Cambridge (CIE) A Level Chemistry

Your notes

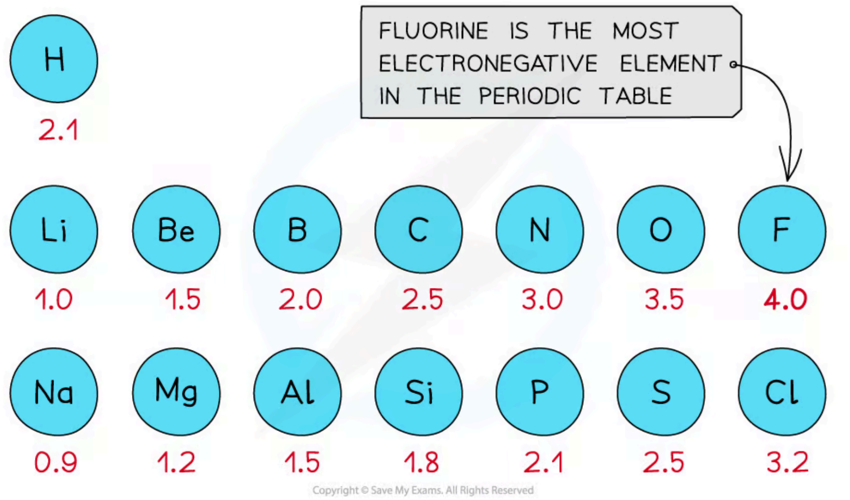
Electronegativity & Bonding

Contents

Defining Electronegativity

- **Electronegativity** is the ability of an atom to attract a pair of electrons towards itself in a covalent bond
- This phenomenon arises from the **positive** nucleus's ability to attract the **negatively** charged electrons, in the outer shells, towards itself
- The **Pauling scale** is used to assign a value of electronegativity for each atom

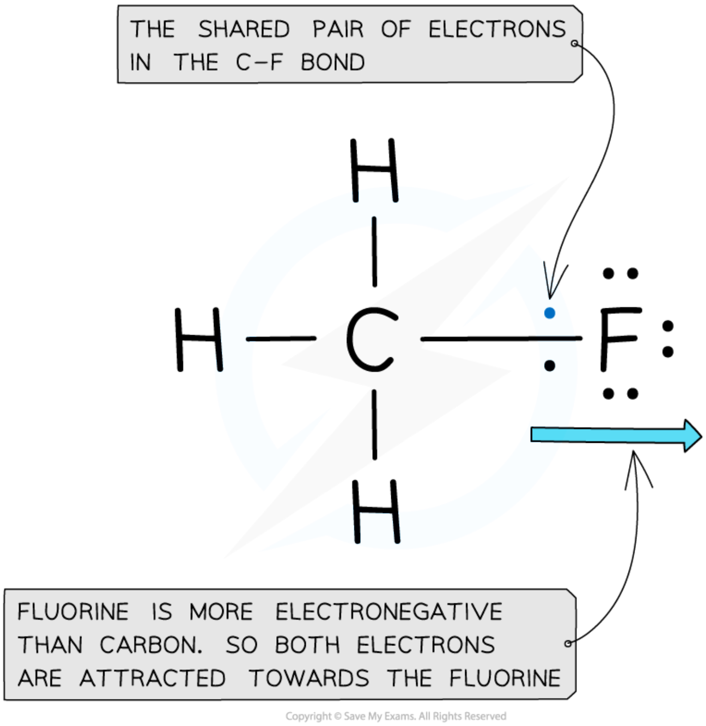
First three rows of the periodic table showing electronegativity values



Electronegativity generally increases across a period and decreases down a group

- Fluorine is the most electronegative atom on the Periodic Table, with a value of 4.0 on the **Pauling Scale**
- It is best at attracting electron density towards itself when covalently bonded to another atom

Electron distribution in the C-F bond of fluoromethane



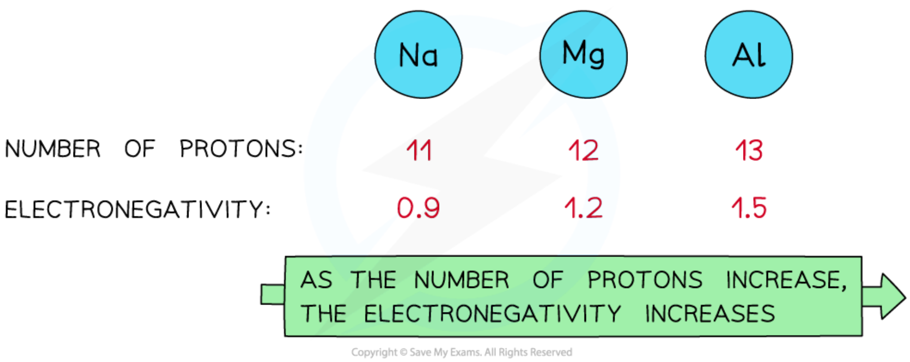
The electrons in the C–F bond are closer to the fluorine due to its electronegativity

Factors Affecting Electronegativity

Nuclear charge

- **Attraction** exists between the positively charged **protons** in the nucleus and negatively charged **electrons** found in the energy levels of an atom
- An **increase** in the number of **protons** leads to an **increase** in **nuclear attraction** for the electrons in the outer shells
- Therefore, an **increased nuclear charge** results in an **increased electronegativity**

Relating electronegativity values to number of protons



As the nuclear charge increases, the electronegativity of an element increases as well

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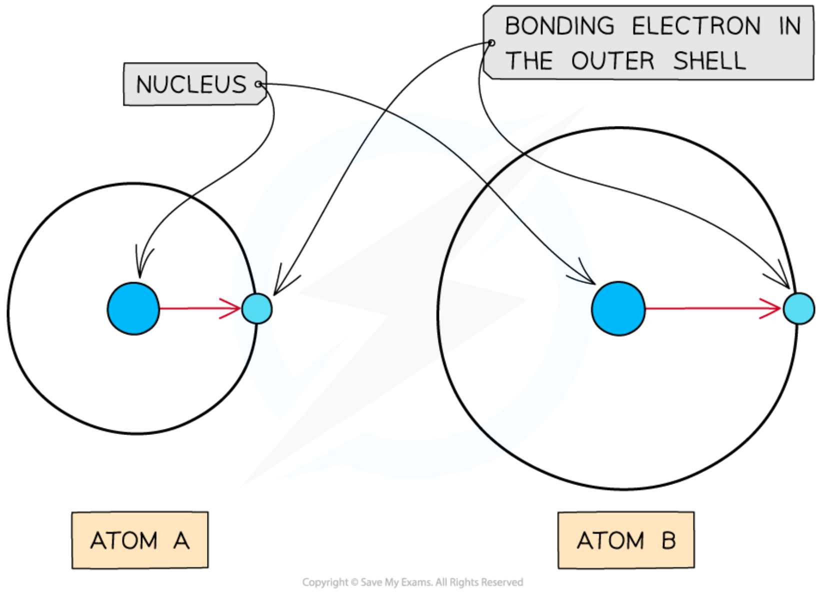
Atomic radius

- The **atomic radius** is the distance between the nucleus and electrons in the **outermost shell**
- Electrons **closer** to the nucleus are more **strongly** attracted towards its positive **nucleus**
- Those electrons **further away** from the nucleus are **less strongly** attracted towards the **nucleus**
- Therefore, an **increased atomic radius** results in a **decreased electronegativity**

How the distance from the nucleus to the outer electrons affects electronegativity

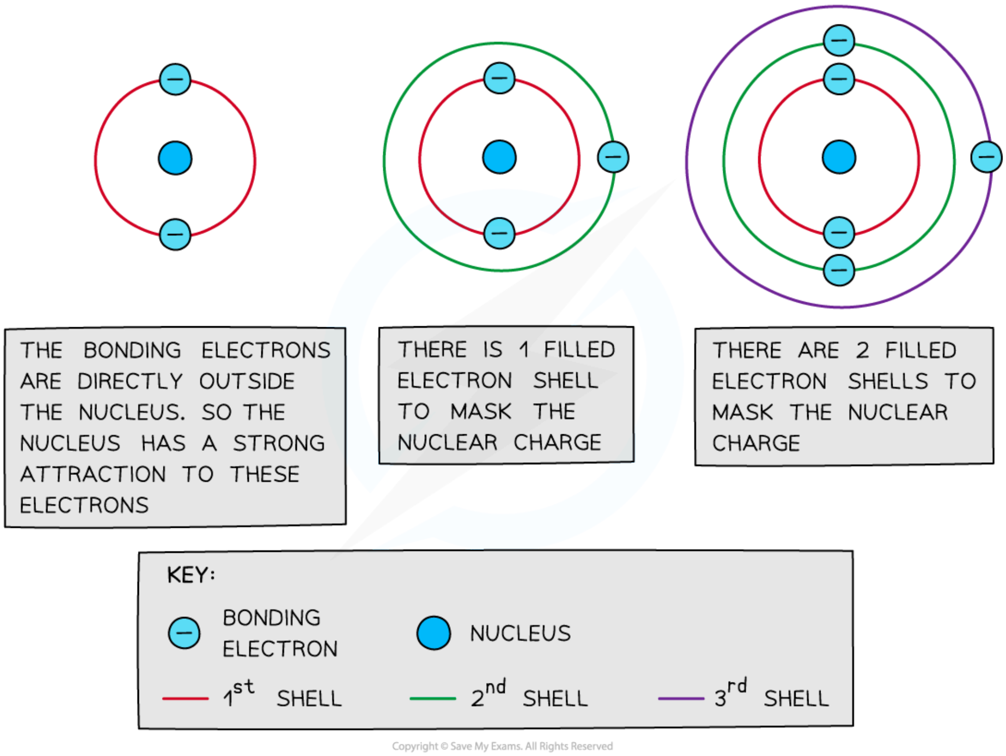


As the atomic radius increases, the nucleus has less of an attraction for the bonding electrons causing atom A to have a higher electronegativity than atom B

Shielding

- **Filled** energy levels can **shield** (mask) the effect of the nuclear charge causing the outer electrons to be **less attracted** to the nucleus
- Therefore, the addition of extra **shells and subshells** in an atom will cause the outer electrons to experience **less** of the attractive force of the nucleus
  - Sodium (Period 3, Group 1) has a higher **electronegativity** than caesium (Period 6, Group 1) as it has fewer shells and therefore the outer electrons experience less shielding than in caesium
- Thus, an increased number of **inner shells and subshells** will result in a **decreased electronegativity**

How shielding affects nuclear attraction



Filled inner energy levels mask the nuclear attraction from the outer bonding electrons

Examiner Tips and Tricks

- The **nuclear charge**, **atomic radius** and **shielding** are all linked to each other.
- As **nuclear charge increases**, the nucleus has a **greater attractive force** on the electrons in shells given that the **shielding** doesn't increase.
- As a result of this, the atomic radius decreases.

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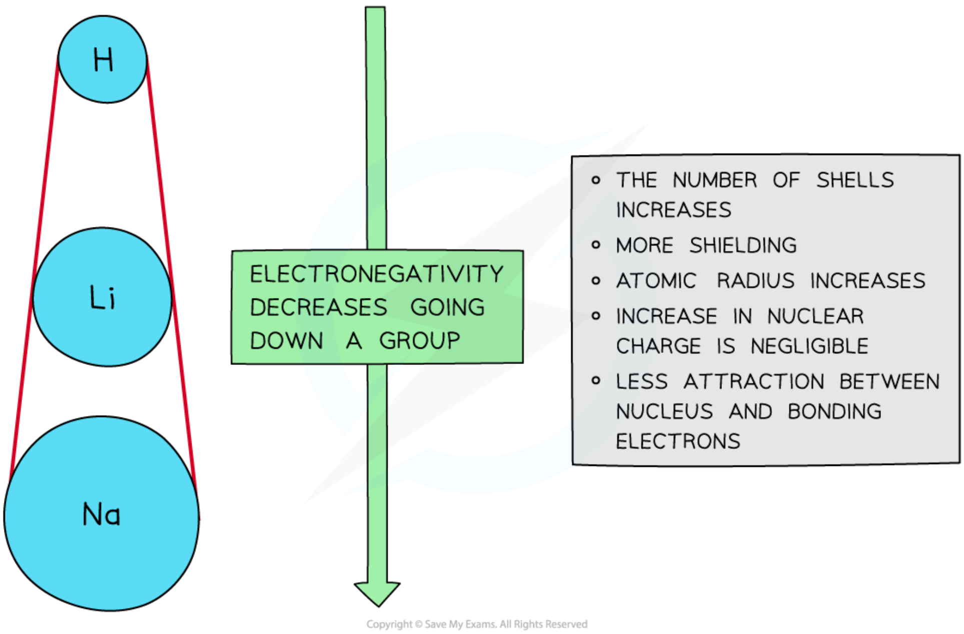
Trends in Electronegativity

- **Electronegativity** varies across **Periods** and down the Groups of the Periodic Table

Down a group

- There is a **decrease** in **electronegativity** going down the Group
- The **nuclear charge increases** as more protons are being added to the nucleus
- However, each element has an extra filled electron shell, which increases the **shielding**
- The addition of the extra shells increases the distance between the nucleus and the outer electrons resulting in **larger atomic radii**
- Overall, there is a decrease in attraction between the nucleus and outer bonding electrons

Why electronegativity decreases down a group



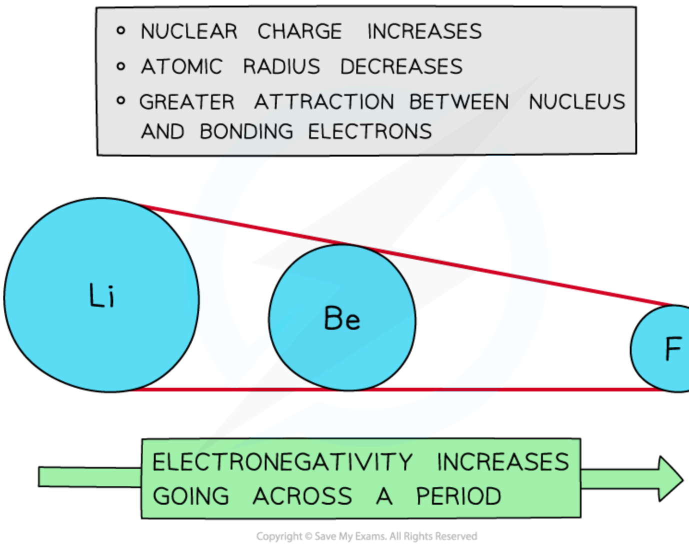
Electronegativity decreases going down the groups of the Periodic Table

Across a period

- Electronegativity **increases** across a Period
- The **nuclear charge increases** with the addition of protons to the nucleus
- **Shielding** remains reasonably the **same** across the Period as no new shells are being added to the atoms

- The nucleus has an increasingly strong attraction for the bonding pair of electrons of atoms across the Period of the Periodic Table
- This results in **smaller atomic radii**

Why electronegativity increases across a period



Electronegativity increases going across the periods of the Periodic Table

Trends down a group & across a period

- **Nuclear charge**
  - Increases down a group
  - Increases across a period
- **Shielding**
  - Increases down a group
  - Reasonably constant across a period
- **Atomic radius**
  - Increases down a group
  - Decreases across a period
- **Electronegativity**
  - Decreases down a group
  - Increases across a period

Examiner Tips and Tricks

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- Remember the general trend is an **increase** in electronegativity towards the top right of the Periodic Table.
- Fluorine is the most electronegative element in the Periodic Table.

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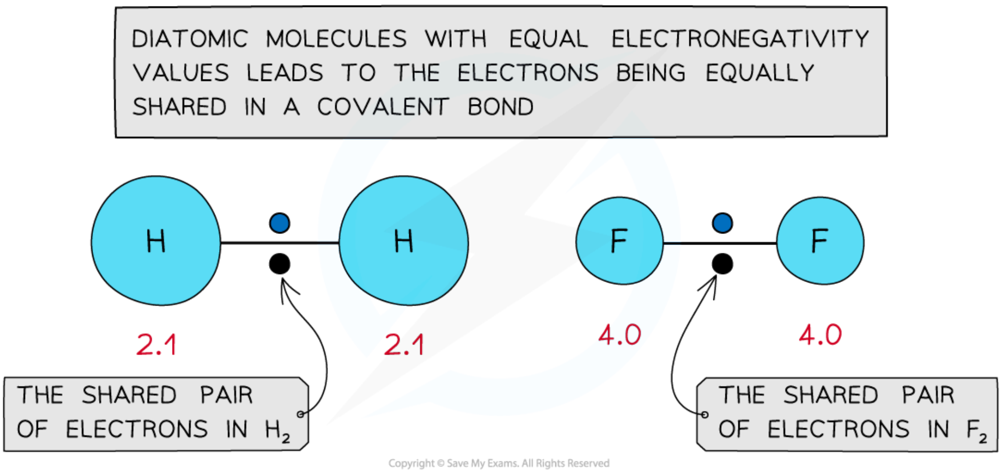
Using Electronegativity to Predict Bond Formation

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

Electronegativity & covalent bonds

- Single **covalent bonds** are formed by **sharing** a pair of electrons between two atoms
- 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 the formation of a **covalent** bond
- The equal distribution leads to a **non-polar molecule**

Electronegativity and non-polar molecules



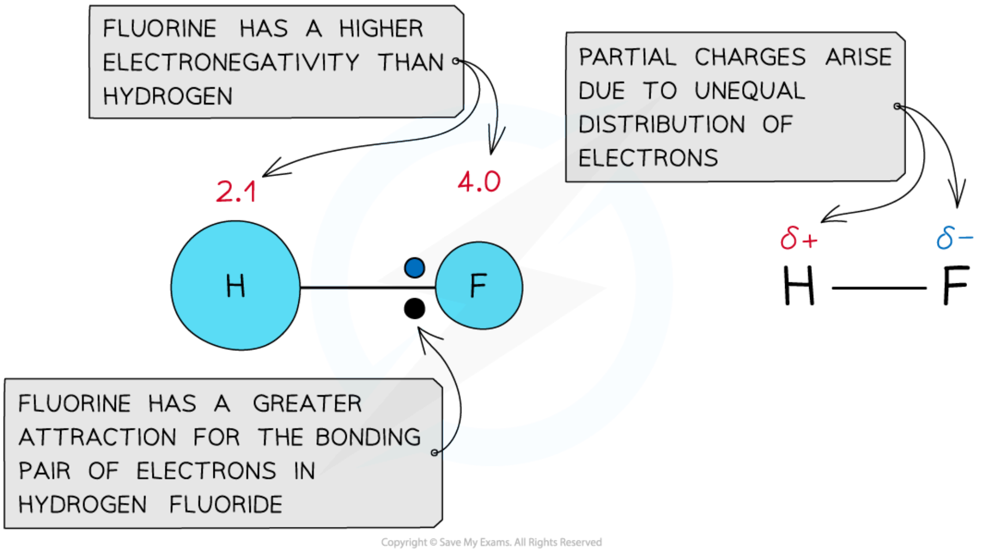
The electronegativity values are equal resulting in the formation of a nonpolar covalent bond

Electronegativity & polar covalent bonds

- When atoms of **different** electronegativities form a molecule, the shared electrons are not equally distributed in the bond
- The **more electronegative** atom (the atom with the higher value on the Pauling scale) will draw the bonding pair of electrons towards itself
- A molecule with partial charges forms as a result
- The **more electronegative** atom will have a **partial negative** charge (delta negative, **δ^-**)

- The **less electronegative** atom will have a partial positive charge (delta positive, **δ⁺**)
- This leads to a **polar covalent molecule**

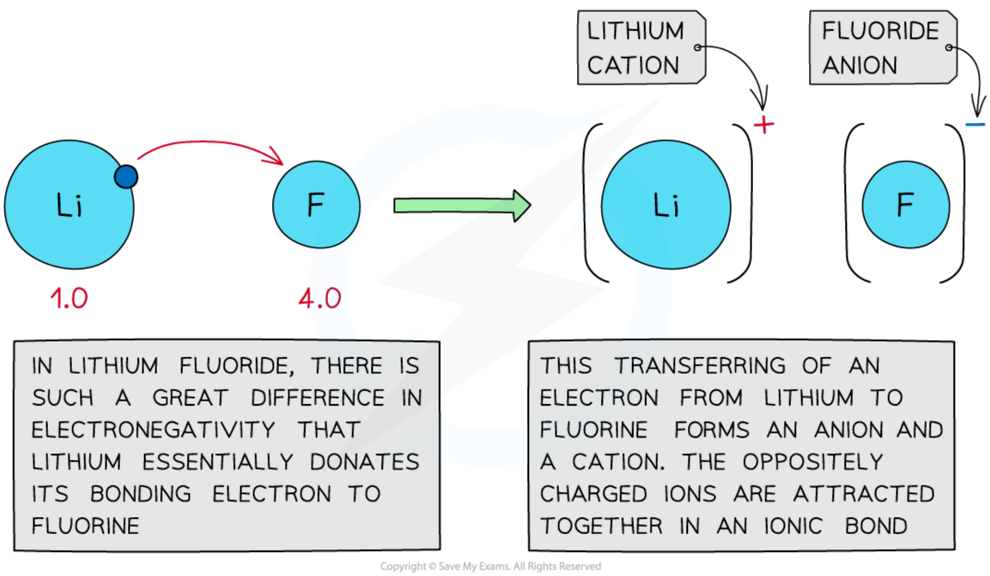
Electronegativity and polar covalent molecules



The electronegativity values are not equal so a polar bond forms

- If there is a **large** difference in electronegativity of the two atoms in a molecule, the least electronegative atom's electron will **transfer** to the other atom
- This in turn leads to an **ionic bond** – one atom transfers its electron and the other gains that electron
  - The **cation** is a **positively charged species** which has **lost** (an) electron(s)
  - The **anion** is a **negatively charged species** which has **gained** (an) electron(s)

Electronegativity and ionic bonding



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Large differences in electronegativity values lead to the formation of ionic bonds

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Examiner Tips and Tricks

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