WJEC Chemistry A-Level

C1.4: Bonding

Detailed Notes

English Specification

Ionic Bonding

Ionic bonding occurs between a **metal and a non-metal**. Electrons are **transferred** from the metal to the non-metal to achieve full outer shells.

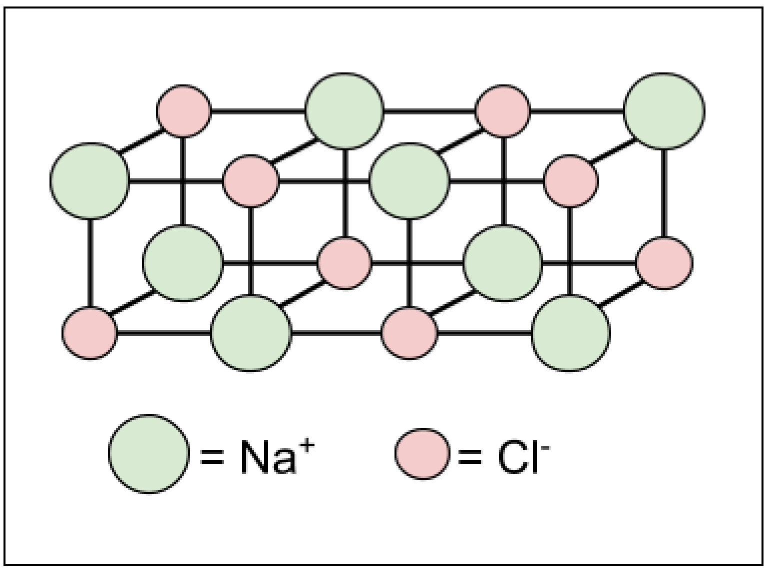
When the electrons are transferred, it creates charged particles called **ions**. Oppositely charged ions attract through **electrostatic forces** to form a **giant ionic crystal lattice**.

Example:

Sodium chloride is an ionic compound formed from Na^+ and Cl^- ions.

Sodium loses an electron to form Na^+ ions and chlorine gains an electron to produce Cl^- ions, each with a full outer electron shell.

These then attract electrostatically to form into an ionic crystal lattice:



Common ionic compound ions include:

Sulfate - SO_4^{2-}
Hydroxide - OH^-
Nitrate - NO_3^-
Carbonate - CO_3^{2-}
Ammonium - NH_4^+

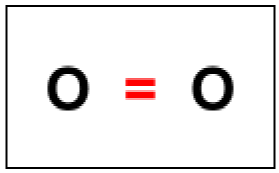
Covalent and Coordinate Bonding

Covalent Bonding

Covalent bonds form between **two non-metals**. Electrons are **shared** between the two outer shells in order to achieve a full outer shell. The electrostatic forces within the molecule must be **balanced** so that there is **no overall attractive or repulsive force**.

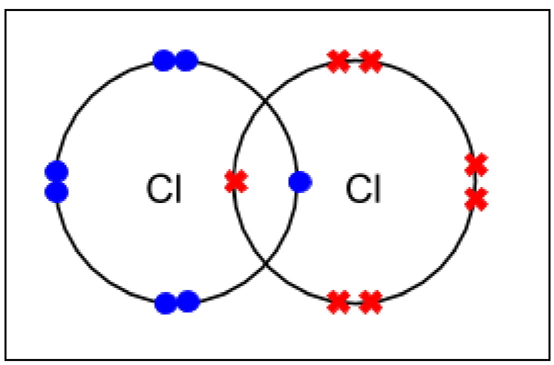
Multiple electron pairs can be shared to produce multiple covalent bonds between atoms. This is seen in molecules of oxygen (O_2) where the two atoms are joined by a **double covalent bond**.

Example:



The shared electron pairs can be represented using **dot and cross diagrams** (below) and a covalent bond shown with a **straight line** when drawing molecule representations (above).

Example:

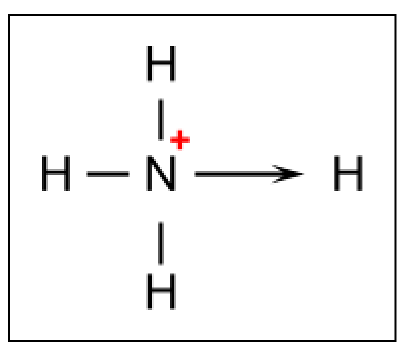


Coordinate Bonding

Coordinate (or dative) bonds form when both of the electrons in the shared pair are supplied from the **same atom**. It is indicated using an **arrow** from the lone electron pair.

Example:

Ammonia (NH_3) has a lone electron pair that can form a dative bond with a H^+ ion to produce an ammonium ion (NH_4^+).



Once a dative bond has formed, it is treated as a **standard covalent bond** as it reacts in exactly the same way.

Bond Polarity and Character

The negative charge around a covalent bond is **rarely spread evenly** around the orbitals of the bonded atoms. Therefore, **charge varies** around a molecule meaning a covalent molecule can have **ionic characteristics**.

Electronegativity

Every atom has electronegativity, which is defined as:

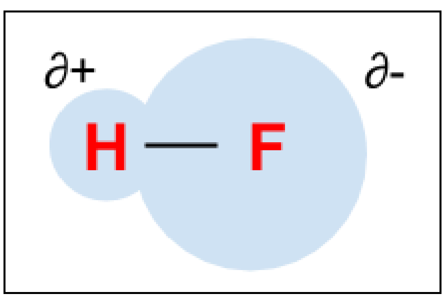
The power of an atom to attract negative charge towards itself within a covalent bond.

This ‘power’ is different for every atom depending on its **size and nuclear charge**. Electronegativity **increases along a period,** as atomic radius decreases. Electronegativity **decreases down a group,** as shielding increases.

Permanent Dipole

If two atoms that are covalently bonded have different electronegativities, a **polar bond** forms. The more electronegative atom draws more of the **negative charge towards itself** and away from the other atom, producing a **∂- region** and a **∂+ region**. This is a permanent dipole.

Example:

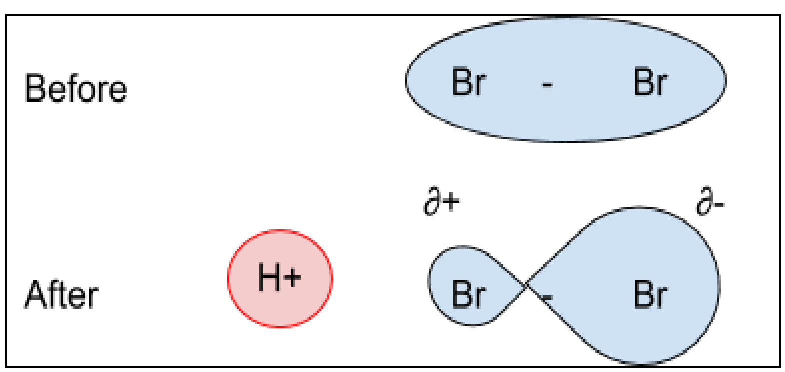


Hydrogen fluoride is a polar molecule as fluorine is a lot more electronegative than hydrogen so electrons are drawn to the fluorine atom.

Polar molecules with a permanent dipole can align to form a **lattice of molecules** similar to an ionic lattice.

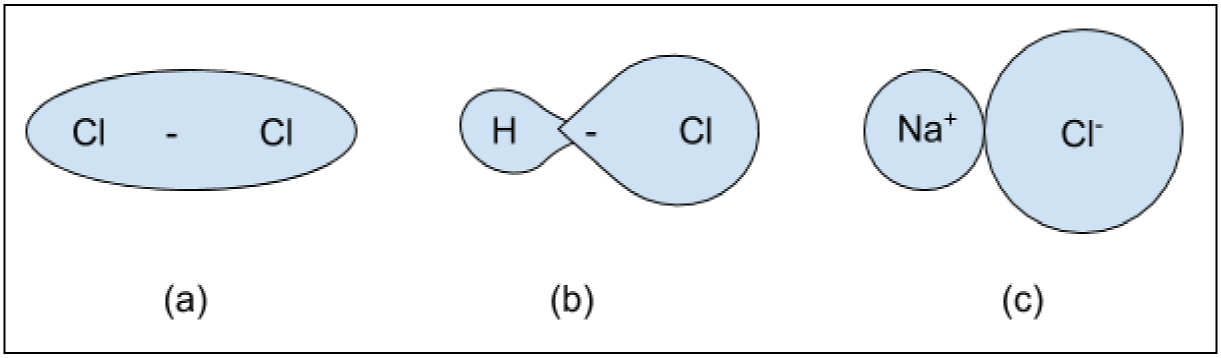
Induced Dipole

An induced dipole can form when the electron orbitals around a molecule are **influenced by another charged** particle.



Bond Character

Due to electronegativity and these dipole effects, bond character can vary from **ionic to covalent** with a form of **intermediate bond** in between.



(a) **Covalent Bond** - electrons **shared equally** between atoms producing **no overall charge**.

(b) **Intermediate Bond** - electrons **shared unequally** meaning there are **partial charges** on the atoms.

(c) **Ionic Bond** - complete **transfer** of electrons between atoms producing **ions** with full charges.

These differing bond characteristics can affect properties such as **boiling point**. They allow for a **greater attraction** between molecules, so more energy is required to break down the intermolecular forces of attraction. They can also affect **solubility**. More polar molecules will be more soluble as the partial charges will **attract to the opposite partial charges** on molecules such as water, allowing them to be 'held' in solution.

Intermolecular Forces

Forces act between the molecules to **hold them in crystal structures**. There are three main types of these forces, which are all much weaker than ordinary ionic or covalent bonds. They **vary in strength** and differ in the types of molecules they act between.

Van der Waals Forces

This is the **weakest** type of intermolecular force that acts as an **induced dipole** between molecules.

The strength of Van der Waals forces varies depending on the **Mr** of the molecule and its shape. The larger the Mr of the molecule, the stronger the intermolecular forces. Straight chain molecules experience stronger van der waals forces than branched chain molecules as they can **pack closer together**. This **reduces the distance** over which the force acts, making it stronger.

Permanent Dipole

This type of intermolecular force acts between molecules with a **polar bond**. The ∂+ and ∂- regions attract each other and hold the molecules together in an **electrostatic lattice-like structure**.

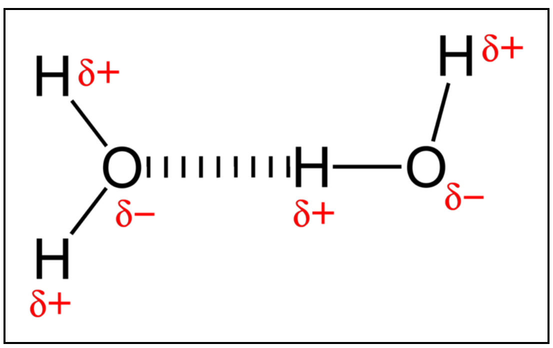
Example:

∂+		∂-	∂+		∂-	∂+		∂-
H	-	F	H	-	F	H	-	F

Hydrogen Bonding

This is the **strongest** type of intermolecular force. Hydrogen bonds are only present if there is a **hydrogen** atom bonded to one of the three most electronegative atoms: **nitrogen, oxygen or fluorine**. The lone pair on these atoms forms a bond with a ∂+ hydrogen atom from another molecule, shown as a dotted line.

Example:



(https://commons.wikimedia.org/wiki/File:Hydrogen-bonding-in-water-2D.png)
Benjah-bmm27 / CC BY-SA 3.0

Molecules held together with hydrogen bonds have much **higher melting and boiling points** compared to similar sized molecules without hydrogen bonding. This shows how the type of intermolecular force heavily **influences the physical properties** of a substance.

Shapes of Molecules

The shape of a molecule is determined by the **number of electron pairs** around the central atom. Each electron pair **naturally repels** each other so that the **largest bond angle** possible exists between the covalent bonds.

The VSEPR Principle

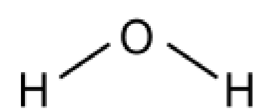
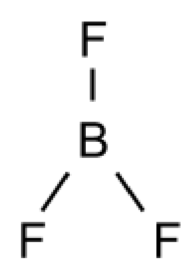
This is the **valence shell electron pair repulsion principle**. Any lone pairs present around the central atom provide **additional repulsive forces**, which changes the bond angle and the shape of the molecule. For every lone pair present, the bond angle between covalent bonds is reduced by **2.5°**.

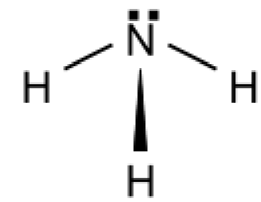
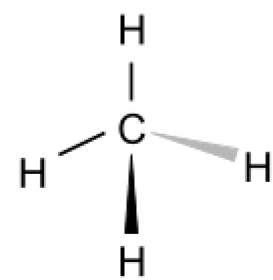
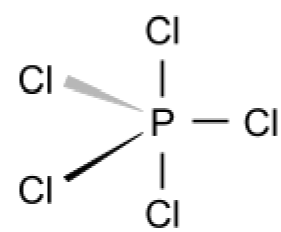
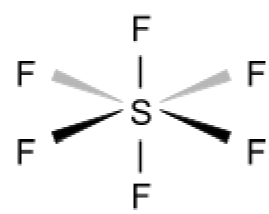
Standard Molecule Shapes

The shape of a molecule can be determined by considering the **type and number of electron pairs**:

1. Find the number of electron pairs.
2. Determine how many of the pairs are bonding pairs and how many are lone pairs.
3. Bonding pairs indicate the basic shape and lone pairs indicate any additional repulsion.

This table shows some **common molecule shapes** and the types of electron pairs present that lead to the shape:

Name	Bonding e^- Pairs	Lone e^- Pairs	Bond Angle (°)	Example
Linear	2	0	180	Cl - Be - Cl
V - Shaped (Bent)	2	2	104.5	
Trigonal Planar	3	0	120	

Triangular Pyramid	3	1	107	
Tetrahedral	4	0	109.5	
Trigonal Bipyramid	5	0	180 and 120	
Octahedral	6	0	90	

Larger ions can also have similar shapes that are similarly affected by electron pairs. The same rules of the VSEPR principle still apply and the shapes produced are the same or very similar to the ones above.