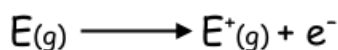
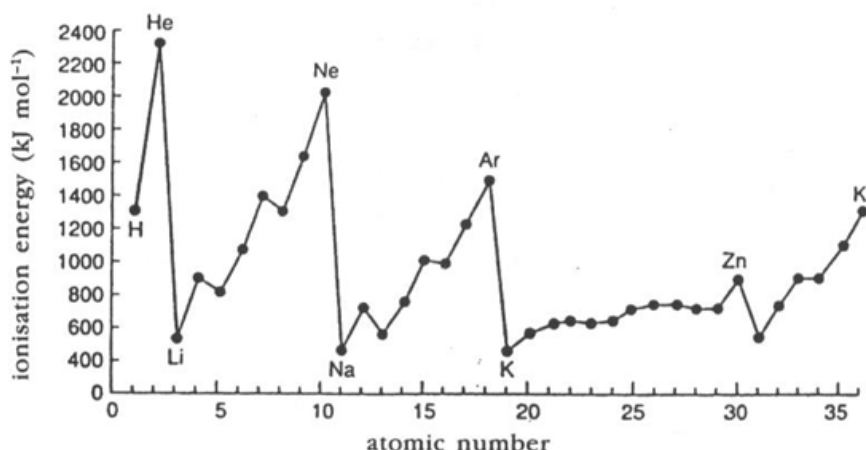


Ionisation Energy

Enthalpy of 1st Ionisation:



- Peak points are all Noble Gases
- All the lowest points are Alkali Metals



Factors Affecting Ionisation Energy

a) Atomic Size

- The greater the atomic radius of the atom, the further the distance of the outermost electrons
- This lowers the ionisation energy as the electrons are less attracted to the nucleus the further away they are

b) Nuclear Charge

- The greater the number of protons, the greater the attraction to the outer electrons and harder it is to remove outer electrons
- Elements across a period have decreasing atomic size and therefore increasing 1st ionisation energy as the increasing positive charge holds onto the outer electrons

c) Screening Effect

- Inner electron shells shield the outer electron shell from the attractive forces from the nucleus
- The larger the atom, the more inner electron shells the atom has and the bigger the screening effect and the lower the ionisation energy.

Down A Group:

1st Ionisation Energy DECREASES

Across A Period:

1st Ionisation Energy INCREASES

However, the above graph shows some irregularities where these trends are generally followed but there is the occasional exception. We can now explain these irregularities using our new knowledge on electronic structure.

Why has Beryllium a higher 1st ionisation energy than the trend would suggest?

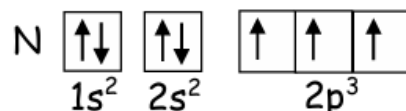
Element	Symbol	Electronic Configuration	1st Ionisation Energy
Lithium	Li	1s ² 2s ¹	$\Delta H_i^\circ = +526 \text{ kJ mol}^{-1}$
Beryllium	Be	1s ² 2s ²	$\Delta H_i^\circ = +905 \text{ kJ mol}^{-1}$
Boron	B	1s ² 2s ² 2p ¹	$\Delta H_i^\circ = +807 \text{ kJ mol}^{-1}$
Carbon	C	1s ² 2s ² 2p ²	$\Delta H_i^\circ = +1090 \text{ kJ mol}^{-1}$

- Removing an electron from Be breaks a full subshell (2s)
- Breaking full subshells requires more energy and this increases the 1st ionisation energy
- Removing an electron from B empties the 2p subshell and makes the 2s² the outer electron subshell.
- Creating a full outer subshell is favoured and this is why the 1st ionisation energy of B is lower than Be

Why has Nitrogen got a higher 1st ionisation energy than the trend would suggest?

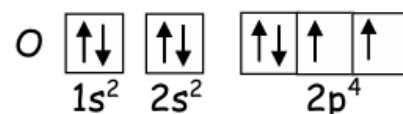
Element	<u>Symbol</u>	Electronic Configuration	1st Ionisation Energy
Carbon	C	1s ² 2s ² 2p ²	$\Delta H_i^\circ = +1090 \text{ kJ mol}^{-1}$
Nitrogen	N	1s ² 2s ² 2p ³	$\Delta H_i^\circ = +1410 \text{ kJ mol}^{-1}$
Oxygen	O	1s ² 2s ² 2p ⁴	$\Delta H_i^\circ = +1320 \text{ kJ mol}^{-1}$
Fluorine	F	1s ² 2s ² 2p ⁵	$\Delta H_i^\circ = +1690 \text{ kJ mol}^{-1}$

- Removing an electron from N breaks a relatively stable half-filled 2p subshell



- Breaking half-full subshells requires more energy and this increases the 1st ionisation energy

- Removing an electron from O creates a half-filled 2p subshell and this is more stable



- Creating a half-full subshell is favoured and this is why the 1st ionisation energy of O is lower than N