Microscopic World I

Unit 5  Atomic structure
Unit 6  The periodic table
Unit 7  Ionic and metallic bonds
Unit 8  Covalent bonds
Unit 9  Relating the properties of substances to structures and bonding
Summary

1. All elements are made of atoms.
2. Chemists use symbols to represent elements.
3. At room temperature and pressure, elements exist in different states (solid, liquid or gas).
4. Elements can be classified into three main groups — metals, metalloids and non-metals.

```
Elements
<table>
<thead>
<tr>
<th>metals</th>
<th>metalloids</th>
<th>non-metals</th>
</tr>
</thead>
</table>
```

5. An atom consists of three types of subatomic particles: protons, neutrons and electrons.

   The nucleus contains protons and neutrons. Electrons move around the nucleus in shells.

6. Atomic number of an element = number of protons in an atom of that element = number of electrons in a neutral atom of that element

7. Mass number = number of protons + number of neutrons

```
atomic number → Z
mass number → A
symbol of an atom ←
```

8. Isotopes are different atoms of an element which have the same number of protons but a different number of neutrons. For example, chlorine has two isotopes: Cl-35 and Cl-37.

9. The relative isotopic mass of a particular isotope of an element is the relative mass of one atom of that isotope on the $^{12}\text{C} = 12.00$ scale.
10 The relative atomic mass of an element is the weighted average relative isotopic mass of all the naturally occurring isotopes of that element on the $^{12}$C = 12.00 scale.

11 The way in which electrons are arranged in an atom is called its electronic arrangement.

12 An orbital is the region in which there is a high probability of finding an electron.

\[ \text{Answer} \]
\[ 87.6 = \frac{86 \times p + 87 \times (100 - p - 76.5) + 88 \times 76.5}{100} \]
\[ p = 16.5 \]
\[ q = 100 - 16.5 - 76.5 \]
\[ q = 7.0 \]

\[ \text{Remarks} \]
- Isotopes of an element do NOT have the same mass.
- Questions may give the relative abundance of the isotopes of an element and ask students to calculate the relative atomic mass of the element.

**Example**

A sample of strontium (Sr) consists of three isotopes $^{86}$Sr, $^{87}$Sr and $^{88}$Sr. The relative abundance of each isotope is shown in the table below:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>$^{86}$Sr</th>
<th>$^{87}$Sr</th>
<th>$^{88}$Sr</th>
</tr>
</thead>
<tbody>
<tr>
<td>Relative abundance (%)</td>
<td>$p$</td>
<td>$q$</td>
<td>76.5</td>
</tr>
</tbody>
</table>

The relative atomic mass of the sample of strontium is 87.6. Calculate the values of $p$ and $q$. (3 marks)
Unit 6
The periodic table

6.1 How to group elements together?
6.2 The periodic table
6.3 Patterns across the periodic table
6.4 Group I elements — alkali metals
6.5 Group II elements — alkaline earth metals
6.6 Group VII elements — halogens
6.7 Group 0 elements — noble gases
6.8 Predicting the chemical properties of unfamiliar elements
6.9 From atoms to ions
6.10 Predicting the charge on an ion

Summary

1 In the periodic table, all the elements are arranged in order of increasing atomic number.
2 The vertical columns in the periodic table are called groups. Groups are numbered from I to VII, followed by Group 0 (or Group VIII).
   Group number of an element = number of outermost shell electrons in an atom of the element
3 The horizontal rows in the periodic table are called periods.
   Period number of an element = number of occupied electron shells in an atom of the element
4 Across a period in the periodic table, the elements change from metals through metalloids to non-metals.
5 a) Elements in the same group have the same number of outermost shell electrons in their atoms and thus they have similar chemical properties.
   b) There is usually a gradual change in the properties of elements as we move down a group.
6 Group I elements — alkali metals
   a) They all have relatively low melting and boiling points when compared with other metals.
   b) They are all soft and can be cut with a knife.
   c) They all have low densities — lithium, sodium and potassium float on water.
   d) They are all reactive metals and must be stored in paraffin oil to prevent them from reacting with air.
   e) They all react vigorously with water to give hydrogen gas and an alkaline solution.
   f) They all react with non-metals to form compounds called salts.
   g) The reactivity of these elements increases as we move down the group.
7 Group II elements — alkaline earth metals
   a) They all have relatively low melting and boiling points when compared with other metals (except Group I metals).
   b) They all have low densities.
   c) They are all reactive metals and react readily with dilute hydrochloric acid to give hydrogen gas.
   d) They all react with non-metals to form compounds called salts.
   e) Group II elements are less reactive than Group I elements. The reactivity increases as we move down the group.
8. Group VII elements — halogens
   a) They are all poisonous and smelly.
   b) They are all non-metals.
   c) They all react with metals to form compounds called salts.
   d) The reactivity of these elements decreases as we move down the group.

9. The following diagram shows the trends of some physical properties and reactivity of Groups I, II, and VII elements.

- melting and boiling points increasing
- reactivity decreasing

10. Group 0 elements — noble gases
    a) They are all colourless gases at room temperature and pressure.
    b) They all have very low melting and boiling points.
    c) They are all very unreactive.

11. The octet rule suggests that atoms become stable by having eight electrons (an octet structure) in their outermost shells (or two electrons, a duplet structure, in the case of some smaller atoms).

12. Atoms can obtain the stable electronic arrangements of atoms of noble gases by gaining or losing electrons.

13. a) Positive charge(s) on an ion formed from the atom of a metal
    = group number of the metal

   b) Negative charge(s) on an ion formed from the atom of a non-metal
    = 8 – group number of the non-metal

<table>
<thead>
<tr>
<th></th>
<th>Group I</th>
<th>Group II</th>
<th>Group III</th>
<th>Group IV</th>
<th>Group V</th>
<th>Group VI</th>
<th>Group VII</th>
<th>Group 0</th>
</tr>
</thead>
<tbody>
<tr>
<td>Period 2</td>
<td>Li⁺</td>
<td>Be²⁺</td>
<td>N³⁻</td>
<td>O²⁻</td>
<td>F</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Period 3</td>
<td>Na⁺</td>
<td>Mg²⁺</td>
<td>Al³⁺</td>
<td>S²⁻</td>
<td>Cl⁻</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Period 4</td>
<td>K⁺</td>
<td>Ca²⁺</td>
<td></td>
<td>Br⁻</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Exam tips

♦ In the periodic table the elements are arranged in increasing order of atomic number, NOT neutron number or relative atomic mass.

✔ Questions often ask about trends when going across a period of the periodic table.
   e.g. Going across the second period of the periodic table (from lithium to fluorine),
   – the elements show a gradual change from having metallic property to having non-metallic property;
   – the elements have the same number of occupied electron shells in their atoms;
   – the atomic size of elements decreases;
   – the melting points and boiling points of elements do NOT vary gradually across a period.

♦ The atomic radius of the elements increases down a group as electrons are added to successive electron shells and thus further away from the nucleus.

♦ The melting point, boiling point, density and hardness of Group I elements are all very much lower than the average values for metals.

♦ Below are some safety precautions for handling Group I elements in the laboratory:
   – wear safety glasses;
   – use forceps;
   – use a safety screen.

♦ Questions may ask students to deduce the atomic number of an element from the known atomic number of another element in the same group.
   e.g. X and Y are alkaline earth metals. The atomic number of X is x. Then the atomic number of Y could be x + 18.

♦ When Barium (a Group II element) is added to a trough of water containing phenolphthalein,
   – the piece of barium sinks in the water;
   – a colourless gas is liberated;
   – an alkaline solution forms and the phenolphthalein turns pink.

♦ The reactivity of elements in Group I / II increases down the group.

♦ Questions may ask students to compare potassium and calcium.
   – The reducing power of potassium is stronger than that of calcium.
   – The hardness of calcium is higher than that of potassium.
   – The density of calcium is higher than that of potassium.
Not all atoms of noble gases have an octet structure in the outermost shell. Atoms of helium have a duplet structure.

Questions may give the charge carried by a stable ion of an unknown element and ask the possible atomic number of the element.

e.g.
An element forms a stable $X^{2+}$ ion. The atomic number of $X$ may be 20 or 38.

Questions often ask about the conversion of an atom of a non-metal to form an anion.

e.g.
When an iodine atom is converted to an iodide ion,
- a reduction is involved;
- the iodide ion has the same number of occupied electron shells as the iodine atom;
- the iodide ion and the iodine atom have the same number of protons.

Questions may describe the number of subatomic particles in the ion of an unknown element and ask students to deduce information about the element / its compound (e.g. the group or period to which the element belongs) or vice versa.

e.g.
Suppose $X^{2+}$ ion has an electronic arrangement $2,8,8$. Thus, an atom of $X$ has an electronic arrangement $2,8,8,2$. $X$ is calcium.

The question may ask students to deduce information about the carbonate of $X$:
- it is a white solid;
- it decomposes on heating;
- it produces a brick-red flame in flame test.

**Example**

The electronic arrangements of the atoms of three elements are shown in the table below:

<table>
<thead>
<tr>
<th>Element</th>
<th>Electronic arrangement of atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>$X$</td>
<td>$2,7$</td>
</tr>
<tr>
<td>$Y$</td>
<td>$2,8,5$</td>
</tr>
<tr>
<td>$Z$</td>
<td>$2,8,p,q$</td>
</tr>
</tbody>
</table>

a) Elements $X$ and $Z$ belong to the same group of the periodic table.

i) Name the group to which they belong. (1 mark)

ii) The atomic number of element $Z$ is 35. What are the values of $p$ and $q$? (1 mark)

iii) Explain, in terms of bonding and structure, why the boiling point of element $Z$ is higher than that of element $X$. (3 marks)

b) Draw an electron diagram of a compound formed between elements $X$ and $Y$, showing electrons in the outermost shells only. (1 mark)

**Answer**

a) i) Halogens (1)

ii) $p = 18$ (0.5)

$q = 7$ (0.5)

iii) The boiling point of a substance depends on the strength of its intermolecular attractions. (1)

The intermolecular attractions between molecules of Group VII elements are van der Waals’ forces. (1)

A molecule of element $Z$ has more electrons than a molecule of element $X$. / Element $Z$ has a bigger molecular size than element $X$. (1)

Thus, the van der Waals’ forces between molecules of element $Z$ are stronger than those between molecules of element $X$. (1)

b) $X\ Y\ X$

**Remarks**

- Questions often ask about the halogens.
  - Bromine is the only non-metal which is a liquid at room temperature and pressure.
  - Iodine vapour is purple in colour.
  - The boiling point of chlorine, bromine and iodine decreases down the group.
  - The reactivity of fluorine, chlorine and bromine decreases with relative atomic mass.
  - Fluorine, chlorine and bromine all react with sodium sulphite solution.

- Astatine is a halogen below iodine.
  - It is soluble in heptane.
  - The oxidizing power of chlorine is stronger than that of astatine.

- Questions often ask about phosphorus $^{31}$P.
  - Phosphorus belongs to Group V and Period 3 of the periodic table.
  - It is a solid at room temperature and pressure.
  - Its atom contains 16 neutrons.
  - It is a non-metal. It forms covalent compounds with non-metals.
  - It forms ionic compounds with metals.
Unit 7

Ionic and metallic bonds

7.1 Conductors, electrolytes and non-conductors
7.2 Evidence of ions from electrolysis of molten lead(II) bromide
7.3 Chemical bonds
7.4 Ionic bonds
7.5 Compounds containing polyatomic ions
7.6 Names of ions
7.7 Naming ionic compounds
7.8 Colours of ionic compounds
7.9 Chemical formulae of ionic compounds
7.10 Metallic bonds in metals

Summary

1. Substances can be classified as conductors, electrolytes and non-conductors.

2. An ionic bond is the strong electrostatic forces of attraction between oppositely charged ions.
   a) An ionic bond is formed when one or more electrons are transferred from one atom (or group of atoms) to another.
   b) When a metal and a non-metal combine to form an ionic compound, atoms of the metal release electrons while atoms of the non-metal gain electrons.

3. The table below shows electron diagrams of some common ionic compounds.

<table>
<thead>
<tr>
<th>Ionic compound</th>
<th>Electron diagram</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium chloride</td>
<td>![Na]^+ [Cl]^-</td>
</tr>
<tr>
<td>Magnesium fluoride</td>
<td>![Mg]^{2+} [F]^-</td>
</tr>
<tr>
<td>Lithium oxide</td>
<td>![Li]^+ [O]^{2-}</td>
</tr>
</tbody>
</table>

4. A polyatomic ion is formed from a group of atoms, instead of a single atom.
5 a) Naming positive ions:
   i) If a metal forms only one kind of positive ion, the name of the ion is the same as the metal.
   ii) Some metals can form more than one kind of positive ion. For example, we use the name of copper(I) ion for Cu\(^{+}\), and copper(II) ion for Cu\(^{2+}\).

   b) Naming negative ions:
      i) Simple negative ions have names ending in ‘-ide’.
      ii) Polyatomic ions containing oxygen have names ending in -ite or -ate.
      iii) The polyatomic ion with less oxygen is named -ite, and that with more oxygen is named -ate.
      iv) A polyatomic ion formed from an oxygen atom and a hydrogen atom is called a hydroxide ion.

6 When naming an ionic compound, name the positive ion first, followed by the negative ion.

7 If an ionic compound has colour, the colour may arise from either the negative or positive ion, or even from both ions.

<table>
<thead>
<tr>
<th>Ion</th>
<th>Chemical formula</th>
<th>Colour</th>
</tr>
</thead>
<tbody>
<tr>
<td>Iron(II)</td>
<td>Fe(^{2+})</td>
<td>pale green</td>
</tr>
<tr>
<td>Iron(III)</td>
<td>Fe(^{3+})</td>
<td>yellow-brown</td>
</tr>
<tr>
<td>Copper(II)</td>
<td>Cu(^{2+})</td>
<td>blue or green</td>
</tr>
<tr>
<td>Permanganate</td>
<td>MnO(_4)(^{2-})</td>
<td>purple</td>
</tr>
<tr>
<td>Dichromate</td>
<td>CrO(_4)(^{2-})</td>
<td>orange</td>
</tr>
<tr>
<td>Chromium(III)</td>
<td>Cr(^{3+})</td>
<td>green</td>
</tr>
<tr>
<td>Nickel(II)</td>
<td>Ni(^{2+})</td>
<td>green</td>
</tr>
<tr>
<td>Manganese(II)</td>
<td>Mn(^{2+})</td>
<td>very pale pink (or colourless)</td>
</tr>
</tbody>
</table>

8 The chemical formula of an ionic compound shows the types of ions present and the ratio of one type of ion to the other.

9 A metallic bond is a type of bond in which positive metal ions are held together by a ‘sea’ of mobile electrons.

   The following diagram shows the metallic bonding in a piece of metal.

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**Exam tips**

- A piece of freshly cut lithium turns dull when placed in air. Lithium oxide is formed.
- It is **WRONG** to draw the electron diagram of lithium oxide in the following way:
- Students should be able to describe the structure and bonding in metal with the help of a diagram.
- The binding forces in both metallic bond and ionic bond are non-directional.
- Metallic bond differs from covalent bond in that metal atoms do not form separate molecules. However, the atomic cores are bound together by being attracted to the electrons between them in both metallic bond and covalent bond.

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### Example

X, Y and Z are three different elements. X\(^{2+}\) ion and Y\(^{3+}\) ion and neon atom have the same electronic arrangement:

a) What are elements X and Y? (1 mark)

b) Draw an electron diagram of the compound formed from elements X and Y, showing electrons in the outermost shells only. (1 mark)

c) Element Z belongs to the same period as element X and the same group as element Y in the periodic table. What is element Z? (1 mark)

**Answer**

a) Element X is magnesium. (0.5)

   Element Y is nitrogen. (0.5)
Unit 8  Covalent bonds

8.1 Covalent bonds
8.2 Covalent bonds in non-metallic elements
8.3 Covalent compounds
8.4 Writing chemical formulae of covalent compounds
8.5 Predicting the formation of ionic and covalent compounds
8.6 Dative covalent bonding
8.7 Bonding in polyatomic ions
8.8 Relative molecular mass and formula mass

Remarks

Questions often ask about the chemical formula and electron diagram of magnesium nitride (Mg₃N₂).
Summary

1. A covalent bond is formed when one or more pairs of outermost shell electrons are shared between two atoms.
2. A covalent bond is the strong electrostatic forces of attraction between the shared electrons and the two positively charged nuclei of the bonded atoms.
3. Atoms of non-metallic elements can join together to form groups called molecules.
4. The table below shows the electron diagrams and models of molecules of some common non-metals.

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Electron diagram</th>
<th>Ball-and-stick model</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>bond pair electrons</td>
<td>H—H</td>
</tr>
<tr>
<td>Chlorine</td>
<td>lone pair electrons</td>
<td>Cl—Cl</td>
</tr>
<tr>
<td>Oxygen</td>
<td></td>
<td>O=O</td>
</tr>
<tr>
<td>Nitrogen</td>
<td></td>
<td>N≡N</td>
</tr>
<tr>
<td>Hydrogen chloride</td>
<td>hydrogen chloride molecule (H—Cl)</td>
<td>H—Cl</td>
</tr>
<tr>
<td>Tetrachloromethane</td>
<td>tetrachloromethane molecule (Cl—C—Cl)</td>
<td>Cl—C—Cl</td>
</tr>
<tr>
<td>Water</td>
<td></td>
<td>H—O—H</td>
</tr>
<tr>
<td>Ammonia</td>
<td></td>
<td>H—N—H</td>
</tr>
<tr>
<td>Carbon dioxide</td>
<td></td>
<td>O=C=O</td>
</tr>
</tbody>
</table>

5. Atoms of different non-metallic elements share electrons with each other to form covalent compounds.
7 a) When a metal combines with a non-metal, an ionic compound forms.

b) When non-metals combine, a covalent compound forms.

- Metal and non-metal combine to form ionic compound
- Non-metals combine to form covalent compound

- Particles present: ions (metal and non-metal)
- Examples: magnesium oxide, sodium chloride

- Particles present: molecules (non-metals)
- Examples: ammonia, carbon dioxide

- Particles present: atoms (a network of covalent bonds linking atoms)
- Examples: silicon dioxide, argon (exists as monoatomic molecules), chlorine, hydrogen

- Particles present: molecules (a network of covalent bonds linking atoms)
- Examples: carbon, silicon

8 A dative covalent bond is a covalent bond in which the bond pair electrons are provided by the same atom.

The following diagram shows the formation of a dative covalent bond in an \( \text{NH}_3\text{BF}_3 \) molecule.

- The bond is formed between nitrogen and boron, with the lone pair of electrons from the nitrogen atom.

9 In substances containing polyatomic ions, the cations and anions are held together by ionic bonding, but each polyatomic ion is a group of atoms held together by covalent bonding.

The following diagram shows the formation of a dative covalent bond in an ammonium ion.

- The bond is formed between nitrogen and hydrogen, with the lone pair of electrons from the nitrogen atom.

10 Relative molecular mass of an element or compound = sum of relative atomic masses of all atoms present in one molecule of the element or compound

11 Formula mass of an ionic compound = sum of the relative atomic masses of all atoms making up one formula unit of the compound

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**Exam tips**
- Questions often ask why noble gases exist as monoatomic molecules or why they are unreactive.
- Chlorine exists as diatomic molecules, NOT 'bimolecular'.
- Questions may ask students to draw electron diagrams of compounds formed from two halogens.
  - e.g. [diagram of F and Cl molecules]
- Hydrogen peroxide (\( \text{H}_2\text{O}_2 \)) is an oxide of hydrogen.
  - [diagram of H₂O₂ molecule]
- Notice that both ionic and covalent bonds exist in ammonium nitrate.

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**Example**

X, Y and Z are three consecutive elements in the periodic table, with the sum of their atomic numbers equal to 54. X forms a stable anion \( X^- \) while Z forms a stable cation \( Z^+ \).

a) What are elements X, Y and Z? (1 mark)

b) P is a compound formed from element X and oxygen. Draw an electron diagram of P, showing electrons in the outermost shells only. (1 mark)

c) Q is a compound formed for element Z and oxygen. Draw an electron diagram of Q, showing electrons in the outermost shells only. (1 mark)

d) Compare the melting points of compounds P and Q. Explain your answer. (3 marks)

**Answer**

a) X: chlorine
Y: argon
Z: potassium

b) [diagram of Cl and O molecules]
d) The melting point of compound Q is higher than that of compound P. (1)

To melt compound Q, a lot of heat is needed to overcome the strong ionic bonds between the ions. (1)

The attractive forces between the molecules of compound P are weak van der Waals’ forces. Little heat is needed to separate the molecules. (1)

Remarks

➤ Questions often ask students to explain the difference between melting points of a substance with a giant ionic structure and a substance with a simple molecular structure.
### Summary

1. The following table summarizes the properties of substances with the four types of structures: giant ionic, giant covalent, simple molecular and giant metallic.

<table>
<thead>
<tr>
<th>Type of structure</th>
<th>Giant ionic</th>
<th>Giant covalent</th>
<th>Simple molecular</th>
<th>Giant metallic</th>
</tr>
</thead>
<tbody>
<tr>
<td>Where this type of structure is found</td>
<td>compounds formed between metals and non-metals</td>
<td>Group IV elements and some of their compounds</td>
<td>some non-metallic elements and some compounds formed between non-metals</td>
<td>metals</td>
</tr>
<tr>
<td>Examples</td>
<td>sodium chloride, magnesium oxide</td>
<td>diamond, graphite, silicon dioxide</td>
<td>hydrogen, iodine, ammonia, dry ice</td>
<td>aluminium, copper, magnesium</td>
</tr>
<tr>
<td>Particles present</td>
<td>ions</td>
<td>atoms</td>
<td>small molecules</td>
<td>positive ions surrounded by a ‘sea’ of electrons</td>
</tr>
<tr>
<td>Example:</td>
<td>NaCl</td>
<td>diamond</td>
<td>dry ice</td>
<td>magnesium</td>
</tr>
<tr>
<td>Solubility in water and non-aqueous solvents</td>
<td>usually soluble in water but insoluble in non-aqueous solvents</td>
<td>insoluble in water and non-aqueous solvents</td>
<td>usually insoluble in water but soluble in non-aqueous solvents</td>
<td>usually insoluble in water* and non-aqueous solvents</td>
</tr>
<tr>
<td>Electrical conductivity</td>
<td>good conductors when molten or in aqueous solution</td>
<td>non-conductors (except graphite)</td>
<td>non-conductors</td>
<td>good conductors</td>
</tr>
</tbody>
</table>

* Some metals react with water to give hydrogen.

2. Allotropes are two (or more) forms of the same element in which the atoms or molecules are arranged in different ways.

3. From the group numbers of elements that make up a compound, we can tell the type of bonding present and the structure of the compound. We can then predict the properties of the compound from its structure.

4. Information such as state (at room temperature and pressure), melting point and electrical conductivity of a substance can help us to predict its structure.

### Exam tips

- Students should be able to give the meaning of the term ‘allotrope’.
- Silicon and carbon belong to the same group. They have similar structures.
- The types of bonding and structure in silicon are covalent bond and giant covalent network, but NOT ‘giant covalent bonds’.
- Although carbon and silicon belong to the same group, they form oxides with different structures. In quartz, attractions between the atoms are covalent bonds. In carbon dioxide, attractive forces between the molecules are weak van der Waals’ forces.
- When considering substances with simple molecular structures (e.g. carbon dioxide and iodine), it is important to distinguish between the covalent bonds holding atoms within the molecules and the weak van der Waals’ forces between the molecules.
- Questions often ask about the structure of iodine.
- When iodine sublimes, it absorbs energy as energy is required to overcome the attractive forces between the iodine molecules.
- Molten sulphur contains mobile sulphur molecules. Molten sulphur does NOT conduct electricity.
**Example**

a) Draw a three-dimensional diagram for the structure of each solid substance below.
   
i) Diamond
   
ii) Graphite
   
iii) Sodium chloride (3 marks)

b) With reference to the bonding and structure of diamond, explain why diamond is so hard. (2 marks)

c) With reference to the bonding and structure of graphite, explain why
   
i) graphite used to make pencil cores can be easily detached to form markings on paper; (2 marks)
   
ii) graphite can be used to make electrodes. (2 marks)

d) Explain whether solid sodium chloride is an electrical conductor. (2 marks)

**Answer**

a) i)

![Diagram of Diamond]

key: carbon atom

(1)

ii)

![Diagram of Graphite]

van der Waals' forces

key: carbon atom

(1)

b) Diamond has a giant covalent structure.

Each atom is bonded to other atoms by strong covalent bonds. Relative motion of the atoms is restricted.

(1)

c) i) There are weak van der Waals' forces between the layers of carbon atoms in graphite.

When graphite is pressed onto a piece of paper, the layers can slide over each other and detach to form markings on paper.

(1)

ii) In graphite, each carbon atom uses three electrons in forming covalent bonds with three other carbon atoms.

The remaining outermost shell electron of each carbon atom is delocalized between the layers of carbon atoms.

Hence graphite can conduct electricity.

(1)

d) Solid sodium chloride contains ions. The ions are held together by strong ionic bonds.

The ions are not mobile.

Hence solid sodium chloride cannot conduct electricity.

(1)

**Remarks**

- Questions often ask about the allotropes of carbon: diamond, graphite and buckminsterfullerene.

- Students should be able to draw clearly the three-dimensional structures of diamond and graphite.

- Questions often ask students to explain the properties of graphite / diamond / buckminsterfullerene based on their bonding and structures.

- Questions often ask students to explain why an ionic compound is an insulator of electricity in the solid state, but it conducts electricity in the molten state.

Notice that a molten ionic compound conducts electricity due to the presence of mobile ions, NOT due to the movement of delocalized electrons.