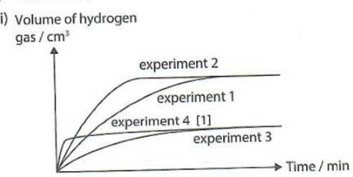
   (ii) Lead [1]

   (iii) $Pb(s) + H_2SO_4(aq) \longrightarrow PbSO_4(s) + H_2(g)$ [1]

   Lead is less reactive than both iron and zinc. Thus, it reacted more slowly and less vigorously than iron and zinc. [1] Less hydrogen gas was produced as the insoluble layer of lead(II) sulfate that formed on the strip prevented the lead from reacting further. [1]

   (c) (i) Calcium [1]

   (ii)



5. (a) No. Based on the information on experiments 1 and 2 [1], when the concentration of hydrochloric acid was increased from 0.100 mol/dm³ to 0.200 mol/dm³ and the concentration of sodium thiosulfate was kept constant, the time taken for the cross to "disappear" remained at 102 s. [1]

   (b) (i) Yes. Based on the information on experiments 1 and 3, when the concentration of sodium thiosulfate was increased from 0.100 mol/dm³ to 0.200 mol/dm³ and the concentration of hydrochloric acid was kept constant, the time taken for the cross to "disappear" decreased to 51 s. [1] This showed that increasing the concentration of sodium thiosulfate increased the rate of reaction. [1]

   (ii) When the concentration of sodium thiosulfate was increased, there were more reactant particles present per unit volume. [1] As the particles were closer to one another, the frequency of collisions and hence the frequency of effective collisions increased, thus increasing the rate of reaction. [1]

   (c) 25.5 s [1]

   (d) By observing the change in colour of the reaction mixture from brown to colourless as iodine is reduced to iodide ions during the reaction [1]

Chapter 15

Checkpoint 15.1

1. (a) Carbon and hydrogen

   (b) Fossil fuels cannot be readily replaced by natural means as they take millions of years to form.

   (c) No. Combustion of fossil fuels releases carbon dioxide. Carbon dioxide is a greenhouse gas which contributes to global warming.

2. (a) 40–75 °C

   (b) No. The petrol fraction boils over a range of temperatures and does not have a fixed boiling point.

3. (a) The petroleum fractions are miscible and have different boiling points.

   (b) **X**, **Y**, **Z**. Based on the positions of the fractions in the fractionating column, it can be inferred that **X** has the smallest molecular size while **Z** has the largest molecular size. The fraction with the smallest molecular size contains the weakest intermolecular forces of attraction, which require the least energy to overcome. Thus, the boiling point of **X** is the lowest.

   (c) Each petroleum fraction does not have a fixed boiling point. This shows that it is a mixture, not a pure compound.

Test Station

1. **B**

Option 1 is true. Fossil fuels are fuels obtained from the fossils of organisms that are buried over millions of years.

Option 2 is not true. Natural gas, which is a type of fossil fuel, consists mainly of methane gas.

Option 3 is not true. Fossil fuels are not renewable as they cannot be replaced at the same rate at which they are used.

Option 4 is true. Fossil fuels contain carbon and hydrogen. Thus, they are known as hydrocarbons.

2. **B**

Since **P** is collected above **Q** in the fractionating column, this shows that **P** has a smaller molecular size with fewer carbon atoms and weaker intermolecular forces of attraction than **Q**. Hence, **P** burns more easily and has a lower boiling point than **Q**.

3. **B**

Since **Z** flows less easily than **X**, it has a larger molecular size and stronger intermolecular forces of attraction. Thus, **Z** has a higher boiling point than **X**.

Since **Y** burns more easily than **X** and **Z**, it has a smaller molecular size with fewer carbon atoms and weaker intermolecular forces of attraction. Thus, **Y** has a lower boiling point than **X** and **Z**.

Since **W** burns less easily than **Z**, it has a larger molecular size with more carbon atoms and stronger intermolecular forces of attraction. Thus, **W** has a higher boiling point than **Z**.

Hence, the hydrocarbons arranged in order of decreasing boiling points is **W, Z, X, Y**.

4. (a) Boiling point. [1] Small hydrocarbons with lower boiling points condense at lower temperatures and are collected at the top of the fractionating column. [1] Large hydrocarbons with higher boiling points condense at higher temperatures and are collected at the bottom of the fractionating column. [1]

(b) Diesel oil, lubricating oil and bitumen. [1] The demand for these fractions is lower than the supply. Hence, the excess supply of these fractions can be used in cracking to form petrol. [1]

5. (a) Fermentation [1]

(b) Combustion of petrol releases carbon dioxide, which contributes to global warming. The carbon dioxide released during the combustion of bioethanol is absorbed by crops during photosynthesis. [1] The crops are in turn used as feedstock in the production of ethanol. [1]

(c) Growing crops for the manufacture of biofuels requires the clearing of land. This may destroy habitats and affect ecosystems. [1]

Chapter 16

Checkpoint 16.1

1. (a) Alkanes

(b) Carbon dioxide and water

(c) (i) *x* is 19 and *y* is 38.

(ii) Propene

2. (a) $2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$

(b) Steam was produced during the combustion of ethane. It then condensed on the glass base of the beaker.

(c) Number of moles of $C_2H_6 = \frac{5}{(2 \times 12) + (6 \times 1)}$
$= 0.16667$ mol

From the equation, mole ratio of $C_2H_6 : O_2 = 2 : 7$

Number of moles of $O_2 = \frac{7}{2} \times 0.16667 = 0.58333$ mol

Volume of O_2 required $= 0.58333 \times 24 = 14\ dm^3$

Checkpoint 16.2

1. (a) Temperature of 500 °C to 700 °C, pressure of 1 atm and aluminium oxide catalyst

(b) Hydrocarbon **P** has a higher boiling point than compound **Q** because it has a larger molecular size and stronger intermolecular forces of attraction. Thus, more energy is required to overcome the intermolecular forces of attraction between the molecules of hydrocarbon **P**.

(c) Compounds **Q** and **R** are likely to be an alkane and an alkene. Aqueous bromine is used to test for unsaturation. Cracking can convert a long-chain hydrocarbon into a short-chain alkane and a short-chain alkene.

2. (a) Unsaturated fats refer to fats that contain hydrocarbon chains with carbon–carbon double bonds in each chain.

(b) Add a sample of saturated fats and a sample of unsaturated fats to aqueous bromine separately. Saturated fats will not cause any change in the brown colour of aqueous bromine. Unsaturated fats will cause aqueous bromine to be decolourised.

(c) A diet containing fats made from oleic acid is healthier. Fats made from oleic acid are unsaturated as oleic acid contains carbon–carbon double bonds and fewer hydrogen atoms than stearic acid.

Test Station

1. **A**

The organic compound is an unsaturated hydrocarbon which can undergo combustion and an addition reaction.

2. **C**

Cracking breaks down long-chain hydrocarbons to shorter-chain hydrocarbons.

3. **C**

Based on Figure 16.2, the hydrocarbon that has a boiling point close to 0 °C contains four carbon atoms. The hydrocarbons in options **B** and **C** have four carbon atoms, but only the one in option **C** is an alkane.

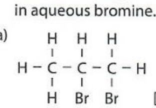
4. (a) (i) **W** and **Y** [1]

(ii) Ultraviolet light [1]

(b) **X**. It contains a carbon–carbon double bond which allows it to undergo an addition reaction with bromine. [1]

(c) Add aqueous bromine to compounds **X** and **Y** separately. [1] **X** will cause aqueous bromine to be decolourised. [1] **Y** will not cause any colour change in aqueous bromine. [1]

5. (a)



[1]

(b) Addition polymerisation [1]

(c) $C_8H_{18} \longrightarrow C_5H_{12} + C_3H_6$ [1]

(d) Cracking produces shorter-chain alkanes such as petrol and shorter-chain alkenes. Petrol is used as fuel while shorter-chain alkenes are used as the starting materials in some industrial processes. [1] Cracking also produces hydrogen gas, which is used as fuel. [1]

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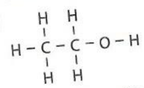
Chapter 17

Checkpoint 17.1

1. (a) (i) Ethanol, C_2H_5OH

(ii) 

(b) $CH_3CH_2OH + O_2 \longrightarrow CH_3COOH + H_2O$

(c) Yeast is added to a solution of glucose in a stoppered conical flask. The temperature of the mixture is kept at 37 °C. The conical flask is connected to a test tube with limewater using a delivery tube to prevent oxygen in the air from entering the conical flask.

(d) **X** can undergo combustion to produce heat for powering vehicles.

2. (a) It contains carbon–carbon double bonds.

(b) (i) Effervescence is observed.

(ii) Brown aqueous bromine is decolourised.

(c) $C_6H_{12}O_6 \longrightarrow 2C_2H_5OH + 2CO_2$

Test Station

1. **B**

The oxidation of propanol forms propanoic acid which contains three carbon atoms and a –COOH group.

2. **A**

Compound **Y** contains two hydroxyl (–OH) groups which are joined to the carbon atoms at the end of its carbon chain. Thus, it undergoes oxidation by potassium manganate(VII), which is decolourised.

Compound **Y** also contains a carbon–carbon double bond. Thus, it causes aqueous bromine to be decolourised.

Compound **Y** is not acidic as it does not contain any –COOH group. Thus, it does not turn blue litmus paper red.

3. **B**

The organic compound is butanoic acid. It can be formed by the oxidation of butanol. It has a lower pH value than aqueous ammonia. It does not turn red litmus paper blue as it is acidic. It does not undergo an addition reaction with bromine as it does not contain any carbon–carbon double bond.

4. (a) Fractional distillation [1]

(b) This prevents the oxidation of ethanol to ethanoic acid. [1]

(c) Ethanol is obtained from the fermentation of glucose which uses sugarcane as the starting material. Sugarcane is a renewable resource as it can be replaced within a short period of time. [1]

5. (a) The amount of ethanol used as a fuel increased from 2000 to 2009 and decreased from 2009 to 2015. [1]

(b) (i) M_r of $C_2H_5OH = (2 \times 12) + (5 \times 1) + 16 + 1 = 46$ [1]

(ii) Mass of C_2H_5OH in 20 g of petrol $= 25\% \times 20 = 5$ g [1]

Number of moles of $C_2H_5OH = \frac{5}{46} = 0.1087$ mol

From the equation, mole ratio of $C_2H_5OH : CO_2 = 1 : 2$

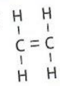
Number of moles of $CO_2 = 2 \times 0.1087 = 0.2174$ mol [1]

Volume of CO_2 produced $= 0.2174 \times 24 = 5.22$ dm^3 [1]

(c) Petroleum is a finite and non-renewable source. The use of biofuels helps to conserve petroleum. [1]

Chapter 18

Checkpoint 18.1

1. (a) (i) 

(ii) The carbon–carbon double bond in each ethene monomer breaks under high temperature and pressure. The monomers are joined together by carbon–carbon single bonds to form poly(ethene).

(b) Ethene causes brown aqueous bromine to be decolourised, but poly(ethene) does not.

(c) Relative molecular mass of each repeat unit, C_2H_4

$= (2 \times 12) + (4 \times 1)$

$= 28$

Number of ethene molecules in the sample

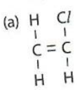
$= \frac{56\,000}{28}$

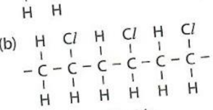
$= 2000$

Since each ethene molecule contains two carbon atoms, number of carbon atoms in 2000 ethene molecules

$= 2000 \times 2$

$= 4000$

2. (a) 

(b) 

(c) Hydrogen chloride

Checkpoint 18.2

1. (a) Percentage of plastic waste recycled in 2018

$= \frac{718}{2558} \times 100\%$

$= 28\%$

(b) Recycling plastics can reduce land pollution and the amount of land required for the disposal of plastics. It can also reduce pollution of water bodies. It can conserve the energy and raw materials used for making plastics.

(c) Poly(ethene) waste undergoes cracking at a high temperature and in the presence of a catalyst to form short chains of alkanes and alkenes.

2. (a) Recycling plastics may cause environmental issues if the wastewater generated from the process is discharged into water bodies without being treated. Recycling can be expensive as costs are incurred for collecting, transporting, sorting and cleaning the plastic waste.

(b) Recycling and awareness programmes can be introduced to the public to teach them the correct way of recycling.

Test Station

1. **B**

Only the compound in option **B** contains a carbon–carbon double bond which is required for addition polymerisation to take place.

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2. **A**

Options **A** and **D** show the correct repeat unit of the addition polymer. However, option **A** shows two repeat units while option **D** shows three repeat units.

3. **B**

Option **A** is true. Polystyrene is obtained through addition polymerisation of an alkene.

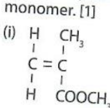
Option **B** is not true. The structure of polystyrene contains only one type of repeat unit. Thus, the polymer is formed from one monomer only.

Option **C** is true. Most plastics are non-biodegradable.

Option **D** is true. The monomer used for addition polymerisation usually contains a carbon–carbon double bond or a carbon–carbon triple bond.

4. (a) A Perspex molecule has strong covalent bonds between the atoms and strong intermolecular forces of attraction which require a large amount of energy to overcome. [1] The monomer molecules have weaker intermolecular forces of attraction which require a smaller amount of energy to overcome. Thus, Perspex has a higher melting point than its monomer. [1]

(b) (i)

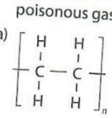


[1]

(ii) The carbon–carbon double bond in each monomer of Perspex breaks under high temperature and pressure. [1] The monomers are joined together by carbon–carbon single bonds to form Perspex. [1]

(c) Carbon monoxide is produced during incomplete combustion of Perspex. Carbon monoxide is a poisonous gas which can cause death. [1]

5. (a)



[1]

(b) Dibromoethane [1], $C_2H_4Br_2$ [1]

(c) $C_2H_4 + Br_2 \longrightarrow C_2H_4Br_2$ [1]

Chapter 19

Checkpoint 19.1

1. (a) A red-brown solid formed on the iron wool. The water level in the boiling tube rose.

(b) Volume of oxygen in the boiling tube = 21 % × 100 = 21 cm^3

Volume of air that remained in the boiling tube = 100 − 21 = 79 cm^3

2. (a) More fossil fuels were burnt during winter to generate electricity for powering heaters.

(b) Sulfur dioxide reacts with oxygen in the air to form acidic compounds which dissolve in rainwater. This results in the formation of acid rain, which corrodes buildings made from calcium carbonate and metals.

Checkpoint 19.2

1. (a) Methane: 1745 ppb, 0.471 %

Nitrous oxide: 312 ppb

(b) (i) Decay of dead plant and animal matter

(ii) Combustion of carbon-containing fuels

(c) The atmosphere on Earth traps heat so that the temperature on Earth is high enough for living things to survive.

2. (a) The combustion of fossil fuels increases the amount of carbon dioxide in the air. This disrupts the carbon cycle as there are not enough plants to remove the increased amount of carbon dioxide through photosynthesis.

(b) Biodiesel is a renewable fuel as the crops used to make biodiesel can be replaced in a short time. Combustion of biodiesel releases lower levels of air pollutants such as carbon monoxide and sulfur dioxide.

(c) Carbon dioxide is a greenhouse gas which traps heat on Earth. Thus, an increased amount of carbon dioxide in the atmosphere will cause the average temperature on Earth to increase and lead to global warming.

Test Station

1. **C**

Air consists of 21 % oxygen. Oxygen reacted with the iron powder to form iron(III) oxide.

Volume of oxygen in the syringe

= 21 % × 30

= 6.3 cm^3

Volume of remaining air in the glass tube

= 30 − 6.3

= 23.7 cm^3

≈ 24 cm^3

2. **B**

Methane gas is produced from the decay of organic matter and causes global warming. Carbon monoxide and ozone do not lead to the formation of acid rain. Sulfur dioxide is not a greenhouse gas and thus does not cause global warming.

3. **C**

Acid rain is formed when sulfur dioxide and oxides of nitrogen react with oxygen in the air to form acidic compounds which dissolve in rainwater.

4. (a) An idling engine has a low temperature. Thus, a very small amount of nitrogen monoxide is formed and released. [1]

(b) When more air is supplied, the fuel undergoes complete combustion, thus producing more carbon dioxide and less carbon monoxide. [1]

(c) (i) Carbon monoxide binds irreversibly with haemoglobin in red blood cells and prevents it from carrying oxygen to various parts of the body. This can cause death. [1]

(ii) Unburnt hydrocarbons can cause irritation of the eyes and respiratory tract. [1]

5. (a) Increased burning of fossil fuels in industries / Increased burning of petrol in vehicles / Large-scale clearing of forests / Increased farming activities (Any two) [1 mark for each correct answer]

(b) Increased greenhouse gas emissions can result in global warming, which leads to climate change. [1] Climate change causes changes in rainfall pattern, such as extreme rain and droughts. This can result in the loss of lives and affect the growth of crops. [1]

Examination Station

Paper 1

1. **D**

A thermometer is used to measure changes in temperature. A stopwatch is used to measure time intervals.

2. **B**

At t_1 min, compound **G** exists as a gas in which its particles are far apart and move randomly in all directions.

At t_4 min, forces of attraction exist between the particles of compound **G**, which is a liquid.

Since compound **G** is a liquid at room temperature and pressure, it changes from a gas to a liquid at between t_2 min and t_3 min.

3. **A**

A chlorine-35 atom and a chlorine-37 atom can combine to form a chlorine molecule with a molecular mass of 72.

Both isotopes have the same number of protons but different numbers of neutrons. They have the same reactivity as they have the same number of electrons.

The chemical formula of a compound formed between magnesium and chlorine-37 is $MgCl_2$ as the charge of one Mg^{2+} ion is balanced by the total charge of two Cl^- ions.

4. **B**

Sodium chloride is insoluble in hexane. Thus, it can be separated from iodine dissolved in hexane by filtration. The purity of sodium chloride can be determined by checking whether it boils or melts at a fixed temperature.

5. **D**

Element **Q** has a valency of 2 while element **R** has a valency of 3. Element **Q** loses two electrons to form the Q^{2+} ion while element **R** gains three electrons to form the R^{3-} ion. In order for the total charge of the positive and negative ions to be balanced, Q^{2+} ions combine with R^{3-} ions in the ratio of 3 : 2. Thus, the ionic compound formed has a chemical formula of Q_3R_2.

6. **C**

Number of moles of 4 g of $O_2 = \frac{4}{2 \times 16} = 0.125$ mol

1 mol of O_2 molecules has 2 mol of atoms.

Number of atoms in 4 g of $O_2 = 2 \times 0.125 \times 6.02 \times 10^{23}$
$= 1.505 \times 10^{23}$ atoms

Number of moles of 1 g of $H_2 = \frac{1}{2 \times 1} = 0.500$ mol

1 mol of H_2 molecules has 2 mol of atoms.

Number of atoms in 1 g of $H_2 = 2 \times 0.500 \times 6.02 \times 10^{23}$
$= 6.02 \times 10^{23}$ atoms

Number of moles of 14 g of $CO = \frac{14}{12 + 16} = 0.500$ mol

1 mol of CO molecules has 2 mol of atoms.

Number of atoms in 14 g of $CO = 2 \times 0.500 \times 6.02 \times 10^{23}$
$= 6.02 \times 10^{23}$ atoms

Number of moles of 16 g of $Cu = \frac{16}{64} = 0.250$ mol

Number of atoms in 16 g of $Cu = 0.250 \times 6.02 \times 10^{23}$
$= 1.505 \times 10^{23}$ atoms

Number of moles of 20 g of $Ar = \frac{20}{40} = 0.500$ mol

Number of atoms in 20 g of $Ar = 0.500 \times 6.02 \times 10^{23}$
$= 3.01 \times 10^{23}$ atoms

16 g of Cu has an equal number of atoms as 4 g of O_2.

7. **C**

$AgNO_3(aq) + KCl(aq) \longrightarrow KNO_3(aq) + AgCl(s)$

$AgCl$ is a precipitate. Thus, it does not ionise in water.

Molar mass of $AgNO_3 = 108 + 14 + (3 \times 16) = 170$ g/mol

Number of moles of 2 g of $AgNO_3 = \frac{2}{170} = 0.0118$ mol

Number of moles of KCl present $= \frac{20}{1000} \times 0.4$
$= 0.00800$ mol

From the equation, mole ratio of $AgNO_3 : KCl = 1 : 1$

Since only 0.00800 mol of KCl is present, it is the limiting reactant.

KCl was used up during the reaction, leaving behind KNO_3 and excess $AgNO_3$ in the reaction mixture.

Thus, the ions present in the reaction mixture at the end of the reaction are Ag^+, K^+ and NO_3^-.

8. **D**

The oxide of **Q** can react with both an alkali and an acid. This shows that it is an amphoteric oxide. Thus, the oxide could be ZnO, PbO or Al_2O_3. Since **Q** is in Period 3 of the periodic table, the oxide is Al_2O_3.

9. **B**

NaOH reacts with HCl, CO_2 and ZnO to form salts and water, which are neutral. NaOH reacts with NH_4Cl to form a salt, water and NH_3 gas. NH_3 can dissolve in water to form an alkaline solution.

10. **D**

Potassium chloride and hydrochloric acid ionise in aqueous solutions. The mobile ions can act as charge carriers. Thus, aqueous potassium chloride and dilute hydrochloric acid are able to conduct electricity.

The valence electrons in the structure of silver metal are mobile. Thus, silver metal is able to conduct electricity.

The ions that make up solid zinc chloride are held in position by strong electrostatic forces of attraction and are not able to move freely. Thus, solid zinc chloride is a poor electrical conductor.

11. **D**

Hydrochloric acid and sulfuric acid are the limiting reactants in experiments I and II respectively.

Experiment I:

$MgCO_3 + 2HCl \longrightarrow MgCl_2 + CO_2 + H_2O$

Number of moles of HCl reacted $= \frac{20}{1000} \times 0.400$
$= 0.008$ mol

Mole ratio of $HCl : CO_2 = 2 : 1$

Number of moles of CO_2 produced $= \frac{0.008}{2} = 0.004$ mol

Experiment II:

$MgCO_3 + H_2SO_4 \longrightarrow MgSO_4 + CO_2 + H_2O$

Number of moles of H_2SO_4 reacted $= \frac{20}{1000} \times 0.400$
$= 0.008$ mol

Mole ratio of $H_2SO_4 : CO_2 = 1 : 1$

Number of moles of CO_2 produced = 1 × 0.008 = 0.008 mol

Experiment I produced half the volume of CO_2 that was produced in experiment II.

$HCl \longrightarrow H^+ + Cl^-$

$H_2SO_4 \longrightarrow 2H^+ + SO_4^{2-}$

A lower concentration of H^+ ions reacted with magnesium carbonate in experiment I than in experiment II. Thus, the reaction in experiment II was faster and the graph representing experiment II has a steeper gradient. Hence, experiment II is represented by graph X while experiment I is represented by graph Z.

12. **B**

From the graph, the mass of the precipitate decreased when more reagent **E** was added. This shows that the precipitate dissolved in excess reagent **E**.

When aqueous sodium hydroxide is added to calcium ions and iron(II) ions, a precipitate which is insoluble in excess sodium hydroxide is formed. When aqueous ammonia is added to aluminium ions, a white precipitate which is insoluble in excess ammonia is formed. Only copper(II) ions react with aqueous ammonia to form a precipitate which dissolves in excess aqueous ammonia. Thus, reagent **E** is aqueous ammonia and the cation in sample **F** is Cu^{2+}.

13. **C**

Chlorine is more reactive than iodine. Thus, chlorine displaces iodine from potassium iodide and the iodine produced causes the solution to turn brown.

Chemical equation:

$Cl_2 + 2KI \longrightarrow 2KCl + I_2$

To obtain the ionic equation, write the chemical formulae of the constituent ions of the reactant and product that are in aqueous state:

$Cl_2 + 2K^+ + 2I^- \longrightarrow 2K^+ + 2Cl^- + I_2$

After removing the spectator ions from both sides of the equation, the ionic equation is:

$Cl_2 + 2I^- \longrightarrow 2Cl^- + I_2$

14. **B**

O_2 causes H_2 to be oxidised as the oxidation state of hydrogen increases from 0 in H_2 to +1 in H_2O. Thus, O_2 acts as an oxidising agent.

CuO causes Fe to be oxidised as the oxidation state of iron increases from 0 in Fe to +2 in FeO. Thus, CuO acts as an oxidising agent.

There is no change in the oxidation states of the elements in reaction 2. Thus, CO_2 does not act as an oxidising agent or a reducing agent.

15. **C**

Iron reacts with all the oxygen in the air inside the boiling tube. Air contains 21 % of oxygen.

Thus, volume of oxygen in 50 cm^3 of air that is used for rusting = 21 % × 50 = 10.5 cm^3

Volume of remaining air in the boiling tube = 50 − 10.5 = 39.5 cm^3

16. **C**

Combustion of an alkane gives out heat. Thus, the reaction is exothermic and the temperature of the surroundings increases. The products of the incomplete combustion of an alkane are carbon monoxide and water.

17. **A**

Excessive cattle rearing results in a large amount of methane released into the atmosphere. Methane is a greenhouse gas which causes global warming. Global warming leads to climate change and hence extreme weather and melting of polar ice caps. Methane does not cause acid rain.

18. **A**

Lubricating oil flows less easily than kerosene, which burns less easily than naphtha. Thus, it is likely to be found at a lower level of the fractionating column and contain more carbon atoms than kerosene and naphtha. It is also likely to have a large molecular mass. Since the crude oil fractions have different molecular masses, they have different densities.

19. **B**

The alkane undergoes a substitution reaction in which one chlorine molecule substitutes two hydrogen atoms in the alkane. Thus, the chemical equation for the reaction is: $C_xH_y + 2Cl_2 \longrightarrow C_xH_{y-2}Cl_2 + 2HCl$

20. **A**

Ethanol is oxidised by acidified potassium manganate(VII) to form ethanoic acid. Thus, acidified potassium manganate(VII) changes from purple to colourless. Ethane does not react with acidified potassium manganate(VII). Thus, no observation is made.

Paper 2

Section A

1. (a) $2CO(g) + 2NO(g) \longrightarrow 2CO_2(g) + N_2(g)$ [1]

(b) Carbon monoxide is oxidised as the oxidation state of carbon increases from +2 in carbon monoxide to +4 in carbon dioxide. [1] Nitrogen monoxide is reduced as the oxidation state of nitrogen decreases from +2 in nitrogen monoxide to 0 in nitrogen. [1]

2. (a) **A**: Iron(II) nitrate

B: Copper(II) carbonate

C: Iron(II) hydroxide

E: Ammonia

F: Copper(II) nitrate

G: Carbon dioxide

[1 mark for 2 correct answers]

(b) $Fe(OH)_2(s) \longrightarrow FeO(s) + H_2O(g)$ [1]

(c) Neutralisation. [1] **D**, which is iron(II) oxide, is a base and reacts with nitric acid to form a salt and water. [1]

3. (a) Reaction I: $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$ [1]

Reaction III: $Cl_2(g) + 2KI(aq) \longrightarrow 2KCl(aq) + I_2(aq)$ [1]

(b) Reaction I: Combustion [1]

Reaction II: Fermentation [1]

Reaction III: Displacement reaction [1]

(c) Reaction IV, which is rusting, can be prevented by using a protective layer to keep iron away from oxygen and water. [1] This can be done by painting and oiling [1].

4. (a) Compound **Y** is an ionic compound which exists as a giant ionic crystal lattice. [1] The oppositely charged ions are held in the lattice by strong electrostatic forces of attraction which require a large amount of energy to overcome. [1]

Ammonia is a covalent compound which exists as simple molecules. [1] The weak intermolecular forces of attraction between the ammonia molecules require a small amount of energy to overcome. [1]

(b) Hydrogen bromide particles and ammonia particles are far apart and move randomly at high speed. [1] When solid **Y** is formed, the particles become very close together and vibrate about fixed positions. [1]

(c) Compound **Y** can change from the solid state to the gaseous state directly. [1]

5. (a) No more effervescence will be observed. [1]

(b) $MgCO_3 + 2HNO_3 \longrightarrow Mg(NO_3)_2 + CO_2 + H_2O$ [1]

Number of moles of magnesium carbonate added

$= \frac{2}{[24 + 12 + (3 \times 16)]}$

$= 0.0238$ mol

Number of moles of nitric acid present

$= \frac{40}{1000} \times 1.0$

$= 0.0400$ mol

From the equation, mole ratio of $MgCO_3 : HNO_3 = 1 : 2$

Number of moles of nitric acid needed to react with $MgCO_3 = 2 \times 0.0238 = 0.0476$ mol [1]

Since only 0.0400 mol of nitric acid is present, it is the limiting reactant. Thus, magnesium carbonate is in excess. [1]

(c) From the equation, mole ratio of $HNO_3 : Mg(NO_3)_2 = 2 : 1$

Number of moles of $Mg(NO_3)_2 = \frac{0.0400}{2}$

$= 0.0200$ mol [1]

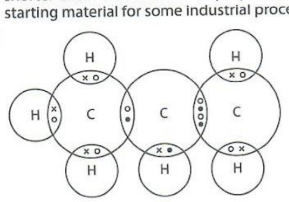
Mass of $Mg(NO_3)_2 = 0.0200 \times [24 + (2 \times 14) + (6 \times 16)]$

$= 2.96$ g [1]

(d) The reaction between magnesium hydroxide and nitric acid is exothermic. Thus, the student can measure the temperature of the reaction mixture to determine whether the reaction is complete. [1] The reaction is complete when the highest temperature is reached. [1]

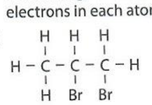
6. (a) Cracking. [1] It converts long-chain alkanes into shorter-chain alkenes such as propene, which is the starting material for some industrial processes. [1]

(b)

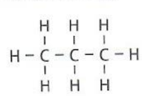


[1 mark for correct number of electrons involved in bonding. 1 mark for correct number of valence electrons in each atom.]

(c)



compound **B** [1]



compound **C** [1]

7. (a) No. **X** contains two pigments, **A** and **C**. Since it contains more than one pigment, it is not a pure substance. [1]

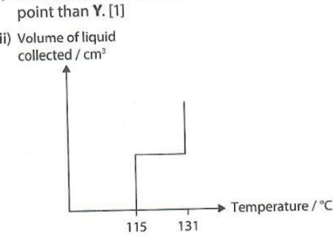
(b) **X**. It does not contain pigment **D**. [1]

(c) Pigment **C**. It travelled the furthest up the chromatogram compared to the other pigments. [1]

(d) No. The pigments are likely to have different solubilities in different solvents. Thus, the chromatogram run using water would be different from the chromatogram run using methanol. [1]

(e) (i) **X** would be collected first as it has a lower boiling point than **Y**. [1]

(ii)



[1 mark for increasing volume at 115 °C and 131 °C. 1 mark for constant volume between 115 °C and 131 °C. 1 mark for correctly labelled graph.]

Section B

8. (a) (i) Calcium [1]

(ii) Potassium reacts very violently / explosively with water in the solution. [1] Hence, the reaction cannot be carried out in the laboratory. [1]

(b) Copper, iron, zinc, magnesium, **Q**/calcium. [1] A more reactive metal is able to displace another metal from its salt solution. [1]

(c) (i) $3Mg(s) + 2Fe^{3+}(aq) \longrightarrow 3Mg^{2+}(aq) + 2Fe(s)$ [1]

(ii) Magnesium is oxidised as its oxidation state increases from 0 in Mg to +2 in Mg^{2+}. [1] Iron is reduced as its oxidation state decreases from +3 in Fe^{3+} to 0 in Fe. [1]

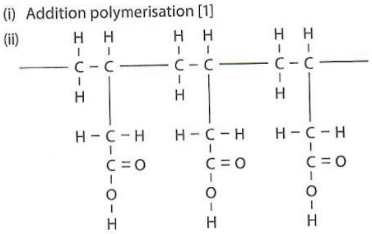
(d) **Q** is a reactive metal. [1] It reacted readily with water in zinc chloride solution to form hydrogen gas. [1]

9. (a) Aqueous bromine was decolourised when added to each compound [1] since both compounds **D** and **E** contain a C=C bond. [1] Hence, the student would not be able to distinguish between the two compounds using aqueous bromine.

(b) Acidified potassium manganate(VII) and heat [1]

(c) (i) Addition polymerisation [1]

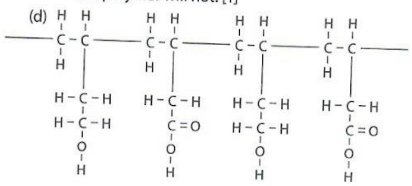
(ii)



[1 mark for correct structure showing all the bonds. 1 mark for showing three repeat units.]

(iii) Add aqueous bromine to the polymer and compound **E** separately. [1] Compound **E** will cause aqueous bromine to be decolourised, but the polymer will not. [1]

(d)



[1 mark for correct structure showing all the bonds and 1 mark for showing two repeat units]

10. (a) $\mathbf{M}CO_3(s) + H_2SO_4(aq) \longrightarrow \mathbf{M}SO_4(aq) + CO_2(g) + H_2O(l)$ [1]

(b) Volume of carbon dioxide produced when the reaction is complete = 50 cm^3

Volume of carbon dioxide produced when the reaction is half complete $= \frac{50}{2} = 25\ cm^3$ [1]

Based on the graph, time taken to produce 25 cm^3 of carbon dioxide = 1.5 min [1]

(c) As the reaction proceeded, more reactants were converted into products and the concentration of reactants decreased. [1] This resulted in a decrease in frequency of collisions between reactant particles and hence a decrease in frequency of effective collisions. [1] Thus, the rate of reaction decreased. [1]

(d) Number of moles of CO_2 produced during the reaction $= \frac{50}{1000} \div 24 = 0.00208$ mol [1]

Based on the equation, mole ratio of $CO_2 : \mathbf{M}CO_3 = 1 : 1$

Number of moles of $\mathbf{M}CO_3$ used = 0.00208 mol [1]

M_r of $\mathbf{M}CO_3 = \frac{0.175}{0.00208} = 84$ [1]

A_r of $\mathbf{M} = 84 - 12 - (3 \times 16) = 24$

M is magnesium. [1]

Examination Station

Paper 1

1. **B**

A gas syringe can be used to measure the volume of hydrogen gas. A stopwatch can be used to measure the time taken for 20 cm^3 of hydrogen gas to be produced.

2. **B**

Sand is insoluble in water. Thus, filtration can be used to separate the two substances. The purity of water can be determined by checking whether it boils at 100 °C.

3. **C**

Iron, which is magnetic, is attracted to the magnet. Thus, sulfur, which is non-magnetic, is left behind.

4. **B**

When ammonium chloride sublimes, it changes from the solid state to the gaseous state. When it is in the solid state, its particles vibrate about fixed positions. When it is in the gaseous state, its particles move about more quickly and become very far apart.

5. **B**

Deuterium has one electron like hydrogen. Thus, it can lose the electron to form the D^+ ion and combine with chlorine to form a compound with the chemical formula DCl. Isotopes have the same number of protons and different numbers of neutrons. Thus, it is not possible for deuterium to have an atomic number of 2 since the atomic number of hydrogen is 1.

6. **D**

Element **Q** has a valency of 2 while element **R** has a valency of 3. Element **Q** loses two electrons to form the $\mathbf{Q}^{2+}$ ion while element **R** gains three electrons to form the $\mathbf{R}^{3-}$ ion. In order for the total charge of the positive and negative ions to be balanced, $\mathbf{Q}^{2+}$ ions combine with $\mathbf{R}^{3-}$ ions in the ratio of 3 : 2. Thus, the ionic compound formed has a chemical formula of $\mathbf{Q}_3\mathbf{R}_2$.

7. **D**

Diagram I shows molecules of two elements. Diagram II shows molecules of two compounds. Diagram III shows molecules of only one compound.

8. **D**

A metal oxide that reacts with an alkali is amphoteric.

9. **A**

The reaction between $AgNO_3$ and HCl results in the formation of HNO_3, which is acidic.

10. **C**

Atoms of different sizes that are present in the structure of an alloy disrupt the regular arrangement of the atoms of a metal and prevent the atoms of similar size from sliding over one another. Thus, an alloy is stronger than a pure metal.

11. **C**

The valency of the chromium(III) ion is 3 while the valency of the oxide ion is 2. In order for the total charge of the positive and negative ions to be balanced, Cr^{3+} ions combine with O^{2-} ions in the ratio of 2 : 3. Thus, chromium(III) oxide has a chemical formula of Cr_2O_3.

12. **B**

Both copper and zinc can form alloys with other metals. Copper(II) oxide is basic, while zinc oxide is amphoteric. Zinc is more reactive than copper. Thus, it reacts with acids and steam while copper does not.

13. **C**

Fluorine is the most reactive element in Group 17. Rubidium is more reactive than lithium, sodium and potassium. Thus, fluorine will react most vigorously with rubidium compared to the other pairs of elements.

14. **A**

Oxides of nitrogen are produced in car engines and in the presence of lightning. They react with oxygen in the air to form acidic compounds, which dissolve in rainwater to form acid rain. Acid rain corrodes metals and damages forests.

15. **C**

The iron strips in test tubes 1 and 3 would not rust as there is no air in test tube 1 and no water and air in test tube 3. The iron strips in test tubes 2 and 4 are exposed to oxygen in the air and water vapour from the damp cotton wool and the air. Thus, they would rust.

16. **C**

Reducing the amount of non-biodegradable waste in landfills is an advantage of recycling plastics.

17. **C**

1 mol of ethane and 1 mol of ethene have different relative molecular masses and hence different masses. Since they have different masses, they also have different densities. Both ethane and ethene contain two carbon atoms.

18. **A**

Aqueous bromine can be used to test for unsaturation. Vegetable oil is unsaturated and hence will cause aqueous bromine to be decolourised. Margarine is saturated and does not react with aqueous bromine. Thus, aqueous bromine will remain brown.

19. **C**

An alkene can undergo chlorination, hydrogenation and combustion. However, it does not undergo a substitution reaction like an alkane.

20. **D**

Number of moles of 2 g of $H_2 = \frac{2}{1+1} = 1$ mol

Number of moles of 2 g of He $= \frac{2}{4} = 0.5$ mol

Number of moles of 4 g of $CH_4 = \frac{4}{12 + (4 \times 1)} = 0.25$ mol

Number of moles of 8.5 g of $NH_3 = \frac{8.5}{14 + (3 \times 1)} = 0.5$ mol

Number of moles of 6 g of C $= \frac{6}{12} = 0.5$ mol

Number of moles of 6 g of Mg $= \frac{6}{24} = 0.25$ mol

Number of moles of 44 g of $CO_2 = \frac{44}{12 + (2 \times 16)} = 1$ mol

Number of moles of 40 g of Ar $= \frac{40}{40} = 1$ mol

From the above, 44 g of carbon dioxide and 40 g of argon have the same number of moles.

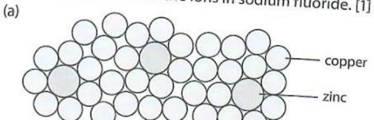
Paper 2

Section A

1. (a) CO [1]
   (b) Fe [1]
   (c) Al_2O_3 [1]
   (d) $CuSO_4$ [1]
2. (a) Ammonia [1]
   (b) The gas turns damp red litmus paper blue. [1]
   (c) Gas syringe [1]
3. (a) No. **X** contains two pigments, **A** and **C**. Since it contains more than one pigment, it is not a pure substance. [1]
   (b) **X**. It does not contain pigment **D**. [1]
   (c) Pigment **C**. It travelled the furthest up the chromatogram compared to the other pigments. [1]
   (d) No. The pigments are likely to have different solubilities in different solvents. Thus, the chromatogram run using water would be different from the chromatogram run using methanol. [1]
4. (a) Gas [1]
   (b) Energy is lost. [1]
   (c) The particles are closer together and move freely throughout the liquid substance. [1]

Section B

5. (a) (i) $Cl_2(g) + 2Br^-(aq) \longrightarrow 2Cl^-(aq) + Br_2(aq)$ [1]
   (ii) Chlorine is less reactive than fluorine. Thus, it is not able to displace fluorine from sodium fluoride. [1]
   (b) (i) Iodine, bromine, chlorine and fluorine [1]
   (ii) The reactivity of the halogens decreases down Group 17. The size of the atom of the halogens increases down the group. [1] Thus, it is more difficult for the nucleus in the atom to attract one more electron. [1]
   (c) Fluorine is a covalent molecule and has a simple molecular structure. Weak intermolecular forces of attraction exist between fluorine molecules. [1] Sodium fluoride is an ionic compound and has a giant ionic crystal lattice structure. Strong electrostatic forces of attraction exist between the oppositely charged ions in the lattice. [1] The weak intermolecular forces of attraction between fluorine molecules require a smaller amount of energy to overcome than the strong electrostatic forces of attraction between the ions in sodium fluoride. [1]
6. (a)



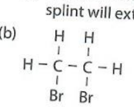
   [1 mark for the disrupted arrangement of atoms. 1 mark for labelling the copper and zinc atoms.]

   (b) (i) Copper did not react with the acid [1] as it is below hydrogen in the reactivity series. [1]
   (ii) Mass of Zn reacted = 10.5 − 6.3 = 4.2 g

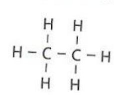
   Number of moles of Zn reacted $= \frac{4.2}{65} = 0.0646$ mol [1]

   (c) As zinc reacted with hydrochloric acid, the amount of H^+ ions in the reaction mixture decreased. Thus, the pH of the reaction mixture increased. [1]
   (d) Similarity: Both isotopes have the same number of protons and electrons. [1]

   Difference: Copper-65 has 2 more neutrons than copper-63. [1]
7. (a) (i) Cracking. [1] It converts long-chain alkanes into shorter-chain alkenes such as ethene, which is the starting material for some industrial processes. [1]
   (ii) Use a burning splint to test the gas. The burning splint will extinguish with a "pop" sound. [1]
   (b)



compound **B** [1]



compound **C** [1]

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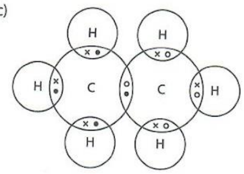
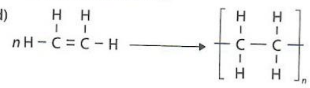
(c)



[1 mark for the correct number of electrons involved in bonding. 1 mark for the correct number of valence electrons in each atom.]

(d)



[1]