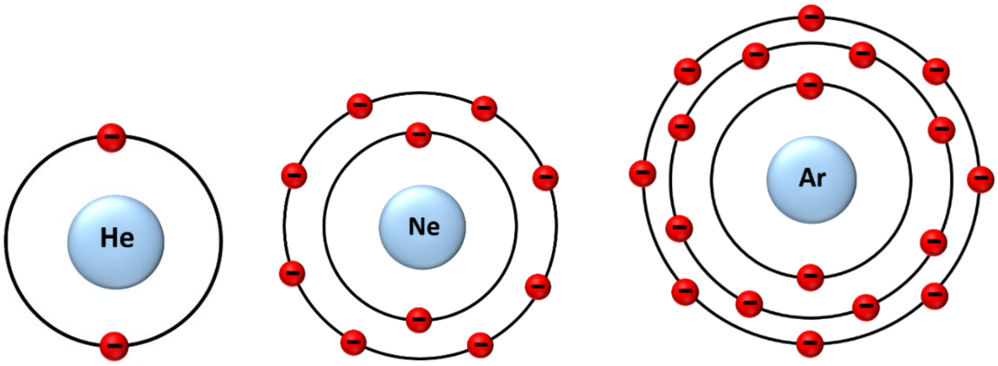
Bonding SL

IB CHEMISTRY SL

25	16		6	2	25
Mn	S	J	C	He	Mn
Manganese	Sulfur		Carbon	Helium	Manganese
54.938045	32.065		12.0107	4.002602	54.938045

The octet rule

- The octet rule states that atoms 'want' to achieve the electron configurations of noble gases.
- The electron configurations of noble gases He, Ne and Ar are shown below.



- Noble gases are stable because they have full outer main energy levels.
- Atoms can achieve the electron configuration of a noble gas by either sharing electrons (covalent bonding) or by losing or gaining electrons (ionic bonding).

Exceptions to the octet rule

- Hydrogen and helium are stable with 2 electrons in their outer energy level.
- Beryllium is stable with 4 electrons in its outer energy level.
- Boron is stable with 6 electrons in its outer energy level.
- Elements in period 3 onwards can hold more than 8 electrons in their outer shells (expanded octets).

Electronegativity and bonding

- Electronegativity is a measure of the attraction of an atom for a bonding pair of electrons.
- It increases across a period due to increasing nuclear charge and decreases down a group as atomic radius increases.

Difference in electronegativity	Type of bond	Example
0	non-polar (pure) covalent bond	Cl-Cl
0.1–0.4	non-polar (weakly polar) covalent bond	C-H
0.5–1.7	polar covalent bond	C-F
≥1.8	ionic	NaCl

4.1 Ionic bonding

Understandings:

- Positive ions (cations) form by metals losing valence electrons.
- Negative ions (anions) form by non-metals gaining electrons.
- The number of electrons lost or gained is determined by the electron configuration of the atom.
- The ionic bond is due to the electrostatic attraction between oppositely charged ions.
- Under normal condition, ionic compounds are usually solid with lattice structures.

Applications and skills:

- Deduction of the formula and name of an ionic compound from its component ions, including polyatomic ions.
- Explanation of the physical properties of ionic compounds (volatility, electrical conductivity and solubility) in terms of their structure.

Guidance:

- Students should be familiar with the following polyatomic ions: NH_4^+ OH^- HCO_3^- CO_3^{2-} SO_4^{2-} PO_4^{3-} NO_3^-

Syllabus checklist

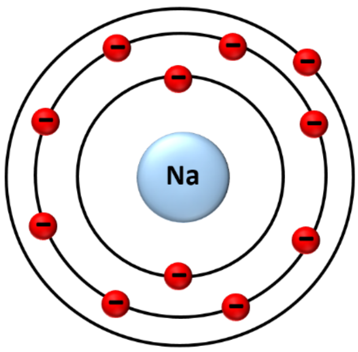
Objective	I am confident with this	I need to review this	I need help with this
Outline the octet rule			
State the exceptions to the octet rule			
Describe the formation of positive and negative ions by either the loss of gain of electrons			
Describe the formation of an ionic bond			
Describe the structure of an ionic compound			
Write formulae for ionic compounds			
Explain the properties of ionic compounds – solubility, electrical conductivity, melting and boiling point			

Ion formation

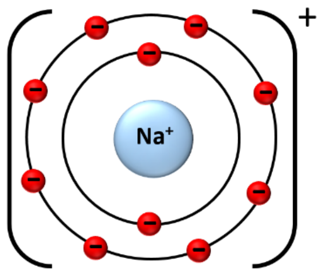
- Neutral atoms have equal numbers of protons and electrons.
- Positive ions (cations) are formed when atoms lose electrons.
- Negative ions (anions) are formed when atoms gain electrons.

Positive ions

- The sodium atom loses its one valence electron to form a positive ion with a 1+ charge, Na^+.
- It now has one less occupied energy level and the same electron configuration as the noble gas neon, Ne.



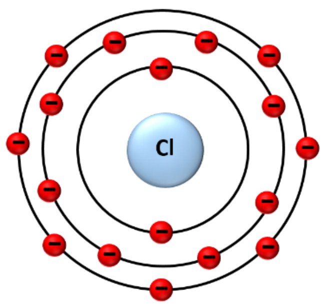
$1s^2\ 2s^2\ 2p^6\ 3s^1$



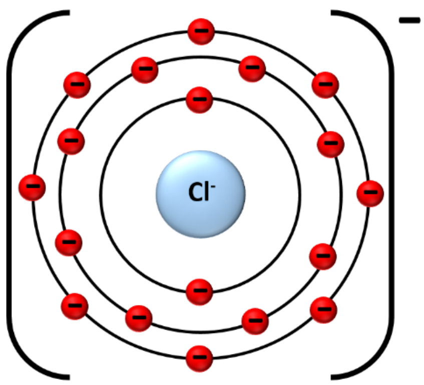
$1s^2\ 2s^2\ 2p^6$

Negative ions

- The chlorine atom gains one electron to form a negative ion with a 1- charge, Cl^-.
- It now has a full outer shell of electrons and the same electron configuration as the noble gas argon, Ar.



$1s^2\ 2^2\ 2p^6\ 3s^2\ 3p^5$



$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Exercises:

1. Explain why the sodium atom has a larger radius than the sodium ion.

2. Explain why the chloride ion has a larger radius than the chlorine atom.

3. Explain why metals tend to lose electrons and form positive ions and non-metals tend to gain electrons and form negative ions.

4. Determine the ions formed by the following atoms:

a) Cs

b) Be

c) Cl

d) N

e) S

f) I

g) Ba

h) Al

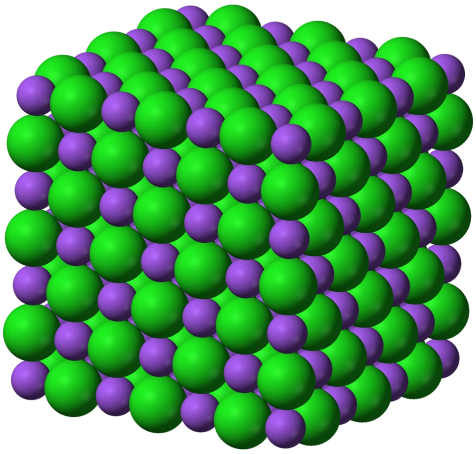
i) O

j) K

Ionic bonding

- An ionic bond is the electrostatic attraction between oppositely charged ions.
- An ionic bond forms between two elements with a difference in electronegativity of equal to or greater than 1.8 units.
- Ionic bonds occur between metal and non-metal elements (those elements on the left and right of the periodic table).

Structure of ionic compounds



Ionic compounds have a lattice structure.
The lattice is held together by the electrostatic attraction between the oppositely charged ions.
Each Na^+ ion is surrounded by six Cl^- ions and each Cl^- ion is surrounded by six Na^+ ions.
Ionic compounds are solids under standard conditions.

Exercises

1. Define an ionic bond.

2. An element on the far left and an element on the far right of the periodic table form a chemical bond. Determine the type of bond they would form and explain your reasoning.

3. Describe the lattice structure of NaCl.

Writing ionic formulae

List of common ions:

Positive ions		Negative ions	
Name	**Symbol**	**Name**	**Symbol**
Hydrogen		Fluoride	
Sodium		Chloride	
Silver		Bromide	
Potassium		Iodide	
Lithium		Hydrogencarbonate	
Ammonium		Hydroxide	
Barium		Nitrate	
Calcium		Oxide	
Copper(II)		Sulfate	
Magnesium		Carbonate	
Zinc		Phosphate	
Mercury(I)		Nitride	
Lead		Sulfide	
Iron(II)		Phosphide	
Iron(III)		Nitrite	
Aluminium		Sulfite	

Exercise: Write the formulae for the following ionic compounds:

1) Potassium bromide
2) Calcium fluoride
3) Beryllium sulfide
4) Strontium Iodide
5) Magnesium nitride
6) Aluminium oxide
7) Sodium carbonate
8) Copper(II) phosphide
9) Zinc phosphate
10) Ammonium nitrate

11) Ammonium sulfate

12) Iron(III) sulfite

13) Copper(II) nitrite

14) Potassium hydrogencarbonate

15) Aluminium sulfate

16) Mercury(I) nitride

17) Iron(II) nitrite

18) Barium nitrate

19) Iron(II) phosphide

20) Calcium hydrogencarbonate

Step 1 Write the ions side by side:

Mg^{2+} Cl^{-}

Step 2 Draw arrows that cross each other:

Mg^{2+} Cl^{-}

Step 3 Write the **charges** at the arrow ends:

Mg^{2+} Cl^{-}

1 2

Step 4 Write the formula as follows:

a) write the positive ion without its charge:

Mg

b) write the number as a subscript unless it is 1:

Mg

c) write the negative ion without its charge:

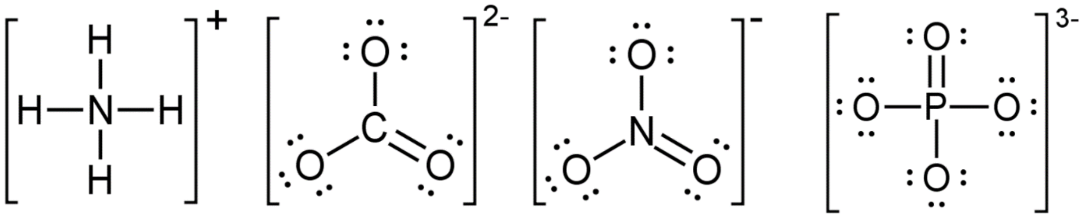
MgCl

d) write the number as a subscript unless it is 1:

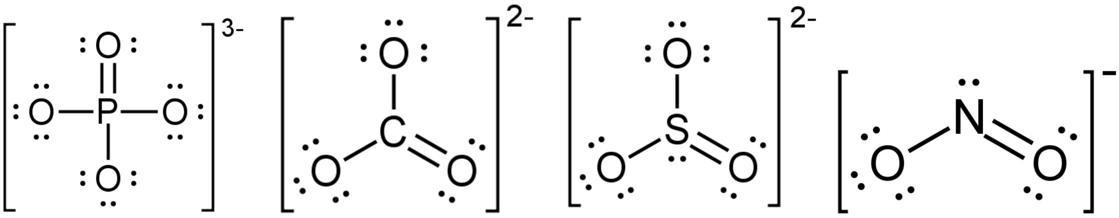
$MgCl_2$ … all done!

Polyatomic ions

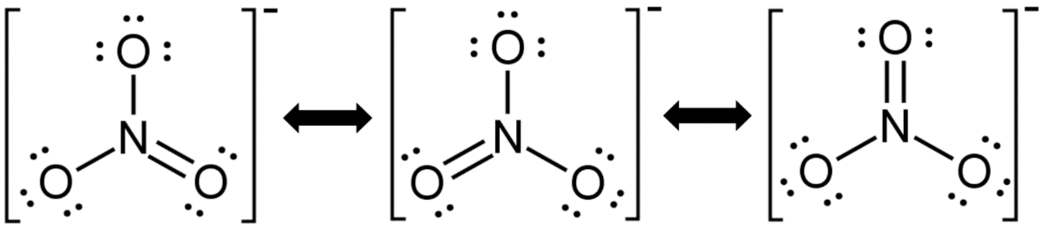
- Polyatomic ions (molecular ions) are ions that consist of two or more atoms bonded together with covalent bonds.



- The atoms in a polyatomic ion are bonded with covalent bonds.
- The bonding between the ions in a compound that contains a polyatomic ion is ionic.
- The geometry of a polyatomic ion depends on the number of electron domains around the central atom.



- Polyatomic ions with more than one position for a multiple bond exist as resonance structures.



- The N-O bonds are equal length and equal strength – intermediate in length and strength between a single and a double bond.

Exercise

State the two types of bonding in an ionic compound containing a polyatomic ion.

Structure and properties of ionic compounds

Electrical conductivity

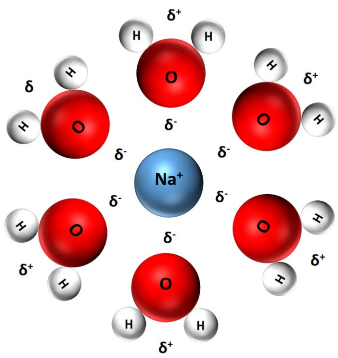
- Ionic compounds do not conduct electricity when solid, because the ions are held in fixed positions by the electrostatic attractions between the ions.
- They only conduct electricity when molten (melted) or dissolved in water (aqueous).
- When molten or dissolved, the ions are free to move and conduct the electric current.

Melting and boiling point

- Ionic compounds have high melting and boiling points due to the strong electrostatic attractions between the oppositely charged ions (NaCl has a melting point of 800 ^{o}C).
- The greater the charge and smaller the ionic radius of the ions, the stronger the electrostatic attraction and the higher the melting point.

Solubility

- Ionic compounds are soluble in polar solvents.
- The ions are separated from the lattice structure by the polar water molecules.
- The ions become surrounded by water molecules (hydration) as shown below.



Exercises:

1. Explain the conductivity of ionic compounds when molten or dissolved and when solid.

2. Explain the high melting point of ionic compounds.

3. Explain why NaF has a higher melting point than KF.

4.2 Covalent bonding

Understandings:

- A covalent bond is formed by the electrostatic attraction between a shared pair of electrons and the positively charged nuclei.
- Single, double and triple covalent bonds involve one, two and three shared pairs of electrons respectively.
- Bond length decreases and bond strength increases as the number of shared electrons increases.
- Bond polarity results from the difference in electronegativities of the bonded atoms.

Applications and skills:

- Deduction of the polar nature of a covalent bond from electronegativity values.

Guidance:

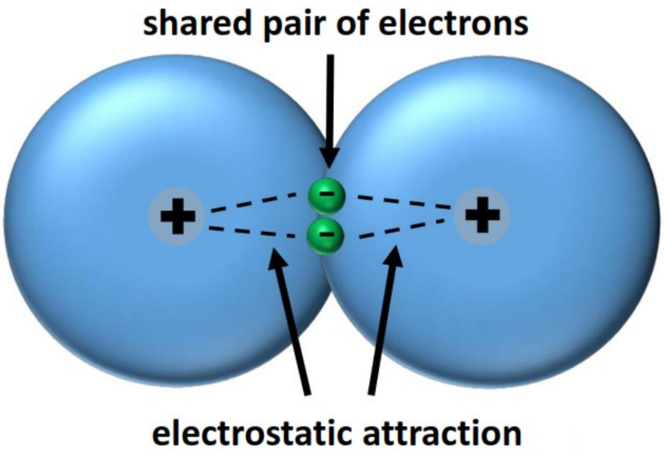
- Bond polarity can be shown either with partial charges, dipoles or vectors.
- Electronegativity values are given in the data booklet in section 8.

Syllabus checklist

Objective	I am confident with this	I need to review this	I need help with this
Describe the formation of a covalent bond			
State the explain the relationship between bond length and bond strength			
Deduce the polarity of a covalent bond based on the difference in electronegativity between the bonding atoms			
Deduce the polarity of a molecules based on its geometry and the presence of polar bonds			
Describe the formation of a coordinate covalent bond			
State examples of molecules that have coordinate covalent bonds			

Covalent bonding

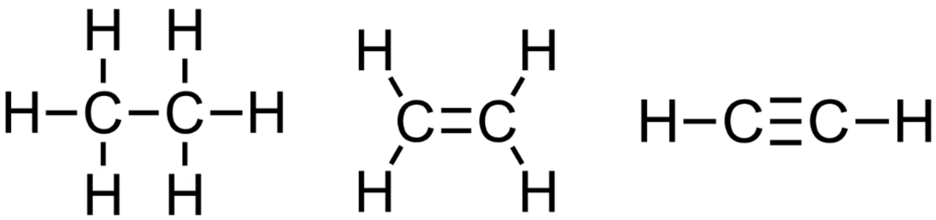
- A covalent bond is the electrostatic attraction between positive nuclei and a shared pair of bonding electrons.
- Covalent bonding occurs between non-metal elements only.
- The electronegativity difference between atoms is between 0.0 and 1.7 units.



- Both nuclei are attracted to the shared pair of bonding electrons.
- The electrostatic attraction between the nuclei and shared pair of bonding electrons forms the covalent bond.

Single, double and triple bonds

- Single bonds are longer and weaker, triple bonds are shorter and stronger.



- The carbon to carbon bonds in C_2H_6 are longer and weaker than those in C_2H_2

Bond	Number of shared electrons	C to C bond strength (kJ mol^{-1})	C to C bond length (10^{-12}m)
Single	2	347	153
Double	4	614	134
Triple	6	839	120

Bond order

- Bond order is the number of bonds between a pair of atoms.
- Single bonds have a bond order of 1, double bonds have a bond order of 2 and triple bonds have a bond order of 3.
- The higher the bond order, the stronger (and shorter) the bond.
- Polyatomic ions such as CO_3^{2-} and NO_3^- and molecules such as benzene, C_6H_6, can have fractional bond orders.
- Bond order can be calculated by dividing the sum of the individual bond orders by the number of bonding groups in the molecule or ion.

Example: calculate the bond order of the bonds in the nitrate ion, NO_3^-.

Polar covalent bonds

- Polar covalent bonds occur between atoms that have a difference in electronegativity.

Difference in electronegativity	Polar or non-polar covalent bond	Example
0	non-polar (pure) covalent bond	Cl-Cl
0.1– 0.4	non-polar (weakly polar) covalent bond	C-H
0.5 –1.7	polar covalent bond	C-F

- Bond polarity can be shown by a vector arrow or by δ^+ and δ^- as shown below.

Exercises:

1. Define a covalent bond.

2. Where on the periodic table would you find elements that bond by covalent bonding?

3. Arrange the following in order of increasing carbon to carbon bond strength (weakest first).

$$C_2H_4, C_2H_6, C_2H_2$$

4. Arrange the following in order of decreasing carbon to carbon bond length (longest first),

$$C_2H_4, C_2H_6, C_2H_2$$

5. State the relationship between the number of electrons in a covalent bond and the length and strength of the bond.

6. Draw a diagram showing the bonding polarity in HF (hydrogen fluoride).

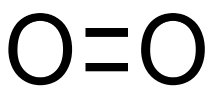
7. For each bond, find the difference in electronegativity and classify as either pure covalent, non-polar covalent or polar covalent. Draw bond dipoles or partial charges on each atom.

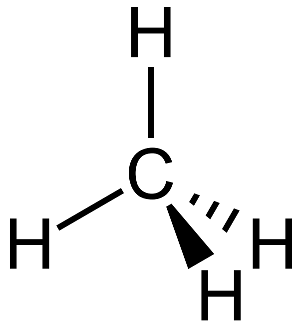
Bond	**Pure covalent, non-polar covalent or polar covalent**	**Bond dipoles**
Br–Br		
C–Cl		
C–I		
C–O		
N–F		
H–F		

Polar and non-polar molecules

The polarity of a molecule depends on two factors:

- The presence of polar bonds within the molecule.
- The shape (geometry) of the molecule.

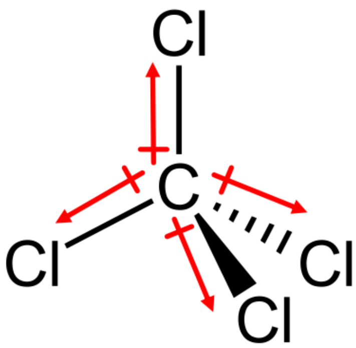




- O_2, N_2 and CH_4 are non-polar molecules.
- O_2 and N_2 are diatomic; both atoms have the same electronegativity.
- CH_4 has non-polar bonds (C-H).

Non-polar molecule with polar bonds

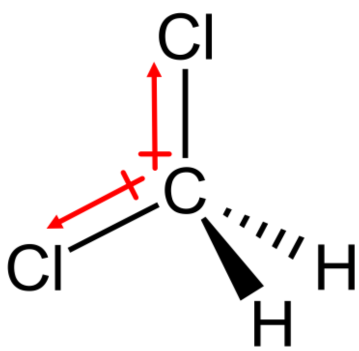
- CCl_4 is a non-polar molecule that has polar bonds.



- The bond polarities cancel out, therefore, the molecule has no net-dipole moment.

Polar molecules

- CH_2Cl_2 is a polar molecule.



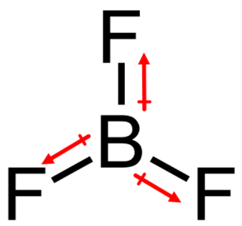
- The bond polarities do not cancel out, therefore, it has a net-dipole moment.

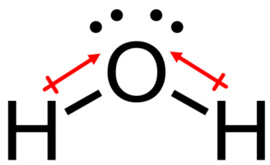
Exercises:

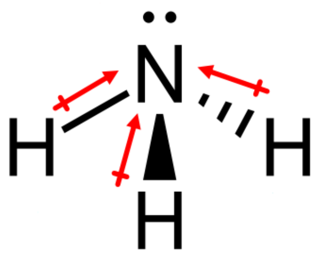
1. Outline the two factors that determine the polarity of a molecule.

2. Explain how a molecule can have polar bonds but overall have no net-dipole moment (be a non-polar molecule).

3. Determine if the molecules below are polar or non-polar, giving a reason.

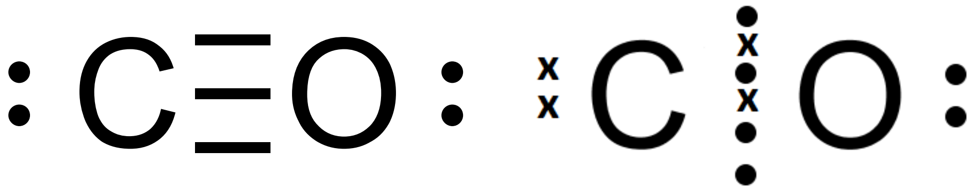






Coordinate covalent bonds

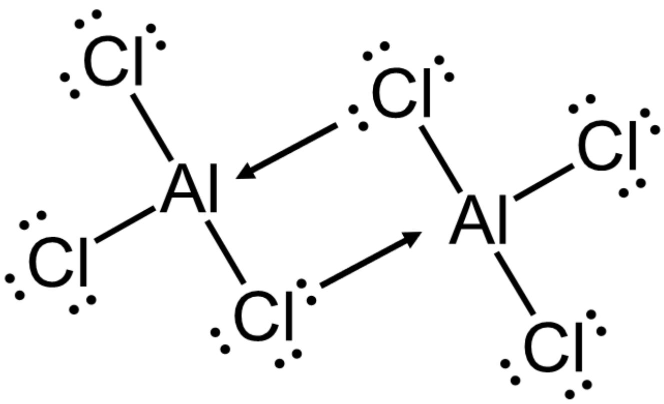
- In a coordinate covalent bond, one atom contributes both the bonding electrons to the bond.



- Draw the Lewis structure of CO, H_3O^+ and NH_4^+ in the boxes below.

CO	H_3O^+	NH_4^+

- Al_2Cl_6 – the dimer formed between two molecules of $AlCl_3$.



4.3 Covalent structures

Understandings:

- Lewis (electron dot) structures show all the valence electrons in a covalently bonded species.
- The "octet rule" refers to the tendency of atoms to gain a valence shell with a total of 8 electrons.
- Some atoms, like Be and B, might form stable compounds with incomplete octets of electrons.
- Resonance structures occur when there is more than one possible position for a double bond in a molecule.
- Shapes of species are determined by the repulsion of electron pairs according to VSEPR theory.
- Carbon and silicon form giant covalent/network covalent structures.

Applications and skills:

- Deduction of Lewis (electron dot) structure of molecules and ions showing all valence electrons for up to four electron pairs on each atom.
- The use of VSEPR theory to predict the electron domain geometry and the molecular geometry for species with two, three and four electron domains.
- Prediction of bond angles from molecular geometry and presence of non-bonding pairs of electrons.
- Prediction of molecular polarity from bond polarity and molecular geometry.
- Deduction of resonance structures, examples include but are not limited to CO_3^{2-}, NO_3^-
- Explanation of the properties of giant covalent compounds in terms of their structures

Guidance:

- The term "electron domain" should be used in place of "negative charge centre".
- Electron pairs in a Lewis (electron dot) structure can be shown as dots, crosses, a dash or any combination.
- Allotropes of carbon (diamond, graphite, graphene, C_{60} buckminsterfullerene) and SiO_2 should be covered.
- Coordinate covalent bonds should be covered.

Syllabus checklist

Objective	I am confident with this	I need to review this	I need help with this
Deduce Lewis structures for molecules and ions with up to four electron pairs on each atom			
Predict the electron domain geometry and molecular geometry of molecules and ions using VSEPR theory			
State and explain the properties of molecular covalent compounds			
State and explain the properties of giant covalent compounds			
Outline the concept of resonance structures			
State and explain the properties of the allotropes of carbon – diamond, graphite, C_{60} and graphene			

Lewis structures

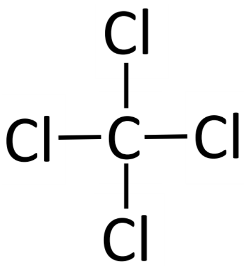
- Lewis structures show all the valence electrons in a molecule; the bonding electrons and the lone pairs of electrons (non-bonding electrons).

How to determine a Lewis structure

1. Calculate number of valence electrons in the molecule.
2. Calculate the number of electrons each atom needs to complete its octet.
3. Subtract 1 from 2 – this will give you the number of bonding electrons in the molecule.
4. Draw the skeletal structure of the molecule with the least electronegative atom at the center.
5. Complete the octets of the atoms in the molecule.

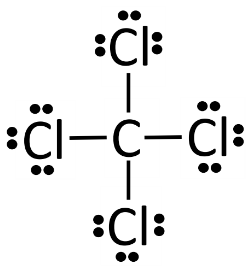
Example – CCl_4 (carbon tetrachloride)

1. Total number of valence electrons = 4 + (4 × 7) = 32
2. Number of valence electrons needed for each atom to complete its octet (5 × 8) = 40
3. 40 - 32 = 8 bonding electrons
4. Skeletal formula:



One single bond = 2 electrons (total 8 bonding electrons).

5. Complete the octets of all the atoms in the molecule.



Do a final check that all electrons are accounted for.
It should match with the number in part 1 (32).

Exercises:

1) Draw Lewis structures for the following molecules:

a) Cl_2

b) HCl

c) CH_3Cl

d) CF_4

e) PF_3

f) BF_3

g) $BeCl_2$

h) N_2H_4

i) H_2O

j) CH_3OH

k) CO

l) C_2H_4

j) HCN

2) Draw Lewis structures for the following ions:

a) Cl^-

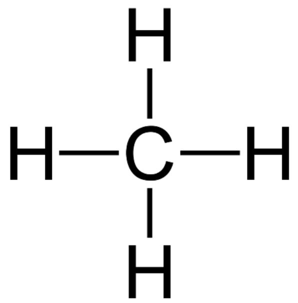
b) NH_4^+

c) OH^-

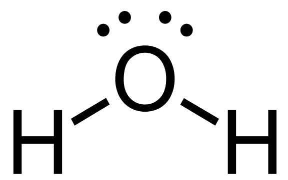
3) Which molecules contain an incomplete octet of electrons?

VSEPR theory

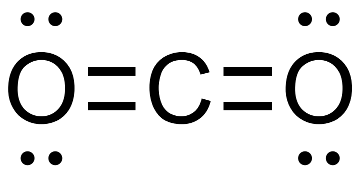
- Valence shell electron pair repulsion theory is used to predict the geometry (shape) of molecules.
- Electron pairs (bonds or lone pairs) repel each other and spread apart as far as possible.
- The term **electron domain** is used to refer to bonds or lone pairs of electrons (non-bonding electrons) around an atom in a molecule.
- Single bonds, double bonds, triple bonds and lone pairs of electrons (non-bonding electrons) count as one electron domain.



How many electron domains are there around the carbon atom?



How many electron domains are there around the oxygen atom?



How many electron domains are there around the carbon atom?

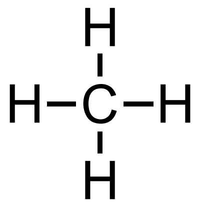
The order of repulsion between electron domains is as follows:

lone pair – lone pair > lone pair – bonding domain > bonding domain – bonding domain

most repulsion **least repulsion**

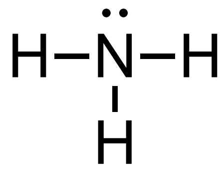
Molecules with four electron domains

- **Four bonding domains, zero non-bonding domains**



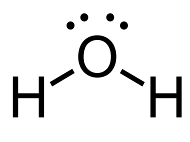
electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
4	4	0	tetrahedral	tetrahedral	109.5°

- **Three bonding domains, one non-bonding domain**



electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
4	3	1	tetrahedral	trigonal pyramidal	107.8°

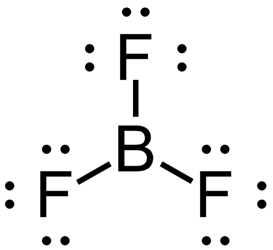
- **Two bonding domains, two non-bonding domains**



electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
4	2	2	tetrahedral	bent	104.5°

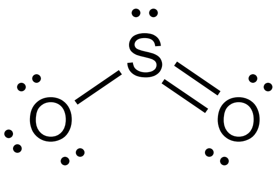
Molecules with three electron domains

- **Three bonding domains, zero non-bonding domains**



electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
3	3	0	trigonal planar	trigonal planar	120°

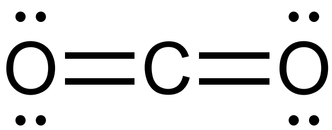
- **Two bonding domains, one non-bonding domain**



electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
3	2	1	trigonal planar	bent	<120°

Molecules with two electron domains

- **Two bonding domains, zero non-bonding domains**



electron domains	bonding domains	non-bonding domains	electron domain geometry	molecular geometry	bond angle
2	2	0	linear	linear	180°

Exercises:

Determine the shape and bond angles in the following molecules:

a) HCN

b) PH_3

c) CH_2Cl_2

d) Cl_2O

e) C_2H_2

f) BF_3

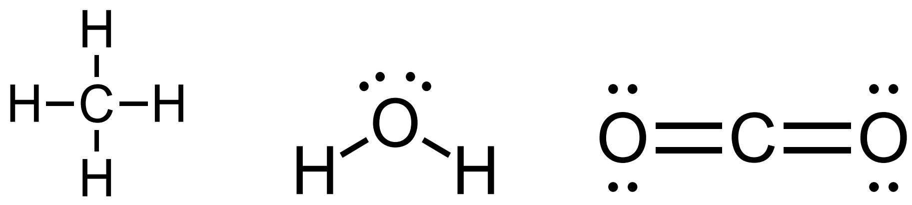
g) BeH_2

h) H_2S

Structure and properties of covalent compounds

- Covalent compounds have two types of structure; simple molecular and giant covalent (network covalent).
- Each type of structure has different physical properties, such as melting and boiling point, electrical conductivity, solubility and hardness.

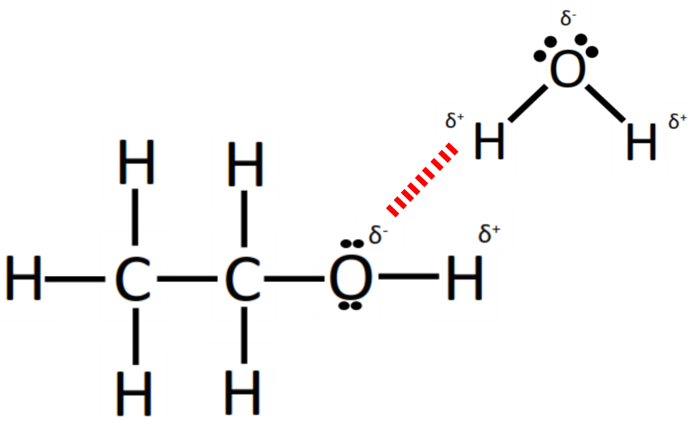
Simple molecular substances



- Simple molecular compounds are mostly liquids and gases due to the weak intermolecular forces between molecules.

Solubility

- Non-polar molecules are soluble in non-polar solvents.
- Polar molecules are soluble in polar solvents (C_2H_5OH is soluble in H_2O)
- The phrase '**like dissolves like'** is useful to remember the solubility.
- Hydrogen bonding between a molecule of C_2H_5OH (ethanol) and a molecule of H_2O is shown below.



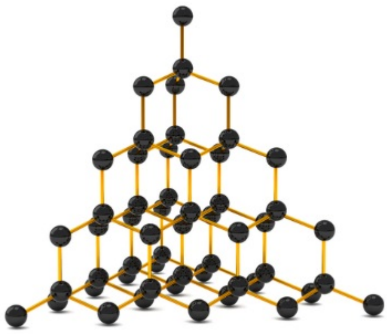
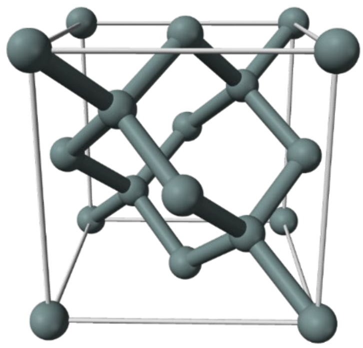
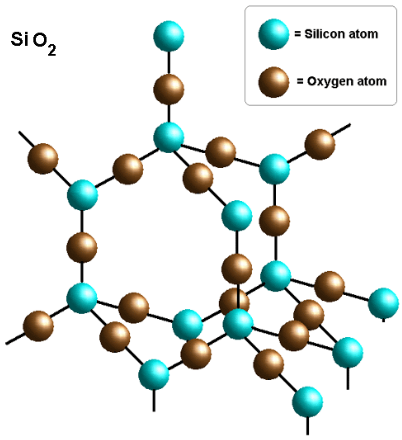
- Ethanol is soluble in water as it is able to form hydrogen bonds between itself and water molecules.

Electrical conductivity

- Simple molecular substances are poor conductors of electricity as they have no free moving charged particles (neither ions nor delocalised electrons).

Giant covalent structures

- Giant covalent structures do not form discrete molecules.
- Examples include diamond, Si and SiO_2

- Giant covalent structures do not conduct electricity (no free moving electrons).
- They are insoluble in polar and non-polar solvents.
- They have high melting and boiling points due to the strong covalent bonds between atoms.
- They are also very hard substances.

Exercises:

1. Classify the following covalent compounds as simple molecular or giant covalent based on their properties:

Property	Simple molecular	Giant covalent
High melting and boiling point		
Insoluble in polar and non-polar solvents		
Liquids and gases under standard conditions		
Do not form discrete molecules		
Poor electrical conductors		
Soluble in polar or non-polar solvents		
Form discrete molecules		
Very hard substances		

2. Explain the following properties of the covalent compounds.

a. Pentane (C_5H_{12}) is insoluble in water but soluble in hexane (C_6H_{14}).

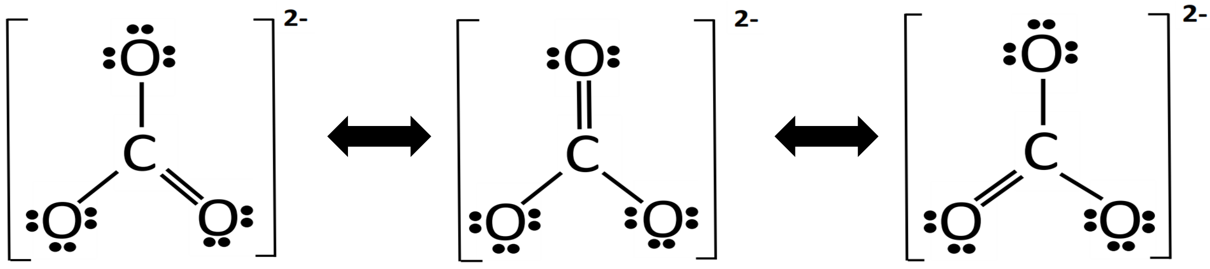
b. Diamond does not conduct electricity.

c. CO_2 is a gas under standard conditions.

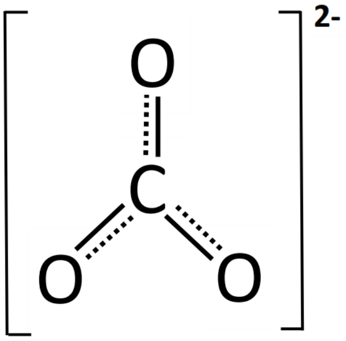
d. SiO_2 has a melting point of 1600^oC

Resonance structures

- Resonance structures occur when there is more than one position for a double (or triple) bond in a molecule.
- The three resonance structures for the carbonate ion (CO_3^{2-}) are shown below.



- The actual structure is a resonance hybrid structure.



The bond lengths and bond strengths are identical - intermediate between a single and a double bond.

Exercise: Draw the resonance structures for the following molecules and ions:

a) Carbonate ion – CO_3^{2-}

b) Nitrate ion - NO_3^-

c) Ozone – O_3

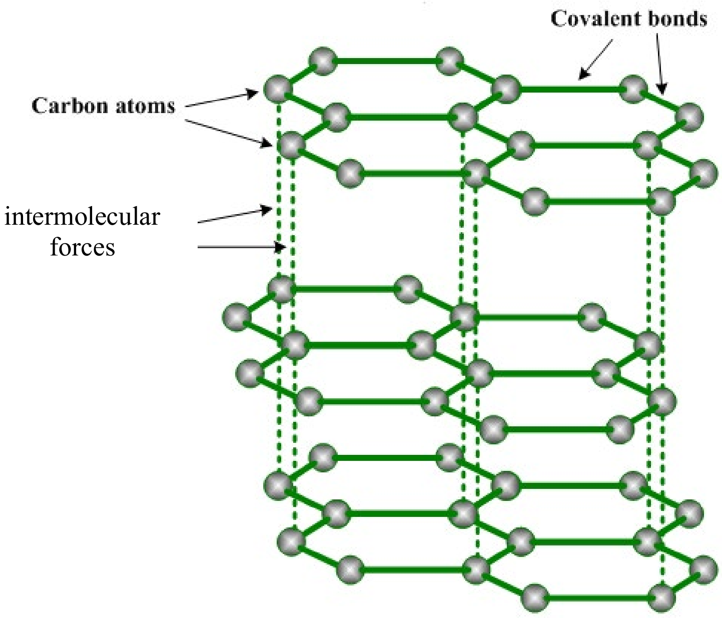
d) Ethanoate ion – CH_3COO^-

e) Benzene – C_6H_6

Allotropes of carbon

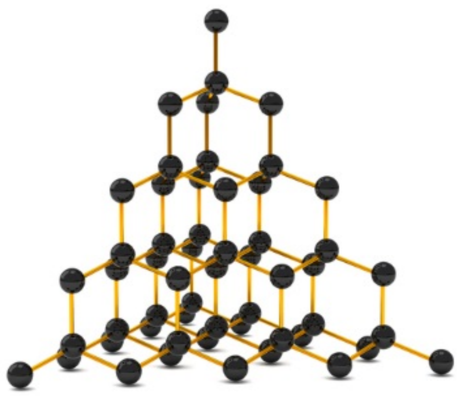
- Allotropes are different forms of the same element in the same physical state.
- Carbon has 4 allotropes – graphite, diamond, Fullerene C_{60} and graphene.
- Different bonding within the structures gives the allotropes different properties.

Graphite



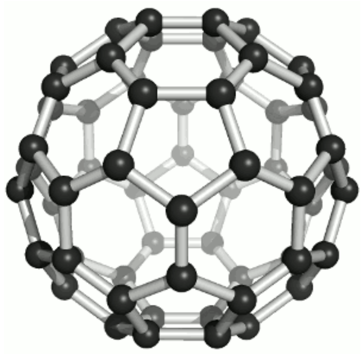
- Graphite has a layered structure.
- The layers are held together by weak intermolecular forces.
- The layers can slide over one another (because of weak intermolecular forces).
- Each carbon atom is bonded to 3 other carbon atoms.
- The bond angle between carbon atoms is 120^o, trigonal planar.
- Graphite is a good conductor of electricity because of the delocalised electrons that are free to move within the structure.

Diamond



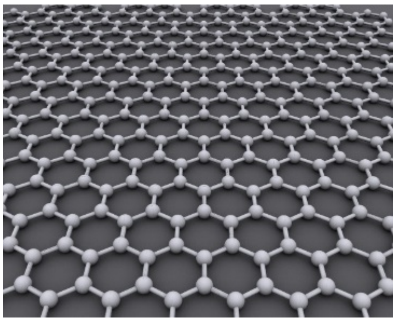
- Giant covalent structure.
- High melting and boiling point.
- Very hard (strong covalent bonds between atoms).
- Each carbon is bonded to 4 other carbon atoms.
- Bond angle is 109.5^o, tetrahedral.
- Does not conduct electricity (no delocalised electrons).
- Used in jewellery and for cutting glass.

Fullerene C_{60}



- Each carbon atom is bonded to 3 other carbon atoms.
- Structure consists of 12 pentagons and 20 hexagons.
- Poor electrical conductor (better than diamond but worse than graphite).

Graphene



- Each carbon atom is bonded to 3 other carbon atoms.
- Bond angle between carbon atoms is 120^o
- Very good heat and electrical conductivity.
- Very thin (one layer thick) but very strong.

Allotropes of carbons and their uses

Allotrope	Uses
Graphite	dry lubricant, electrode rods, pencils
Diamond	jewellery, tools for cutting glass
Fullerene C_{60}	lubricant, nanotubes
Graphene	lightweight, thin, flexible, yet durable display screens, electric/photonics circuits, solar cells

Exercises:

1. Outline why graphite is a good conductor of electricity but diamond is not.

2. Give one reason for the high melting and boiling point of diamond.

3. State and explain the bond angles of graphite and diamond.

4. Explain why graphite is used for making electrodes.

5. Explain why graphite is a very soft substance.

4.4 Intermolecular forces

Understandings:

- Intermolecular forces include London (dispersion) forces, dipole-dipole forces and hydrogen bonding.
- The relative strengths of these interactions are London (dispersion) forces < dipole-dipole forces < hydrogen bonds.

Applications and skills

- Deduction of the types of intermolecular force present in substances, based on their structure and chemical formula.
- Explanation of the physical properties of covalent compounds (volatility, electrical conductivity and solubility) in terms of their structure and intermolecular forces.

Guidance:

- The term "London (dispersion) forces" refers to instantaneous induced dipole-induced dipole forces that exist between any atoms or groups of atoms and should be used for non-polar entities. The term "van der Waals" is an inclusive term, which includes dipole–dipole, dipole-induced dipole and London (dispersion) forces.

Syllabus checklist

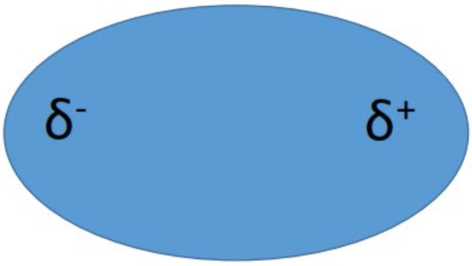
Objective	I am confident with this	I need to review this	I need help with this
Outline the formation of the three types of intermolecular force			
Outline the features of a molecule that determine the type of intermolecular forces that exist between the molecules			
Explain the physical properties of covalent compounds (volatility, electrical conductivity and solubility) based on the type of intermolecular forces between the molecules			

Intermolecular forces

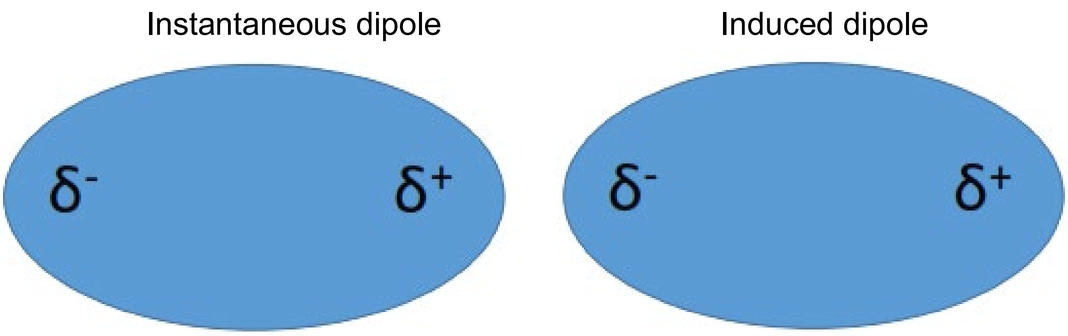
- Intermolecular forces are forces of attraction **between** molecules.
- The three types of intermolecular force are:
  London dispersion forces, dipole-dipole forces and hydrogen bonding.
- London dispersion forces and dipole-dipole forces are collectively known as van der Waals forces.

London dispersion forces

- London dispersion forces are the weakest type of intermolecular force.
- They exist between all atoms or molecules (both polar and non-polar).
- Due to the constant motion of electrons, an atom or molecule can develop a temporary (instantaneous) dipole.



- An instantaneous dipole in one molecule can cause an induced dipole in a nearby molecule.



- London dispersion forces have a significant effect on the physical properties of molecules, such as boiling point.

Exercises:

1. State the types of intermolecular forces that are van der Waals forces.

2. Outline the formation of London dispersion forces.

Boiling point of the halogens

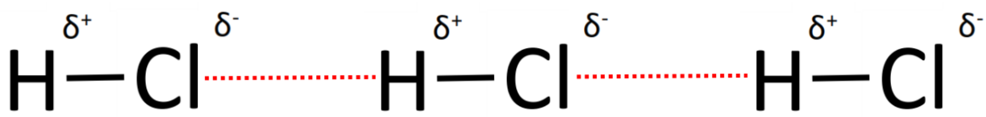
	Molar mass (g mol^{-1})	Boiling point (°C)
F_2	30.8	-188
Cl_2	70.9	-34.0
Br_2	160	58.0
I_2	254	193

- The boiling points of the halogens increase as their molar masses increase.
- The strength of London dispersion forces increases with increasing molar mass.
- Stronger London dispersion forces result in a higher boiling point (more energy is required to overcome the attractive forces between molecules).

Exercise: Explain why the boiling points of the halogens increase down the group.

Dipole-dipole forces

- Dipole-dipole forces occur between polar molecules that have a permanent dipole.
- They are the second strongest type of intermolecular force.
- There is an electrostatic attraction between the partial positive charge in one molecule and the partial negative charge on another.
- The dipole-dipole attractions between HCl molecules are shown below.



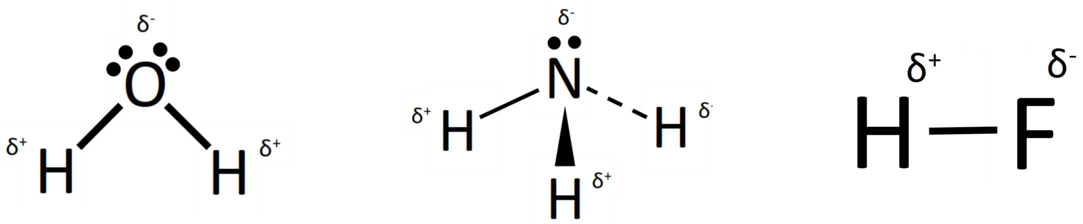
Exercises:

1. State the type of molecules that dipole-dipole forces occur between.

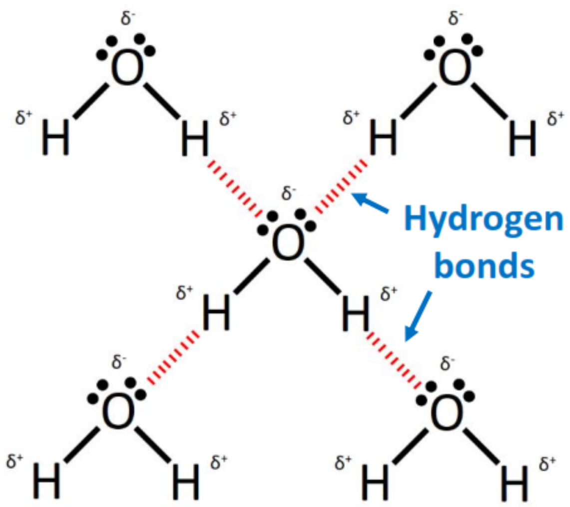
2. Outline the formation of the dipole-dipole forces in HCl

Hydrogen bonding

- Hydrogen bonding occurs when a hydrogen atom is bonded to either a nitrogen, oxygen or fluorine atom.
- It is the strongest type of intermolecular force.
- Example of compounds that have hydrogen bonding between their molecules are H_2O, NH_3 and HF.



- Hydrogen bonds between water molecules are shown below. These are responsible for the high boiling point of water.



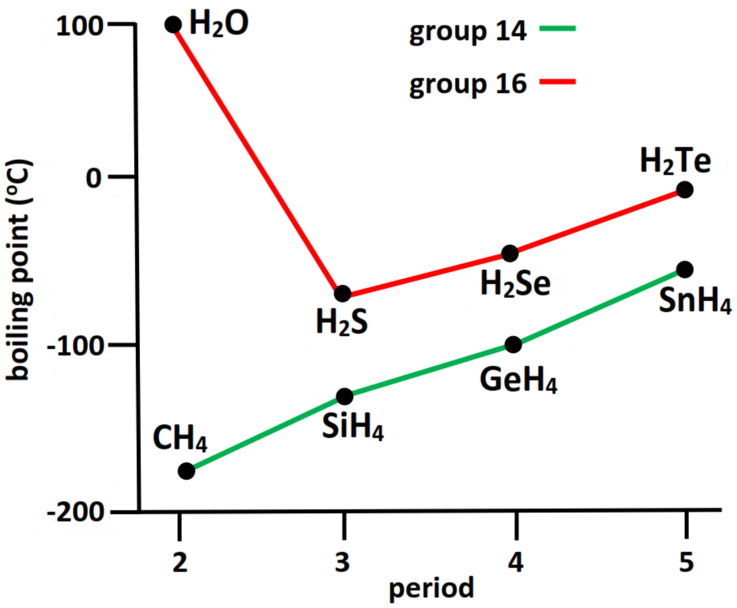
- The hydrogen bond is between the partial positive charge on the hydrogen atom and the partial negative charge on the oxygen atom
- Water has a much higher boiling point compared to other molecules with similar molar masses because of the effect of hydrogen bonding.

Exercises:

1. Which groups of atoms are necessary for the formation of a hydrogen bond?

2. Outline the formation of a hydrogen bond between water molecules.

Boiling points of group 14 and group 16 hydrides



Exercise: From the above graph, describe and explain the trend in boiling point of the group 14 and group 16 hydrides.

Exercises:

1. Arrange the following in terms of increasing strength:

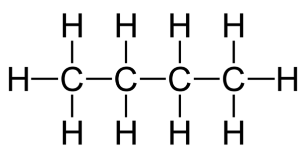
   Hydrogen bonding   London dispersion forces   Dipole-dipole forces

2. Identify the **strongest** type of intermolecular forces in the following molecules:

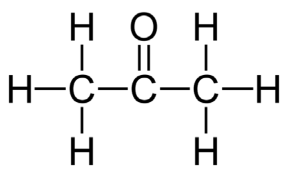
   a) Cl_2
   b) HCl
   c) HF
   d) CH_4
   e) CCl_4
   f) CH_3Cl
   g) H_2O
   h) CH_3OH
   i) C_2H_6
   j) NH_3

3. Explain why, at room temperature, F_2 and Cl_2 are gases, Br_2 is a liquid and I_2 is a solid.

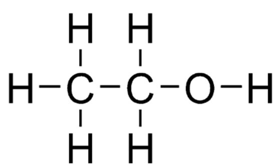
4. The structures of butane, propanone, and ethanol can be seen below. They have similar molar masses, but different boiling points. Explain the difference in boiling point between the three molecules.



butane B.P. -1^{o}C



propanone B.P 56^{o}C



ethanol B.P 78^{o}C

4.5 Metallic bonding

Understandings:

- A metallic bond is the electrostatic attraction between a lattice of positive ions and delocalized electrons.
- The strength of a metallic bond depends on the charge of the ions and the radius of the metal ion.
- Alloys usually contain more than one metal and have enhanced properties.

Applications and skills:

- Explanation of electrical conductivity and malleability in metals.
- Explanation of trends in melting points of metals.
- Explanation of the properties of alloys in terms of non-directional bonding.

Guidance:

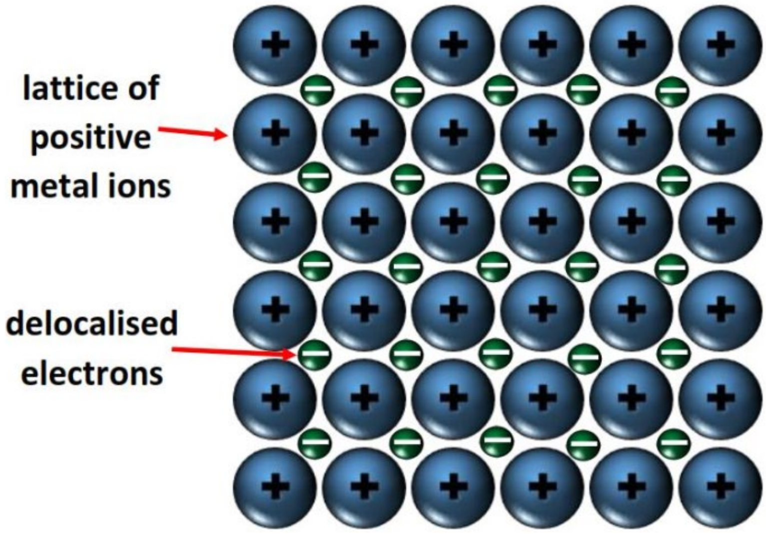
- Trends should be limited to s- and p-block elements.
- Examples of various alloys should be covered.

Syllabus objectives

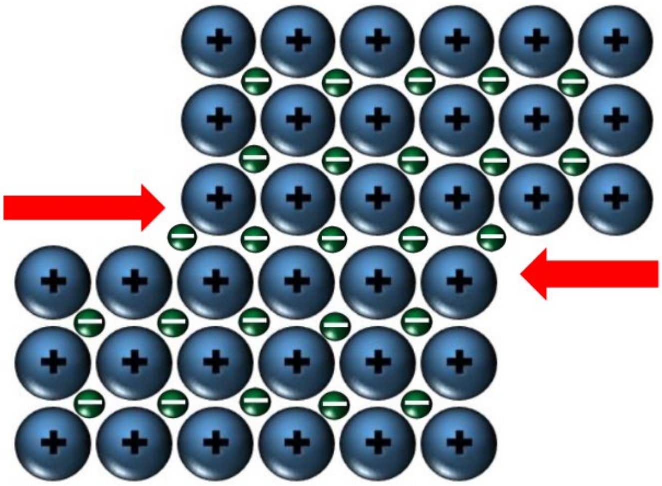
Objective	I am confident with this	I need to review this	I need help with this
Describe the structure of metals and the formation of the metallic bond			
Explain why metals are malleable			
Explain why metals are good electrical conductors			
Explain the factors that affect the strength of the metallic bond			
Explain the enhanced properties of alloys			

Metallic bonding

- The metallic bond is the electrostatic attraction between a lattice of positively charged metal ions and delocalised electrons.



- Metals are good conductors of heat and electricity because of the presence of delocalised electrons.
- Metals are malleable (can be bent into shape) and ductile (can be drawn into wires) because the metallic bond remains intact even if the structure is distorted.
- The layers can slide over each other when metals are bent, hammered, or stretched, without breaking the metallic bond.



- Metals are reflective – the delocalised electrons in the metallic structure reflect light.

Strength of the metallic bond

- The strength of the metallic bond is determined by the charge on the metal ion and the ionic radius of the metal ion.

Ion	charge on ion	ionic radius (× 10^{-12} m)	melting point (°C)
Na^+	1+	102	98
Mg^{2+}	2+	72	650

- Mg has a higher melting point than Na due to its greater ionic charge and smaller ionic radius.
- The higher the charge on the ion, the more delocalised electrons that exist in the metallic structure.
- This results in a stronger electrostatic attraction between the lattice of positive metal ions and the delocalised electrons, and a stronger metallic bond.
- The strength of the metallic bond decreases down a group as the size of the metal cation increases.
- This results in a decreasing melting point down a group, such as group 1.

Exercises:

1. Outline the metallic structure and the formation of the metallic bond.

2. Explain the following properties of metals:

   a. Metals are good conductors of heat and electricity

   b. Metals are malleable and ductile

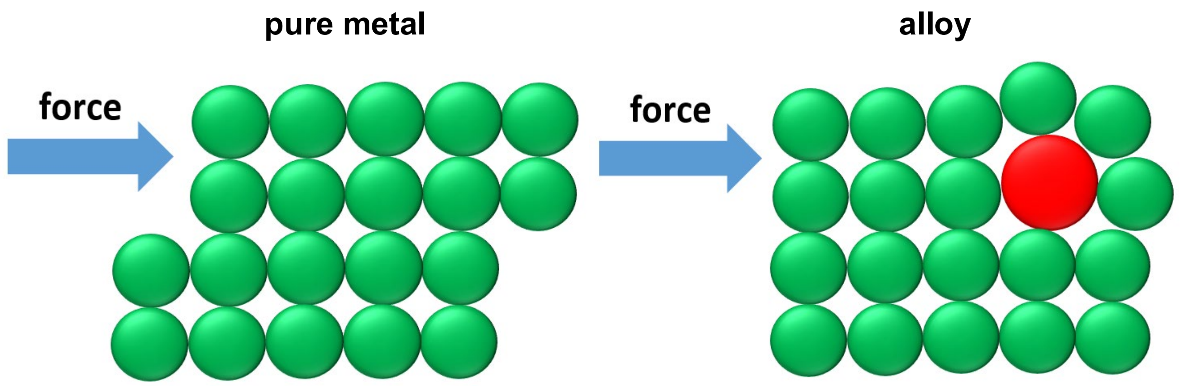
   c. Metals are shiny

3. Explain why aluminium has a higher melting point than sodium.

4. Explain why the melting point of the group 1 metals decreases down the group.

Alloys

- Alloys are materials that are composed of two or more metals or a metal and a non-metal.
- The bonding in metals is non-directional; the force of attraction between the positive metal ions and the delocalised electrons acts in all directions around the fixed metal ions.
- Alloys have enhanced properties (increased tensile strength and increased resistance to corrosion); they have different properties to the metals that they are made from.
- They tend to be harder (less malleable) and have greater tensile strength (stronger).
- The added metal atoms can distort the lattice structure.
- The distortion of the lattice structure makes it more difficult for the layers to slide over each other.



- In a pure metal the layers can slide over each other.
- The presence of different sized metal atoms (ions) means the layers cannot slide over each other as easily as in a pure metal.

Exercise: Explain why alloys are harder than pure metals.

Uses of alloys

Alloy	Component metals	Properties and uses
Steel	iron, carbon	high tensile strength; used in construction
Stainless steel	iron, nickel, chromium	resistant to corrosion; used in cooking implements
Brass	copper and zinc	pipes
Bronze	copper and tin	coins, medals, tools
Pewter	tin, copper, antimony	decorative ornaments
Solder	lead and tin	low melting point; used to join metals in electrical circuits
Nichrome	nickel and chromium	heating elements

Exercises: Suggest an alloy for the following uses with a reason for your choice.

1. To use in electrical circuits.
2. To make water pipes.
3. To use in an ornament.
4. To construct a bridge.
5. To make a saucepan.