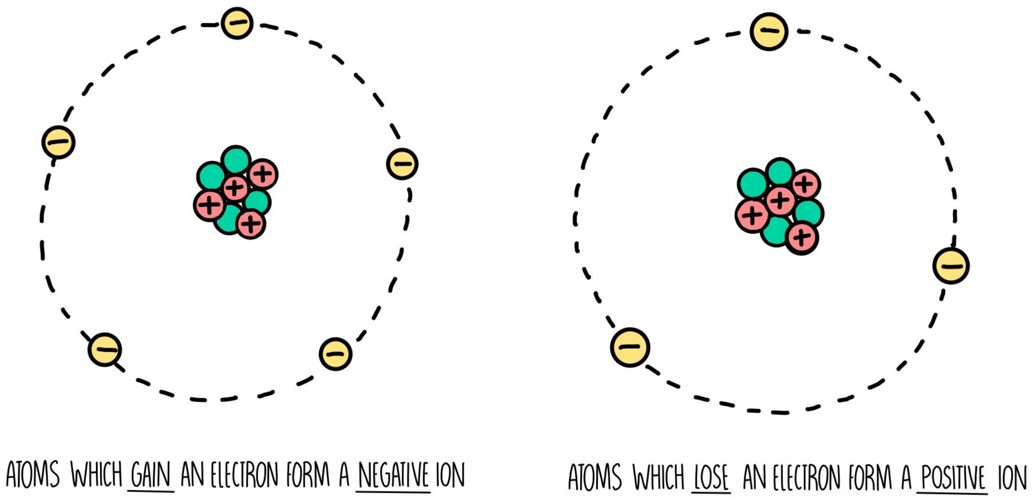
Bonding and Structure

Ionic bonding

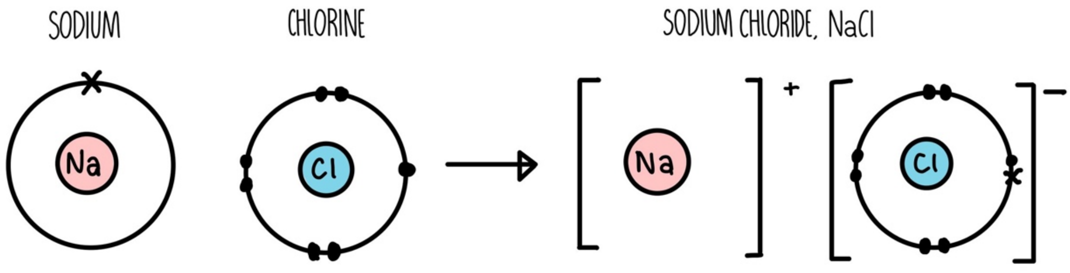
Atoms form **ions** when they **gain or lose electrons** to achieve a complete octet.

- Group 1 and 2 metals tend to lose electrons and form **positive ions**.
- Non-metals in groups 6 and 7 tend to gain electrons and form **negative ions**.

An ionic bond is the **electrostatic attraction between positive and negative ions**.

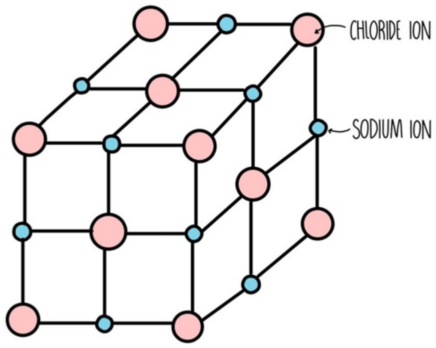


We can show how the metallic atom loses electrons and the non-metallic atom gains electrons by drawing a **dot and cross diagram**.



Ionic compounds can form **giant lattices**, with many positive and negative ions forming numerous ionic bonds.

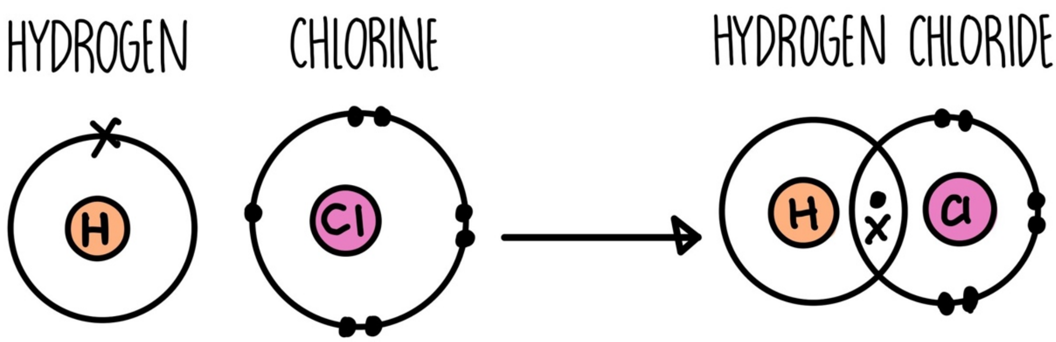
- These structures have **high melting and boiling points** because ionic bonds are **strong** and require a great deal of energy to break.



- This makes ionic compounds solid at room temperature.
- They do not conduct electricity when solid but can conduct electricity when **molten** or as part of an **aqueous solution**. Remember that in order to conduct electricity, there need to be *charged* molecules that are *free to move*. The positive and negative ions can only move when the ionic compound is melted or dissolved in solution.

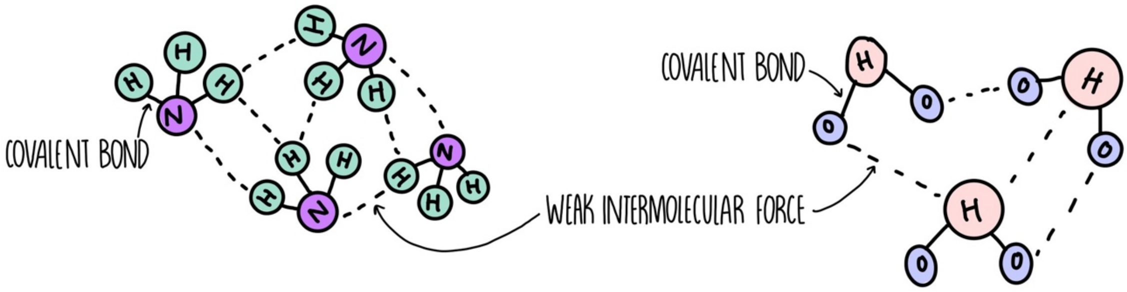
Covalent bonding

- A covalent bond is formed when two atoms **share electrons**. When this happens there is an **electrostatic attraction** between the nuclei (which are positively charged) and the pair of shared electrons (which are negatively charged).
- Covalent bonds are represented with the shared pair of electrons drawn in overlapping circles.



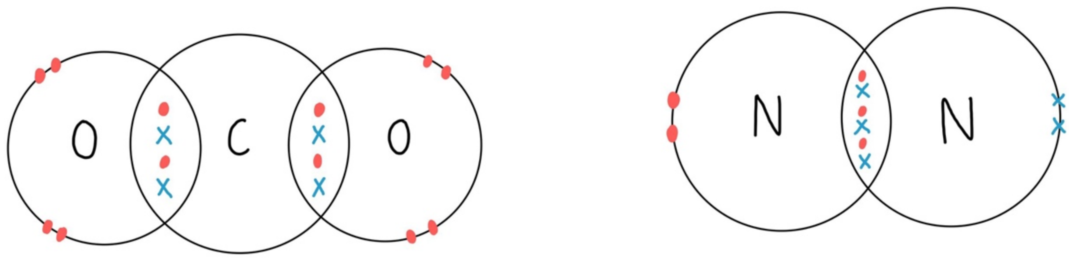
Molecules that consist of just a **few atoms** held together by covalent bonds are called **simple covalent molecules**.

- Examples include water, H_2O, carbon dioxide, CO_2 and ammonia, NH_3.
- They have **low melting and boiling points** because the strong covalent bonds remain intact when these molecules change state. Instead, it is the **weak intermolecular forces** that are broken.

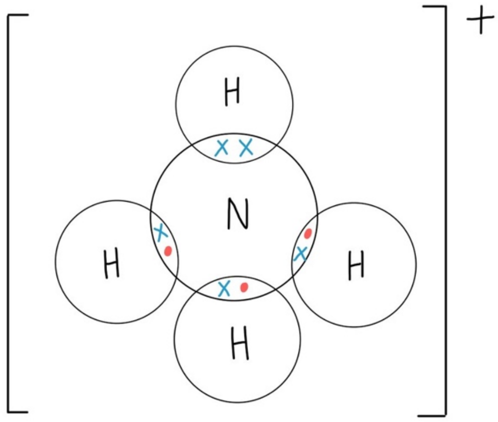


Although all simple covalent molecules will have low melting and boiling points, those with a **large molecular mass** will have a **higher melting and boiling point** than a molecule with a smaller mass. This is because the weak intermolecular forces that are found between molecules form **from electrons**. Heavier molecules have a **larger number of electrons** which means they can form more intermolecular forces, thereby increasing their melting and boiling point.

If a compound forms double or triple bonds, you just need to draw **two or three pairs** of dots and crosses. The image below shows a dot-and-cross diagram for the double bonds in a molecule of carbon dioxide and a triple bond in a molecule of nitrogen.



Dative covalent bonds (aka **coordinate** covalent bonds) are a special type of covalent bond that is formed when **both electrons** come from the **same species**. For example, ammonia has a lone pair of electrons which it can use to form a dative covalent bond with a hydrogen ion, forming an ammonium ion. To make it clear that both electrons have been donated by ammonia, we use the **same symbol** for both electrons in the bond.

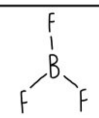
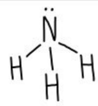
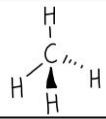
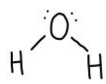
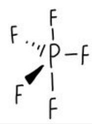
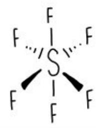


Average bond enthalpies

- The average amount of **energy** required to **break a bond**.
- Bonds with higher bond enthalpies are **stronger** than bonds with lower bond enthalpies.
- Bond enthalpies are **approximations** – the actual energy required to break a particular bond will depend on its **surroundings** (i.e. the other atoms and bonds present in the compound).

Shapes of molecules

- The shapes of molecules depend on **how many bonds** they form and **how many lone pairs** they have.
- Bonding electrons and lone pair electrons will arrange themselves so that they are a **maximum distance apart** to **minimise repulsion** between them.
- **Lone pairs repel more than bonding pairs**, so **lone pair-lone pair bond angles** are the **largest**, followed by **lone pair-bond pair angles** then **bond pair-bond pair angles**.

SHAPE	NO. OF BOND PAIRS	NO. OF LONE PAIRS	BOND ANGLES	EXAMPLE
LINEAR	2	0	180°	Cl–Be–Cl
TRIGONAL PLANAR	3	0	120°	
TRIGONAL PYRAMIDAL	3	1	107°	
TETRAHEDRAL	4	0	109.5°	
BENT	2	2	104.5°	
TRIGONAL BIPYRIMIDAL	5	0	90/120°	
OCTAHEDRAL	6	0	90°	

Bond polarity and electronegativity

- **Electronegativity** = the ability of an atom to **pull an electron pair in a covalent bond towards itself**.
- Fluorine is the most electronegative element.
- Electronegativity can be measured on the **Pauling scale** where each element is assigned a Pauling value. **The higher the Pauling value**, the **more electronegative** the element.
- For two atoms forming a covalent bond, the **greater the difference in their electronegativities,** the **more polar** the bond.
- Although a molecule might have polar bonds, it can still be a **non-polar molecule** if it has a **symmetrical geometry** — the polar bonds will cancel each other out.

Intermolecular forces

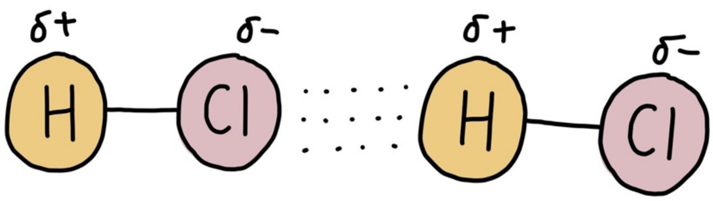
Intermolecular forces are the forces of attraction **between** different molecules. There are two main categories of IMFs:

1. **Permanent** dipole-dipole interactions

2. **Induced** (temporary) dipole-dipole interactions (aka London dispersion forces)

Permanent dipole-dipole interactions

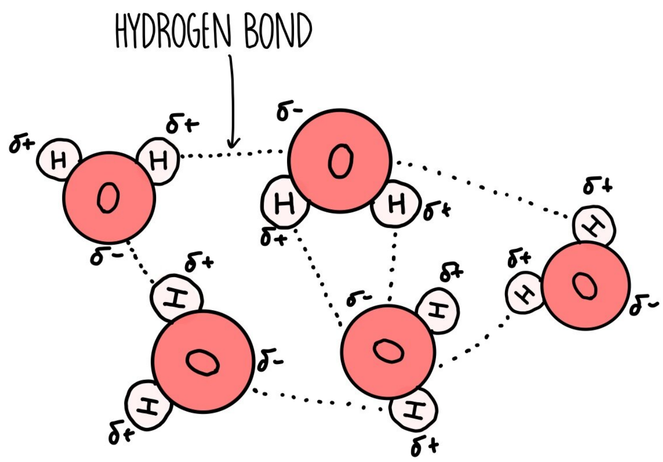
- Formed when there is a polar bond due to an electronegativity difference.
- The slightly positive region on one atom will be attracted to the slightly negative region on another atom of a different molecule.



Hydrogen bonding

- This special type of **permanent dipole-dipole interaction**.
- It's the attraction between **hydrogen** on one molecule with an **electronegative atom** (usually O, N or F) on another molecule. Anything with an –OH (such as water and alcohols) can form hydrogen bonds.
- Hydrogen bonding is the **strongest type of intermolecular force**, which is why water has such a high melting and boiling point compared to similar liquids.

Hydrogen bonding is also responsible for the fact that ice is less dense than liquid water. In the solid state there are more hydrogen bonds between water molecules. Since these bonds are long, the water molecules are held a further distance apart compared to when they're in the liquid state, making ice less dense.



Induced dipole-dipole interactions

- Occur due to the **random and uneven distribution of electrons** within a molecule.
- The uneven distribution of electrons temporarily creates regions that are slightly positive and other regions which are slightly negative. This is called an **instantaneous dipole**.
- The slightly negative region of one molecule will repel the electrons in another molecule. This temporarily creates a dipole in that second molecule — an **induced dipole**. The slightly negative region on one ethane molecule will be attracted to the slight positive region on the other.
- Another name given to induced dipole-dipole interactions is **London dispersion forces**.

Since electrons are responsible for induced dipole-dipole interactions, this means that **ALL molecules** experience these types of forces. The **more electrons** a molecule has, the **stronger** these interactions are. For example, as you go down group 7 boiling point increases because the halogen molecules have more electrons and therefore stronger London dispersion forces. This is why the halogens gradually change state from gaseous (F, Cl) to liquid (Br) to solid (I, At) as you descend the group.