Ionisation Energies

Yim and Mara

First Ionisation Energy

The minimum energy needed to remove a mole of electrons from a mole of gaseous atoms to form a mole of univalent cations in the gaseous state

Single charge

$$X(g) \rightarrow X^{+}(g) + e^{-}$$

- ➔ Will always result in a 1+ ion (no matter what you would expect)
- ➔ Every element has its own unique ionisation energy
- ➔ It's an endothermic reaction - you have to supply energy

Second Ionisation Energy

The minimum energy needed to remove a mole of electrons from a mole of univalent cations in the gaseous state to form a mole of divalent cations in the gaseous state

$$X^{+}(g) \rightarrow X^{2+}(g) + e^{-}$$

Third Ionisation Energy

$$X^{2+}(g) \rightarrow X^{3+}(g) + e^{-}$$

Fourth Ionisation Energy

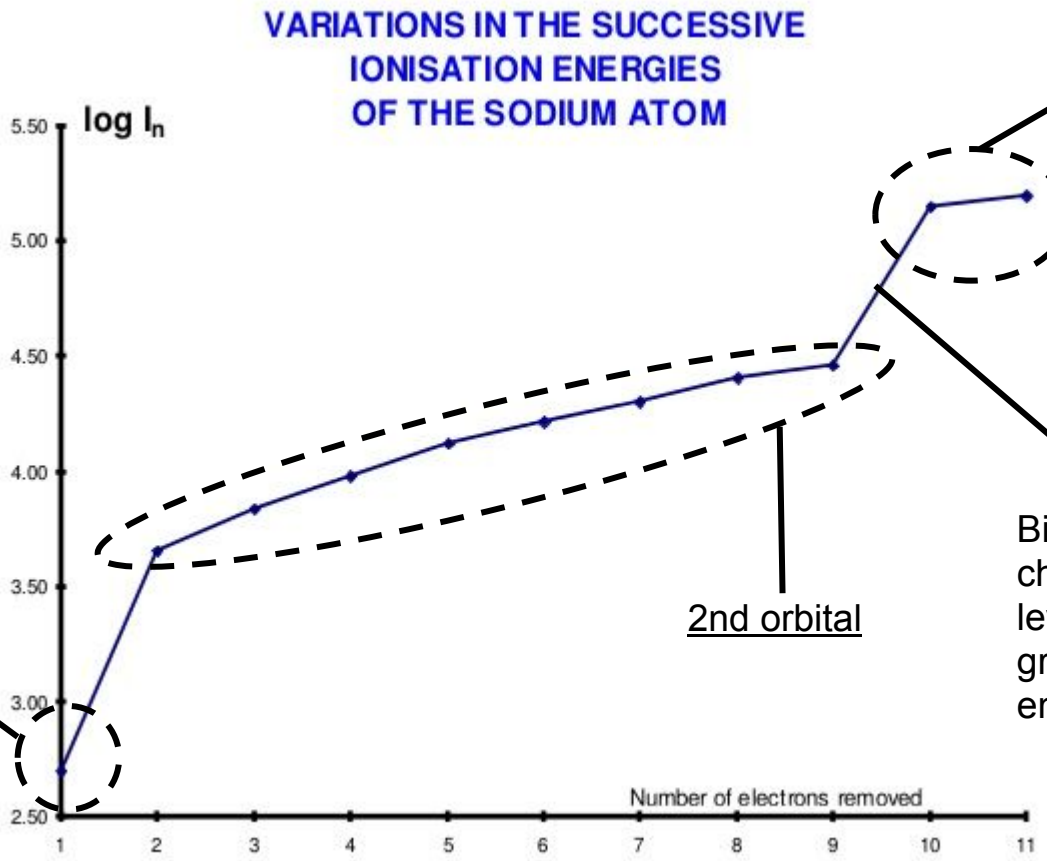
$$X^{3+}(g) \rightarrow X^{4+}(g) + e^{-}$$

And so on…

<u>As you remove electrons, the ionisation energies increase</u> (general rule) because you're trying to remove an electron from an increasing positively charged body.

To remove the electron, you have to overcome the forces of attraction between the (+) nucleus and the (-) electrons.

You require more energy as you go down IE's because the positive charge increases (2+, 3+, 4+…) meaning there is a stronger force of attraction.



1st orbital : highest ionisation energy as **(1)** most (+) charged
(2) Closest to nucleus
(3) There's no *shielding effect*

3rd orbital : lowest ionisation energy as
(1) Atom is neutral (no charge)
(2) It's the furthest away from nucleus (biggest atomic radius)
(3) There's a *shielding effect*

2nd orbital

Big jump in energy as it's changing to a new energy level (closer to nucleus = greater attraction = more energy required)

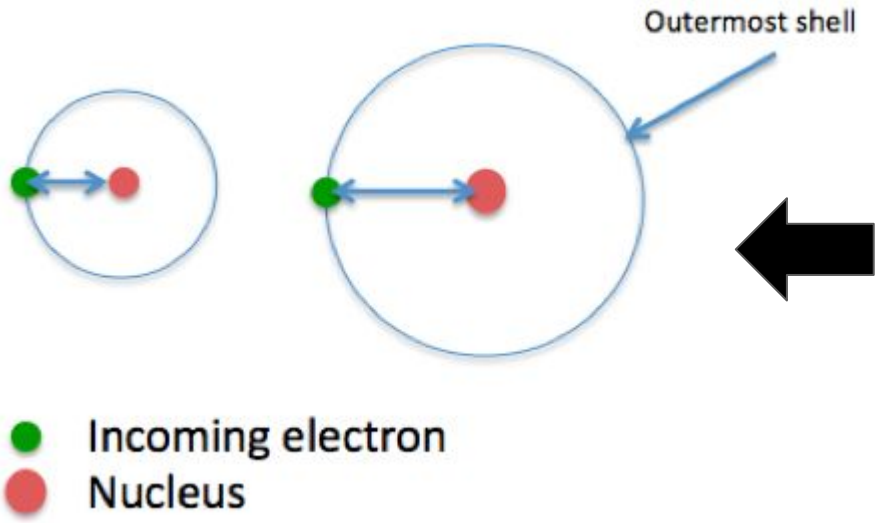
What Factors Affect First Ionisation Energy?

1. Atomic radius *(the size of the atom)*
2. Nuclear charge
3. Shielding effect

Atomic Radius

- The ionisation energy depends on the size of atom
- because with an increase in size, the distance between the nucleus and electrons increases, the force of attraction between (+) nucleus and (-) valence electrons then decreases.
- So, ionisation energy decreases with an increase in size

For example, as the size of an atom increases, the outermost electrons are less attracted to the nucleus. The outermost electron will become easier to remove



The greater the distance between the nucleus and the outer electron of an atom, the less the ionisation energy.

Nuclear Charge

When the nucleus contains more protons than electrons in the energy levels - it becomes ‘more positive’ (the positivity is stronger because there’s a difference/ imbalance of electrons and protons).

So the force of attraction between the (-) electrons and (+) nucleus increases and therefore the IE increases as these forces of attraction must first be overcome before the electron can be removed.

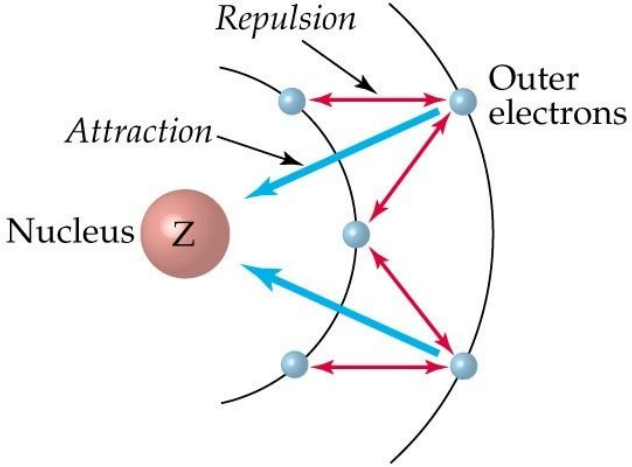
Shielding

Outer electrons (on energy level furthest away from nucleus) are repelled by the other electrons in the atoms

They are also simultaneously being attracted to the (+) nucleus

The repulsion from the other electrons ‘shields’ the outer electrons from this attraction

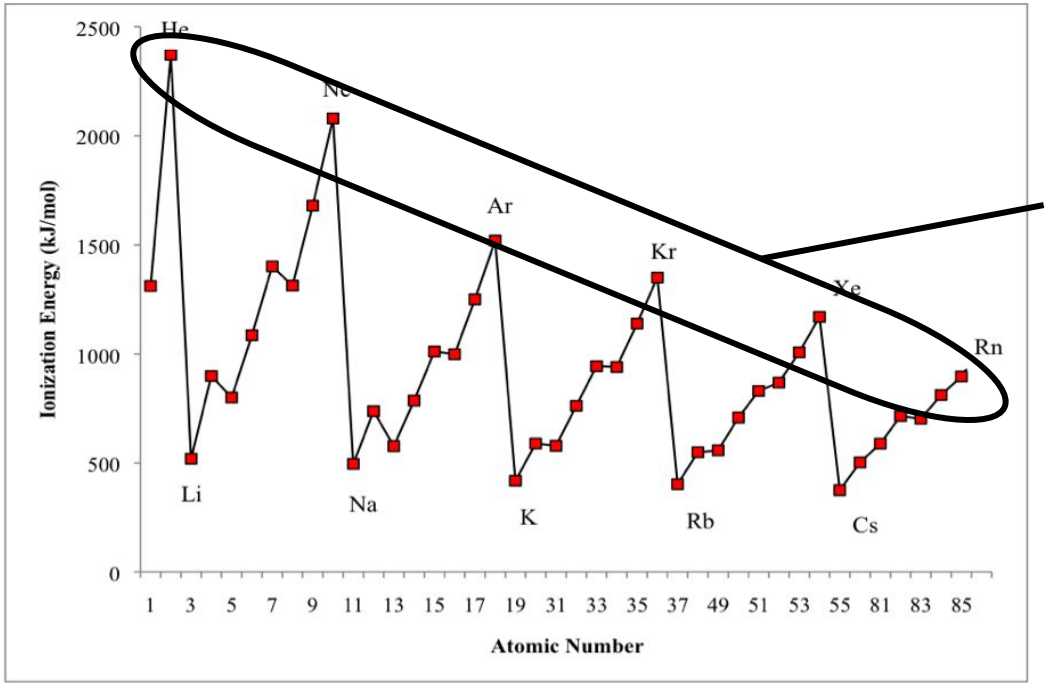
Shielding is more effective if the electrons are closer to the nucleus as when you get farther away - the attraction between the nucleus and electrons decreases anyways



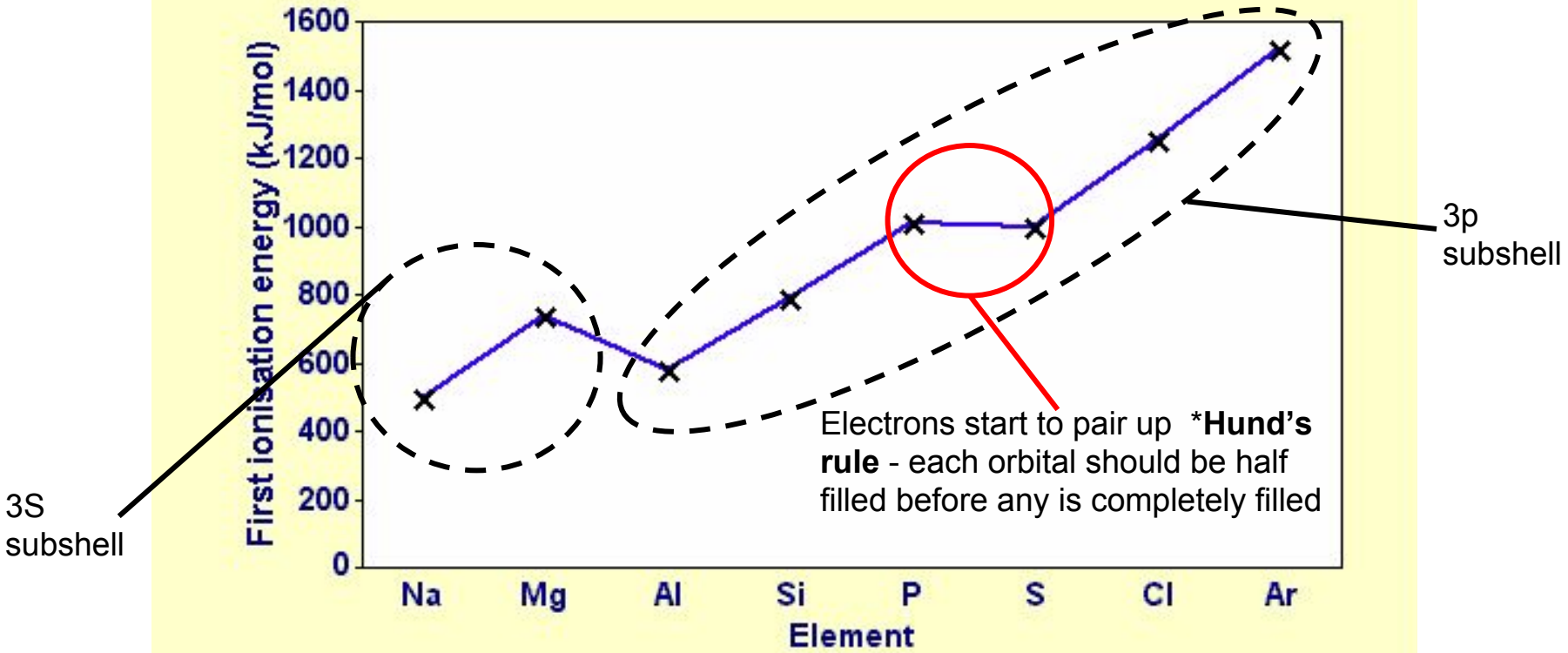
Ionisation Energy for periods 1-6

In the same period, the ionisation energies increase, when a new period starts the energy will fall back down and rise up again

1. The graph line will repeat when a new period starts.



2. In the graph, the last element has higher energy in the group because they have full shell electron.



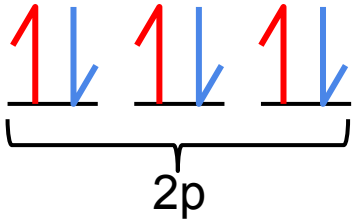
Shows evidence for sub shells in the third energy level

Hunds Rule

States that each orbital should be half filled before any is completely filled

Added first

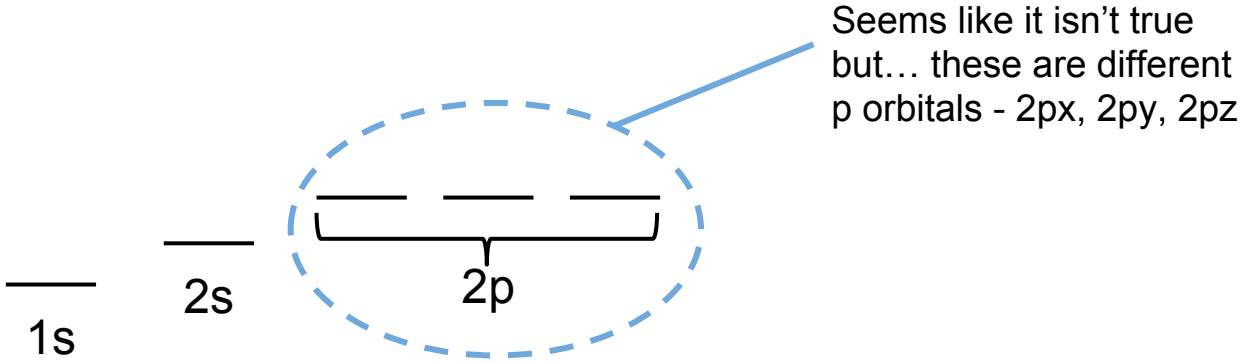
Added after



We can see evidence for this in the graph from the slight fall in energy

Pauli's Exclusion Principle

States that there can only be 2 electrons in each orbital



Aufbau Principle

States that electronic configuration follows an order in which the lower levels fill up first.

The first electron goes into the lowest energy level available, the next electron pairs up with it in the same orbital and the third electron goes into the next orbital (as you can only have 2 electrons in the first energy level) and so on…

The ionisation energies provides evidence for this principle as they have a trend of increasing when you remove electrons….

Energy Levels and Subshells

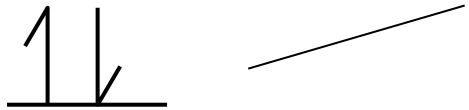
There is a maximum amount of electrons each energy level can hold. The first energy level can hold 2 electrons (first row of periodic table - hydrogen and helium)

1 H

2 He

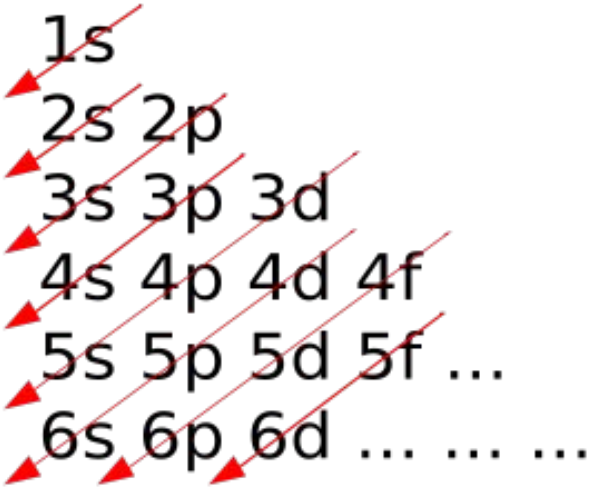
Each row of the periodic table corresponds to one shell filling

First energy level/ orbital - ‘S orbital’



1s

The second orbital can hold 8 electrons, the third orbital can hold 18 electrons, 4 = 32, 5=32...



Electron Configuration Table

H 1																	He 2
1s																	1s
Li 1	Be 2											B 1	C 2	N 3	O 4	F 5	Ne 6
2s												2p					
Na 1	Mg 2											Al 1	Si 2	P 3	S 4	Cl 5	Ar 6
3s												3p					
K 1	Ca 2	Sc 1	Ti 2	V 3	Cr 4	Mn 5	Fe 6	Co 7	Ni 8	Cu 9	Zn 10	Ga 1	Ge 2	As 3	Se 4	Br 5	Kr 6
4s		3d										4p					
Rb 1	Sr 2	Y 1	Zr 2	Nb 3	Mo 4	Tc 5	Ru 6	Rh 7	Pd 8	Ag 9	Cd 10	In 1	Sn 2	Sb 3	Te 4	I 5	Xe 6
5s		4d										5p					
Cs 1	Ba 2	La* 1	Hf 2	Ta 3	W 4	Re 5	Os 6	Ir 7	Pt 8	Au 9	Hg 10	Tl 1	Pb 2	Bi 3	Po 4	At 5	Rn 6
6s		5d										6p					
Fr 1	Ra 2	+Ac 1	Rf 2	Ha 3													
7s		6d															

Ce 1	Pr 2	Nd 3	Pm 4	Sm 5	Eu 6	Gd 7	Tb 8	Dy 9	Ho 10	Er 11	Tm 12	Yb 13	Lu 14
4f													
Th 1	Pa 2	U 3	Np 4	Pu 5	Am 6	Cm 7	Bk 8	Cf 9	Es 10	Fm 11	Md 12	No 13	Lr 14
5f													

Reminders

The period (or row) number tells us how many energy levels are occupied and the group number tells us how many electrons are in the valence shells

Eg. Phosphorus is in group 5 and in period 3, so it has 3 occupied shells with 5 electrons on the outer shell

TO WRITE ELECTRONIC STRUCTURE

➔ Look for noble gas 1 above period and add on the ‘leftover electrons’

Eg. Calcium = 2,8,8,2 OR [Ar], 2 (Ar = 2,8,2)

There are Different Types of Orbitals in an Atom...

1. S orbital

➔ Can only hold 2 electrons; every shell has a single s orbital

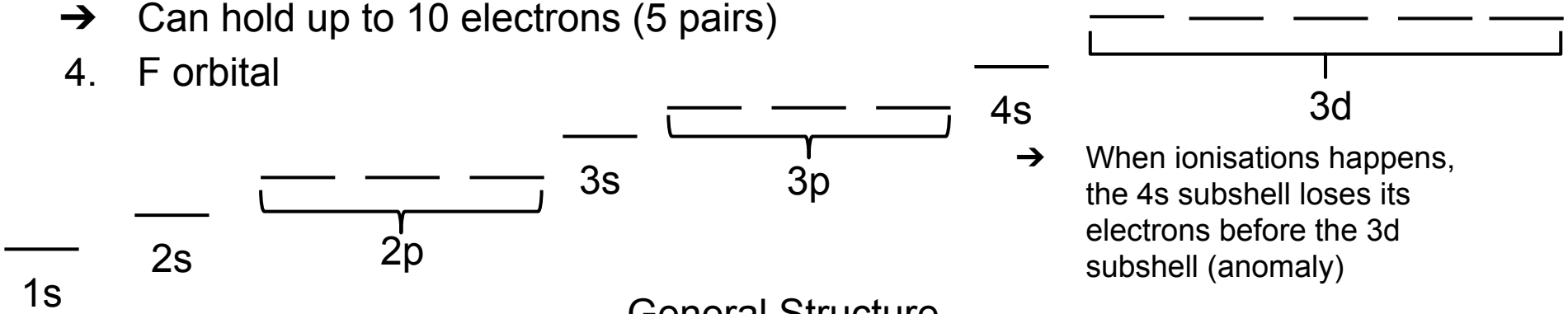
2. P orbital

➔ Fills up after s orbital is filled; can have up to 6 electrons (3 pairs); the first shell (n=1) doesn't have a p orbital

3. D orbital

➔ Can hold up to 10 electrons (5 pairs)

4. F orbital



General Structure

➔ When ionisations happens, the 4s subshell loses its electrons before the 3d subshell (anomaly)