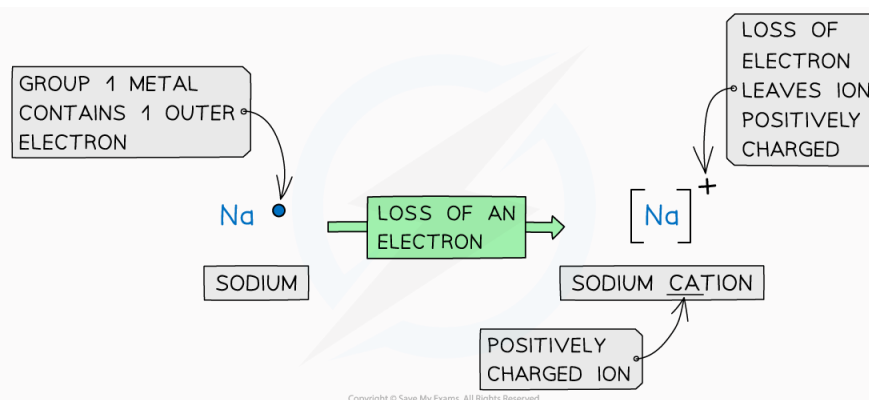
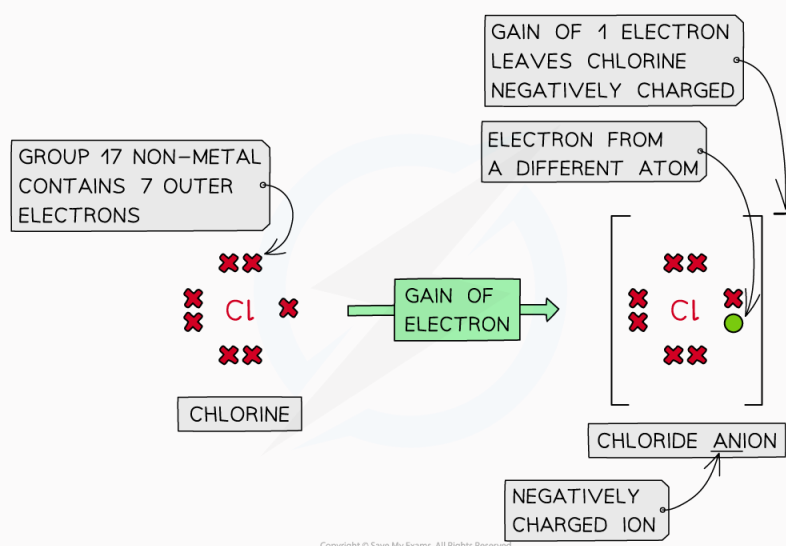


Forming Ions

- As a general rule, metals are on the left of the Periodic Table and non-metals are on the right-hand side
- Ionic bonds involve the transfer of electrons from a metallic element to a non-metallic element
- Transferring electrons usually leaves the metal and the non-metal with a full outer shell
- Metals lose electrons from their valence shell forming positively charged cations
- Non-metal atoms gain electrons forming negatively charged anions
- Once the atoms become ions, their electronic configurations are the same as a noble gas.
 - A sodium ion (Na^+) has the same electronic configuration as neon: [2,8]
 - A chloride ion (Cl^-) also has the same electronic configuration as argon: [2,8,8]

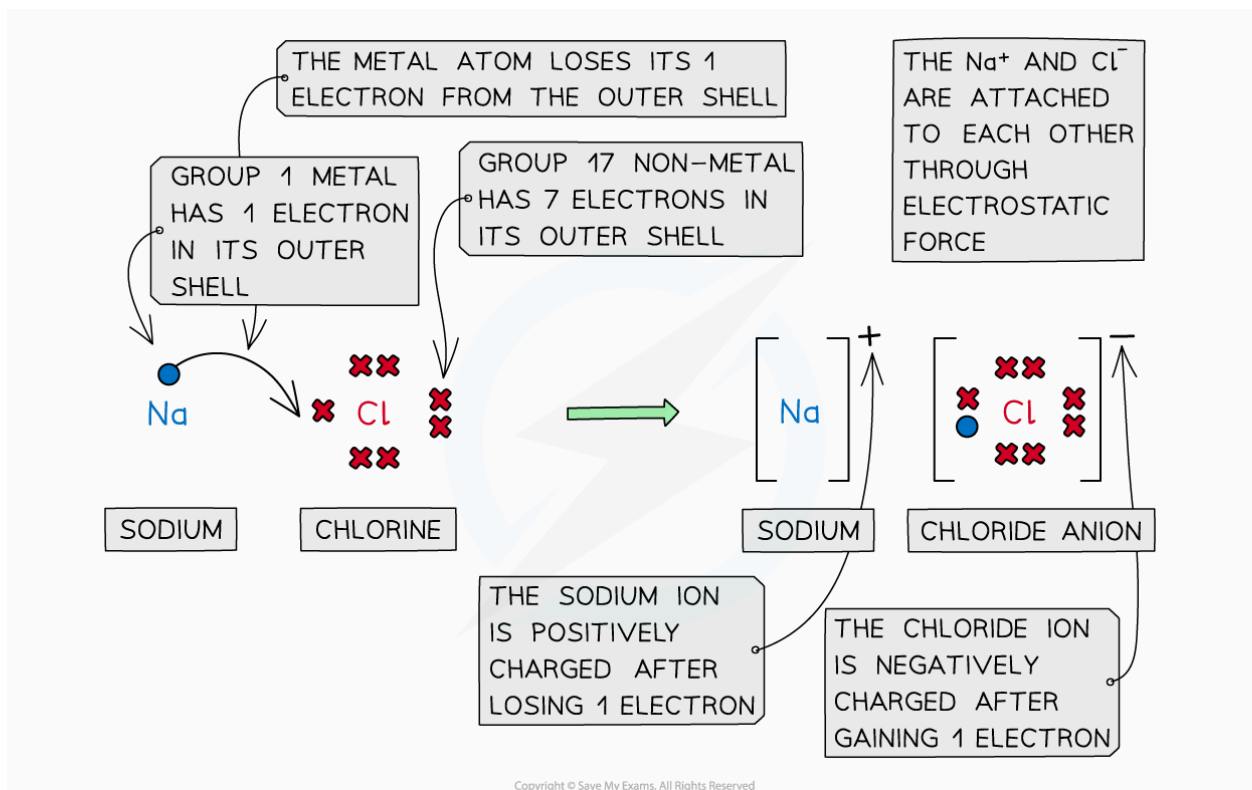


Forming cations by the removal of electrons from metals



Forming anions by the addition of electrons to nonmetals

- Cations and anions are oppositely charged and therefore attracted to each other
- Electrostatic attractions are formed between the oppositely charged ions to form ionic compounds
- This form of attraction is very strong and requires a lot of energy to overcome
 - This causes high melting points in ionic compounds



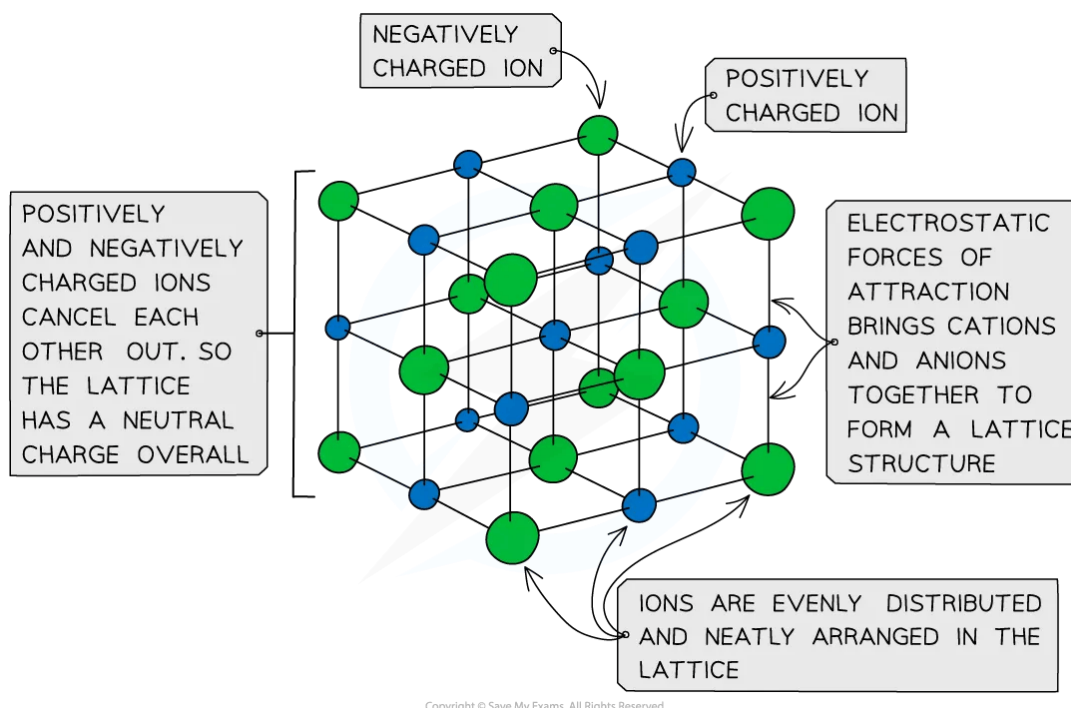
Cations and anions bond together using strong electrostatic forces, which require a lot of energy to overcome

Exam Tip

Metals usually lose all electrons from their outer valence shell to become cations. You can make use of the groups on the periodic table to work out how many electrons an atom is likely to lose or gain by looking at the group an atom belongs to.

Ionic Lattices

- The ions form a lattice structure which is an evenly distributed crystalline structure
- Ions in a lattice are arranged in a regular repeating pattern so that positive charges cancel out negative charges
- Therefore the final lattice is overall electrically neutral



Ionic solids are arranged in lattice structures

Properties of Ionic Compounds

- Different types of structure and bonding have different effects on the physical properties of substances such as their melting and boiling points, electrical conductivity and solubility

Ionic bonding & giant ionic lattice structures

- Ionic compounds are strong
 - The strong electrostatic forces in ionic compounds keep the ions held strongly together
- They are brittle as ionic crystals can split apart
- Ionic compounds have high melting and boiling points
 - The strong electrostatic forces between the ions in the lattice act in all directions and keep them strongly together
 - Melting and boiling points increase with the charge density of the ions due to the greater electrostatic attraction of charges
 - $Mg^{2+}O^{2-}$ has a higher melting point than $Na^{+}Cl^{-}$
- Ionic compounds are soluble in water as they can form ion-dipole bonds
- Ionic compounds only conduct electricity when molten or in solution
 - When molten or in solution, the ions can freely move around and conduct electricity
 - As a solid, the ions are in a fixed position and unable to move around

Table comparing the characteristics of giant ionic lattices with other structure types

	Giant ionic	Giant metallic	Simple covalent	Giant covalent
Melting / boiling point	High	Moderately high to high	Low	Very high
Electrical conductivity	Only when molten or in solution	When solid or liquid	Do not conduct electricity	Do not conduct electricity (except graphite)
Solubility	Soluble	Insoluble but some may react	Usually insoluble unless they are polar	Insoluble
Hardness	Hard, brittle	Hard, malleable	Soft	Very hard (diamond and silica) or soft (graphite)
Physical state at room temperature	Solid	Solid	Solid, liquid or gas	Solid
Forces	Electrostatic attraction between ions	Delocalised electrons attracting positive ions	Weak intermolecular forces and covalent bonds within a molecule	Electrons in covalent bonds between atoms
Particles	Ions	Positive ions in a sea of electrons	Small molecules	Atoms

Examples	NaCl	Copper	Br ₂	Graphite, silicon(IV) oxide
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Worked example

The table below shows the physical properties of substances X, Y and Z.

Substance	Melting point (°C)	Electrical conductivity when molten	Solubility in water
X	839	Good	Soluble
Y	95	Very poor	Almost insoluble
Z	1389	Good	Insoluble

Which one of the following statements about X, Y and Z is completely true?

Statement 1: X has a giant ionic structure, Y has a giant molecular structure, Z is a metal

Statement 2: X is a metal, Y has a simple molecular structure, Z has a giant molecular structure

Statement 3: X is a metal, Y has a simple molecular structure, Z has a giant ionic structure

Statement 4: X has a giant ionic structure, Y has a simple molecular structure, Z is a metal

Answer:

- Compound X has a relatively high melting point, is soluble in water and conducts electricity when molten

- This suggests that X has a giant ionic structure
- Compound Y has a low melting point which suggests that little energy is needed to break the lattice
 - This suggests that Y is a simple molecular structure
 - This is further supported by its low electrical conductivity and it being almost insoluble in water
- Compound Z has a very high melting point, which is characteristic of either metallic, giant ionic lattices or giant covalent / molecular lattices
 - However since it is insoluble in water, compound Z must be a metal
- Therefore, the correct answer is Statement 4

Formulae & Names of Ionic Compounds

- Ionic compounds are formed from a metal and a nonmetal bonded together
- Ionic compounds are electrically neutral; the positive charges equal the negative charges

Charges on positive ions

- All metals form positive ions
 - There are some non-metal positive ions such as ammonium, NH_4^+ , and hydrogen, H^+
- The metals in Group 1, Group 2 and Group 13 have a charge of 1+ and 2+ and 3+ respectively
- The charge on the ions of the transition elements can vary which is why Roman numerals are often used to indicate their charge
- This is known as Stock notation after the German chemist Alfred Stock
- Roman numerals are used in some compounds formed from transition elements to show the charge (or oxidation state) of metal ions
 - Eg. in copper (II) oxide, the copper ion has a charge of 2+ whereas in copper (I) nitrate, the copper has a charge of 1+

Non-metal ions

- The non-metals in group 15 to 17 have a negative charge and have the suffix 'ide'
 - Eg. nitride, chloride, bromide, iodide
- Elements in group 17 gain 1 electron so have a 1- charge, eg. Br^-
- Elements in group 16 gain 2 electrons so have a 2- charge, eg. O^{2-}
- Elements in group 15 gain 3 electrons so have a 3- charge, eg. N^{3-}
- There are also more polyatomic or compound negative ions, which are negative ions made up of more than one type of atom

GROUP							18	
1	2	H^+	13	14	15	16	17	NONE
Li^+	Be^{2+}					O^{2-}	F^-	NONE
Na^+	Mg^{2+}		Al^{3+}			S^{2-}	Cl^-	NONE
K^+	Ca^{2+}	TRANSITION ELEMENTS	Ga^{3+}				Br^-	NONE
Rb^+	Sr^{2+}						I^-	NONE

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The charges of simple ions depend on their position in the Periodic Table

- There are seven polyatomic ions you need to know for IB Chemistry:

Formulae of Polyatomic Ions Table

Ion	Formula and Charge
Ammonium	NH_4^+
Hydroxide	OH^-
Nitrate	NO_3^-
Sulfate	SO_4^{2-}
Carbonate	CO_3^{2-}
Hydrogen carbonate	HCO_3^-
Phosphate	PO_4^{3-}

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Worked example

Determine the formulae of the following ionic compounds

- magnesium chloride
- aluminium oxide
- ammonium sulfate

Answer:

Answer 1: Magnesium chloride

- Magnesium is in group 2 so has a charge of 2+
- Chlorine is in group 17 so has a charge of 1-
- Magnesium needs two chlorine atoms for each magnesium atom to be balanced so the formula is MgCl_2

Answer 2: Aluminium oxide

- Aluminum is in group 13 so the ion has a charge of 3+
- Oxygen is in group 16 so has a charge of 2-
- The charges need to be equal so 2 aluminium to 3 oxygen atoms will balance electrically, so the formula is Al_2O_3

Answer 3: Ammonium sulfate

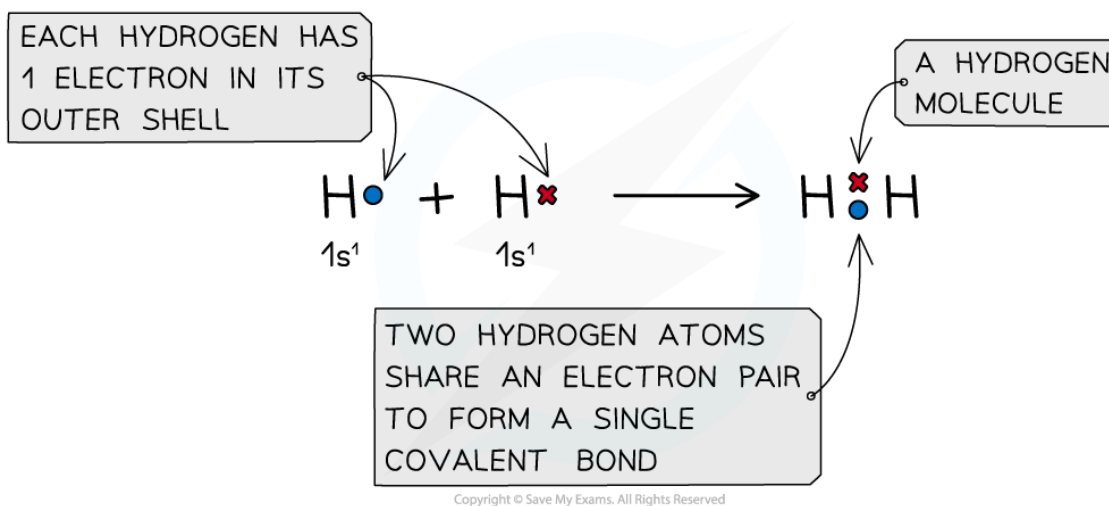
- Ammonium is a polyatomic ion with a charge of 1+
- Sulfate is a polyatomic ion and has a charge of 2-
- The polyatomic ion needs to be placed in a bracket if more than 1 is needed
- The formula of ammonium nitrate is $(\text{NH}_4)_2\text{SO}_4$

Exam Tip

Remember: polyatomic ions are ions that contain more than one type of element, such as OH^-

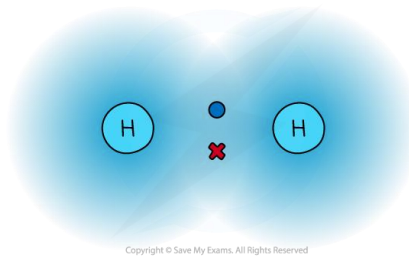
Covalent Bonds

- Covalent bonding occurs between two non-metals
- A covalent bond involves the electrostatic attraction between nuclei of two atoms and the electrons of their outer shells
- No electrons are transferred but only shared in this type of bonding
- When a covalent bond is formed, two atomic orbitals overlap and a molecular orbital is formed
- Covalent bonding happens because the electrons are more stable when attracted to two nuclei than when attracted to only one



The positive nucleus of each atom has an attraction for the bonding electrons shared in the covalent bond

- In a normal covalent bond, each atom provides one of the electrons in the bond. A covalent bond is represented by a short straight line between the two atoms, H-H
- Covalent bonds should not be regarded as shared electron pairs in a fixed position; the electrons are in a state of constant motion and are best regarded as charge clouds



A representation of electron charge clouds. The electrons can be found anywhere in the charge clouds

- Non-metals are able to share pairs of electrons to form different types of covalent bonds
- Sharing electrons in the covalent bond allows each of the 2 atoms to achieve an electron configuration similar to a noble gas
 - This makes each atom more stable
- In some instances, the central atom of a covalently bonded molecule can accommodate more or less than 8 electrons in its outer shell
 - Being able to accommodate more than 8 electrons in the outer shell is known as 'expanding the octet rule'
 - Accommodating less than 8 electrons in the outer shell means that the central atom is 'electron deficient'

- Some examples of this can be found in the section on Lewis structures

Exam Tip

Covalent bonding takes place between two nonmetal atoms. Remember to use the periodic table to decide how many electrons are in the outer shell of a nonmetal atom.

Predicting Covalent Bonding

- The differences in Pauling electronegativity values can be used to predict whether a bond is covalent or ionic in character

Electronegativity & covalent bonds

- In diatomic molecules the electron density is shared equally between the two atoms
 - Eg. H_2 , O_2 and Cl_2
- Both atoms will have the same electronegativity value and have an equal attraction for the bonding pair of electrons leading to formation of a covalent bond
- A difference of less than around 1.0 in electronegativity values will be associated with covalent bonds, although between 1.0 and 2.0 can be considered polar covalent:

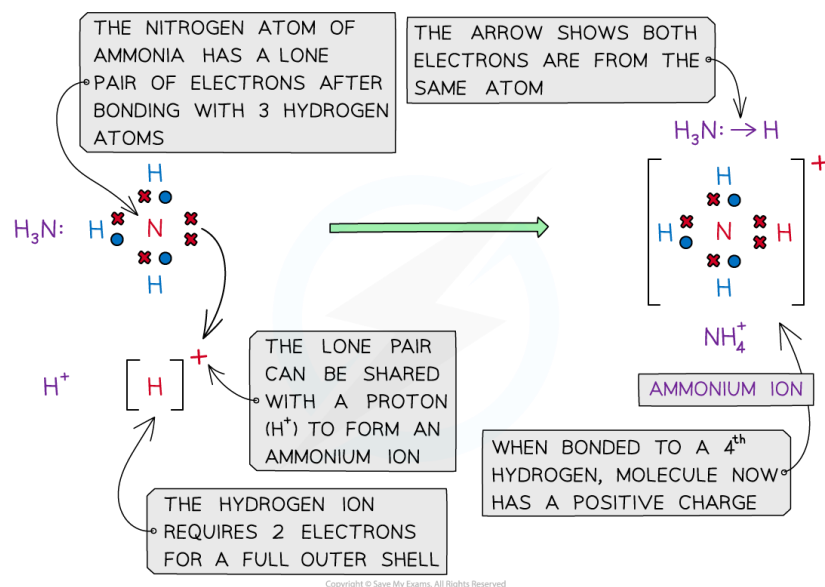
You can use the Pauling scale to decide whether a bond is polar or nonpolar:

Difference in Electronegativity	Bond Type
< 1.0	Covalent
1.0 – 2.0	Polar Covalent
> 2.0	Ionic

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Coordinate Bonds

- In simple covalent bonds the two atoms involved share electrons
- Some molecules have a lone pair of electrons that can be donated to form a bond with an electron-deficient atom
 - An electron-deficient atom is an atom that has an unfilled outer orbital
- So both electrons are from the same atom
- This type of bonding is called dative covalent bonding or coordinate bond
- An example of a dative bond is in an ammonium ion
 - The hydrogen ion, H^+ is electron-deficient and has space for two electrons in its shell
 - The nitrogen atom in ammonia has a lone pair of electrons which it can donate to the hydrogen ion to form a dative covalent bond



Ammonia (NH_3) can donate a lone pair to an electron-deficient proton (H^+) to form a charged ammonium ion (NH_4^+)

- More examples of coordinate bonding can be found in the section on Lewis Structures

Multiple Bonds

- Non-metals are able to share more than one pair of electrons to form different types of covalent bonds
- Sharing electrons in the covalent bond allows each of the 2 atoms to achieve an electron configuration similar to a noble gas
 - This makes each atom more stable
- It is not possible to form a quadruple bond as the repulsion from having 8 electrons in the same region between the two nuclei is too great

Covalent Bonds & Shared Electrons Table

Type of covalent bond	Number of electrons shared
Single (C – C)	2
Double (C = C)	4
Triple (C ≡ C)	6

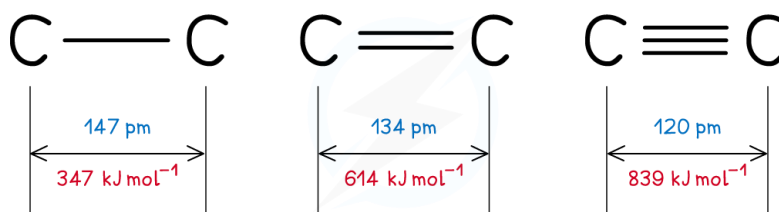
Bond Length & Strength

Bond energy

- The bond energy is the energy required to break one mole of a particular covalent bond in the gaseous states
 - Bond energy has units of kJ mol^{-1}
- The larger the bond energy, the stronger the covalent bond is

Bond length

- The bond length is internuclear distance of two covalently bonded atoms
 - It is the distance from the nucleus of one atom to another atom which forms the covalent bond
- The greater the forces of attraction between electrons and nuclei, the more the atoms are pulled closer to each other
- This decreases the bond length of a molecule and increases the strength of the covalent bond
- Triple bonds are the shortest and strongest covalent bonds due to the large electron density between the nuclei of the two atoms
- This increase the forces of attraction between the electrons and nuclei of the atoms
- As a result of this, the atoms are pulled closer together causing a shorter bond length
- The increased forces of attraction also means that the covalent bond is stronger



Triple bonds are the shortest covalent bonds and therefore the strongest ones

- Test your knowledge of covalent bonding:

Worked example

Which molecules react together to form a dative covalent bond?

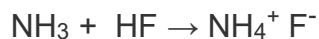
- Cl_2 and HF
- C_2H_2 and Cl_2
- NH_3 and HF
- CH_4 and NH_3

Answer:

The correct option is C.

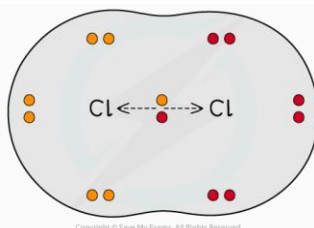
- To form a dative covalent bond one species must have a lone pair of electrons and the other must be electron deficient.

- NH_3 has a lone pair and HF splits into H^+ (electron deficient) and F^-



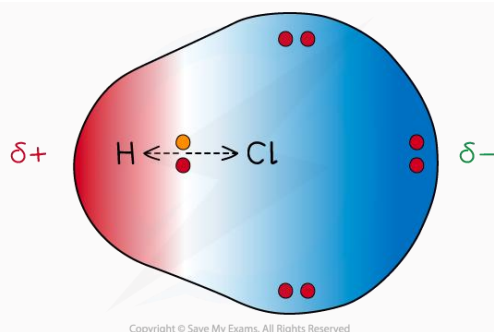
Bond Polarity

- When two atoms in a covalent bond have the same electronegativity the covalent bond is nonpolar



The two chlorine atoms have identical electronegativities so the bonding electrons are shared equally between the two atoms

- When two atoms in a covalent bond have different electronegativities the covalent bond is polar and the electrons will be drawn towards the more electronegative atom
- As a result of this:
 - The negative charge centre and positive charge centre do not coincide with each other
 - This means that the electron distribution is asymmetric
 - The less electronegative atom gets a partial charge of δ^+ (delta positive)
 - The more electronegative atom gets a partial charge of δ^- (delta negative)
- The greater the difference in electronegativity the more polar the bond becomes

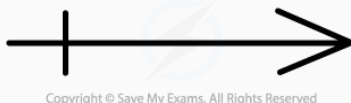


Cl has a greater electronegativity than H causing the electrons to be more attracted towards the Cl atom which becomes delta negative and the H delta positive

Dipole moment

- The dipole moment is a measure of how polar a bond is

- The direction of the dipole moment is shown by the following sign in which the arrow points to the partially negatively charged end of the dipole:



The sign shows the direction of the dipole moment and the arrow points to the delta negative end of the dipole

Worked example

The electronegativity values of four elements are given.

C = 2.6 N = 3.0 O = 3.4 F = 4.0

What is the order of increasing polarity of the bonds in the following compounds?

- A. $\text{CO} < \text{OF}_2 < \text{NO} < \text{CF}_4$
- B. $\text{NO} < \text{OF}_2 < \text{CO} < \text{CF}_4$
- C. $\text{CF}_4 < \text{CO} < \text{OF}_2 < \text{NO}$
- D. $\text{CF}_4 < \text{NO} < \text{OF}_2 < \text{CO}$

Answer:

The correct option is B.

- You have to calculate the difference in electronegativity for the bonds and then rank them from smallest to largest:

NO ($3.4 - 3.0 = 0.4$)

OF₂ ($4.0 - 3.4 = 0.6$)

CO ($3.4 - 2.6 = 0.8$)

CF₄ ($4.0 - 2.6 = 1.4$)

Lewis Structures

- Lewis structures are simplified electron shell diagrams and show pairs of electrons around atoms.
- A pair of electrons can be represented by dots, crosses, a combination of dots and crosses or by a line. For example, chlorine can be shown as:



Different Lewis Structures for chlorine molecules

- Note: Cl–Cl is not a Lewis structure, since it does not show all the electron pairs.
- The “octet rule” refers to the tendency of atoms to gain a valence shell with a total of 8 electrons

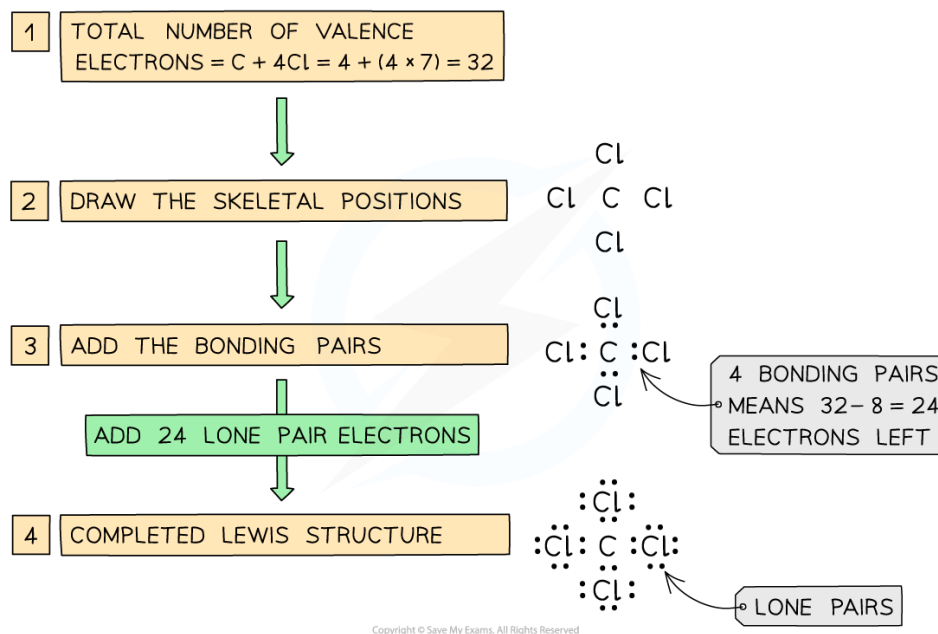
Steps for drawing Lewis Structures

1. Count the total number of valence
2. Draw the skeletal structure to show how many atoms are linked to each other.
3. Use a pair of crosses or dot/cross to put an electron pair in each bond between the atoms.
4. Add more electron pairs to complete the octets around the atoms (except H which has 2 electrons)
5. If there are not enough electrons to complete the octets, form double/triple bonds.
6. Check the total number of electrons in the finished structure is equal to the total number of valence electrons

Worked example

Draw a Lewis structure for CCl_4

Answer:



Steps in drawing the Lewis Structure for CCl_4

Further examples of Lewis structures

- Follow the steps for drawing Lewis structures for these common molecules

Molecule	Total number of valence electrons	Lewis structure
CH ₄	C + 4H 4 + (4 × 1) = 8	$\begin{array}{c} \text{H} \\ \\ \text{H}:\text{C}:\text{H} \\ \\ \text{H} \end{array}$
NH ₃	N + 3H 5 + (3 × 1) = 8	$\begin{array}{c} \text{H}:\ddot{\text{N}}:\text{H} \\ \\ \text{H} \end{array}$
H ₂ O	2H + O (2 × 1) + 6 = 8	$\text{H}:\ddot{\text{O}}:\text{H}$
CO ₂	C + 2O 4 + (2 × 6) = 16	$:\ddot{\text{O}}:\text{C}:\ddot{\text{O}}:$
HCN	H + C + N 1 + 4 + 5 = 10	$\text{H}:\text{C}::\text{N}:$

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Incomplete Octets

- For elements below atomic number 20 the octet rule states that the atoms try to achieve 8 electrons in their valence shells, so they have the same electron configuration as a noble gas
- However, there are some elements that are exceptions to the octet rule, such as H, Li, Be, B and Al
 - H can achieve a stable arrangement by gaining an electron to become 1s², the same structure as the noble gas helium
 - Li does the same, but losing an electron and going from 1s²2s¹ to 1s² to become a Li⁺ ion
 - Be from group 2, has two valence electrons and forms stable compounds with just four electrons in the valence shell
 - B and Al in group 13 have 3 valence electrons and can form stable compounds with only 6 valence electrons
- There are two examples of Lewis structures with incomplete octets you should know, BeCl₂ and BF₃:

Incomplete Octets Examples

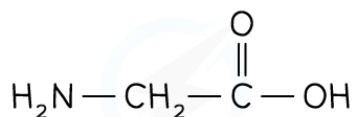
Molecule	Total number of valence electrons	Lewis structure
BeCl_2	$\text{Be} + 2\text{Cl} =$ $2 + (2 \times 7) = 16$	$:\ddot{\text{Cl}}:\text{Be}:\ddot{\text{Cl}}:$
BF_3	$\text{B} + 3\text{F} =$ $3 + (3 \times 7) = 24$	$:\ddot{\text{F}}:\text{B}:\ddot{\text{F}}:$ $:\ddot{\text{F}}:$

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- Test your understanding of Lewis diagrams in the following example:

Worked example

How many electrons are in the 2-aminoethanoic acid molecule?



2-AMINOETHANOIC ACID

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- A. 18
- B. 20
- C. 28
- D. 30

Answer:

The correct option is D.

- You must count the lone pairs on N and O as well as the bonding pairs. There are 5 'hidden' pairs of bonding electrons in the OH, CH_2 and NH_2 groups. Hydrogen does not follow the octet rule.

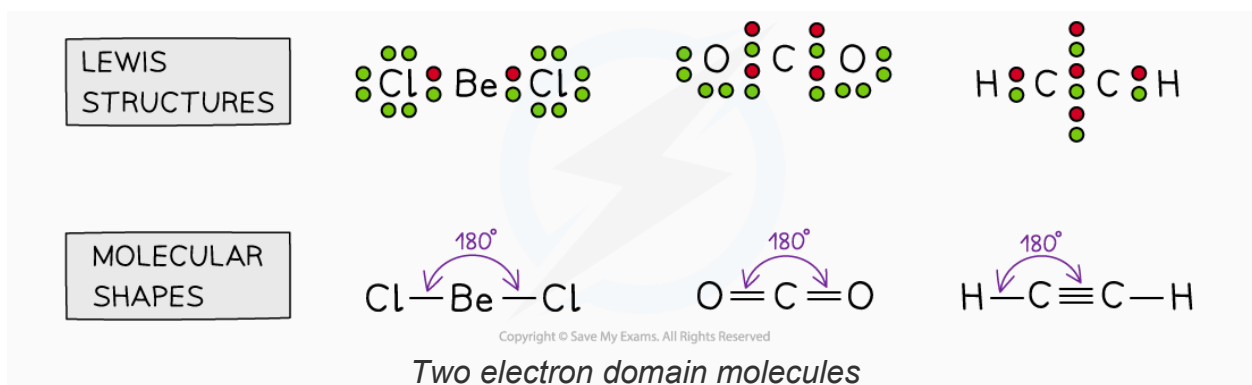
Shapes of Molecules

- When an atom forms a covalent bond with another atom, the electrons in the different bonds and the non-bonding electrons in the outer shell all behave as negatively charged clouds and repel each other
- In order to minimise this repulsion, all the outer shell electrons spread out as far apart in space as possible
- Molecular shapes and the angles between bonds can be predicted by the valence shell electron pair repulsion theory known by the abbreviation VSEPR theory
- VSEPR theory consists of three basic rules:

1. All electron pairs and all lone pairs arrange themselves as far apart in space as is possible.
 2. Lone pairs repel more strongly than bonding pairs.
 3. Multiple bonds behave like single bonds
- These three rules can be used to predict the shape of any covalent molecule or ion, and the angles between the bonds
 - The regions of negative cloud charge are known as domains and can have one, two or three pairs electrons

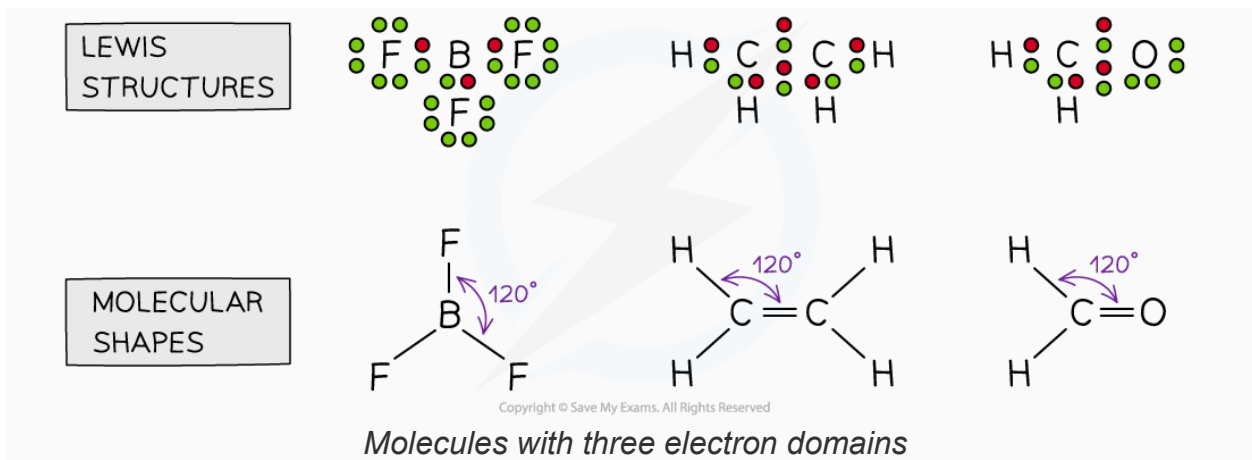
Two electron domains

- If there are two electron domains on the central atom, the angle between the bonds is 180°
- Molecules which adopt this shape are said to be LINEAR
- Examples of linear molecules include BeCl_2 , CO_2 , and $\text{HC}\equiv\text{CH}$

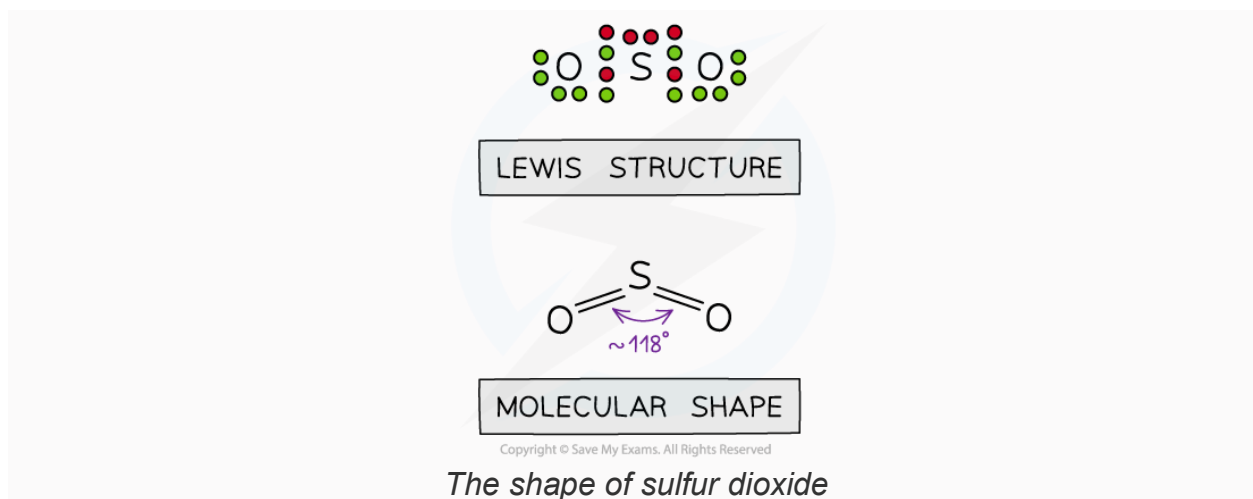


Three electron domains

- If there are three electron domains on the central atom, the angle between the bonds is 120°
- Molecules which adopt this shape are said to be TRIANGULAR PLANAR or TRIGONAL PLANAR
- Examples of three electrons domains which are all bonding pairs include BF_3 and CH_2CH_2 and CH_2O



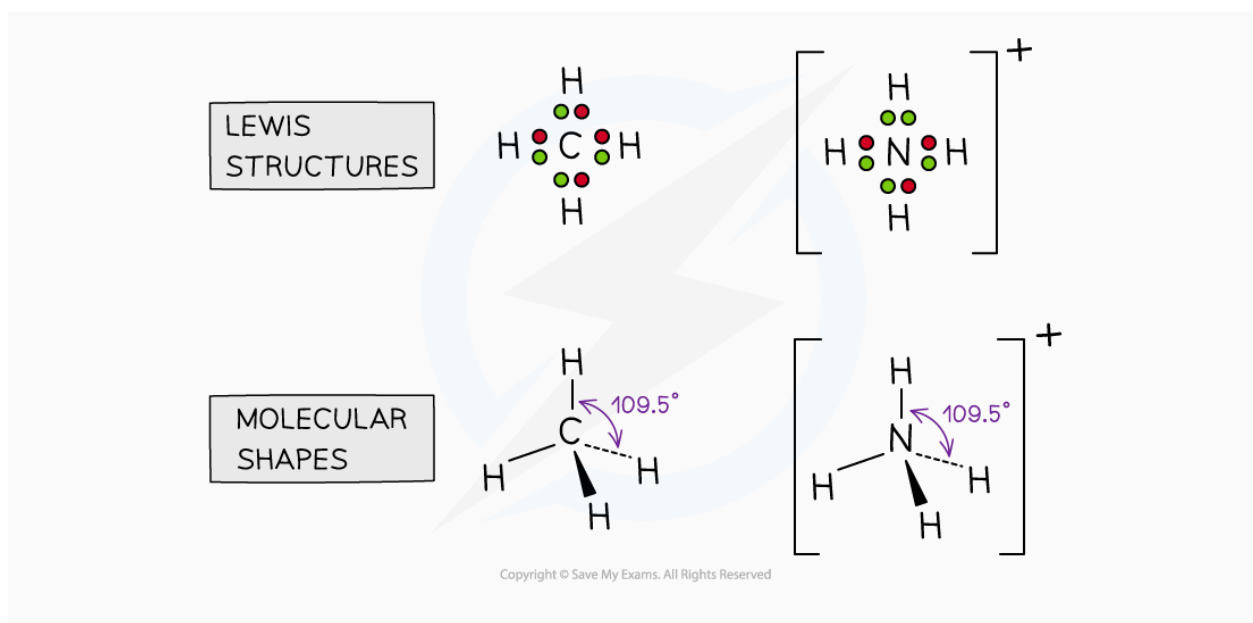
- If one of these electron domains is a lone pair, the bond angle is slightly less than 120° due to the stronger repulsion from lone pairs, forcing the bonding pairs closer together. E.g. SO_2
- The bond angle is approximately = 118°



- Sulfur dioxide is an example of a molecule that 'expands the octet' as you will see there are 10 electrons around the sulfur atom which is possible for 3rd period elements and above
- This shape is no longer called triangular planar as the shape names are only based on the atoms present, this molecule is BENT LINEAR

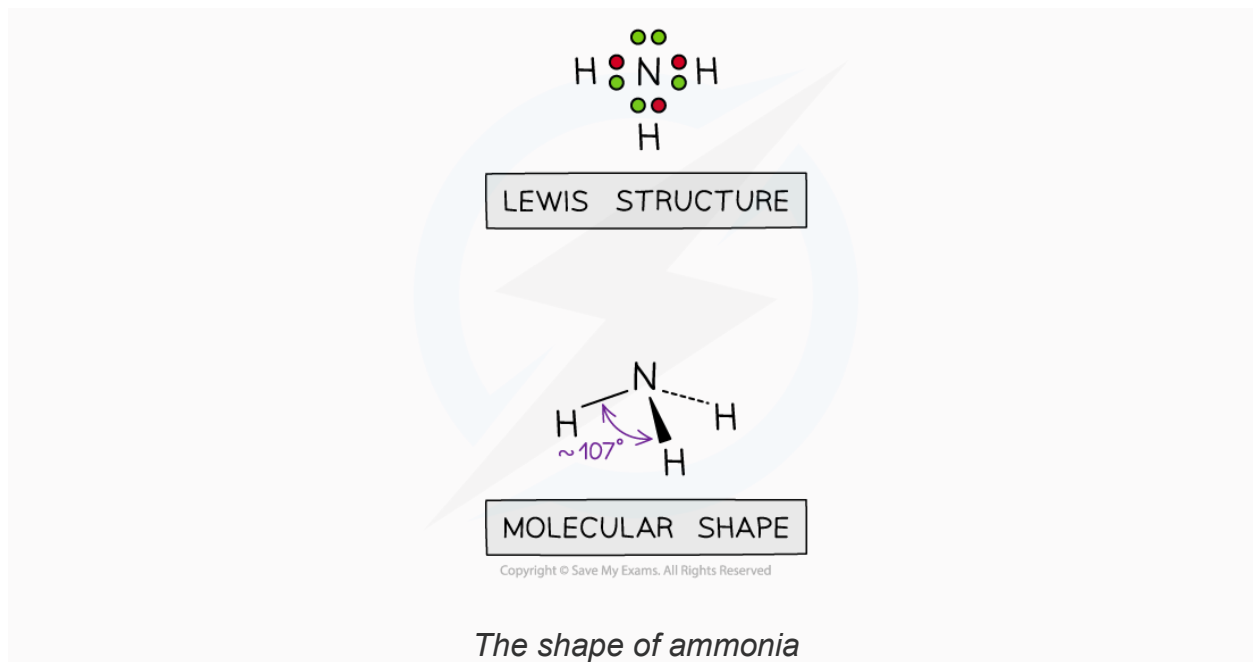
Four electron domains

- If there are four electron domains on the central atom, the angle between the bonds is approx 109° . E.g. CH_4 , NH_4^+
- Molecules which adopt this shape are said to be TETRAHEDRAL

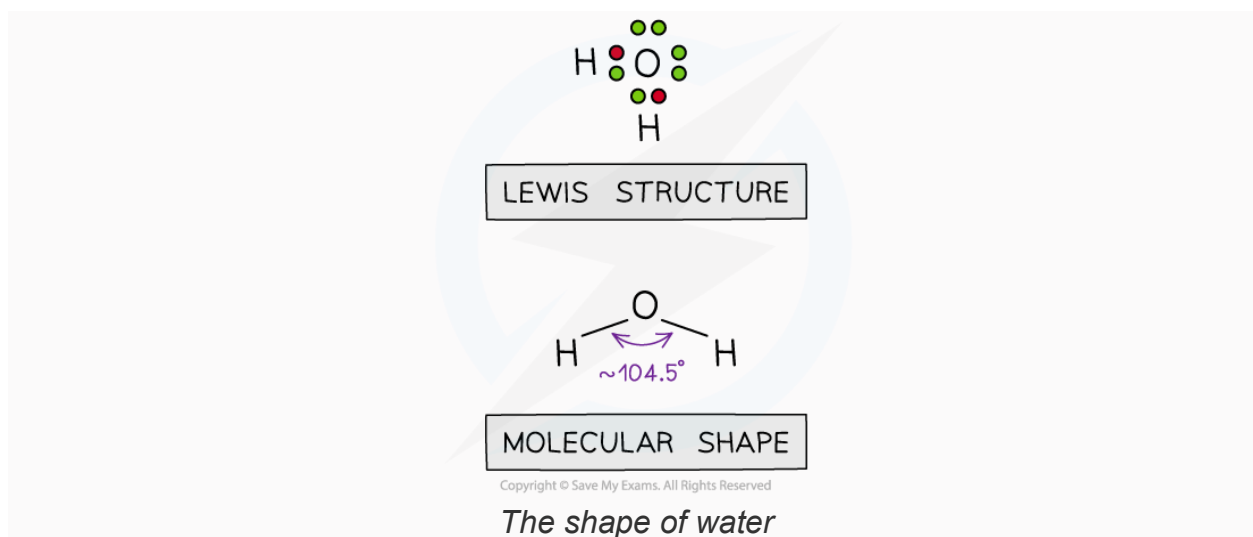


Molecules with four electron domains

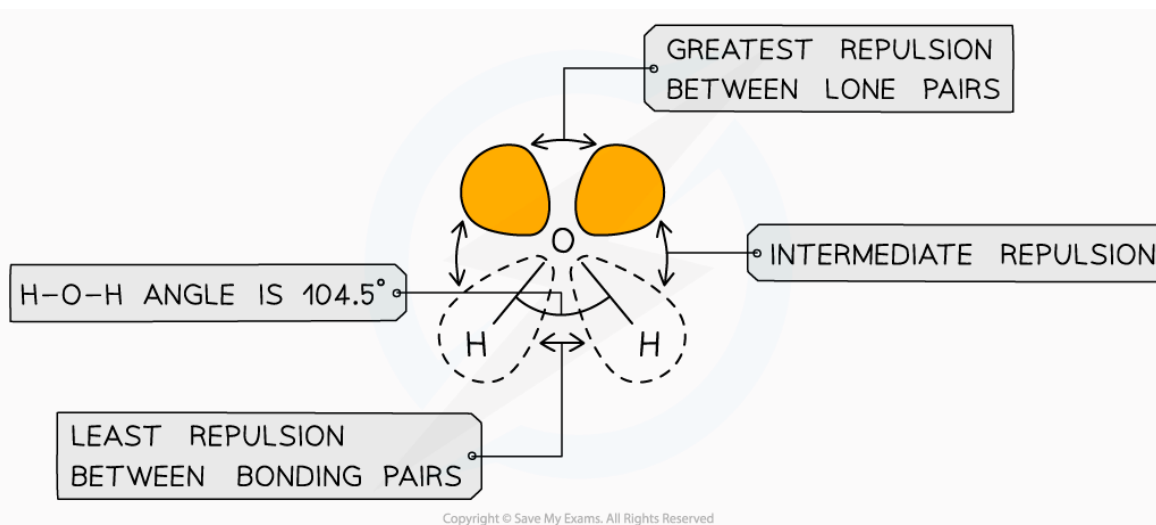
- If one of the electron domains is a lone pair, the bond angle is slightly less than 109° , due to the extra lone pair repulsion which pushes the bonds closer together (approx 107°). E.g. NH_3 ,



- Molecules which adopt this shape are said to be TRIANGULAR PYRAMIDAL or TRIGONAL PYRAMIDAL
- If two of the electron domains are lone pairs, the bond angle is also slightly less than 109° , due to the extra lone pair repulsion (approx 104°). E.g. H_2O
- Molecules which adopt this shape are said to be BENT or ANGULAR or BENT LINEAR or V-shaped (when viewed upside down)



- Lone pairs are pulled more closely to the central atoms so they exert a greater repulsive force than bonding pairs



Different types of electron pairs have different repulsive forces

Summary table of electron domains and molecular shapes

- These are the domains and molecular geometries you need to know for Standard Level:

Bonding pairs	Lone pairs	Total pairs	Domain geometry	Molecular geometry	Bond angle
2	0	2	linear	linear	180°
3	0	3	trigonal planar	trigonal planar	120°
2	1	3	trigonal planar	bent linear	118°
4	0	4	tetrahedral	tetrahedral	109.5°
3	1	4	tetrahedral	trigonal pyramid	107°
2	2	4	tetrahedral	bent linear	104.5°

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Exam Tip

Be careful to distinguish between molecular shape and electron domain shape as it can be easy to confuse the two. Sometimes they are the same as is the case of methane, but other times they can be different like ammonia which has a tetrahedral domain shape, but triangular pyramid molecular shape. Always draw the Lewis structure before you attempt to deduce the shape and bond angle as you could easily miss some lone pairs

Predicting Shapes & Bond Angles

- Before you predict the shape of any molecule work out the Lewis structure to determine the number of bonding and lone pairs
- Apply the VSEPR rules and you should be successful in deducing the correct shape and bond angle

Worked example

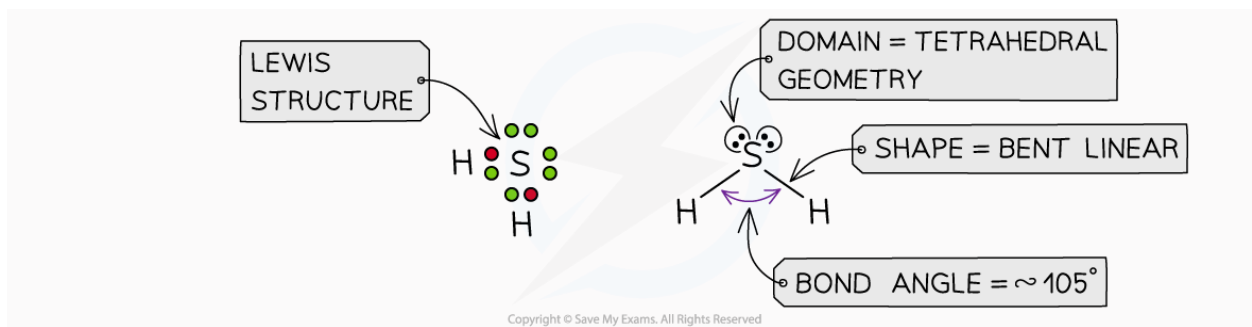
Predict the domain geometry, shape and bond angle in the following molecules or ions:

1. H₂S
2. NH₂Cl
3. NO₂⁺
4. ClF₂⁺

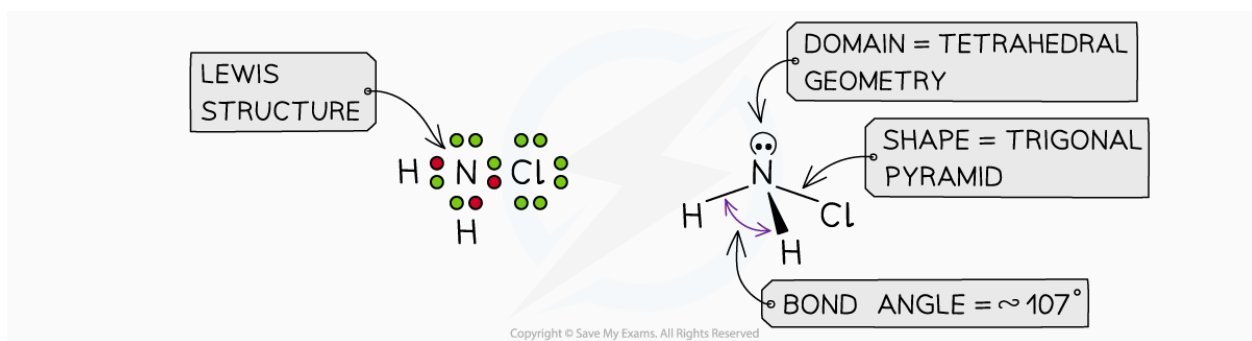
Answers:

Answer 1: The total number of valence electrons in H_2S is $1 + 1 + 6 = 8$, so there are four pairs of electrons around S

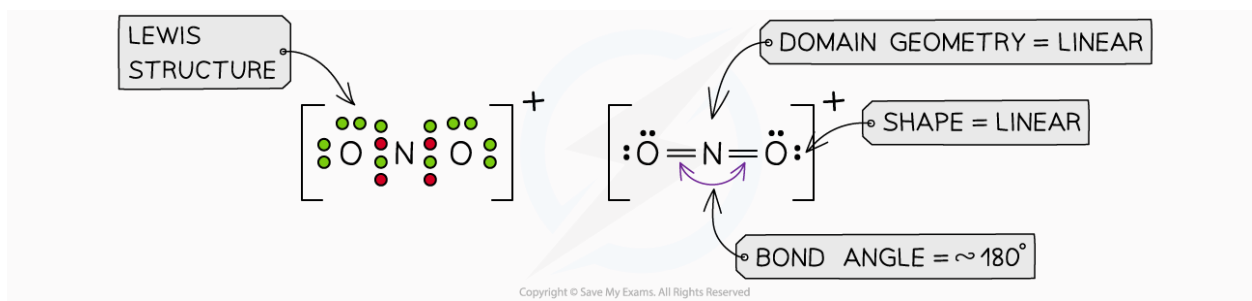
Hydrogen only forms one bond, so there are two bonding pairs and two lone pairs:



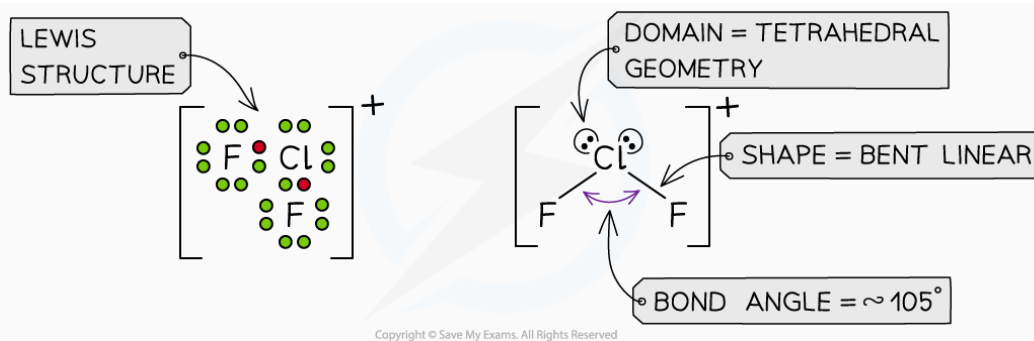
Answer 2: The total number of valence electrons in NH_2Cl is $5 + 1 + 1 + 7 = 14$



Answer 3: The total number of valence electrons in NO_2^+ is $5 + 6 + 6 - 1 = 16$ (subtracting one for the positive charge)



Answer 4: The total number of valence electrons in ClF_2^+ is $7 + 7 + 7 - 1 = 20$ (subtracting one for the positive charge)



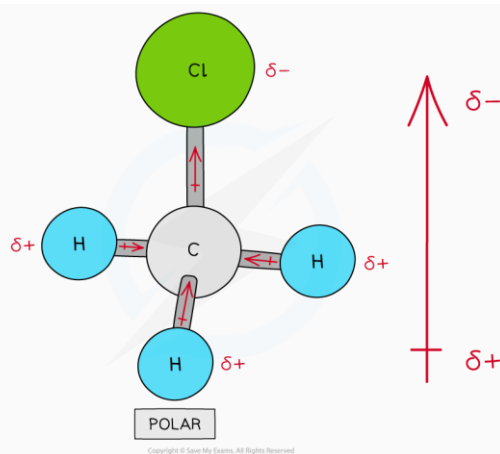
Exam Tip

For Standard Level Chemistry you are only required to know the shape of molecules up to four electron domains.

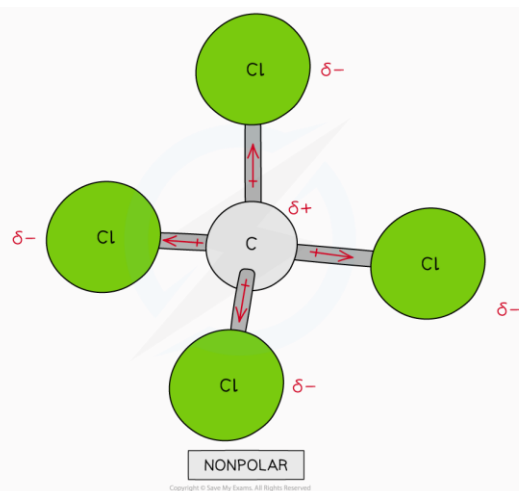
Molecular Polarity

Assigning polarity to molecules

- There is a difference between bond polarity and molecular polarity
- To determine whether a molecule is polar, the following things have to be taken into consideration:
 - The polarity of each bond
 - How the bonds are arranged in the molecule
- Some molecules have polar bonds but are overall not polar because the polar bonds in the molecule are arranged in such way that the individual dipole moments cancel each other out

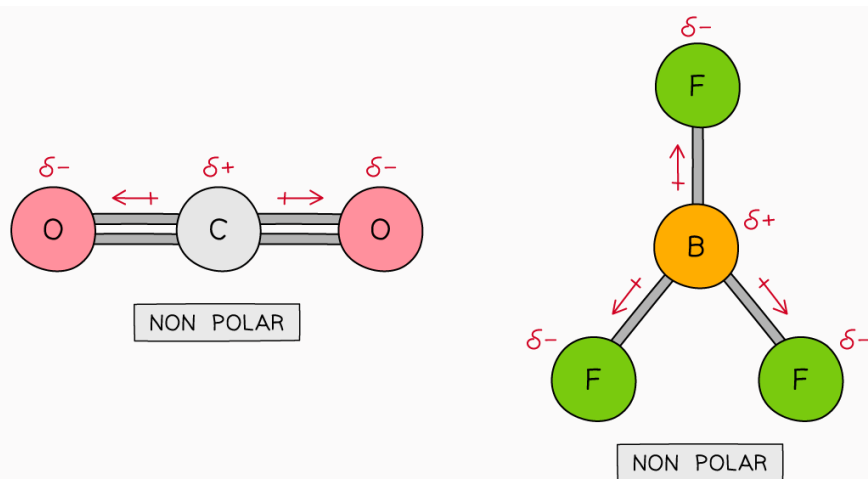


There are four polar covalent bonds in CH_3Cl which do not cancel each other out causing CH_3Cl to be a polar molecule; the overall dipole is towards the electronegative chlorine atom



Though CCl_4 has four polar covalent bonds, the individual dipole moments cancel each other out causing CCl_4 to be a nonpolar molecule

- Further examples of molecules with no net dipole:



Carbon dioxide and boron trifluoride have polar bonds but no net dipole

- Try your hand at this polarity question:

Worked example

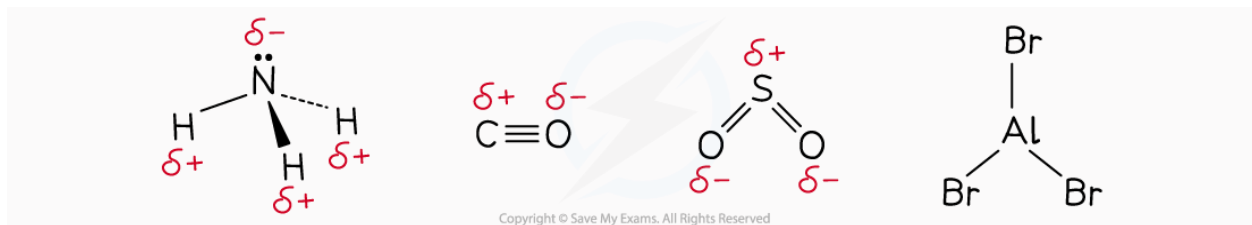
Which molecule is non-polar?

- A. NH_3
- B. CO
- C. SO_2
- D. AlBr_3

Answer:

The correct option is D.

- The shapes and polarity of the molecules are as follows:



Although the Al-Br bonds are polar, the trigonal planar molecule is symmetrical so the dipoles cancel out leaving a non-polar molecule

Exam Tip

One of the clues about molecular polarity is to look at the symmetry of the molecule

Molecules which are symmetrical are unlikely to be polar

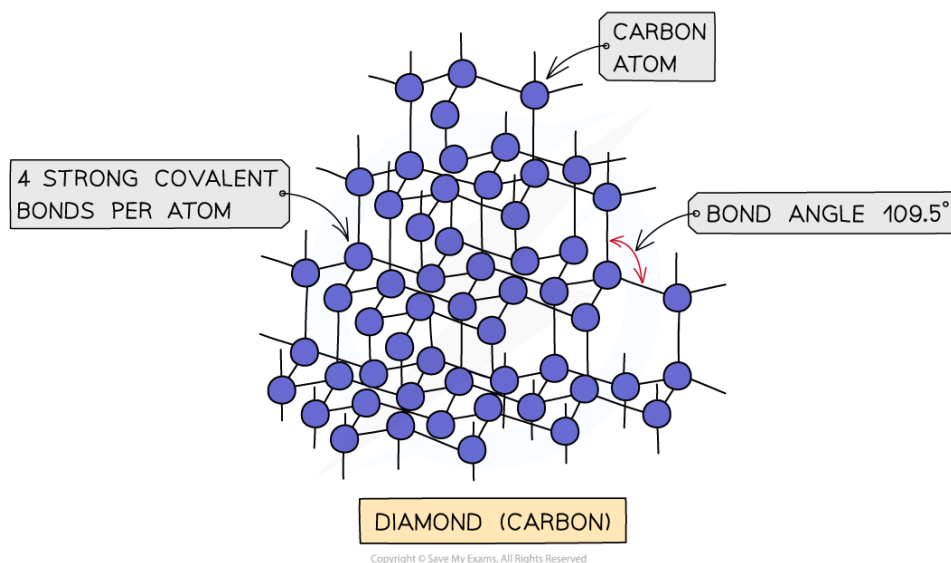
Giant Covalent Structures

Covalent lattices

- Covalent bonds are bonds between nonmetals in which electrons are shared between the atoms
- In some cases, it is not possible to satisfy the bonding capacity of a substance in the form of a molecule; the bonds between atoms continue indefinitely, and a large lattice is formed. There are no individual molecules and covalent bonding exists between all adjacent atoms
- Such substances are called giant covalent substances, and the most important examples are C and SiO₂
- Graphite, diamond, buckminsterfullerene and graphene are allotropes of carbon

Diamond

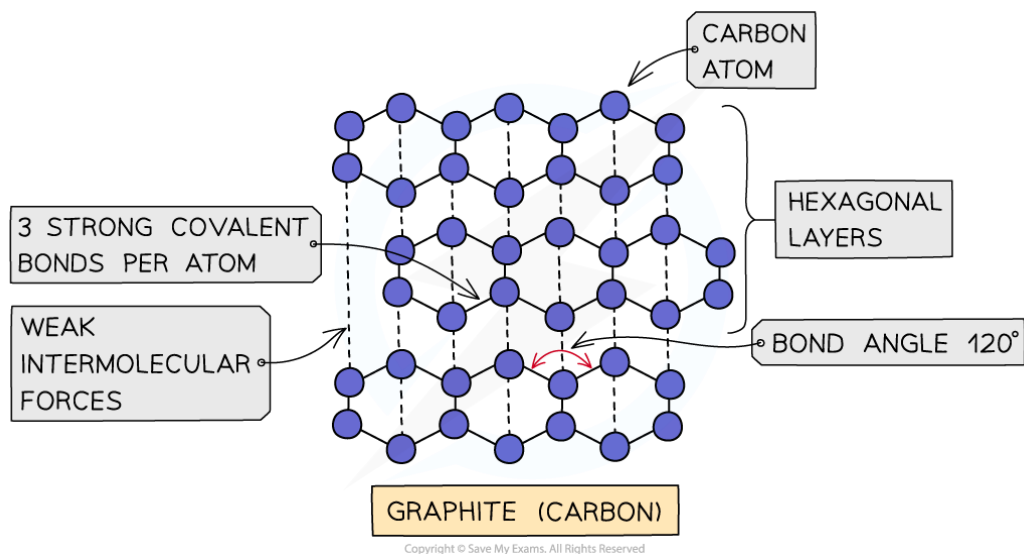
- Diamond is a giant lattice of carbon atoms
- Each carbon is covalently bonded to four others in a tetrahedral arrangement with a bond angle of 109.5°
- The result is a giant lattice with strong bonds in all directions
- Diamond is the hardest substance known
 - For this reason it is used in drills and glass-cutting tools



The structure of diamond

Graphite

- In graphite, each carbon atom is bonded to three others in a layered structure
- The layers are made of hexagons with a bond angle of 120°
- The spare electron is delocalised and occupies the space in between the layers
- All atoms in the same layer are held together by strong covalent bonds, and the different layers are held together by weak intermolecular forces

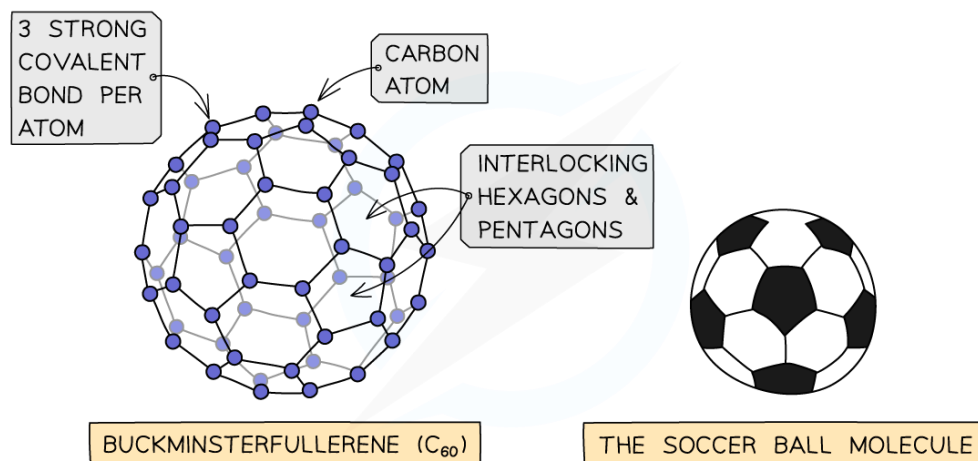


The structure of graphite

Buckminsterfullerene

- Buckminsterfullerene is one type of fullerene, named after Buckminster Fuller, the American architect who designed domes like the Epcot Centre in Florida

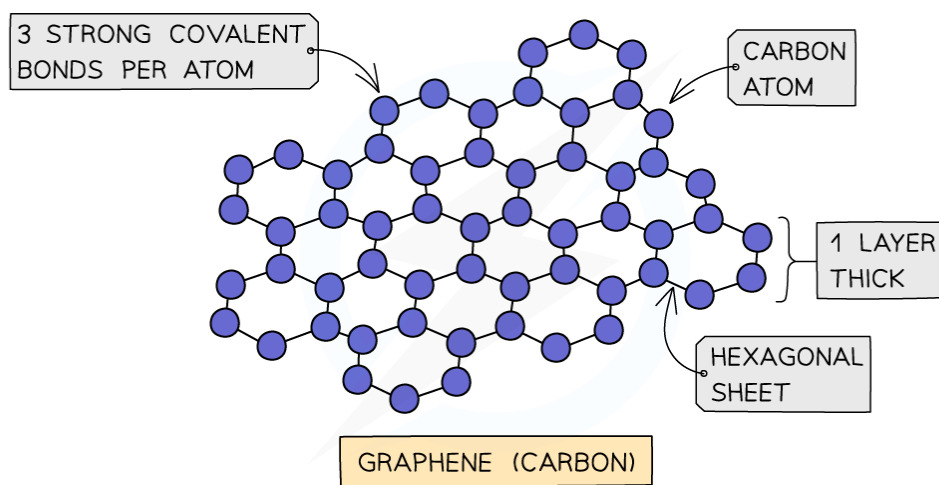
- It contains 60 carbon atoms, each of which is bonded to three others by single covalent bonds
- The fourth electron is delocalised so the electrons can migrate throughout the structure making the buckyball a semi-conductor
- It has exactly the same shape as a soccer ball, hence the nickname the football molecule



The structure of buckminsterfullerene

Graphene

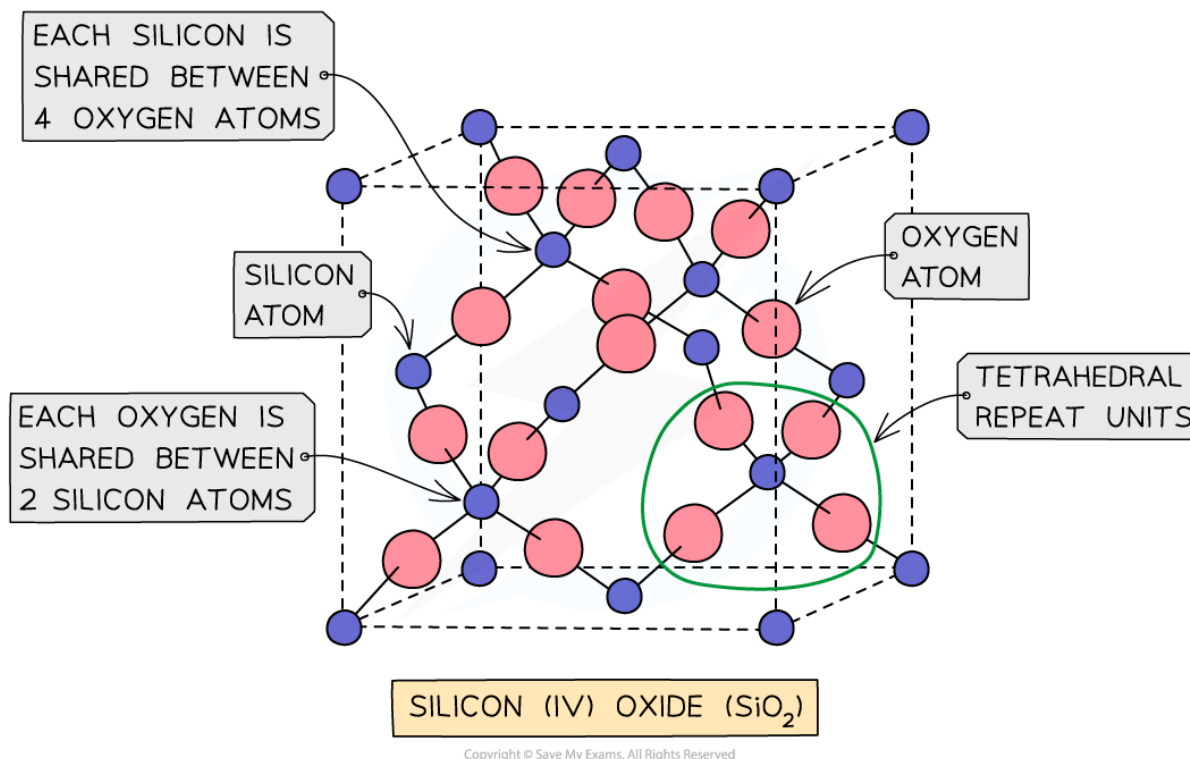
- Some substances contain an infinite lattice of covalently bonded atoms in two dimensions only to form layers. Graphene is an example
- Graphene is made of a single layer of carbon atoms that are bonded together in a repeating pattern of hexagons
- Graphene is one million times thinner than paper; so thin that it is actually considered two dimensional



The structure of graphene

Silicon(IV)oxide

- Silicon(IV)oxide is also known as silicon dioxide, but you will be more familiar with it as the white stuff on beaches!
- Silicon(IV)oxide adopts the same structure as diamond - a giant structure made of tetrahedral units all bonded by strong covalent bonds
- Each silicon is shared by four oxygens and each oxygen is shared by two silicons
- This gives an empirical formula of SiO_2



The structure of silicon dioxide

Properties of Giant Structures

- Different types of structure and bonding have different effects on the physical properties of substances such as their melting and boiling points, electrical conductivity and solubility

Covalent bonding & giant covalent lattice structures

- Giant covalent lattices have very high melting and boiling points
 - These compounds have a large number of covalent bonds linking the whole structure
 - A lot of energy is required to break the lattice
- The compounds can be hard or soft
 - Graphite is soft as the forces between the carbon layers are weak
 - Diamond and silicon(IV) oxide are hard as it is difficult to break their 3D network of strong covalent bonds

- Graphene is strong, flexible and transparent which it makes it potentially a very useful material
- Most compounds are insoluble with water
- Most compounds do not conduct electricity however some do
 - Graphite has delocalised electrons between the carbon layers which can move along the layers when a voltage is applied
 - Graphene is an excellent conductors of electricity due to the delocalised electrons
 - Buckminsterfullerene is a semi-conductor
 - Diamond and silicon(IV) oxide do not conduct electricity as all four outer electrons on every carbon atom is involved in a covalent bond so there are no free electrons available

Characteristics of Giant Covalent Structures Table

	Diamond	Graphite	Buckminster-fullerene	Graphene	Silicon dioxide
Melting and boiling point	Very high	Very high	Low	Very high	Very high
Electrical Conductivity	Non-conductor	Good	Semi-conductor	Very good conductor	Non-conductor
Appearance	Transparent crystals	Grey-black solid	Yellow solid	Transparent sheets	Transparent crystals
Special Characteristics	Hardest known naturally occurring substance	Soft and slippery	Very light and strong	Very strong and flexible; 100 times stronger than steel	Piezoelectric - produces electric charge from mechanical stress

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Exam Tip

Although buckminsterfullerene is included in this section it is not classified as a giant structure as it has a fixed formula, C₆₀

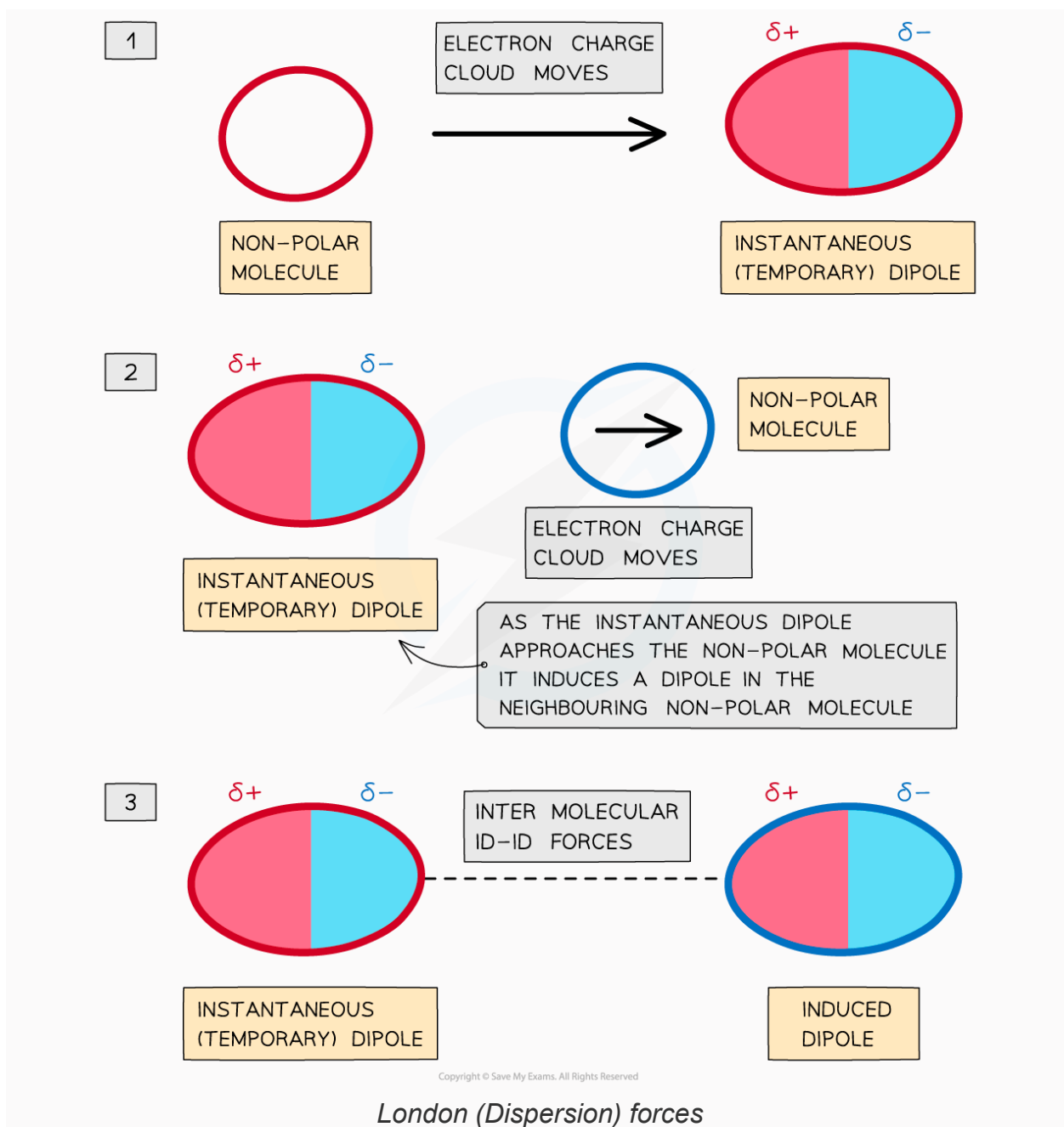
Intermolecular Forces

- There are no covalent bonds between molecules in molecular covalent compounds. There are, however, forces of attraction between these molecules, and it is these which must be overcome when the substance is melted and boiled
- These forces are known as intermolecular forces
- There are three main types of intermolecular forces:
 - London(dispersion) forces

- Dipole-dipole attraction
- Hydrogen bonding

London (dispersion) forces

- The electrons in atoms are not static; they are in a state of constant motion
 - It is therefore likely that at any given time the distribution of electrons will not be exactly symmetrical - there is likely to be a slight surplus of electrons on one side of the atoms



This is known as a temporary dipole

- It lasts for a very short time as the electrons are constantly moving

○ Temporary dipoles are constantly appearing and disappearing
Consider now an adjacent atom. The electrons on this atom are repelled by the negative part of the dipole and attracted to the positive part and move accordingly

This is a temporary induced dipole

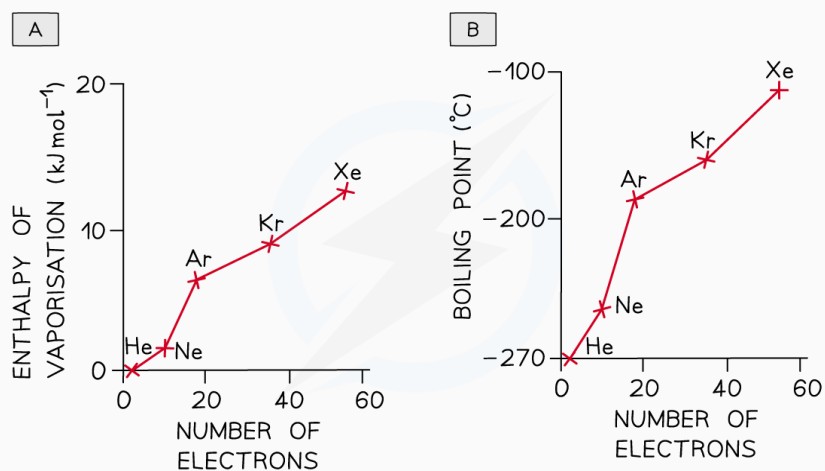
- There is a resulting attraction between the two atoms, and this known as London (dispersion) forces, after the German chemist, Fritz London
- London (dispersion) forces are present between all atoms and molecules, although they can be very weak
- They are the reason all compounds can be liquefied and solidified
- London (dispersion) forces tend to have strengths between 1 kJmol^{-1} and 50 kJmol^{-1} .

The strength of the London(dispersion) forces in between molecules depends on two factors:

- the number of electrons in the molecule
- Surface area of the molecules

Number of electrons

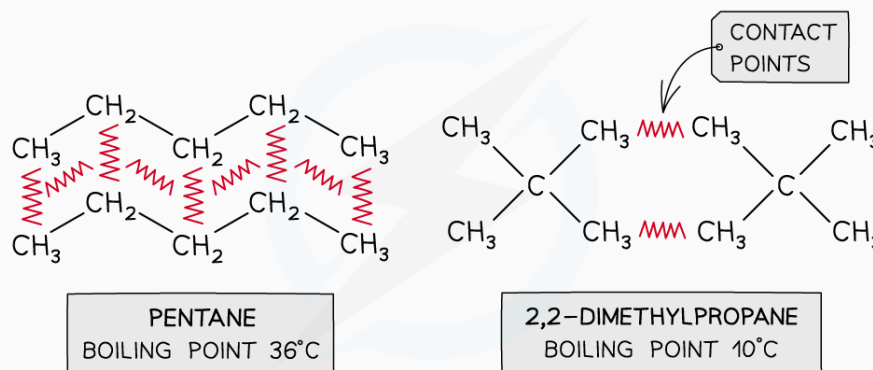
- The greater the number of electrons in a molecule, the greater the likelihood of a distortion and thus the greater the frequency and magnitude of the temporary dipoles
- The dispersion forces between the molecules are stronger and the melting and boiling points are larger
- The enthalpies of vaporisation and boiling points of the noble gases illustrate this factor:



As the number of electrons increases more energy is needed to overcome the forces of attraction between the noble gases atoms

Surface area

- The larger the surface area of a molecule, the more contact it will have with adjacent molecules
- The greater its ability to induce a dipole in an adjacent molecule, the greater the London (dispersion) forces and the higher the melting and boiling points
- This point can be illustrated by comparing different isomers containing the same number of electrons:

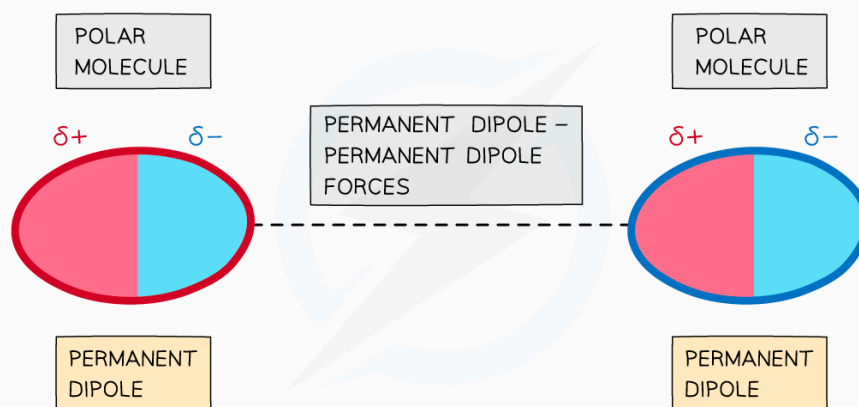


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Boiling points of molecules with the same numbers of electrons but different surface areas

Dipole-dipole attractions

- Temporary dipoles exist in all molecules, but in some molecules there is also a permanent dipole
- In addition to the London (dispersion) forces caused by temporary dipoles, molecules with permanent dipoles are also attracted to each other by permanent dipole-dipole bonding



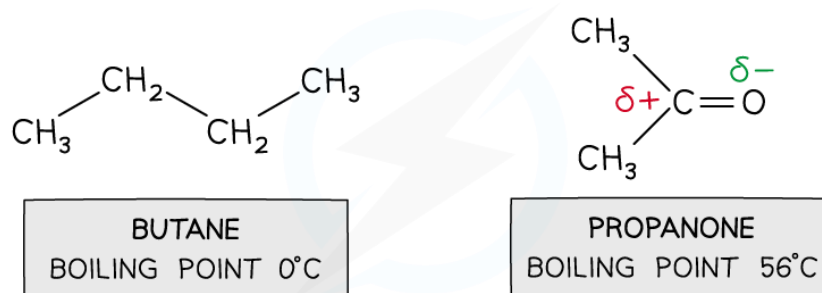
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The delta negative end of one polar molecule will be attracted onwards the delta positive end of a neighbouring polar molecule

- This is an attraction between a permanent dipole on one molecule and a permanent dipole on another.
- Dipole-dipole bonding usually results in the boiling points of the compounds being slightly higher than expected from temporary dipoles alone
 - it slightly increases the strength of the intermolecular attractions
- The effect of dipole-dipole bonding can be seen by comparing the melting and boiling points of different substances which should have London (dispersion) forces of similar strength

Comparing butane and propanone

- For small molecules with the same number of electrons, dipole-dipole attractions are stronger than dispersion forces
 - Butane and propanone have the same number of electrons
 - Butane is a nonpolar molecule and will have only dispersion forces
 - Propanone is a polar molecule and will have dipole-dipole attractions and dispersion forces
 - Therefore, more energy is required to break the intermolecular forces between propanone molecules than between butane molecules
 - The result is that propanone has a higher boiling point than butane

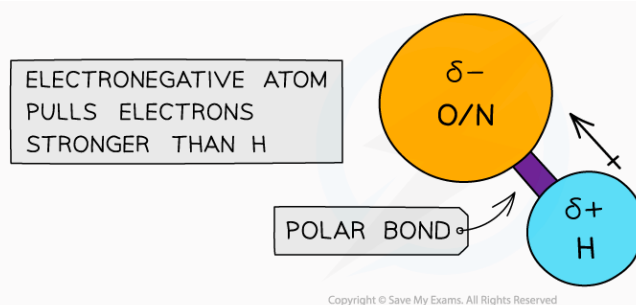


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Comparing substances with permanent and temporary dipoles in smaller molecules with an equal number of electrons

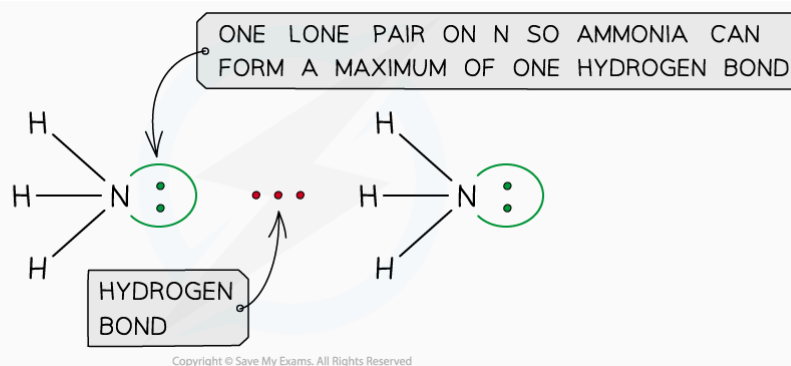
Hydrogen bonding

- Hydrogen bonding is the strongest type of intermolecular force
 - Hydrogen bonding is a special type of permanent dipole – permanent dipole bonding
- For hydrogen bonding to take place the following is needed:
 - A species which has an O or N or F (very electronegative) atom with an available lone pair of electrons
 - A hydrogen attached to the O, N or F
- When hydrogen is covalently bonded to an electronegative atom, such as O or N, the bond becomes very highly polarised
- The H becomes so $\delta+$ charged that it can form a bond with the lone pair of an O or N atom in another molecule

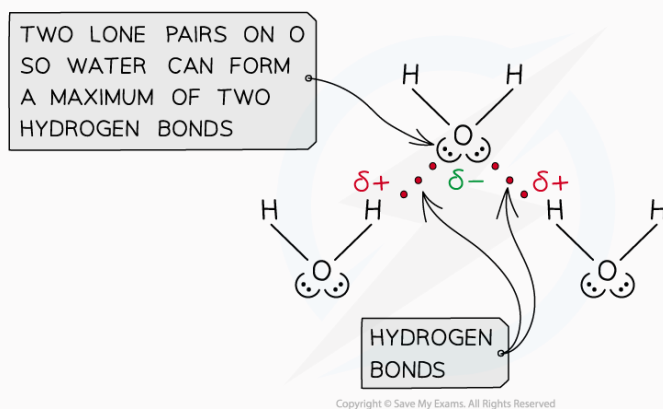


The electronegative atoms O or N have a stronger pull on the electrons in the covalent bond with hydrogen, causing the bond to become polarised

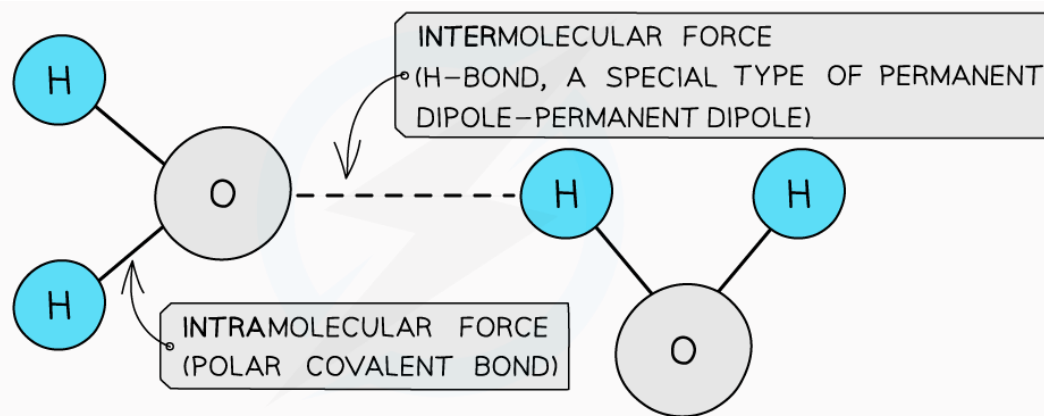
- Hydrogen bonds are represented by dots or dashes between H and the N/O/F element
- The number of hydrogen bonds depends on:
 - The number of hydrogen atoms attached to O or N in the molecule
 - The number of lone pairs on the O or N



Ammonia can form a maximum of one hydrogen bond per molecule



Water can form a maximum of two hydrogen bonds per molecule



The polar covalent bonds between O and H atoms are intramolecular forces and the permanent dipole – permanent dipole forces between the molecules are intermolecular forces as they are a type of van der Waals' force

Exam Tip

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 species. You may be wondering about the term “van der Waals” forces: it is an inclusive term and refers to dipole–dipole, dipole-induced dipole and London (dispersion) forces

Deducing Intermolecular Forces

- In order to deduce the types of intermolecular forces present in molecules you need information about the structure and chemical formula of the molecules
- The chemical formula will tell you about the presence of electronegative elements present in the molecule
 - Any potential polar bonds can be identified
 - If N, O or F is present then hydrogen bonds are potentially possible
- The structure and symmetry of the molecule will enable you to determine if the molecule is polar following the principles laid out in 4.1.10 Molecular Polarity

Worked example

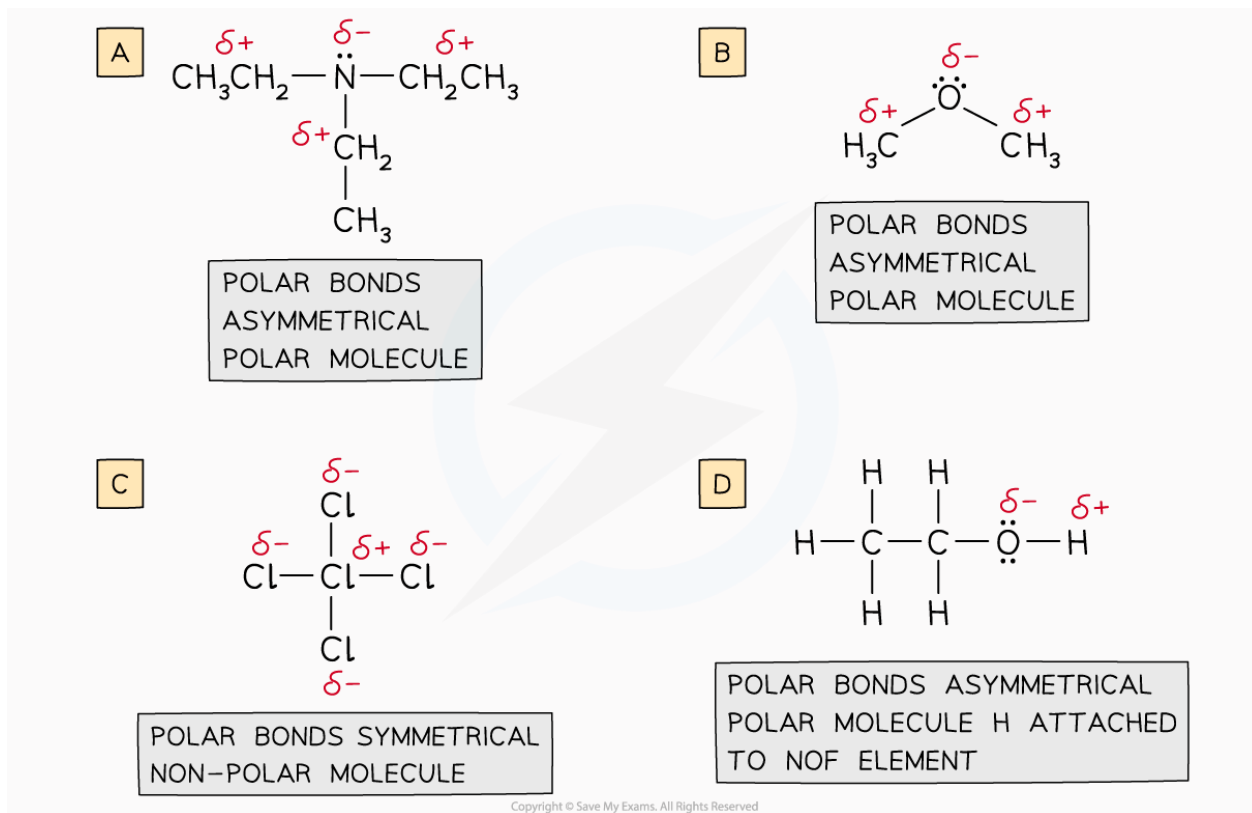
Which of the compounds below can form intermolecular hydrogen bonds in the liquid state?

- A. $(\text{CH}_3\text{CH}_2)_3\text{N}$
- B. CH_3OCH_3
- C. CCl_4
- D. $\text{C}_2\text{H}_5\text{OH}$

Answer:

The correct option is D.

- Draw the displayed structures of the molecules:



- Molecules A and B form dipole-dipole attractions as they are polar molecules
- Molecule C forms dispersion forces as the dipoles cancel out so there is no overall polarity
- Molecule D is the only one capable of forming hydrogen bonds

Exam Tip

Sometimes a question will ask you to name all the IMFs present in molecules and students frequently forget to include dispersion forces which are present in all molecules, since everything contains electrons!

Properties of Covalent Compounds

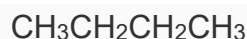
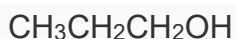
- The physical properties of molecular covalent compounds are largely influenced by their intermolecular forces
- If you know the type of intermolecular forces present you can predict the physical properties like melting and boiling point, solubility, and conductivity

Melting and boiling point

- When covalent molecular substances change state you are overcoming the intermolecular forces
- The stronger the forces the more energy need to break the attraction
- Intermolecular forces are much weaker than covalent bonds, so many covalent substances are liquid or gases at room temperature
- Substance with a low melting and boiling point are said to be very volatile
- The strength of the intermolecular forces increases with
 - the size of the molecule
 - the increase in the polarity of the molecule
 - Drawing the structure of the molecule helps identify and rank molecules according to boiling point as the following example shows:

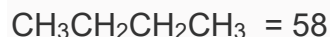
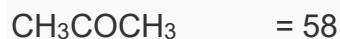
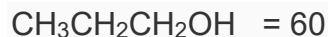
Worked example

Place these three molecules in the correct order from lowest to highest boiling point and explain your reasoning:



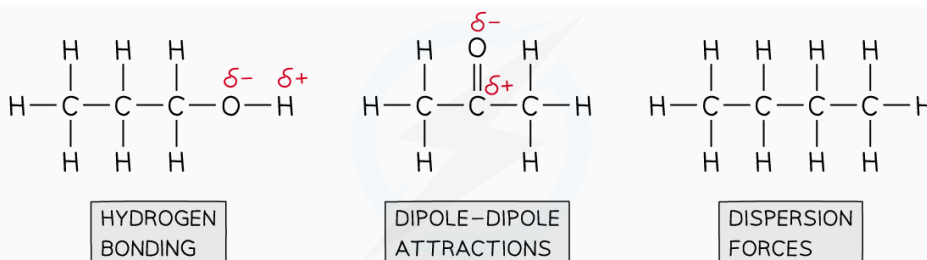
Answer:

Step 1: The first thing to do is find the approximate relative molecular mass:



This tells you the molecules are approximately the same size so the dispersion forces will be similar

Step 2: Draw the structures of the molecules and identify the intermolecular forces present

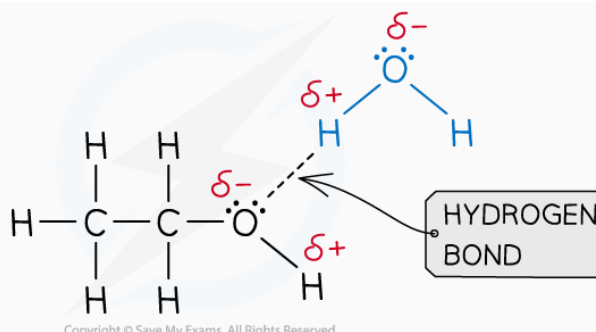


So, the order of boiling from lowest to highest is:



Solubility

- The general principle is that 'like dissolves like' so non-polar substances mostly dissolve in non-polar solvents, like hydrocarbons and they form dispersion forces between the solvent and the solute
- Polar covalent substances generally dissolve in polar solvents as a result of dipole-dipole interactions or the formation of hydrogen bonds between the solute and the solvent
- A good example of this is seen in organic molecules such as alcohols and water:



Hydrogen bonds form between ethanol and water

- As covalent molecules become larger their solubility can decrease as the polar part of the molecule is only a smaller part of the overall structure
 - This effect is seen in alcohols for example where ethanol, C_2H_5OH , is readily soluble but hexanol, $C_6H_{13}OH$, is not
- Polar covalent substances are unable to dissolve well in non-polar solvents as their dipole-dipole attractions are unable to interact well with the solvent
- Giant covalent substances generally don't dissolve in any solvents as the energy needed to overcome the strong covalent bonds in the lattice structures is too great

Conductivity

- As covalent substances do not contain any freely moving charged particles they are unable to conduct electricity in either the solid or liquid state
- However, under certain conditions some polar covalent molecules can ionise and will conduct electricity
- Some giant covalent structures are capable of conducting electricity due to delocalised electrons, as seen in Section 4.1.11 Giant Covalent Structures, but they are exceptions to the general rule

Comparing the Properties of Covalent Compounds Table

	Non-polar covalent substances	Polar covalent substances	Giant covalent substances	Ionic substances
Melting and boiling point	Low	Low	Very high	Very high
Volatility	Highest	High	Low	Low
Solubility in polar solvents	Insoluble	Some solubility depending on molecular size	Insoluble	Soluble
Solubility in non-polar solvents	Soluble	Some solubility depending on molecular size	None	Insoluble
Electrical conductivity	None	None	None – except graphite, graphene	Only when molten or aqueous

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Worked example

Compound X has the following properties:

Melting point	Electrical conductivity	
	1450 °C	solid
poor		poor

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What is the most probable structure of X?

- A. Network covalent
- B. Polar covalent molecule
- C. Ionic lattice
- D. Metallic lattice

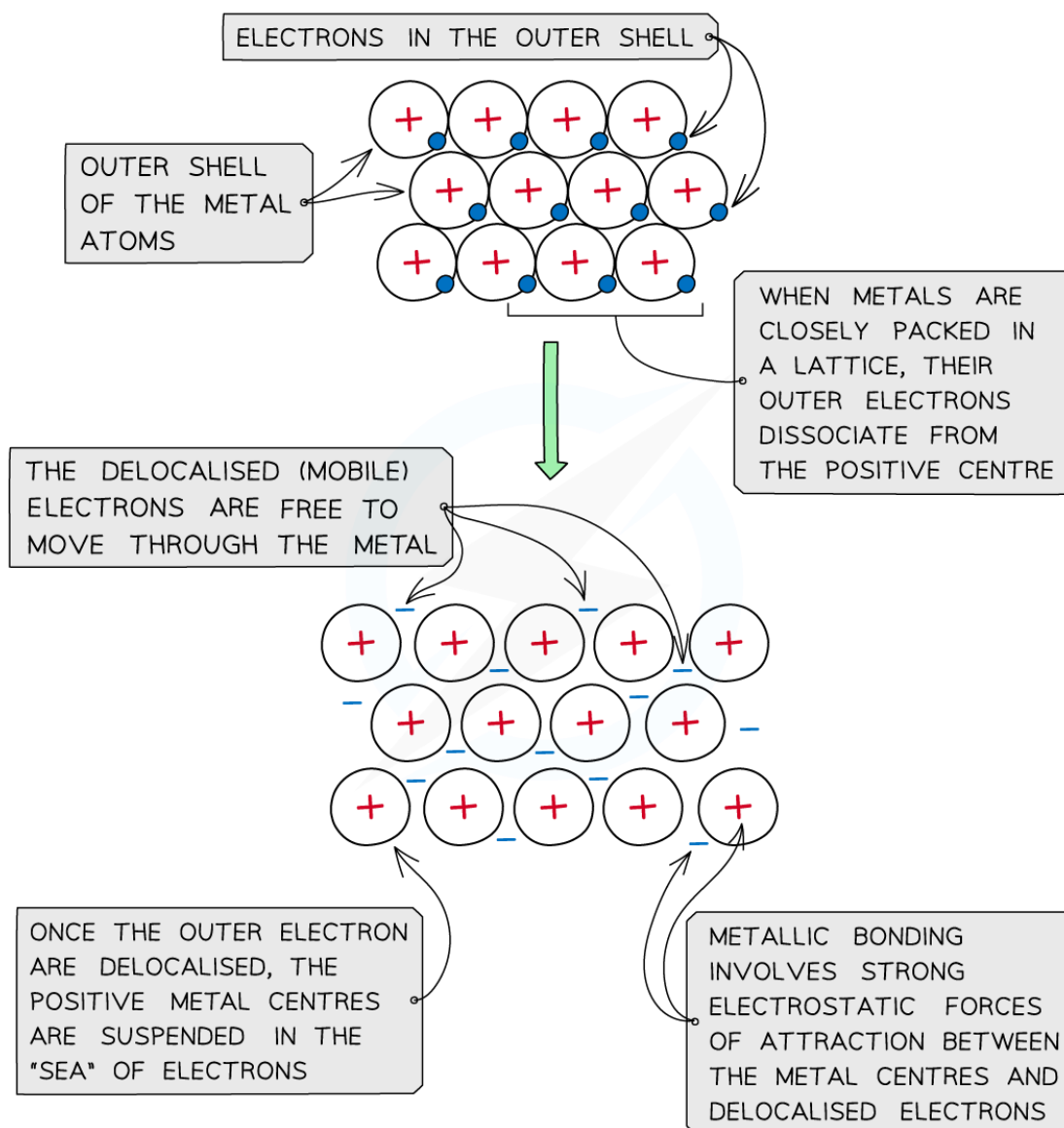
Answer:

The correct option is A

- A high melting point is characteristic of a giant structure, which could be metallic, ionic or covalent
- The poor conductivity as a liquid and solid would match a giant covalent or network covalent structure

Metallic Bonding

- Metal atoms are tightly packed together in lattice structures
- When the metal atoms are in lattice structures, the electrons in their outer shells are free to move throughout the structure
- The free-moving electrons are called 'delocalised' electrons and they are not bound to their atom
- When the electrons are delocalised, the metal atoms become positively charged
- The positive charges repel each other and keep the neatly arranged lattice in place
- There are very strong electrostatic forces between the positive metal centres and the 'sea' of delocalised electrons

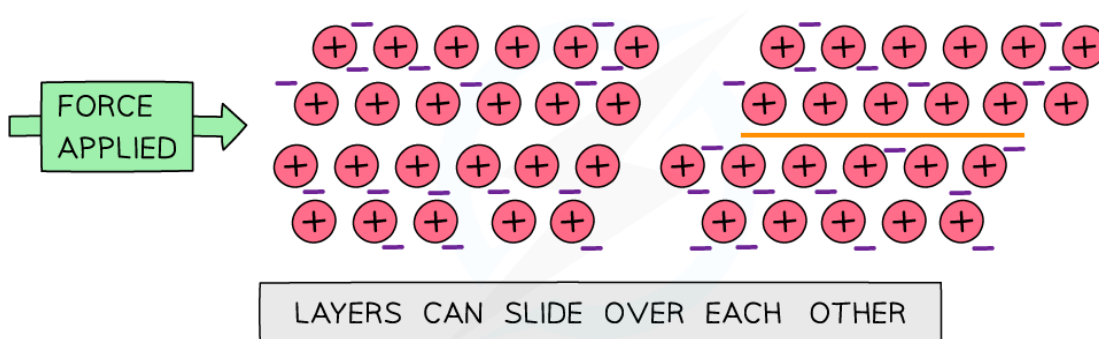


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The positive metal centres are suspended in a 'sea' of delocalised electrons

Properties of Metals

- Metallic compounds are malleable
 - When a force is applied, the metal layers can slide
 - The attractive forces between the metal ions and electrons act in all directions
 - So when the layers slide, the metallic bonds are re-formed
 - The lattice is not broken and has changed shape



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- Metallic compounds are strong and hard
 - Due to the strong attractive forces between the metal ions and delocalised electrons
- Metals have high melting and boiling points

Conductivity

- Metals can conduct electricity when in the solid or liquid state
 - As both in the solid and liquid state there are mobile electrons which can freely move around and conduct electricity
- Since the bonding in metals is non-directional, it does not really matter how the cations are oriented relative to each other
- The metal cations can be moved around and there will still be delocalised electrons available to hold the cations together

Metallic Bonding Summary Table

Property	Explanation
High melting and boiling points	<p>Positive metal atom centres and the delocalised electrons in a metallic lattice have strong electrostatic forces between them</p> <p>Therefore, high energy is needed to overcome the strong forces of attraction</p> <p>As the number of mobile charges increase from left to right of the Periodic Table, the melting and boiling points increase as the electrostatic forces become stronger</p>
Electrical Conductivity	<p>When a potential difference is applied to the metallic lattice, the delocalised electrons repel away from the negative terminal and move towards the positive terminal</p> <p>As the number of valence electrons increases across the Period, the number of delocalised charges also increases:</p> <ul style="list-style-type: none"> ◦ Sodium = 1 valence electron ◦ Magnesium = 2 valence electrons ◦ Aluminium = 3 valence electrons ◦ Etc. <p>Therefore the ability to conduct electricity also increases</p>

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Strength of Metallic Bonds

Not all metallic bonds are equal. There are several factors that affect the strength of a metallic bond:

The charge on the metal ion

- The greater the charge on the metal ion, the greater the number of electrons in the sea of delocalised electrons and the greater the charge difference between the ions and the electrons
- A greater charge difference leads to a stronger electrostatic attraction, and therefore a stronger metallic bond
- This effect can be seen in melting point data across a period, as the charge on the metal ion increases without a significant change in ionic radius:

Group	1	2	3 (13)
Metal	Sodium	Magnesium	Aluminium
Melting point / K	371	923	933

The radius of the metal ion

- Metal ions with smaller ionic radii exert a greater attraction on the sea of delocalised electrons
- This greater attraction means a stronger metallic bond, requiring more energy to break
- This can be seen in data from metals, descending a group, where the charge on the ion remains constant but the ionic radius increases:

Period	3	4	5
Metal	Sodium	Potassium	Rubidium
Melting point / K	371	336	312

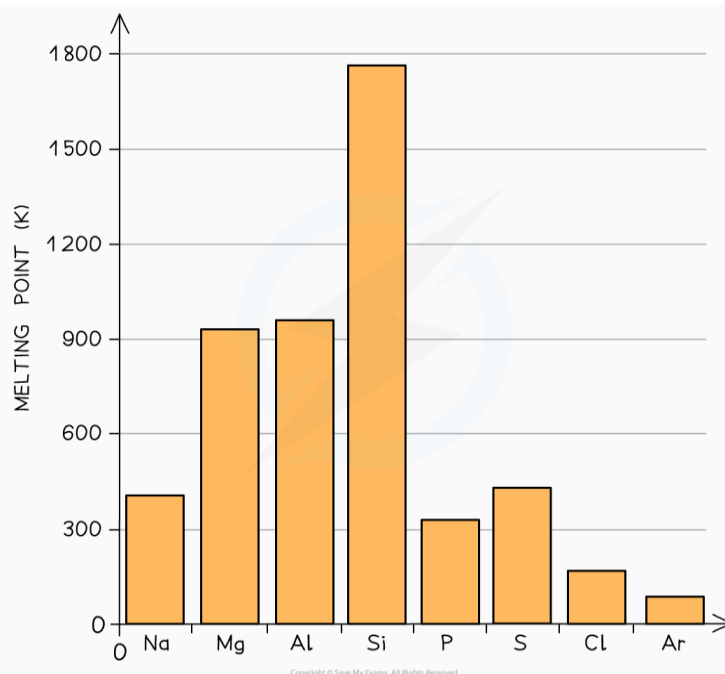
Trends in Melting Points of Metals

- The strength of electrostatic attraction can be increased by:
 - Increasing the number of delocalised electrons per metal atom
 - Increasing the positive charges on the metal centres in the lattice
 - Decreasing the size of the metal ions
- These factors can be seen in the trends across a period and down a group

Melting points of metals across a period

- If you compare the electron configuration of sodium, magnesium and aluminium you can see the number of valence electrons increases
 - Na = $1s^2 2s^2 2p^6 3s^1$
 - Mg = $1s^2 2s^2 2p^6 3s^2$
 - Al = $1s^2 2s^2 2p^6 3s^2 3p^1$

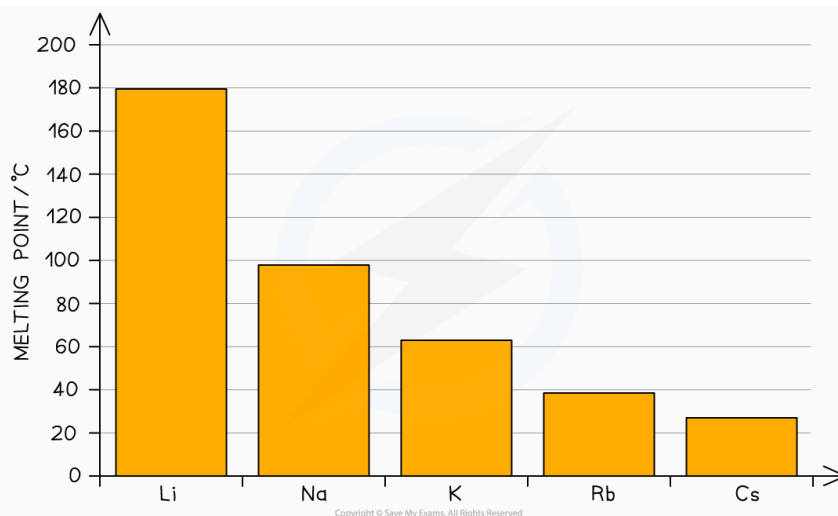
- Aluminium ions are also a smaller size than magnesium ions or sodium ions and these two factors lead to stronger metallic bonding which can be seen in the melting points
- The stronger the metallic bonding, the more energy is needed to break the metallic lattice and so the higher the melting point
- As we go across period 3 we can see the effect of stronger metallic bonding on the metals
 - Remember only the first three elements have metallic bonding in this graph



Melting points as you go across a period. The metallic bonding gets stronger from Na to Al

Melting points of metals down a group

- As you go down the group the size of the cation increases so this decreases the attraction between the valence electrons and the metallic lattice, leading to a reduction of the melting point



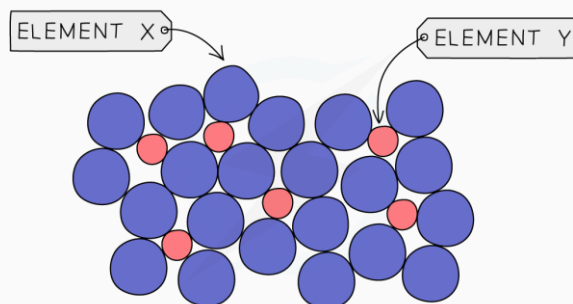
Melting points as you go down a group of metals. The metallic bonding gets weaker from Li to Cs

Exam Tip

You see from the graph that the melting point of aluminium is not that much higher than magnesium. It is a reminder to us that these are trends and not rules about melting points and sometimes there are other factors which can result in subtle differences from what was expected. One factor here is the metal packing structure - this can also influence the melting point, but it is beyond what is required in the IB Chemistry syllabus. You just need to learn and explain the broad trends

Alloys

- Alloys are mixtures of metals, where the metals are mixed together physically but are not chemically combined
- They can also be made from metals mixed with nonmetals such as carbon
- Ions of the different metals are spread throughout the lattice and are bound together by the delocalized electrons
- It is possible to form alloys because of the non-directional nature of the metallic bonds



In a metallic lattice the regular structure of metal cations (shown by Element Y) is disrupted by the presence of another element (Element X)

- Alloys have distinct properties due to the different packing of the cations in the lattice
- Alloys often have properties that can be very different to the metals they contain, for example they can have greater strength, hardness or resistance to corrosion or extreme temperatures
- Alloys contain atoms of different sizes, which distorts the regular arrangements of cations
- This makes it more difficult for the layers to slide over each other, so they are usually much harder than the pure metal
- Below is a table of some common alloys and their uses:

Common Alloys and their Uses Table

Alloy	Elements present	Properties	Uses
Brass	copper and zinc	strong and resistant to corrosion	door handles, hinges, musical instruments
Steel	iron, carbon and other elements like chromium, vanadium, and molybdenum	very strong	construction, bridges, cars
Solder	lead and tin	low melting point	joining metals in electrical circuits and jewellery
Bronze	copper and tin	hard and strong resistant to corrosion	medals, sculptures, ship fittings

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Exam Tip

You don't need to learn the specific alloys, but you should be able to use examples you know to explain why alloys have the properties they do compared to pure metals

Oxides

Oxides across a period

- The acid-base character of the oxides provides evidence of chemical trends in the periodic table
- The broad trend is that oxides change from basic through amphoteric to acidic across a period
- Aluminium oxide is amphoteric which means that it can act both as a base (and react with an acid such as HCl) and an acid (and react with a base such as NaOH)

Acidic & Basic Nature of the Period 3 Oxides

Period 3 oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₄ O ₁₀	SO ₂ , SO ₃
Acid/base nature	Basic	Basic	Amphoteric	Acidic	Acidic	Acidic

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- The acidic and basic nature of the Period 3 elements can be explained by looking at their structure, bonding and the Period 3 elements' electronegativity

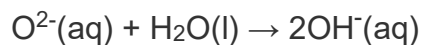
Structure, Bonding & Electronegativity of the Period 3 Elements Table

Period 3 oxide	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₄ O ₁₀	SO ₂ , SO ₃		
Relative melting point	High	High	Very high	Very high	Low	Low		
Chemical bonding	Ionic	Ionic	Ionic (with some degree of covalent)	Covalent	Covalent	Covalent		
Structure	Giant ionic	Giant ionic	Giant ionic	Giant covalent	Simple molecular	Simple molecular		
Element	Na	Mg	Al	Si	P	S	Cl	O
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0	3.5

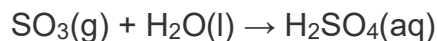
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- The difference in electronegativity between oxygen and Na, Mg and Al is the largest
- Electrons will therefore be transferred to oxygen when forming oxides giving the oxide an ionic bond
- The oxides of Si, P and S will share the electrons with the oxygen to form covalently bonded oxides

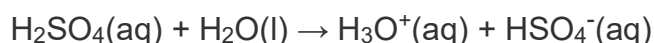
- The oxides of Na and Mg which show purely ionic bonding produce alkaline solutions with water as their oxide ions (O^{2-}) become hydroxide ions (OH^-):



- The oxides of P and S which show purely covalent bonding produce acidic solutions with water because when these oxides react with water, they form an acid which donates H^+ ions to water
 - Eg. SO_3 reacts with water as follows:



- The H_2SO_4 is an acid which will donate an H^+ to water:



Exam Tip

Only examples of general trends across periods and groups are required, but you should be able to link trends in ionization energy, electron affinity and electronegativity with trends in chemical character such as the nature of the oxides and metallic/ non-metallic behaviour.

Oxides reacting with Water

- The pH changes for the reactions of the oxides with water can be explained by reference to the following equations

Reaction of Oxides with Water

Oxide	Chemical equation	pH
Na ₂ O	Na ₂ O(s) + H ₂ O(l) → 2NaOH(aq)	14 (strongly alkaline)
MgO	MgO(s) + H ₂ O(l) → Mg(OH) ₂ (aq)	10 (weakly alkaline)
P ₄ O ₁₀	P ₄ O ₁₀ (s) + 6H ₂ O(l) → 4H ₃ PO ₄ (aq)	2 (strongly acidic)
NO ₂	2NO ₂ (aq) + H ₂ O(l) → HNO ₃ (aq) + HNO ₂ (aq)	1 (strongly acidic)
SO ₂ , SO ₃	SO ₂ (g) + H ₂ O(l) → H ₂ SO ₃ (aq) SO ₃ (g) + H ₂ O(l) → H ₂ SO ₄ (aq)	1 (strongly acidic)

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- The pattern here is that:
 - The metallic oxides form hydroxides when they react with water
 - The non-metallic oxides form oxoacids when they react with water

Exam Tip

You should learn how to construct these equations exactly as they are specifically mentioned in the syllabus

Making Predictions

- The position of an element in the periodic table can be used to predict and explain its metallic and non-metallic behaviour
- This is illustrated by the bonding of the oxides
- Metal and non-metal elements generally form ionic compounds so the elements Na to Al have giant ionic structures
- The oxides become more ionic as you go down the group as the electronegativity decreases
- The oxides become less ionic as you go across a period as the electronegativity increases
- The oxides of non-metals such as S, N and P form molecular covalent compounds
- Sometimes you may be asked to make predictions about oxides that are not specifically mentioned in the syllabus but you should be able to deduce their properties if you understand the patterns outlined above, as the following example shows:

Worked example

Which of these oxides produces the solution with the highest pH when added to water ?

- A. CO_2
- B. SO_3
- C. CaO
- D. Na_2O

Answer:

The correct option is D.

- CO_2 and SO_3 will produce a pH below 7 as they are non-metal oxides. CaO and Na_2O will produce a pH above 7 as they are metal oxides; however the pH is decreasing as you go across a period, so Na_2O will have a higher pH than CaO

Alkali Metals

The group 1 metals

- The group 1 metals are called the alkali metals because they form alkaline solutions with high pH values when reacted with water
- Group 1 metals are lithium, sodium, potassium, rubidium, caesium and francium
- They all end in the electron configuration ns^1

Physical properties of the group 1 metals

- The group 1 metals:
 - Are soft and easy to cut, getting softer and denser as you move down the group
 - Have shiny silvery surfaces when freshly cut
 - Conduct heat and electricity
 - They all have low melting points and low densities and the melting point decreases going down the group as the atomic radius increases and the metallic bonding gets weaker

What would you see when sodium is added to water?

- I. a gas is given off
- II. the temperature of the water increases
- III. a clear, colourless solution is formed

- A. I and II only
- B. I and III only
- C. II and III only
- D. I, II and III

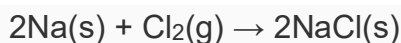
Answer:

The correct option is D.

- Bubbles of hydrogen gas are given off. The sodium melts, so that tells you it is an exothermic reaction and the product, sodium hydroxide, is very soluble so a clear, colourless solution would be formed.

Alkali metals with halogens

- All the alkali metals react vigorously with the halogens in group 17
- The reaction results in an alkali metal halide salt



- The reaction becomes increasingly vigorous going down group 1 because
 - The atoms of each element get larger going down the group
 - This means that the ns_1 electron gets further away from the nucleus and is shielded by more electron shells.
 - The further an electron is from the positive nucleus, the easier it can be lost in reactions

Can you apply what you have just learned? Check out the following example question:

Worked example

Which pair of elements has the most vigorous reaction?

- A. Cs and I
- B. Li and Cl
- C. Cs and F
- D. Li and F

Answer:

The correct option is C.

- You need to choose the lowest element in group 1 and the highest element in group 17 to predict the most vigorous reaction as reactivity increases going down group 1, but decreases going down group 17

Halogens

The halogens

- These are the group 17 non-metals that are poisonous and include fluorine, chlorine, bromine, iodine and astatine
- Halogens are diatomic, meaning they form molecules of two atoms
- All halogens have seven electrons in their outer shell
- They form halide ions by gaining one more electron to complete their outer shells

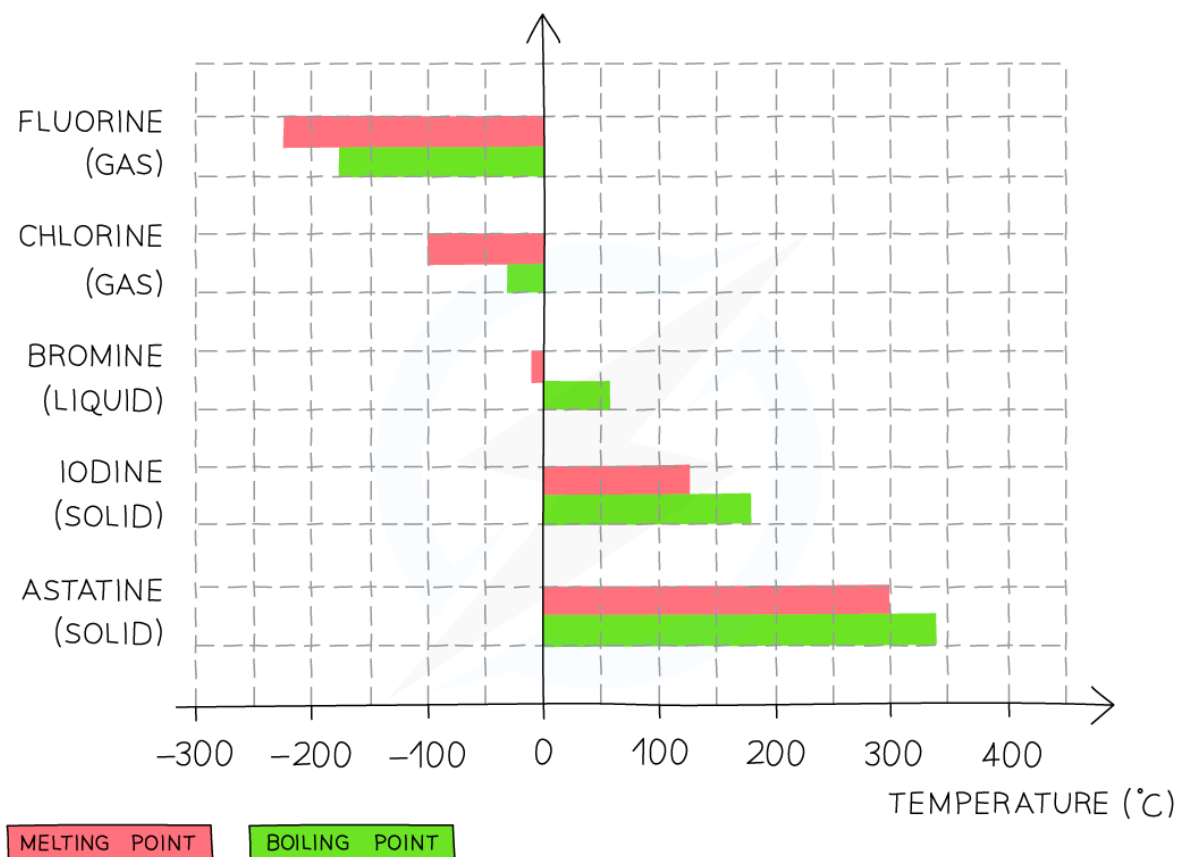
Colours and States at Room Temperature

HALOGEN	PHYSICAL STATE AT ROOM TEMPERATURE	COLOUR	COLOUR IN SOLUTION
FLUORINE	GAS	YELLOW	–
CHLORINE	GAS	PALE GREEN	GREEN-BLUE
BROMINE	LIQUID	RED-BROWN (READILY EVAPORATES TO FORM A BROWN GAS)	ORANGE
IODINE	SOLID	BLACK (SUBLIMES TO FORM A PURPLE GAS)	DARK BROWN

Trends in physical properties of the halogens

Melting point

- The density and melting and boiling points of the halogens increase as you go down the group



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Explaining the trend in reactivity in Group VII

- Reactivity of group 17 non-metals decreases as you go down the group
- The halogens electron configurations all end in ns^2np^5
- Each outer shell contains seven electrons and when they react, they will need to gain one outer electron to get a full outer shell of electrons
- Going down the group, the electron affinity decreases and the atomic radius increases
- As you go down group 17, the number of shells of electrons increases so shielding also increases
- This means that the outer electrons are further from the nucleus so there are weaker electrostatic forces of attraction that attract the extra electron needed
- The electron is attracted less readily, so the lower down the element is in Group 17 the less reactive it is

Reaction of the halogens with halide ions in displacement reactions

- A halogen displacement occurs when a more reactive halogen displaces a less reactive halogen from an aqueous solution of its halide
- The reactivity of group 17 non-metals increases as you move up the group
- Out of the 3 halogens, chlorine, bromine and iodine, chlorine is the most reactive and iodine is the least reactive

Aqueous Solution Colour of Halogens

AQUEOUS SOLUTION	COLOUR
CHLORINE	VERY PALE GREEN, BUT USUALLY APPEARS COLOURLESS AS IT IS VERY DILUTE
BROMINE	ORANGE BUT WILL TURN YELLOW WHEN DILUTED
IODINE	BROWN

Halogen displacement reactions

Chlorine and bromine

- If you add chlorine solution to colourless potassium bromide solution, the solution becomes orange as bromine is formed
- Chlorine is above bromine in group 17 so it is more reactive
- Chlorine will therefore displace bromine from an aqueous solution of a metal bromide



potassium bromide + chlorine → potassium chloride + bromine

Bromine and iodine

- Bromine is above iodine in group 17 so it is more reactive
- Bromine will therefore displace iodine from an aqueous solution of a metal iodide



bromine + sodium iodide → sodium bromide + iodine

Test yourself on halogen displacements:

Worked example

Which of the statements below are correct?

- I. potassium chloride solution will react with fluorine to form chlorine.
- II. sodium chloride solution will react with iodine to form chlorine.
- III. lithium iodide solution will react with bromine to form iodine.

- A. I and II only
- B. I and III only
- C. II and III only
- D. I, II and III

Answer:

The correct option is B.

- Fluorine will displace chlorine as it is higher up in the group and bromine will displace iodine for the same reason.
- Iodine is below chlorine so cannot displace chlorine from sodium chloride