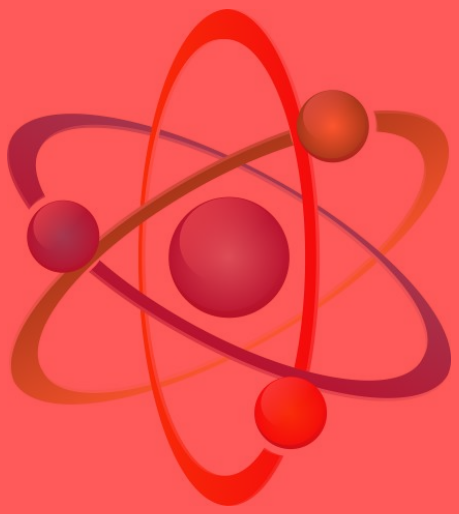
IGCSE CHEMISTRY STUDY NOTES
UNIT 3 ATOMS, ELEMENTS & COMPOUNDS



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3.2.2 Ions and ionic bonds

CORE

Describe the formation of ions by electron loss or gain

- Ions are electrically charged particles.
- They are formed from atoms by electron loss or gain.
- When an atom loses electron/s, a positive ion is formed.
- When an atom gains electron/s, a negative ion is formed.
- The magnitude of the charge on an ion is equal to the number of electrons lost or gained.

Example:
Sodium ion is formed from Sodium atom by the **loss of 1 electron** from the outer shell.

Ion forming equation
$Na \rightarrow Na^{+} + e^{-}$

An Oxide ion is formed from Oxygen atom by the **gain of 2 electrons**.

Ion forming equation
$O + 2e^{-} \rightarrow O^{2-}$

The number of electrons lost or gained is determined by the valency of an element.
Refer to ATOMS, ELEMENTS & COMPOUNDS 3.1 for notes on Valency.

Charges on ions formed by elements of respective groups

1+ 2+ The Periodic Table of Elements 3+ 4+ 3- 2- 1-

Group

I	II											III	IV	V	VI	VII	VIII
							1 H hydrogen 1										2 He helium 4
3 Li lithium 7	4 Be beryllium 9											5 B boron 11	6 C carbon 12	7 N nitrogen 14	8 O oxygen 16	9 F fluorine 19	10 Ne neon 20
11 Na sodium 23	12 Mg magnesium 24											13 Al aluminium 27	14 Si silicon 28	15 P phosphorus 31	16 S sulfur 32	17 Cl chlorine 35.5	18 Ar argon 40
19 K potassium 39	20 Ca calcium 40	21 Sc scandium 45	22 Ti titanium 48	23 V vanadium 51	24 Cr chromium 52	25 Mn manganese 55	26 Fe iron 56	27 Co cobalt 59	28 Ni nickel 59	29 Cu copper 64	30 Zn zinc 65	31 Ga gallium 70	32 Ge germanium 73	33 As arsenic 75	34 Se selenium 79	35 Br bromine 80	36 Kr krypton 84
37 Rb rubidium 85	38 Sr strontium 88	39 Y yttrium 89	40 Zr zirconium 91	41 Nb niobium 93	42 Mo molybdenum 96	43 Tc technetium –	44 Ru ruthenium 101	45 Rh rhodium 103	46 Pd palladium 106	47 Ag silver 108	48 Cd cadmium 112	49 In indium 115	50 Sn tin 119	51 Sb antimony 122	52 Te tellurium 128	53 I iodine 127	54 Xe xenon 131
55 Cs caesium 133	56 Ba barium 137	57–71 lanthanoids	72 Hf hafnium 178	73 Ta tantalum 181	74 W tungsten 184	75 Re rhenium 186	76 Os osmium 190	77 Ir iridium 192	78 Pt platinum 195	79 Au gold 197	80 Hg mercury 201	81 Tl thallium 204	82 Pb lead 207	83 Bi bismuth 209	84 Po polonium –	85 At astatine –	86 Rn radon –
87 Fr francium –	88 Ra radium –	89–103 actinoids	104 Rf rutherfordium –	105 Db dubnium –	106 Sg seaborgium –	107 Bh bohrium –	108 Hs hassium –	109 Mt meitnerium –	110 Ds darmstadtium –	111 Rg roentgenium –	112 Cn copernicium –		114 Fl flerovium –		116 Lv livermorium –		

Key: atomic number; atomic symbol; name; relative atomic mass

***Noble gases (group VII) are chemically inert, hence do not form ions.**

Formulae of commonly encountered ions grouped by valency: 1, 2 and 3

	Valency (oxidation state)					
	1		2		3	
Metals	Lithium	(Li^+)	Magnesium	(Mg^{2+})	Aluminium	(Al^{3+})
	Sodium	(Na^+)	Calcium	(Ca^{2+})	Iron(III)	(Fe^{3+})
	Potassium	(K^+)	Copper(II)	(Cu^{2+})		
	Silver	(Ag^+)	Zinc	(Zn^{2+})		
	Copper(I)	(Cu^+)	Iron(II)	(Fe^{2+})		
			Lead	(Pb^{2+})		
			Barium	(Ba^{2+})		
Non-metals	Fluoride	(F^-)	Oxide	(O^{2-})		
	Chloride	(Cl^-)	Sulfide	(S^{2-})		
	Bromide	(Br^-)				
	Hydrogen	(H^+)				
Groups of atoms	Hydroxide	(OH^-)	Carbonate	(CO_3^{2-})	Phosphate	(PO_4^{3-})
	Nitrate	(NO_3^-)	Sulfate	(SO_4^{2-})		
	Ammonium	(NH_4^+)	Dichromate(VI)	($Cr_2O_7^{2-}$)		
	Hydrogencarbonate	(HCO_3^-)				
	Manganate(VII)	(MnO_4^-)				

Describe the formation of ionic bonds between elements from Groups I and VII

Elements of group I form ions of the type M^{1+}.
Elements of group VII form ions of the type X^{1-}.
Ionic bonds form between oppositely charged ions as a result of electrostatic attraction.

$M \rightarrow M^+ + e^-$

$X + e^- \rightarrow X^-$

$M^+ + X^- \rightarrow MX$

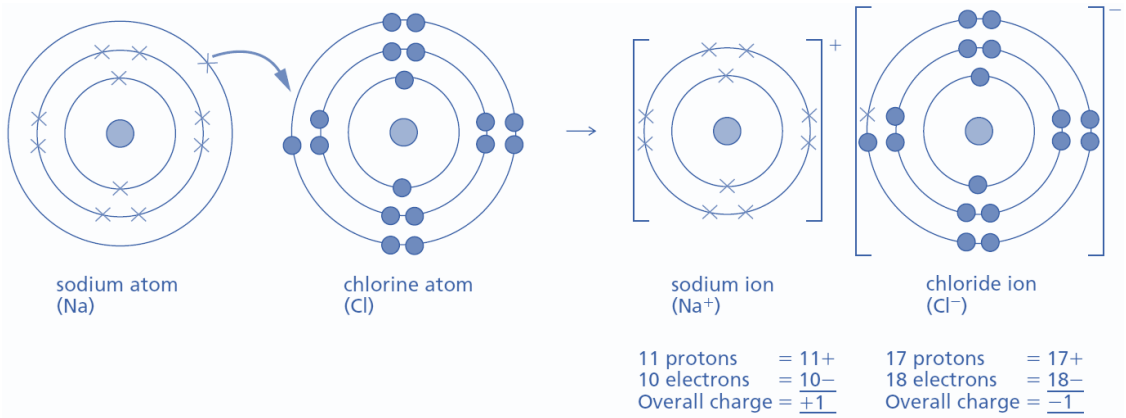
Example:

$Na \rightarrow Na^+ + e^-$

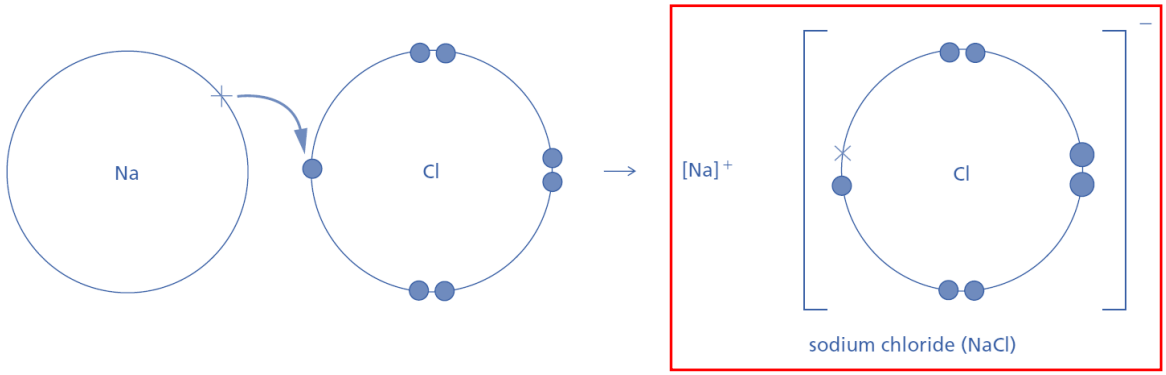
$Cl + e^- \rightarrow Cl^-$

$Na^+ + Cl^- \rightarrow NaCl$

Ionic bonding in Sodium chloride



Simplified diagram showing outer shell electrons only



EXAM TIP:

The diagram above can be used for all Alkali metal halides. The part emphasized in red needs to be drawn. Only the symbols of the elements need to be changed.

3.2.2 Ions and ionic bonds

SUPPLEMENT

Describe the formation of ionic bonds between metallic and non-metallic elements

Metallic elements form ions of the type M^{x+} by losing electrons.
Non-metallic elements form ions of the type N^{y-} by gaining electrons.
Ionic bonds form between oppositely charged ions as a result of electrostatic attraction.

$$M^{x+} + N^{y-} \rightarrow M_yN_x$$

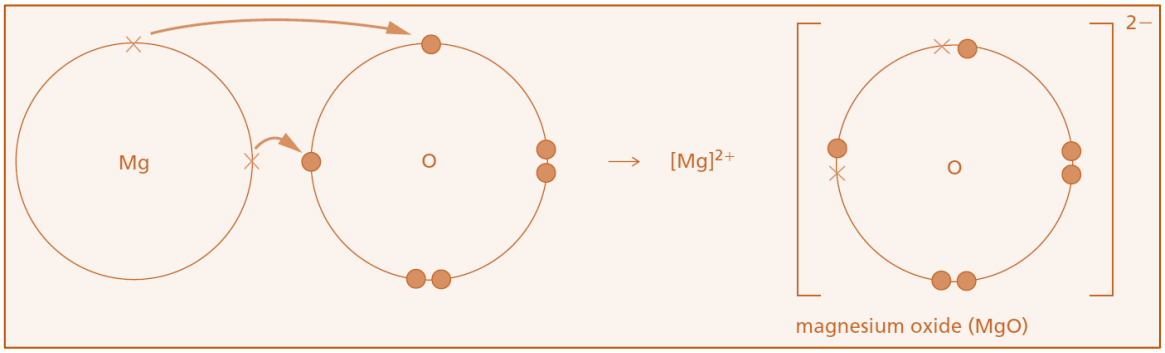
Example:

$Mg \rightarrow Mg^{2+} + 2e^-$

$O + 2e^- \rightarrow O^{2-}$

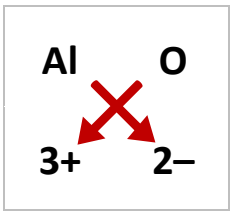
$Mg^{2+} + O^{2-} \rightarrow MgO$

Ionic bonding in Magnesium oxide



$Al \rightarrow Al^{3+} + 3e^-$

$O + 2e^- \rightarrow O^{2-}$



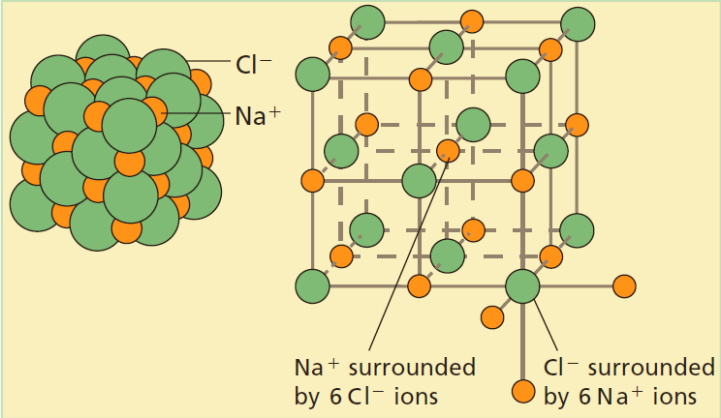
$2Al^{3+} + O^{2-} \rightarrow Al_2O_3$

Describe the lattice structure of ionic compounds as a regular arrangement of alternating positive and negative ions

Lattice: ordered arrangement / regular arrangement

Lattice structure of ionic compounds consists of a regular arrangement of alternating positive and negative ions. **(MUST LEARN!)**
The ions attract each other strongly and are held together in the lattice structure by strong ionic bonds.

Cubic lattice structure of Sodium chloride

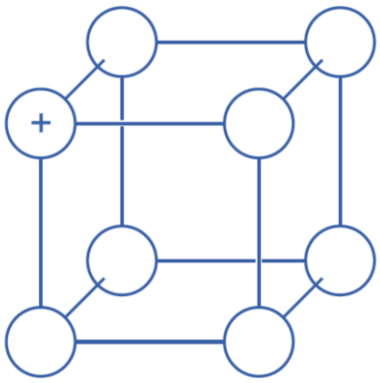


Sodium chloride is a typical ionic compound.

The diagram shows part of a lattice of sodium chloride.

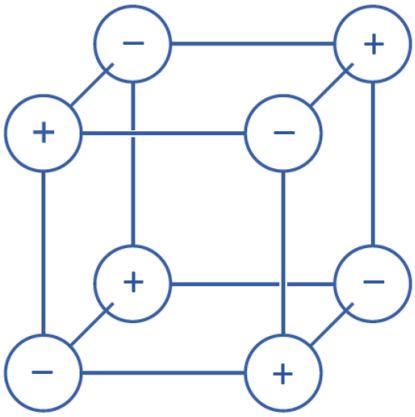
Complete the diagram to show the ions present. Use '+' for Na+ ions and '–' for Cl^- ions.

One ion has been completed for you.



Solution:

Ionic lattice → ordered arrangement of alternating positive and negative ions



3.2.3 Molecules and covalent bonds

CORE

Describe the formation of single covalent bonds in H_2, Cl_2, H_2O, CH_4, NH_3 and HCl as the sharing of pairs of electrons leading to the noble gas configuration

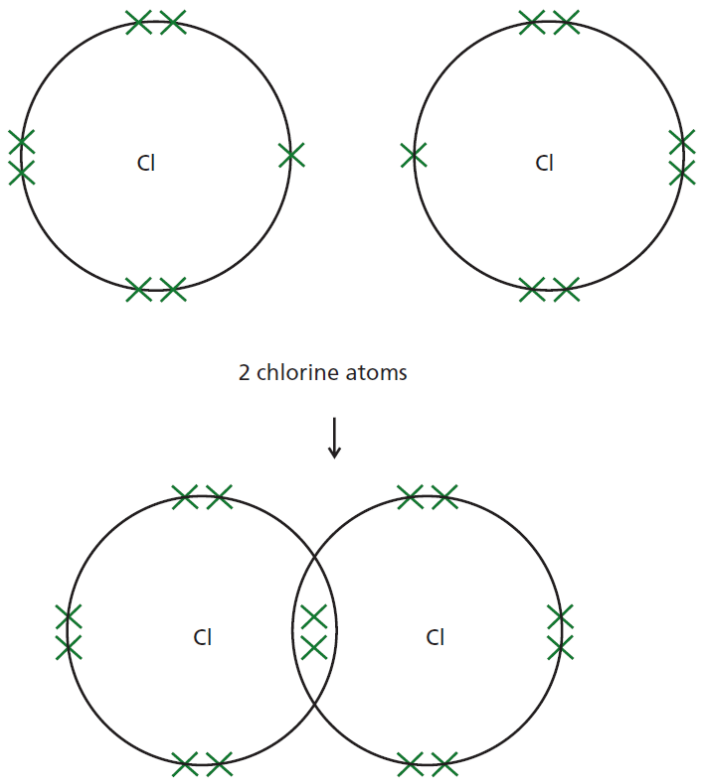
Covalent bonds are formed between atoms of non-metals by sharing of electron pair/s in order to attain a noble gas configuration.

A single covalent bond is formed by sharing of **ONE PAIR** of electrons between two atoms (that may be same of different).
It is generally represented by a single dash —

The number of electrons shared by an atom is (generally) equal to the valency of that atom. **(MUST KNOW!)**

Example:

Chlorine has 7 electrons in its outer shell.
It needs 1 more electron to attain the noble gas configuration (**OCTET RULE**).
The valency of Chlorine is 1.
Two Chlorine atoms therefore share 1 electron each to form a diatomic Chlorine molecule, Cl_2.



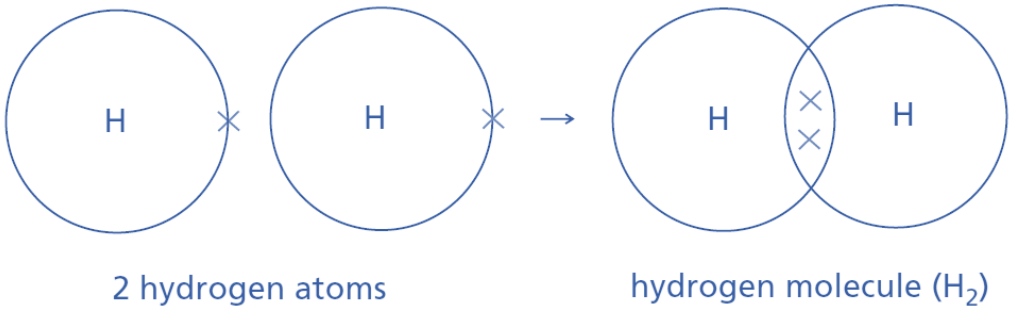
Diatomic molecule: made up of two atoms bonded together by covalent bond/s

Formation of single covalent bonds in H_2, H_2O, CH_4, NH_3 and HC*l*

HYDROGEN MOLECULE, H_2

Valency of Hydrogen = 1

Number of electrons shared by each Hydrogen atom = 1



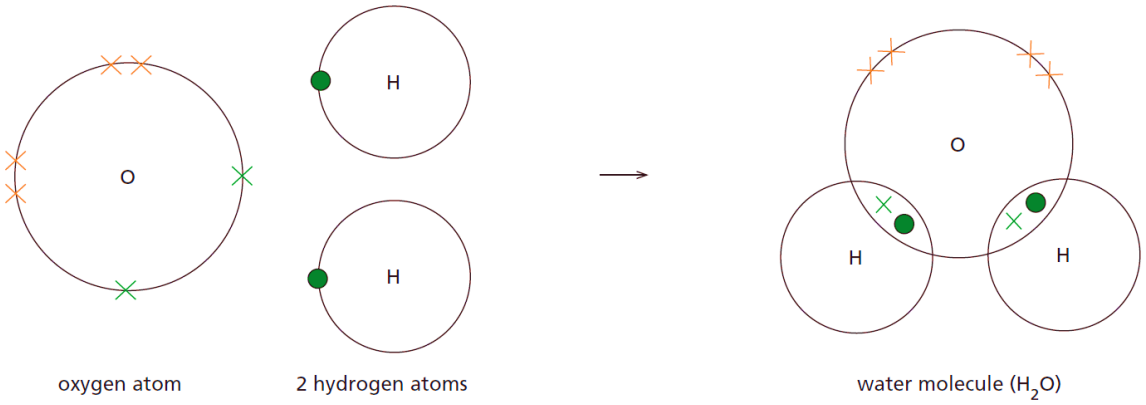
WATER MOLECULE, H_2O

Valency of Hydrogen = 1

Number of electrons shared by each Hydrogen atom = 1

Valency of Oxygen = 2

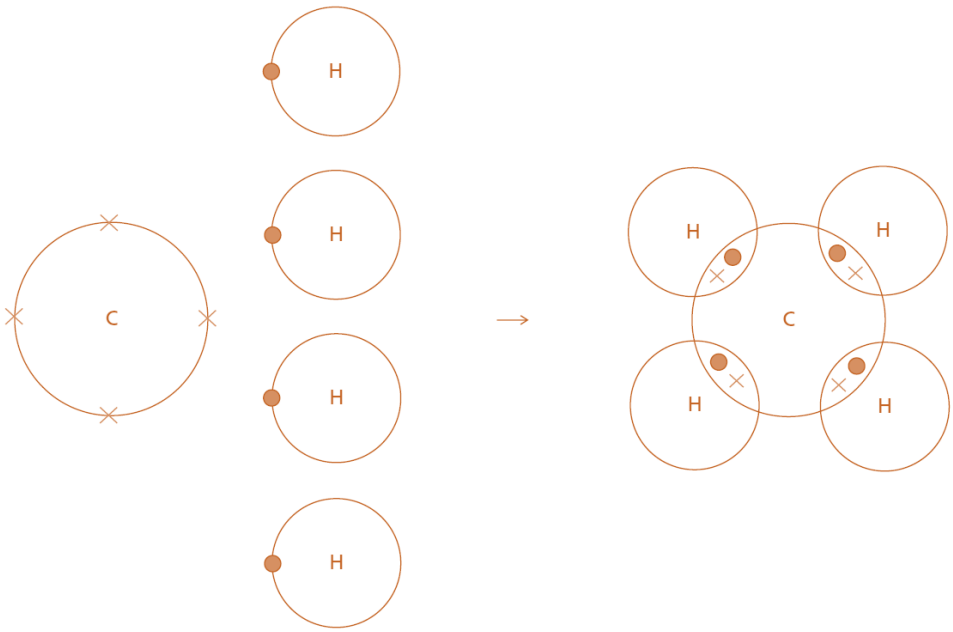
Number of electrons shared by Oxygen atom = 2



METHANE MOLECULE, CH_4

Valency of Hydrogen = 1
Number of electrons shared by each Hydrogen atom = 1

Valency of Carbon = 4
Number of electrons shared by Carbon atom = 4

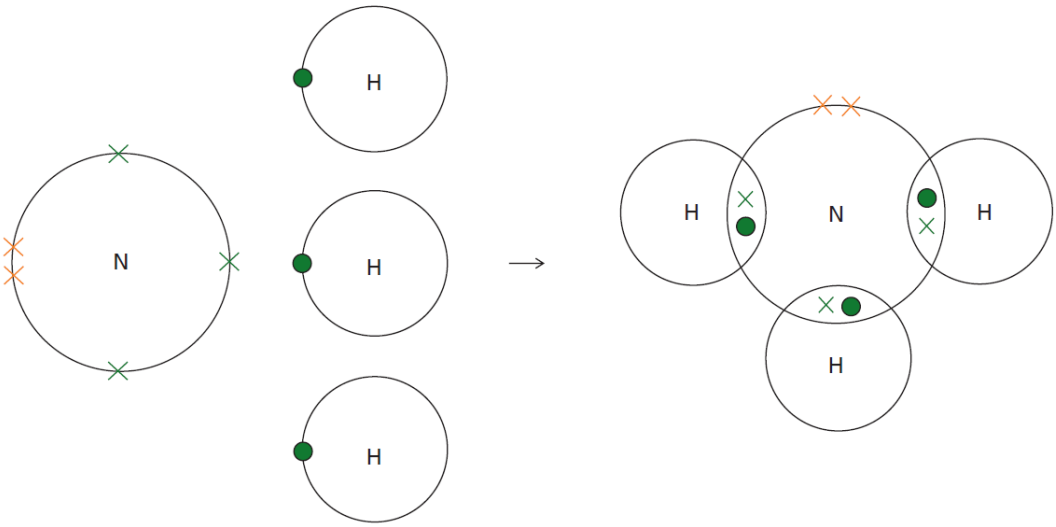


carbon atom 4 hydrogen atoms methane molecule (CH_4)

AMMONIA MOLECULE, NH_3

Valency of Hydrogen = 1
Number of electrons shared by each Hydrogen atom = 1

Valency of Nitrogen = 3
Number of electrons shared by Nitrogen atom = 3



nitrogen atom 3 hydrogen atoms ammonia molecule (NH_3)

Describe the differences in volatility, solubility and electrical conductivity between ionic and covalent compounds

Property	Ionic compounds	Covalent compounds
Volatility (boiling point)	High boiling point, low volatility	Relatively lower boiling points, higher volatility
	Inter-ionic forces of attraction (ionic bonds) are very strong. A large amount of heat energy is required to overcome these forces, hence high melting point / boiling point!	Intermolecular forces in simple covalent compounds are much weaker than ionic bonds and can be overcome easily. Covalent compounds therefore have relatively lower melting points / boiling points. **NOTE:** Compounds that have a giant covalent structure have very high melting and boiling points.
Solubility	Generally water-soluble	Generally soluble in non-aqueous solvents
Electrical conductivity	Conduct electricity in the molten state / aqueous solution but not in the solid state	Do not contain ions / free electrons, hence cannot conduct electricity
	Ions are free to move in the molten / aqueous states but not in the solid state as they are held together by strong ionic bonds	

3.2.3 Molecules and covalent bonds

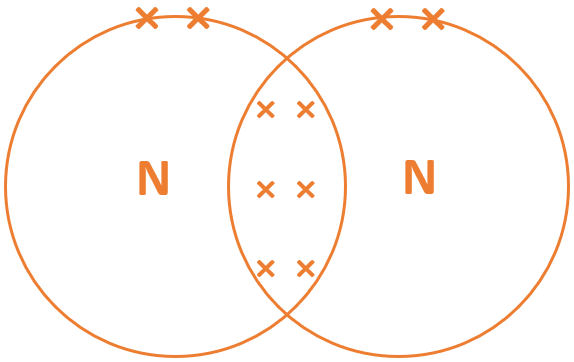
SUPPLEMENT

Describe the electron arrangement in more complex covalent molecules such as N_2, C_2H_4, CH_3OH and CO_2

NITROGEN MOLECULE, N_2

Valency of Nitrogen = 3

Number of electrons shared by each Nitrogen atom = 3



A triple bond (3 pairs of shared electrons) exists between the Nitrogen atoms in Nitrogen molecule, N ≡ N.

ETHENE MOLECULE, C_2H_4

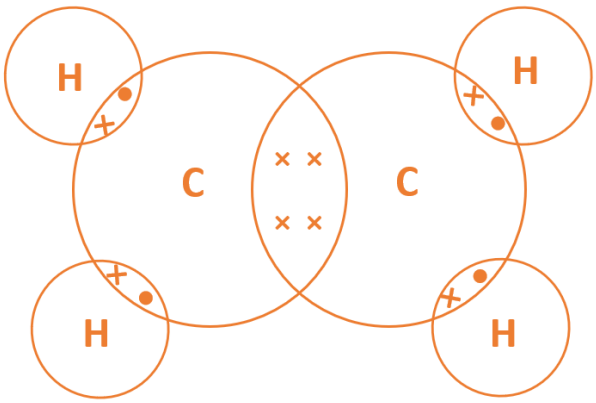
Valency of Carbon = 4

Number of electrons shared by each Carbon atom = 4

Valency of Hydrogen = 1

Number of electrons shared by each Hydrogen atom = 1

C	×
H	•



A double bond (2 pairs of shared electrons) exists between the Carbon atoms in Ethene, C = C.

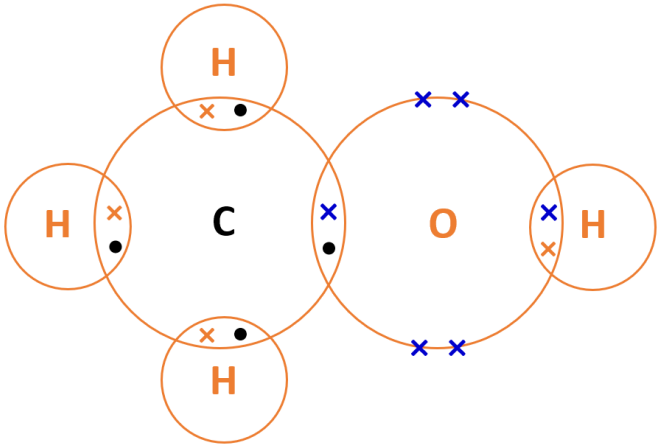
METHANOL MOLECULE

Valency of Carbon = 4
Number of electrons shared by Carbon atom = 4

Valency of Hydrogen = 1
Number of electrons shared by each Hydrogen atom = 1

Valency of Oxygen = 2
Number of electrons shared by Oxygen atom = 2

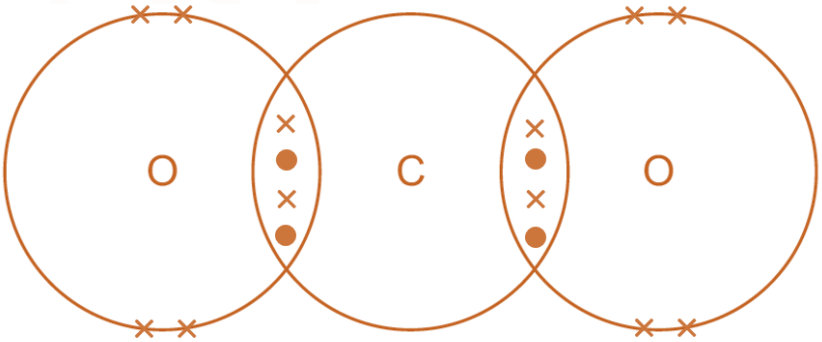
C	•
H	×
O	×



CARBON DIOXIDE MOLECULE, CO_2

Valency of Carbon = 4
Number of electrons shared by Carbon atom = 4

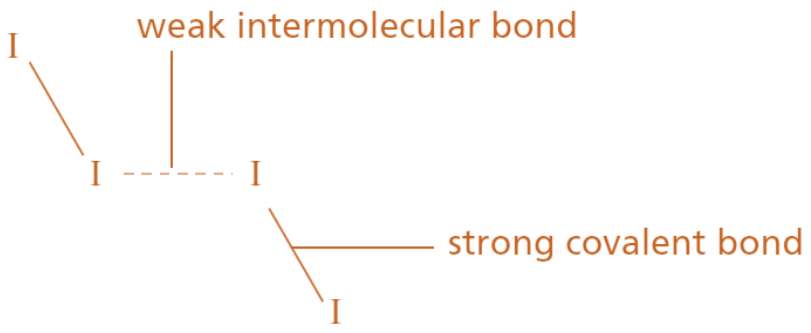
Valency of Oxygen = 2
Number of electrons shared by each Oxygen atom = 2



Double bonds (2 pairs of shared electrons) exists between the Carbon and Oxygen atoms in Carbon dioxide, O = C = O.

Explain the differences in melting point and boiling point of ionic and covalent compounds in terms of attractive forces

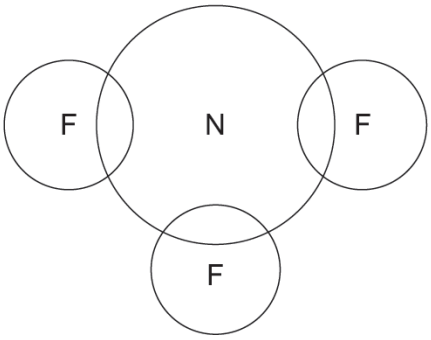
FREQUENTLY ASSESSED CONTENT

Property	Ionic compounds	Covalent compounds
Melting point and boiling point	High melting and boiling points	Relatively lower melting and boiling points
	Inter-ionic forces of attraction (ionic bonds) are very strong. A large amount of heat energy is required to overcome these forces, hence high melting and boiling points	**Intermolecular forces** in simple covalent compounds are much weaker than ionic bonds and can be overcome easily. Covalent compounds therefore have relatively lower melting points / boiling points. **NOTE:** Compounds that have a giant covalent structure have very high melting and boiling points.
		

QUESTION 1

Complete the dot-and-cross diagram to show the electron arrangement in a molecule of NF_3.

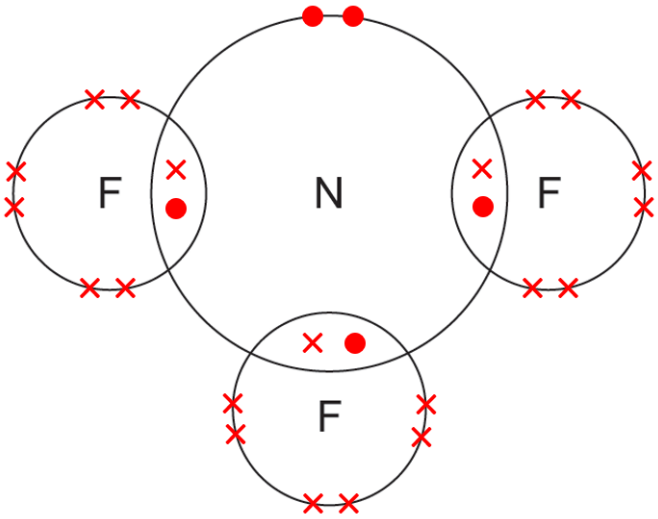
Show outer shell electrons only.



Solution:

	Outer shell electrons
N	5
F	7

***Number of outer shell electrons = Group number**



QUESTION 2

Nitrosyl chloride, NOC*l*, is a gas at room temperature. It has the structure shown.

O=N–Cl

Nitrosyl chloride has a boiling point of –6 °C.

Explain why nitrosyl chloride has a low boiling point.

SOLUTION:

Nitrosyl chloride has a simple molecular structure with weak forces of attraction between the molecules.
Less energy is required to break the weak intermolecular forces, hence low boiling point.

QUESTION-BANK

QUESTIONBANK

ASSESSMENT

ASSESSMENT 3.2:
BONDING: Covalent & Ionic