STRUCTURE AND BONDING

Table of Contents

TYPES OF CHEMICAL BONDS

Organizer

Objectives

By the end of the chapter the learner should be able to:

(a) State the significance of valence electrons in bonding.

(b) Explain qualitatively the formation of ionic, covalent and metallic bonds.

(c) Diagrammatically illustrate ionic, covalent, dative, hydrogen bonds and van der Waals forces.

(d) Predict the bond type and structure of a given substance from its physical properties.

(e) Explain the changes in bond type across period 3.

(f) Select materials for use based on bond types and structure.

(g) Predict the properties of a given substance on the basis of the bonds present.

STRUCTURE AND BONDING

When atoms of the same or different elements combine during a chemical reaction, a mutual force of attraction develops between them.

A **Chemical Bond** is the mutual force of attraction that holds the particles of atoms of the same or different elements together.

Structure is a regular pattern in which chemical bonds hold the particles of a substance together.

Atoms of other elements achieve the stable electron arrangement of noble gases by gaining, losing or sharing the valence electrons.

The attraction between unlike charges and repulsion forces between like charges forms an important basis in bonding.

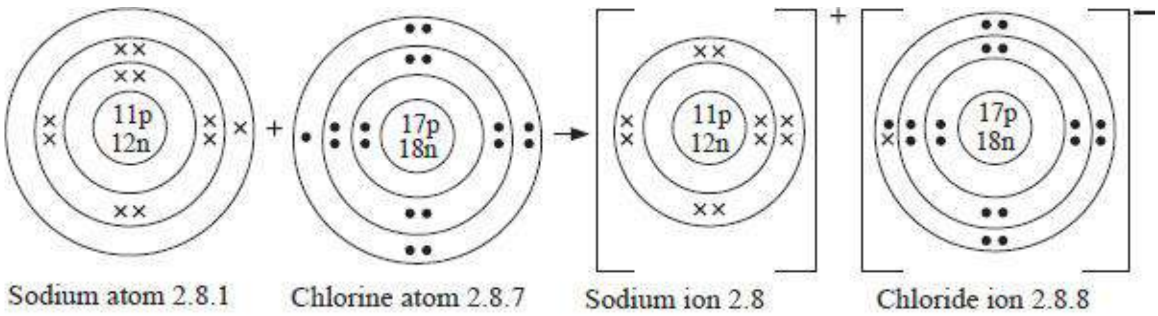
Ionic Bond (electrovalent bond)

An **ionic bond** is the **electrostatic force of attraction** between **ions with opposite charges.**

An ionic bond *is formed when there is* ***complete transfer of valence electrons from one atom to another resulting in two ions with opposite charges*** which mutually attract one another. The resulting compound is known as an **ionic compound**.

Generally, the reaction between metals and non-metals results in the formation of ionic compounds.

For example, sodium reacts with chlorine to form sodium chloride as shown below.



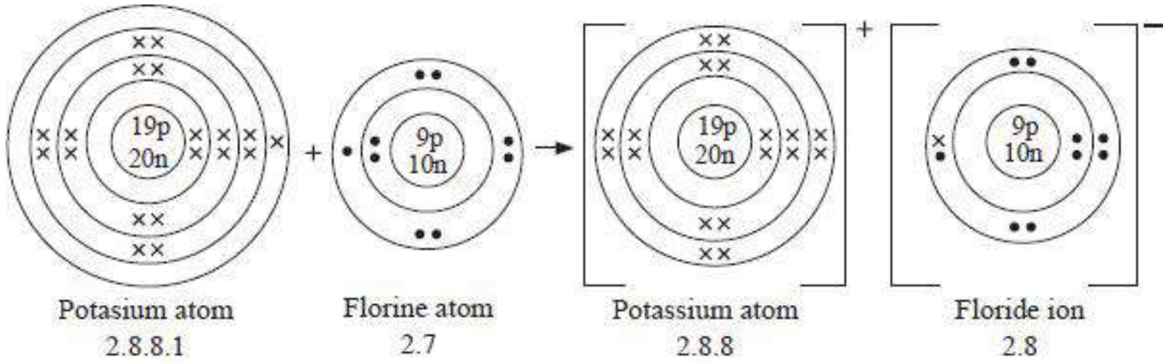
$$Na + Cl \longrightarrow Na^{+}Cl^{-1}$$

The sodium atom reacts by losing its single valence electron to the chlorine atom. The resulting sodium particle has 10 electrons and 11 protons. This results in the formation of a sodium ion with a net positive charge, Na^+. The chlorine atom on the other hand, accepts the electron donated by the sodium atom into its outermost energy level. The resulting particle (a chloride ion) has 18 electrons and 17 protons. Thus the chloride ion formed has net negative charge, Cl^-.

The sodium ion with a net positive charge and the chloride ion with a net negative charge attract each other in an **ionic bond** and the resulting compound is referred to as an **ionic compound**.

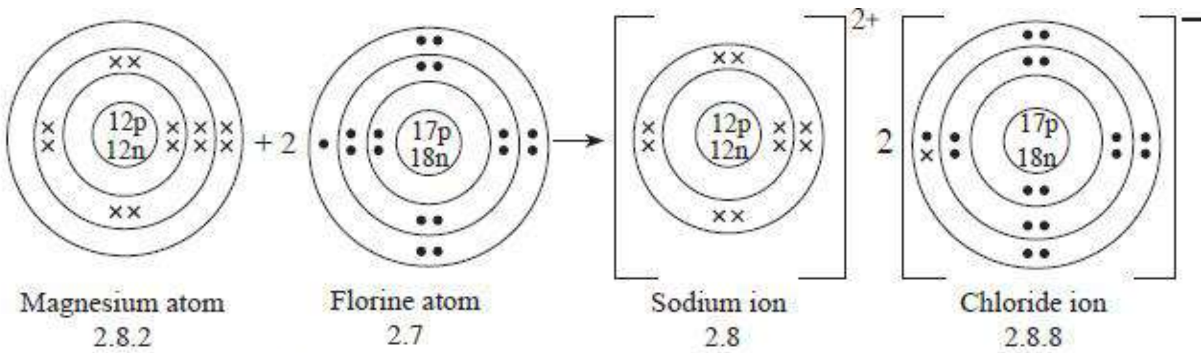
Other examples of ionic compounds include potassium fluoride, magnesium oxide and magnesium chloride.

1. **Potassium flouride**



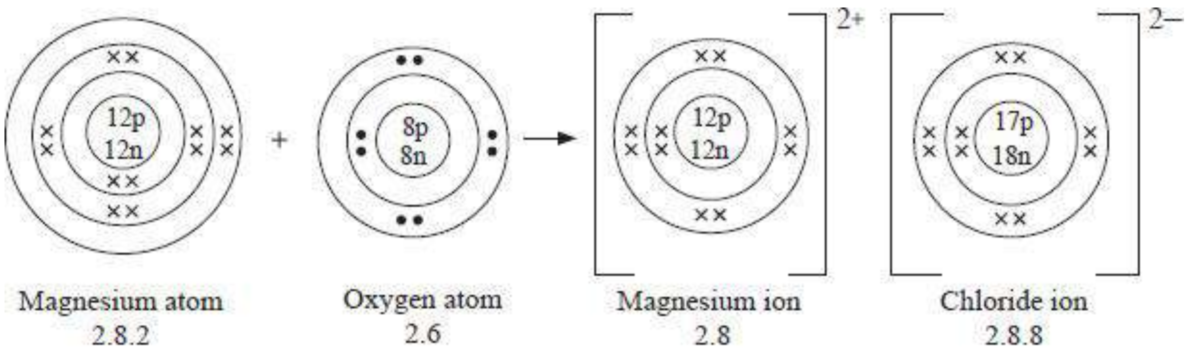
$K + F \longrightarrow KF$

2. Magnesium chloride



$Mg + 2Cl \longrightarrow MgCl_2$

3. **Magnesium oxide**.



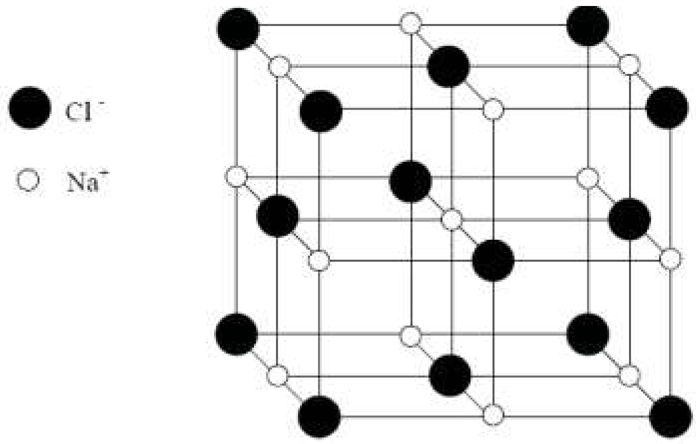
$Mg + O \longrightarrow MgO$

Giant Ionic Structures

Substances with ionic bonds are **crystals** with **giant ionic structures.**

A **crystal** is a solid form of a substance in which the particles are arranged in a definite pattern repeated regularly in three dimensions.

Most ionic substances are crystalline in nature.



For example, the sodium chloride structure shown alongside consists of many sodium ions and chloride ions that are arranged and packed in a regular pattern. Each sodium ion is surrounded by six chloride ions that are equidistant from it. Similarly, each chloride ion is surrounded by six sodium ions. This pattern repeats itself many times in all directions. The result is the formation of a giant ionic structure.

Other examples of ionic substances with giant ionic structures include **potassium nitrate, sodium iodide, potassium bromide and calcium nitrate.**

Physical Properties of Ionic Compounds

Below is a summary of some physical properties of ionic compounds.

Property \ Compound		Sodium fluoride	Sodium chloride	Sodium bromide	Sodium iodine
Solubility in water		Soluble	Soluble	Soluble	Soluble
M.P (°C)		993	801	747	661
B.P. (°C)		1695	1413	1390	1304
Electrical conductivity of:	Solid	Does not	Does not	Does not	Does not
	Melt or solution	Conducts	Conducts	Conducts	Conducts

Discussion Questions

1. **Comment on the solubility of ionic substances in water.**

Most ionic substances **dissolve in water** because **they are made up of oppositely charged ions which are attracted by the polar water molecules.**

2. **Explain why the melting and boiling points of ionic compounds are generally high.**

Ionic substances have high melting and boiling points. This is because **ionic bonds are strong**. Melting and boiling involve breaking these ionic bonds. A lot of heat energy is required to break the bonds.

3. **Explain why ionic compounds conduct electricity in the molten state and in aqueous solution but not in the solid state.**

Ionic compounds **do not conduct electricity in the solid state**. This is because **the ions forming the structure are not mobile but occupy fixed positions in the structure**. In the **molten state or aqueous solutions,** the ions are **mobile within the molten liquid** or solution and therefore **conduct** electricity.

Ionic substances conduct electricity by use of **molten ions.**

2. Covalent and Co-ordinate Bond

Covalent bond

A covalent bond is a bond formed when the combining atoms **share a pair (or pairs) of electrons between them**, **each donating an electron** to the shared pair.

Covalent bonding occurs when atoms of non-metals combine. The combining atoms could be of the same element or from different elements. It is common in **molecules**. Substances with covalent bonds either have **molecular or giant atomic structures.**

A molecule *is a group of atoms (two or more) of the same or different elements that are held together by strong covalent bonds*.

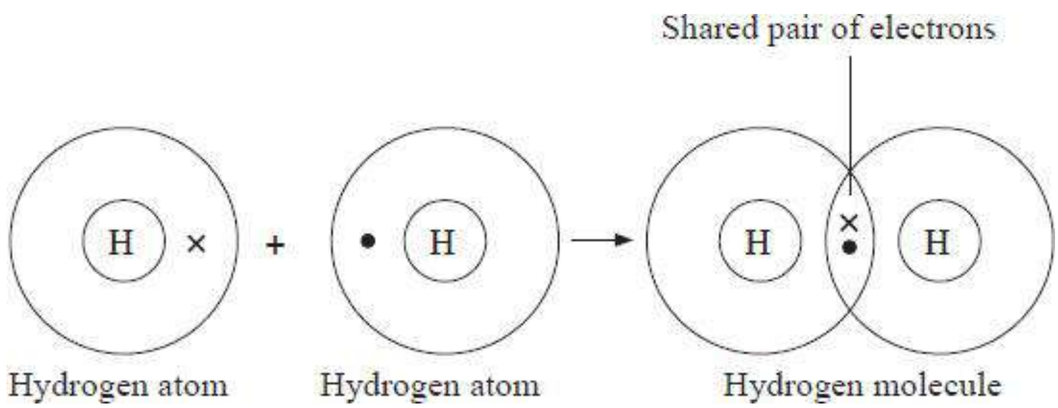
Substances consisting of molecules include **water, most gases, sugar, oils, fats, naphthalene, paraffin wax and sulphur.**

Substances consisting of molecules are referred to as **molecular substances.**

For example, two hydrogen atoms can combine to form a hydrogen molecule.

Each hydrogen atom has only one electron in its occupied energy level. It therefore needs to gain one electron to attain the stable electron arrangement of 2 helium (duplet state).

Since the combing atoms are of the same element, none would readily lose an electron to the other. The atoms therefore end up sharing a pair of electrons, each atom contributing an electron to the shared pair.



The covalent bond formed by the two hydrogen atoms in a hydrogen molecule can also be represented as follows.

$$H \overset{\times}{\bullet} H \quad \text{Or} \quad H-H$$

The **single line (–)** between the hydrogen atoms represents a **covalent bond** consisting of a shared pair of electrons.

A **single shared pair of electrons is represented by a single line (–)** and is called a **single covalent bond e.g. H – H.**

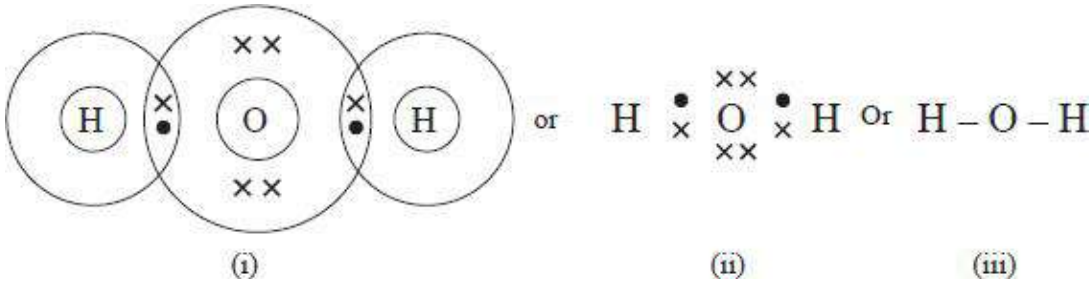
Two shared pairs of electrons are represented by **two lines (=)** and are called a **double covalent bond e.g. O=O**

Three shared pairs of electrons are represented by **three lines thus: (≡)** and are referred to as a **triple covalent bond e.g. N≡N.**

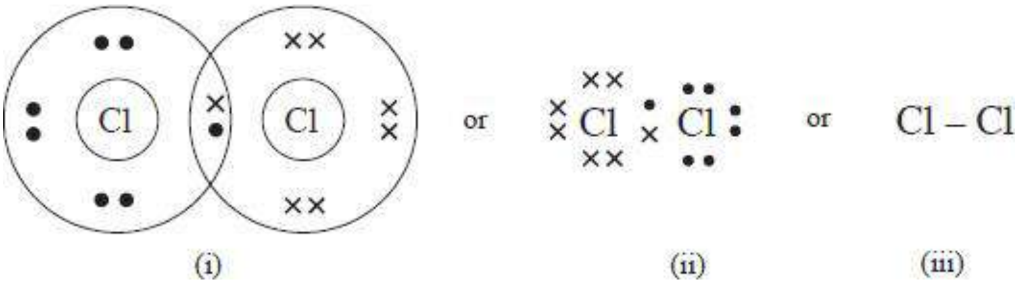
When drawing the structures, only the valence electrons are shown.

Other examples of substances which are covalently bonded include:

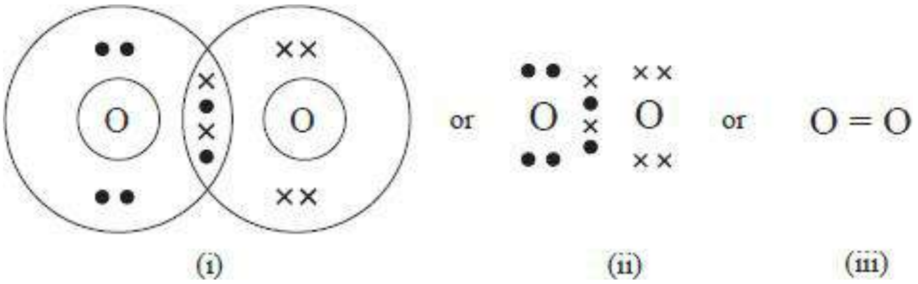
(i) Water molecule, H_2O



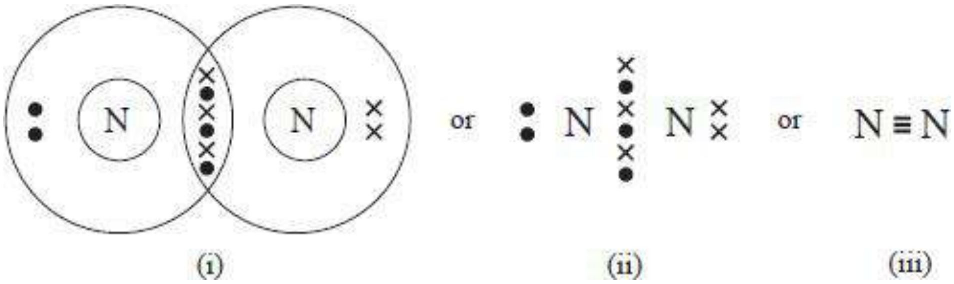
(ii) Chlorine molecule, Cl_2



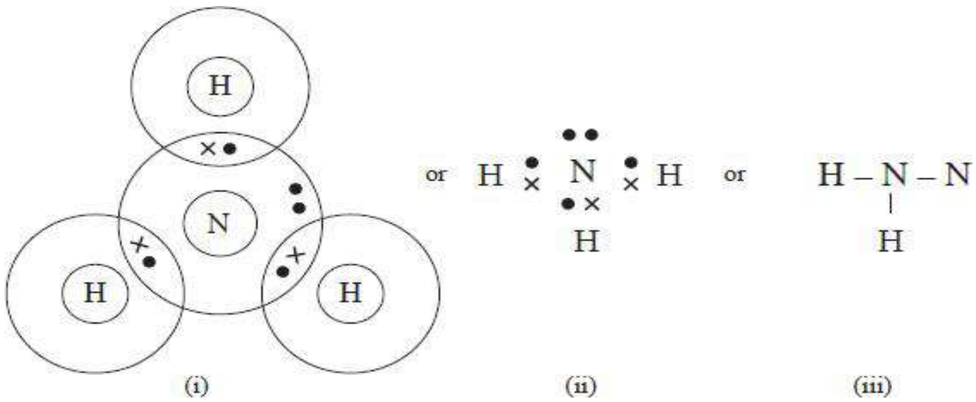
(iii) Oxygen molecule, O2



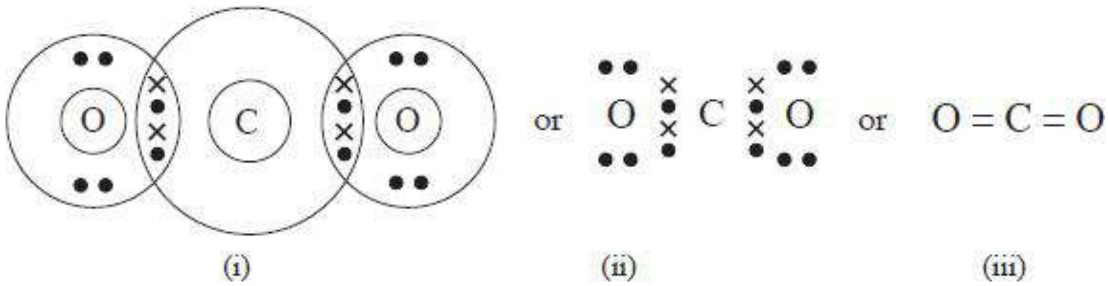
(iv) Nitrogen molecule, N_2



(v) Ammonia, NH_3



(vi) Carbon (IV) oxide, CO_2



Co-ordinate Bond (Dative bond)

A co-ordinate bond *is a type of a covalent bond in which the shared pair of electrons forming the bond is contributed by only one of the atoms forming the bond.*

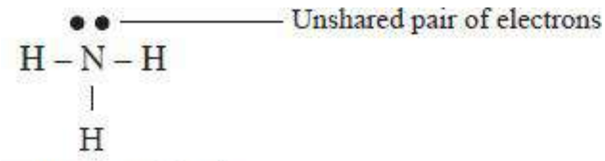
Examples of substances which have coordinate bonds include ammonium ion (NH_4^+) ;carbon (II) oxide, CO, ;hydroxonium ion, H_3O^+ ;Ozone, O_3, aluminum chloride dimer, Al_2Cl_6 and nitric (v) acid, HNO_3

(i) Ammonium ion (NH_4^+).

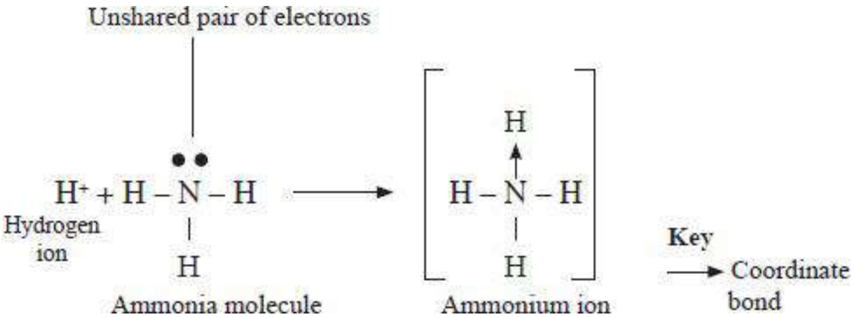
In the case of ammonium ion, the pair of electrons forming the co-ordinate bond is contributed by the ammonia molecule.

A hydrogen ion (H^+) and an ammonia molecule (NH_3) combine to form an ammonium ion (NH_4^+).

The hydrogen ion has no electron around its nucleus. On the other hand, the ammonia molecule has an unshared pair of valence electrons that has not been used in bonding.



Ammonia molecule



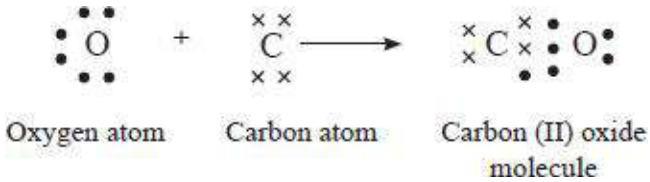
When the hydrogen ion combines with the ammonia molecule, the hydrogen ion accepts the pair of electrons into its empty first energy level.

The total number of protons in the ammonium ion is more than the total number of electrons resulting in a net positive charge.

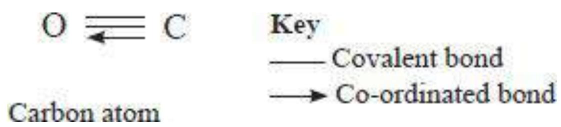
(ii) Carbon (II) oxide

In carbon (II) oxide, Oxygen donates one of its two unshared pairs of electrons to form a co-ordinate bond with the carbon atom.

A carbon atom has four valence electrons and therefore needs to gain four electrons to fill its outermost energy level. Oxygen on the other hand has six valence electrons and needs two electrons to attain the stable configuration of 2.8.

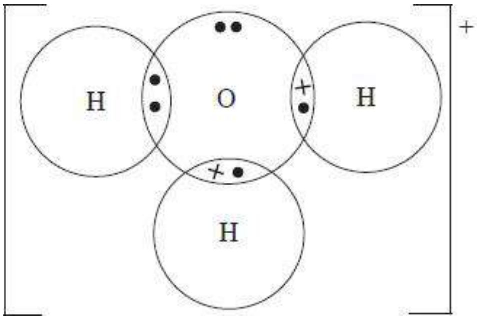


The oxygen atom and the carbon atom in carbon (II) oxide form two covalent bonds and a co-ordinate bond.



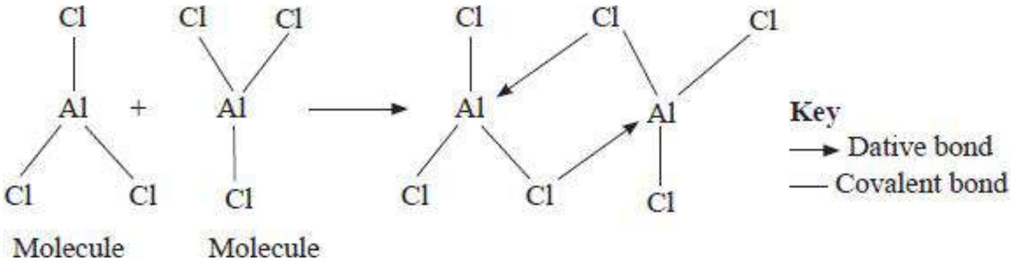
(iii) Hydroxonium ion (H3O+)

In hydroxonium ion, a water molecule donates one of its unshared pair of electrons to form a coordinate bond with the hydrogen ion. The resulting ion has a net positive charge as shown below.

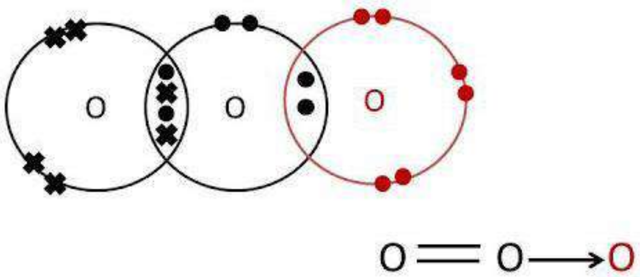


(iv) **Aluminium chloride dimer**.

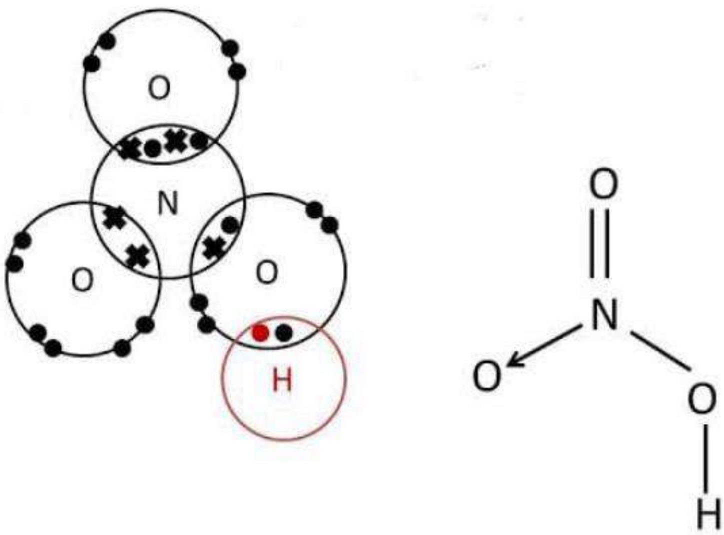
Aluminium combines with chlorine by forming covalent bonds to form a molecule of aluminium chloride (Al_2Cl_3). Two aluminium chloride molecules then combine through coordinate bonds to form a dimer as shown below.



(v) Ozone, O_3



(vi) **Nitric (V) acid, HNO_3**



Molecular Structures

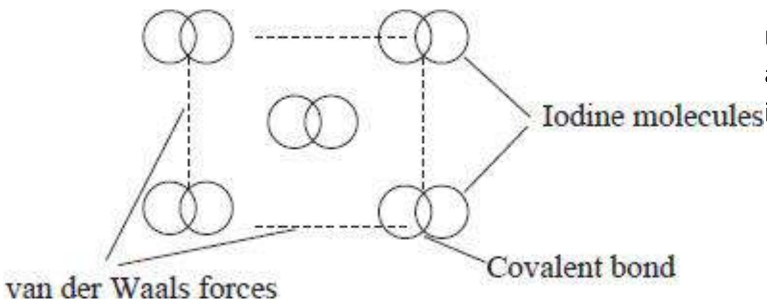
Most molecular substances are **gases or liquids at room temperature.** Some molecular substances include sulphur, sugar, iodine, fats, paraffin wax and naphthalene.

Discussion Question:

Using the structure of iodine, differentiate between a covalent bond and Van der waals forces.

Van der Waals forces are **weak forces of attraction** between **molecules or atoms** which exist **only when the particles are close together.**

In molecules such as iodine, the atoms forming the molecules are held together by strong covalent bonds. The molecules in the solid substance are in turn held together in a regular pattern by weak van der Waals forces. The regular pattern repeats itself many times resulting to a molecular structure as shown below.



The van der Waals forces hold the iodine molecules together in layers. This arrangement explains the flaky nature of iodine crystals.

Physical Properties of Substances with Molecular Structures.

The table below gives the physical properties of molecular substances.

Substance / Property	Sugar (sucrose)	Naphtha-lene	Iodine	Rhombic sulphur	Water	Hydrogen sulphide
Solubility in water	Soluble	Insoluble	Insoluble	Insoluble	–	Slightly soluble
Molecular mass	183	128	186	256	18	34
Melting point (°C)	186	82	113	114	0	–85
Boiling point (°C)	–	218	183	444	100	–60
Electrical conductivity	Poor	Poor	Poor	Poor	Poor	Poor

Discussion Questions

1. **Give the reasons why molecular substances are poor conductors of electricity?**

Molecular substances are **poor conductors of both heat and electricity.** This is because they have structures composed of molecules. There are n**o ions or delocalised electrons in the structures to enable them to conduct electricity.**

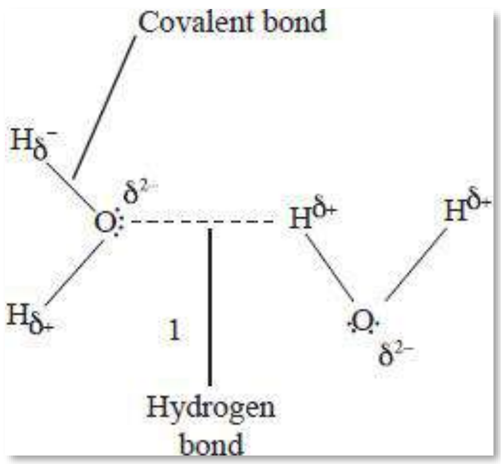
2. **What is the general trend in the melting and boiling points in relation to the molecular masses?**

Molecular substances **have low melting and boiling points**. Although the atoms forming a molecule are held by strong covalent bonds, the **intermolecular forces are usually the weak van der Waals forces.** As melting and boiling involves breaking the weak van der Waals forces, the melting and boiling points are low. Melting and boiling points of molecular substances increase with the increase in relative molecular mass.

3. **Explain the effect of hydrogen bonds on melting and boiling points of molecular substances**

Although the melting point and boiling point of molecular substances increases with increase in molecular mass, some substances however, display a disparity.

For example, water (H_2O) with relative molecular mass of 18 has a higher melting point of (°C) than hydrogen sulphide with a relative molecular mass 34 and a melting point of –85°C. In water the molecules are held together by hydrogen bonds while molecules in H_2S are held by the weak van der Waals forces.



A hydrogen bond *is an intermolecular force in which the electropositive hydrogen atom of one molecule is attracted to an electronegative atom of another molecule*.

In water, the electropositive hydrogen of one molecule is attracted to the electronegative oxygen of another molecule.

A hydrogen bond is **stronger than the van der Waals** forces but **weaker than a covalent bond.**

The influence of hydrogen bonding on the physical properties of molecular substances is also illustrated by **ethanol (C_2H_5OH) and dimethyl ether (C_2H_6O).** Both substances have the same relative mass, 46. However, dimethyl ether boils at –24°C while ethanol boils at 78.4°C. This is because ethanol has hydrogen bonds as the intermolecular forces while dimethyl ether has the weaker van der Waals forces.

4. State and explain the differences in the solubility of molecular substances?

Most molecular substances are insoluble in water. However, some such as sugar, ethanol and ethanoic acid are soluble due to the presence of hydrogen bonding.

Giant Covalent Structure (Giant Atomic structures)

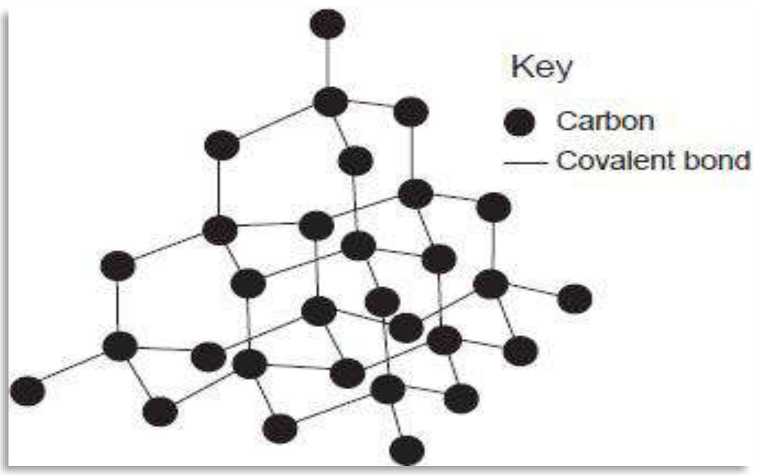
A giant covalent structure is a structure that consists of an indefinite number of atoms which are all covalently bonded together. The pattern repeats itself and extends in all directions.

Substances with giant atomic structures include **diamond, graphite and silicon (IV) oxide.**

Diamond

Diamond is an allotrope of carbon. **Allotropes are different forms of an element in the same physical state.**

In the structure of diamond each carbon atom is bonded to other carbon atoms by strong covalent bonds. The carbon atoms in diamond form an octahedral structure as shown below.



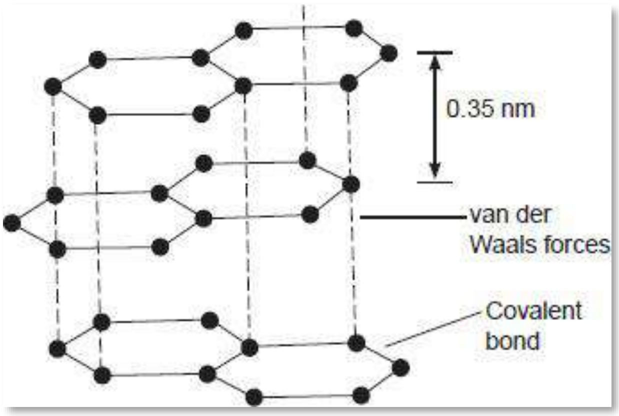
The pattern repeats itself in all directions resulting in a giant atomic structure.

In diamond, all the valence electrons of each carbon atom are used in bonding. There are therefore **no delocalised electrons** in the structure. Thus, diamond is a **poor conductor of heat and electricity.**

Diamond is the **hardest known substance**. This is because **all the bonds in diamond are strong covalent bonds and the atoms in the structure are closely packed**. For this reasons its melting point is **very high (3700°C)**

Graphite

Each carbon atom in graphite is bonded to three other carbon atoms. Because only three of the four valence electrons in each atom are used in bonding, the forth valence electron is **delocalized** in the structure.

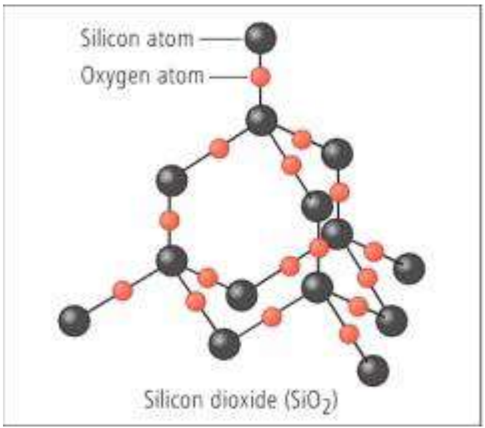


The structure of graphite consists of layers in which the carbon atoms are held together by strong covalent bonds forming hexagonal arrangement as shown below. Each layer is held to the other by van der Waals forces.

The **presence of delocalised electrons** in the structure of graphite explains its electrical **conductivity**.

Layers of graphite are **held together by van der Waals forces, therefore easily slide over each other when pressed and this gives graphite its slippery feel**.

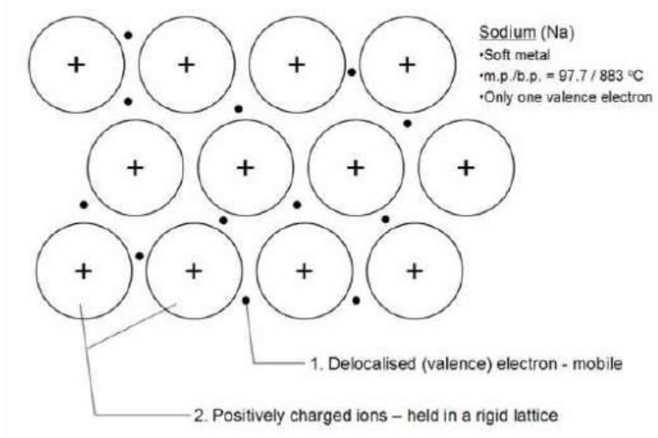
Silicon (IV) Oxide



Silicon has four electrons in its outer occupied energy level. Each silicon atom is bonded covalently to four oxygen atoms forming a giant covalent structure as below. .

The Metallic Bond

Metallic bond is the **electrostatic force of attraction** between the **positive nuclei** and the **delocalised valence electrons** in the metallic structure.



In a metallic structure there are many atoms surrounding any one atom. The valence electrons are therefore mutually attracted to many nuclei. This gives rise to a situation where the positive nuclei appear to be immersed in a sea of mobile electrons. The mobile electrons in the structure are said to be delocalised.

Metallic bonds are strong bonds. The pattern is repeated many times resulting in a giant metallic structure.

Physical Properties of Metals

The table below gives a summary of some physical properties of metals.

Metal	Valency	Melting point (°C)	Boiling point (°C)	Atomic radii (nm)	Electrical conductivity
Lithium	1	180	1330	0.133	Good
Sodium	1	98	890	0.155	Good
Potassium	1	64	774	0.203	Good
Magnesium	2	651	1110	0.136	Good
Aluminum	3	1083	2582	0.125	Good

Discussion Questions

1. **Why are metals good conductors of electricity?**

All metals are good conductors of both heat and electricity. This is because there are **delocalised electrons** in the metallic structure. Thermal and electrical conductivity increases with the increased number of delocalised electrons from each atom in the structure. This explains why aluminium is a better conductor than magnesium.

2. **Explain why Metals have high melting points.**

Metals have relatively high melting and boiling points. This is because the metallic bond is a strong bond.

3. **Explain why potassium has a lower melting point than lithium.**

The melting point of lithium is higher than that of potassium because lithium has a **smaller atomic size** than potassium, therefore it has a stronger metallic bond.

4. **Magnesium has lower melting point than aluminium. Explain.**

Aluminium has a higher melting point and boiling point than magnesium. This is because aluminium has a **smaller atomic size than magnesium** therefore it has stronger metallic bonds. Aluminium also has **more delocalized electrons** than magnesium which also contribute to the strength of the metallic bond.

5. **Aluminium is a better electrical and thermal conductor than magnesium. Explain.**

Thermal and electrical conductivity increases with the increased number of delocalised electrons from each atom in the structure. This explains why aluminium is a better conductor than magnesium.

Types of Bonds Across a Period

Bond types vary from metallic to covalent across a period. The structures also vary from giant metallic, giant covalent to molecular. It is expected that similar compounds of the elements in period 3 will exhibit variation in bond types, structure and properties. This can be illustrated by studying the bond types and properties of oxides and chlorides of elements in period 3.

Oxides of Elements in Period 3

- The elements in period 3 form oxides when they react with oxygen.
- The oxides of sodium and magnesium **dissolve** in water to form **alkaline solutions.**
- Aluminum oxide **does not dissolve** in water.
- The oxides of the **non-metals** dissolve in water to form acidic solutions, however **silicon (IV) oxide does not dissolve in water.**

Oxide	NaO	MgO	Al_2O_3	SiO_2	P_2O_5	SO_2	Cl_2O_7
Physical state	Solid	Solid	Solid	Solid	Solid	Gas	Gas
M.P. (°C)	1193	3075	2045	1728	563	–76	–60
B.P. (°C)	1278	3601	2980	2231	301	–10	–9
Struc-ture	Giant ionic	Giant ionic	Giant ionic	Giant ionic	Molecular	Molecular	Molecular
Type of bonding	Ionic	Ionic	Ionic	Covalent	Covalent	Covalent	Covalent
Nature of oxides	Basic (alkaline)	Basic (Alkaline)	Ampho-teric	Acidic	Acidic	Acidic	Acidic
Reac-tion with acids	Reacts to form salt and water	Reacts to form salt and water	Reacts to form salt and water	No reaction	No reac-tion	No reaction	No reaction

- Oxides of sodium and magnesium react with acids to form a salt and water.
- Aluminium oxide reacts with both acids and alkalis and therefore it is an **amphoteric oxide**.
- Oxides of the non-metals do not react with acids but react with alkalis.
- All the oxides of elements in period 3 **except** those of **sulphur and chlorine** are **solids**. Sulphur (IV) oxide is a **gas** at room temperature. Several gaseous oxides of chlorine do exist, however they cannot be prepared in the laboratory.

Discussion Questions

1. **Explain why sulphur (IV) oxide is a gas at room temperature.**

The atoms of sulphur and oxygen are held together by covalent bonds but the molecules of sulphur (IV) oxide are attracted to each other by weak van der Waals forces.

2. **Why is the melting point of magnesium oxide higher than that of sodium oxide?**

Both sodium oxide and magnesium oxide have a giant ionic structure. However, the melting point of magnesium oxide is higher because the electrostatic forces of attraction between magnesium ions and oxide ions are stronger. This is due to the fact that the magnesium ion has a charge of $^{+}2$ and is smaller in size than the sodium ion.

3. **Explain the high melting and boiling points of silicon (IV) oxide.**

Silicon (IV) oxide has a giant covalent structure. Each silicon atom is attached to four oxygen atoms and each oxygen atom is attached to two silicon atoms. Silicon uses all its valency electrons to form strong covalent bonds with oxygen. The covalent bonds in silicon are extra ordinarily strong. This explains the high melting point of silicon (2231°C).

4. **What is the general trend in the bond types of the oxides of elements across period 3?**

In general the bond types change from ionic to covalent across the period. The structure of the oxides changes from giant ionic, to giant atomic and finally to molecular. This explains the trends in properties of the oxides.

Chlorides of Elements in Period 3

Most elements of period 3 form stable chlorides.

The trend in bond types, structure and properties of chlorides of period 3 elements show variation across the period as shown in the table below.

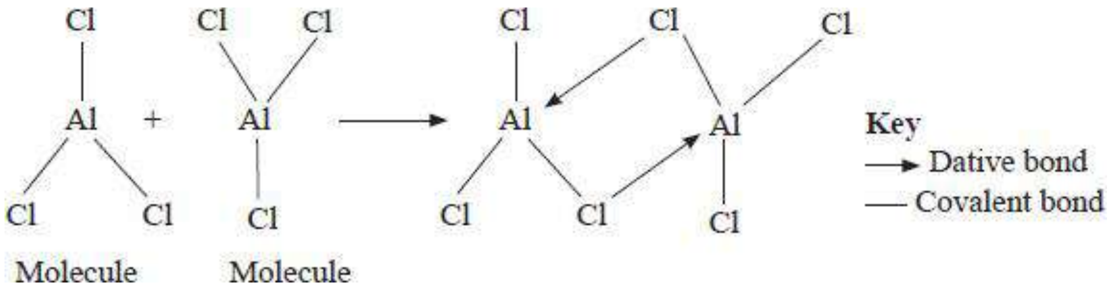
Chloride	NaCl	$MgCl_2$	AlCl	$SiCl_4$	PCl_5	SCl_2
Physical state	Solid	Solid	Solid	Liquid	Liquid	Gas
m.P. (°C)	801	714	180 (s)	–70	–94 (s)	–78
b.P (°C)	1467	1437	–	57	–	59 (d)
Structure	Giant ionic	Giant ionic	Molecular (Dimer)	Molecular	Molecular	Molecular
Bond type	Ionic	Ionic	Ionic/ Covalent	Covalent	Covalent	Covalent
pH of solution	7	7	3	3	3	3
Effect on water	Soluble	Soluble	Hydrolysed	Hydrolysed	Hydro-lysed	Hydrlysed

Key: s – sublimes, d – decomposes

Discussion Questions

Explain how the chlorides of elements in Period 3 dissolve in water.

- Sodium chloride dissolves in water resulting in a **slight drop in temperature.**
  Magnesium chloride dissolves readily with a **small increase in temperature.** Both chlorides form **neutral solutions. These chlorides are ionic and therefore fully dissociate into ions.**
- **Anhydrous aluminium chloride** differs from the other metallic chlorides because **it exists in molecular form.** Aluminium combines with chlorine by forming covalent bonds to form a molecule of aluminium chloride (Al_2Cl_3). Two aluminium chloride molecules then combine through coordinate bonds to form a **dimer** as shown below.



Formation of a dimer in aluminium chloride

Aluminium chloride **is hydrolysed** by water to form an **acidic solution** and therefore **behaves as a covalent rather than an ionic chloride.**

Hydrolysis is a reaction that involves breaking down a substance by water.

Molecular chlorides undergo hydrolysis.

- Silicon (IV) chloride is also hydrolysed by water to form a solution which is acidic. Alot **of heat is given out** and fumes of hydrogen chloride gas are given out.

Silicon (IV) chloride + Water $\longrightarrow$ Silicon (IV) Oxide + Hydrogen chloride

$$SiCl_4(l) + 2H_2O\ (l) \longrightarrow SiO_2(s) + 4HCl(aq)$$

- Phosphorus (III) chloride or phosphorus (V) chloride react vigorously with water to form an acidic solution. A lot of **heat is also evolved** and this makes the temperature of the water to rise.

Phosphorus (III) chloride + Water $\longrightarrow$ Phosphorous acid + Hydrogen chloride

$$PCl_3(s) + 3H_2O(l) \longrightarrow H_3PO_3(s) + 3HCl(g)$$

Phosphorus (V) chloride + Water $\longrightarrow$ Phosphoric acid + hydrogen chloride

$$PCl_5(s) + 4H_2O(l) \longrightarrow H_3PO_4(aq) + 5HCl(aq)$$

Summary of Characteristics of Bonds

Property	Substances with		
	Covalent bonds	Ionic bonds	Metallic bonds
Electrical conductivity	Non-conductors except graphite	Solids do not conduct. Aqueous solutions and molten state conduct	Conduct
Thermal conductivity	Non-conductors, except graphite	Do not conduct	Conduct
Melting point (°C)	Low for molecular High for giant structures	Usually high	Generally high
Boiling point (°C)	Low if molecular High if giant structure	Usually high	Generally high
Solubility	Generally insoluble in water but soluble in organic solvents	Generally soluble in water	Some metals react with water

1. 2006 Q 5 P1
The atomic numbers of elements C and D are 19 and 9 respectively. State and explain the electrical conductivity of the compound CD in:
(a) Solid state (1½marks)
(b) aqueous state. (1½ marks)

2. 2006 Q 20 P1
(a) Distinguish between a covalent bond and a co-ordinate bond. (2 marks)

(b) Draw a diagram to show bonding in an ammonium ion. (1 mark)

3. 2006 Q21 P1

(a) Explain why the metals magnesium and aluminium are good conductors of electricity. (1 mark)

(b) Other than cost, give two reasons why aluminium is used for making electric cables while magnesium is not. (2 marks)

4. 2007 Q 3b P2

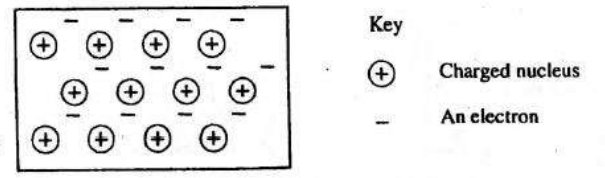
Both chlorine and iodine are halogens.

(a) What are halogens? (1mark)

(b) In terms of structure and bonding, explain why the boiling point of chlorine is lower than that of iodine. (2 marks)

5. 2007 Q 29 P1

The diagram below is a section of a model of the structure of element T.

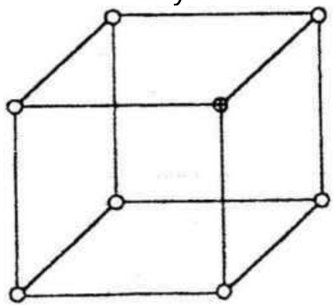


(a) State the type of bonding that exists in T. (1 mark)

(b) In which group of the period table does element T belong? Give a reason. (2 marks)

6. 2007 Q 5a-c P2

(a) The diagram below represents part of the structure of a sodium chloride crystal. The position of one of the sodium ions in the crystal is shown as ⊕

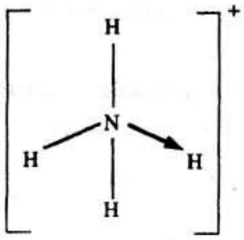


(i) On the diagram, mark the position of the other three sodium ions (2 marks)

(ii) The melting and boiling points of sodium chloride are 801°C and 1413°C respectively. Explain why sodium chloride does not conduct electricity at 25°C, but does so at temperatures between 801° C and 1413°C (2 marks)

(b) Give a reason why ammonia gas is highly soluble in water (2 marks)

(c) The structure of an ammonia ion is shown below:



Name the type of bond represented in the diagram by N→ H (1 mark)

7. 2008 Q 11 P1
The table below gives atomic numbers of elements represented by the letters A, B, C and D.

Element	**A**	**B**	**C**	**D**
Atomic number	15	16	17	20

Use the information to answer the questions that follow.

(a) Name the type of bonding that exists in the compound formed when A and D react (1 mark)

(b) Select the letter which represents the best oxidizing agent. Give a reason for your answer. (2 marks)

8. 2008 Q 2b P2
The table below gives information about elements A_1, A_2, A_3 and A_4

Element	**Atomic Number**	**Atomic Radius (nm)**	**Ionic radius (nm)**
A_1	3	0.134	0.074
A_2	5	0.090	0.012
A_3	13	0.143	0.050
A_4	17	0.099	0.181

(i) In which period of the periodic table is element A2? Give a reason. (2 marks)

(ii) Explain why the atomic radius of:

I. A_1 is greater than that of A_2;

II. A_4 is smaller than its ionic radius (2 marks)

(iii) Select the element which is in the same group as A_3 (1 mark)

(iv) Using dots (.) and crosses(x) to represent outermost electrons, draw a diagram to show the bonding in the compound formed when A_1 reacts with A_4 (1 mark)

9. 2009 Q 3 P1
The atomic number of sulphur is 16. Write the electron arrangement of sulphur in the following (2 marks)

(a) H_2S

(b) SO_3^{2-}

10. 2009 Q 5 P1
In terms of structure and bonding, explain why the melting point of oxygen is much lower than that of sodium. (3 marks)

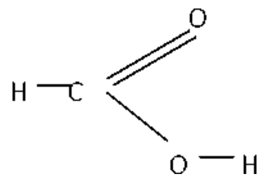
11. 2009 Q 8 P1
Using dots (.) and crosses (X), show bonding in:

(a) The compound formed when nitrogen reacts with fluorine. (Atomic numbers F= 9, N=7) (2 marks)

(b) Sodium oxide. (Atomic numbers Na = 11, O= 8) (1 mark)

12. 2009 Q17 P1

The structure of methanoic acid is



What is the total number of electrons used for bonding in a molecule of methanoic acid? Give reasons. (2 marks)

13. 2009 Q 24 P1

The boiling points of some compounds of hydrogen with some elements in groups 4 and 6 of the periodic table are given below.

Compound	Boiling point(°C)	Compound	Boiling point(°C)
CH_4	-164.0	H_2O	100.0
SiH_4	-112.0	H_2S	-61.0

(a) Which of the compounds $CH_4$and SiH_4 has the stronger intermolecular forces? (1 mark)

(b) Explain why the boiling point of H_2O and H_2S show different trends from that of CH4 and SiH_4. (2 marks)

14. 2009 Q 6a, c P2

(a) Study the table below and complete it.(W^- and X^{4+} are not the actual symbols of the ions)

Ion	Number of Protons	Number of neutrons	Mass Number	Electron Arrangement
W^-		20		2.8.8
X_4^+	14		28	

(b) The atomic numbers of Na and Mg are 11 and 12 respectively. Which of the element has higher ionization energy? Explain. (2 marks)

15. 2010 Q 14 P1

(a) Using electrons in the outermost energy level, draw the dot (.) and cross (x) diagram for the molecules H_2O and C_2H_4. (H = 1, C = 6, O = 8) (2 marks)

(i) H_2O

(ii) C_2H_4

(b) The formula of a complex ion is $[Zn(NH_3)_4]^{2+}$. Name the type of bond that is likely to exist between zinc and ammonia in the complex ion. (1 mark)

16. 2010 Q 27 P1

The atomic numbers of phosphorus, sulphur and potassium are 15, 16 and 19 respectively. The formulae of their ions are P^{3-}, S^{2-} and K^+. These ions have the same number of electrons.

(a) Write the electron arrangement for the ions. (1 mark)

(b) Arrange the ions in the order of increasing ionic radius starting with the smallest. Give a reason for the order. (2 marks)

17. 2010 Q 3 P2

Use the information in the table below to answer the questions that follow.
The letters do not represent the actual symbols of the elements.

Element	Atomic number	Melting point (°C)
R	11	97.8
S	12	650.0
T	15	44.0
U	17	-102
V	18	-189
W	19	64.0

(a) Give the reasons why the melting point of:

(i) S is higher than that of R (1 mark)

(ii) V is lower than that of U (1 mark)

(b) How does the reactivity of W with chlorine compare with that of R with chlorine? Explain (2 marks)

(c) Write an equation for the reaction between T and excess oxygen (1 mark)

(d) When 1.15g of R was reacted with water, 600cm^3 of gas was produced. Determine the relative atomic mass of R. (Molar gas volume = 24000 cm^3) (3 marks)

(e) Give one use of element V (1 mark)

18. 2011 Q 5b(iii-vii) P2

The table below gives the number of electrons, protons and neutrons in particles A, B, C, D, E, F and G.

Particle	Protons	Electrons	Neutrons
A	6	6	6
B	10	10	12
C	12	10	12
D	6	6	8
E	13	10	14
F	17	17	18
G	8	10	8

(i) Which particle is likely to be a halogen? (1 mark)

(ii) What is the mass number of E? (1 mark)

(iii) Write the formula of the compound formed when E combines with G. (1 mark)

(iv) Name the type of bond formed in (iii) above (1 mark)

(v) How does the radii of C and E compare? Give a reason. (2 marks)

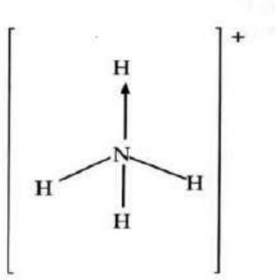
(vi) Draw a dot (.) and cross (x) diagram for the compound formed between A and F. (1 mark)

(vii)Why would particle B not react with particle D? (1 mark)

19. 2011 Q 28 P1

The diagram below shows the bonding between aluminium chloride and ammonia.
(a) Name the types of bonds that exist in the molecule. (1 mark)
(b) How many electrons are used for bonding in the molecule? (1 mark)

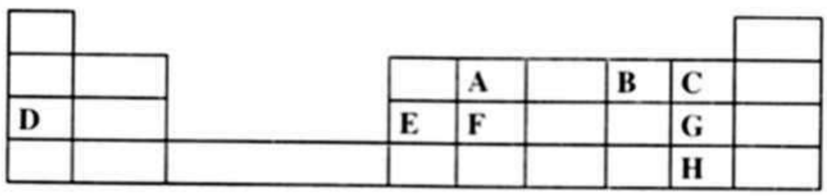
20. 2012 Q7 P1, 2016 Q6 P1
Ammonium ion has the following structure:



Label on the structure:
(a) Covalent bond; (1 mark)
(b) Coordinate (dative) bond. (1 mark)

21. 2012 Q 16 P1
Use the following information on substances S, T and hydrogen to answer the questions that follow:
I. T displaces V from a solution containing V ions
II. hydrogen reacts with the heated oxide of S but has no effects on heated oxide of V.
(a) Arrange substances S, t, V and hydrogen in the order of increasing reactivity. (2 marks)
(b) If T and V are divalent metals, write an ionic equation for the reaction in (i) above. (1 mark)

22. 2012 Q 2 P2
The grid below is part of the periodic table. Use it to answer the questions that follow. (the letters are not the actual symbols of the elements).



(a) Which is the most reactive non-metallic element shown in the table? Explain. (2 marks)

(b) (i) Write the formula of the compound formed when element **A** reacts with element **B** (1 mark)
(ii) Name the bond type in the compound formed in b(i) above (1 mark)

(c) (i) What is the name given to the group of elements where, **C**, **G** and **H** belong?
(ii) Write an equation for the reaction that occurs when C in gaseous form is passed through a solution containing ions of element H.
(1 mark)

(d) The melting points of elements **F** and **G** are 1410 °C and -101 °C respectively. In

terms of structure and bonding, explain why there is a large difference in the melting points of **F** and **G**.
(2 marks)

(e) **D** forms two oxides. Write the formula of each of the two oxides. (1 mark)

(f) **J** is an element that belongs to the 3^{rd} period of the periodic table and a member of the alkaline earth elements. Show the position of **J** in the grid (1 mark)

23. 2013 Q 4 P1
In terms of structure and bonding, explain the following observations:
(a) The melting point of aluminium is higher than that of sodium. (1½mark)
(b) Melting point of chlorine is lower than that of sulphur (1½ mark)

24. 2014 Q21 P1
Given that the atomic number of Y is 13 and that of Z is 9:
(a) Write the electronic arrangement of Y and Z; (1 mark)
(b) Draw the dot (.) and cross(x) diagram for the compound formed by Y and Z. (1 mark)

25. 2016 Q16 P1
The atomic number of sulphur is 16. Write the electron arrangement of sulphur in the following:
(a) H_2S (1 mark)
(b) SO_3^{2-} (1 mark)

26. 2017 Q21 P1.
The atomic numbers of some elements **P**, **Q**, **R** and **S** are 6, 8, 12 and 17 respectively.
(a) Draw the dot (•) and cross (X) diagrams for the compounds formed when:
(i) **R** and **Q** react (1 mark)
(ii) **P** and **S** react. (1 mark)
(b) Explain why the melting point of the compound formed by **P** and **S** is lower than that formed by **R** and **Q**. (1 mark)

27. 2018 P1
In terms of structure and bonding, explain why graphite is used as a lubricant in machines.

(3 marks)

28. 2018 Q11 P1
Element U has atomic number 12 while element V has atomic number 16. How do the melting points of their oxides compare? Explain. (3 marks)

29. 2019 P1 Q8.
Table 1 shows the properties of two chlorides, D and E.
Table 1

Chloride	Melting Points(°C)	Electrical Conductivity(liquid)
D	1074	Good
E	203	Poor

(a) State the type of bond present in:
(i) D...................... (1 mark)

(ii) E....................... (1 mark)

(b) Explain in terms of structure and bonding, the difference in electrical activity of the chlorides D and E.

(1 mark)