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— TUITION —

Phone: 0442081445350

www.chemistryonlinetuition.com

Emil:asherrana@chemistryonlinetuition.com

CHEMISTRY

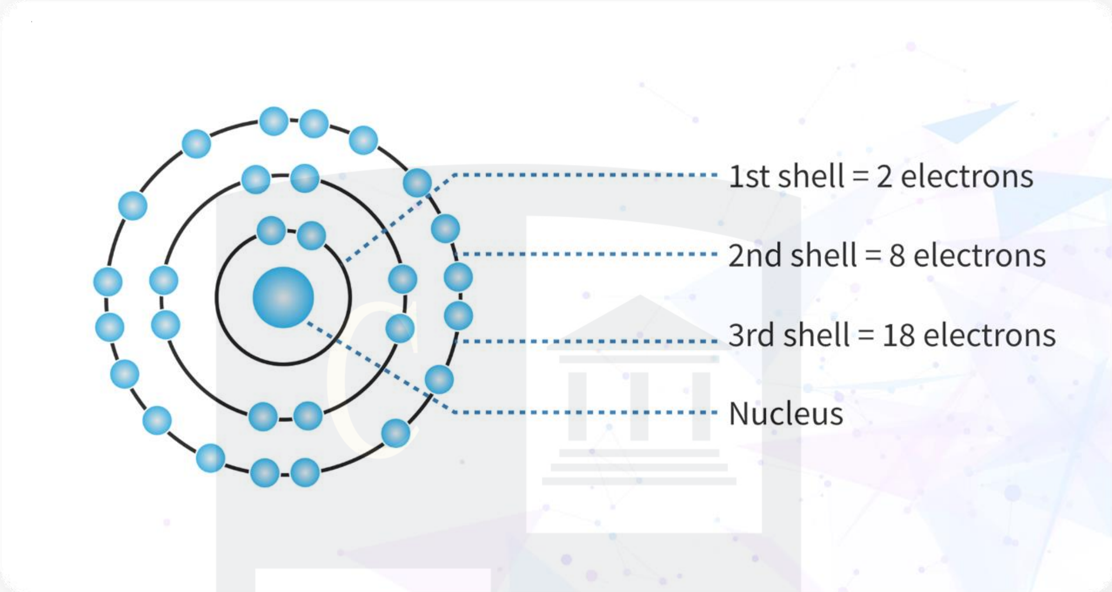
REVISION NOTES

ATOMICS STRUCTURE -2

Electronic Structure & Ionization Energy

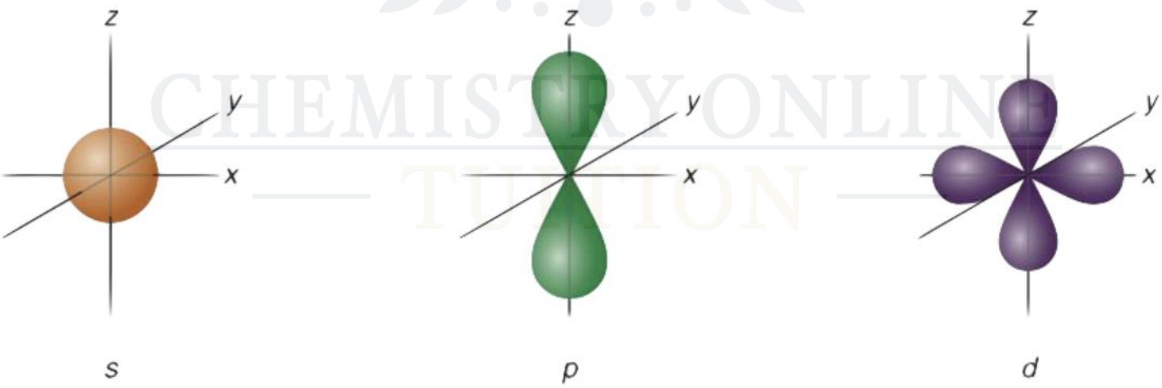
Electron arrangement in an atom is as follows:

Principal energy levels are sequentially numbered as 1, 2, 3, 4, and so forth, known as shells, with the first level (1) being the closest to the nucleus.



The energy level contains **sub-energy levels** labeled:

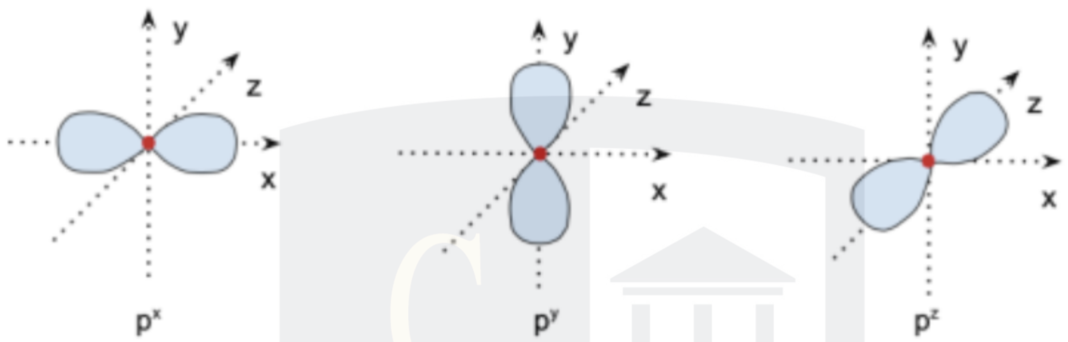
S, **P**, **D**, and **F**.



- S sub-level can accommodate up to 2 electrons.
- P sub-level can accommodate up to 6 electrons.
- D sub-level can accommodate up to 10 electrons, and
- F sub-level can accommodate up to 14 electrons.

These sub-levels consist of **orbitals** capable of holding up to 2 electrons with opposite spins.

Example of P-Orbitals:



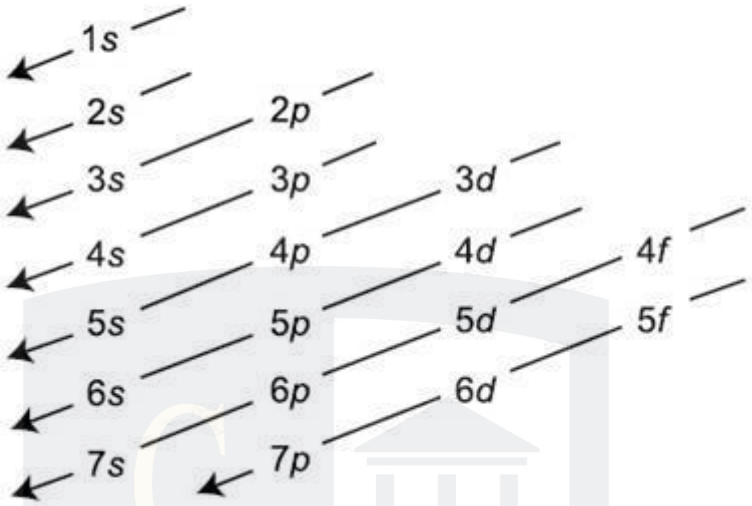
An atom fills sub-shells in order of increasing energy; the 3d sub-shell has higher energy than 4s and gets filled after the 4s.

The order of energy of different orbitals in an atom is given below:

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 4f < 5d < 6p < 7s < 5f < 6d < 7p

Principle Level	1	2	3	4
Sub-Level	1s	2s,2p	3s,3p,3d	4s,4p,4d,4f

Need help remembering this? Try using this trick!

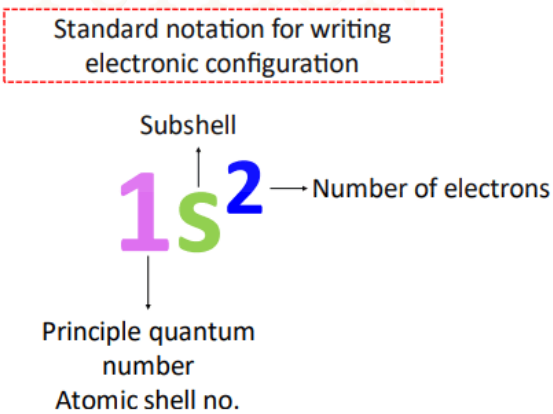


Exam Tip:

Electron configurations of atoms and ions up to Z = 36 in terms of shells and sub-shells (orbitals) s, p, and d might be asked in exams.

Standard notation for writing the electronic configuration of atoms:

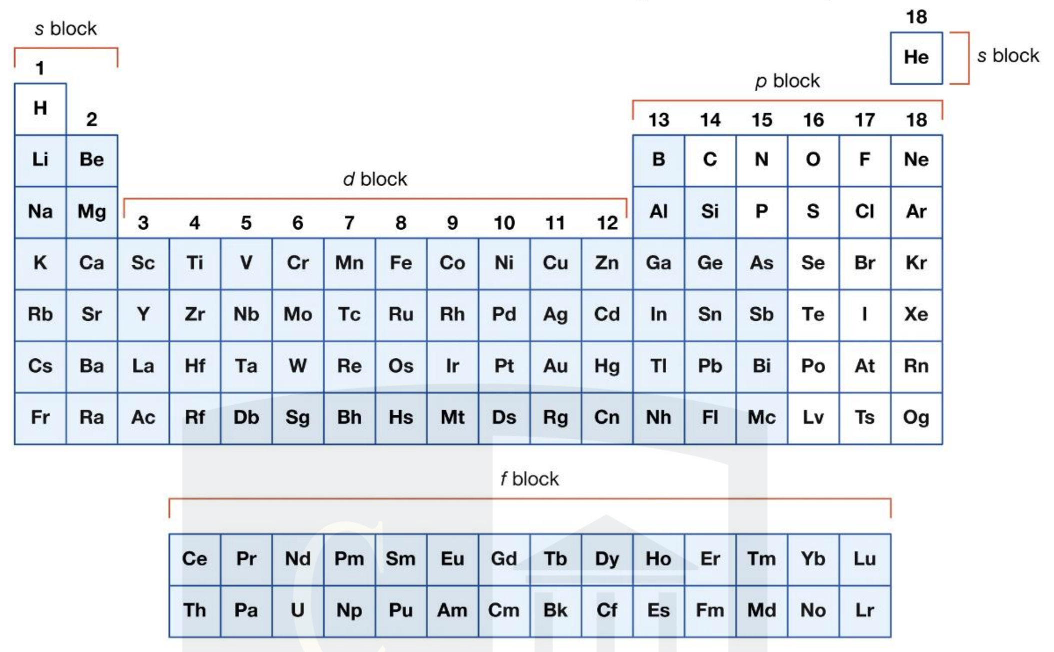
- Principle quantum number (1,2,3,…) to represent the atomic shell number. For example,1 refers to shell no. 1, 2 denotes shell no. 2, and so on and so forth.
- Symbols (s,p,d,f) for representing subshells. For example, 1s represents the first subshell of shell no. 1 (it has only one subshell), 2s stands for the first subshell of shell no. 2 while 2p denotes the second subshell of shell no.2
- Numbers (1,2,3,…) in the superscript next to s,p,d,f denotes the number of electrons present in a specific subshell of an atomic element



Periodic Table

The periodic table of the elements showing the group numbers and the *s*, *p*, *d*, and *f* blocks

Elements in the shaded area are the metals of organometallic chemistry.



The periodic table is organized by dividing it into blocks corresponding to the sub-shells filled by each element's outermost electrons.

S-block element is defined by having the outermost electron filling an s-sub shell, like sodium, with the electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^1$.

P-block element has an outer electron filling a p-sub shell, seen in chlorine with the electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$.

D-block element has its outer electron filling a d-subshell, as in the case of vanadium with the electronic configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^3$

ELECTRONIC CONFIGURATION OF TRANSITION METALS

Transition metals have two unusual rules:

1) An atomic sub-shell that is half-full or full tends to be highly stable in 3d shell.

Chromium and **Copper** exhibit unusual sub-shell filling behavior. They donate one of their 4s electrons to the 3d sub-shell, resulting in a half-full (chromium) or full (copper) d sub-shell that is particularly stable.

Element	Expected configuration	Actual configuration
Chromium	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^4$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$
Copper	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2 3d^9$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$

2) When transition metals form ions, they tend to lose their 4s electrons before losing their 3d electrons.

When atoms other than transition metals form ions, they lose electrons from the 3d sub-shell first, and then the 4s sub-shell as the 3d sub-shell is at a higher energy level. However, transition metals lose electrons from the 4s first.

Element	Expected configuration	Actual configuration
Chromium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
Copper	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$

<u>The table below shows the transition metal's electronic configuration in atom and ion form</u>:

Sc	1s22s22p63s23p6 4s23d1	Sc	$^{3+}$ [Ar] $4s^0 3d^0$
Ti	1s22s22p63s23p6 4s23d2	Ti	$^{3+}$ [Ar] $4s^0$3d1
V	1s22s22p63s23p6 4s23d3	V	$^{3+}$ [Ar] $4s^0 3d^2$
Cr	1s22s22p63s23p6 4s13d5	Cr	$^{3+}$ [Ar] $4s^0 3d^3$
Mn	1s22s22p63s23p6 4s23d5	Mn	$^{2+}$ [Ar] $4s^0 3d^5$
Fe	1s22s22p63s23p6 4s23d6	Fe	$^{3+}$ [Ar] $4s^0 3d^5$
Co	1s22s22p63s23p6 4s23d7	Co	$^{2+}$ [Ar] $4s^0 3d^7$
Ni	1s22s22p63s23p6 4s23d8	Ni	$^{2+}$ [Ar] $4s^0 3d^8$
Cu	1s22s22p 3s23p6 4s13d10	Cu	$^{2+}$ [Ar] $4s^0 3d^9$
Zn	1s22s22p63s23p6 4s23d10	Zn	$^{2+}$ [Ar] $4s^0 3d^{10}$

When Forming Ions

lose 4s before 3d

Exam Question:

What is the electron configuration of V^{2+} in the ground state?

A $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3$

B $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1\ 4s^2$

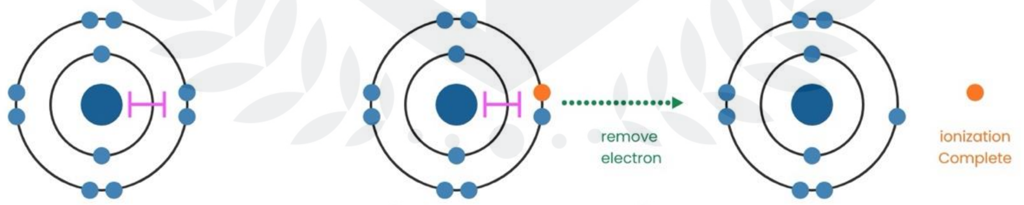
C $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3\ 4s^2$

D $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$

(Please visit the resources section on Chemistryonlinetuition.com for more resources)

IONIZATION ENERGY

Ionization is the process of removing an electron from an atom or a molecule. The process is endothermic because energy is required to break the force of attraction between the electron and the central positive nucleus.

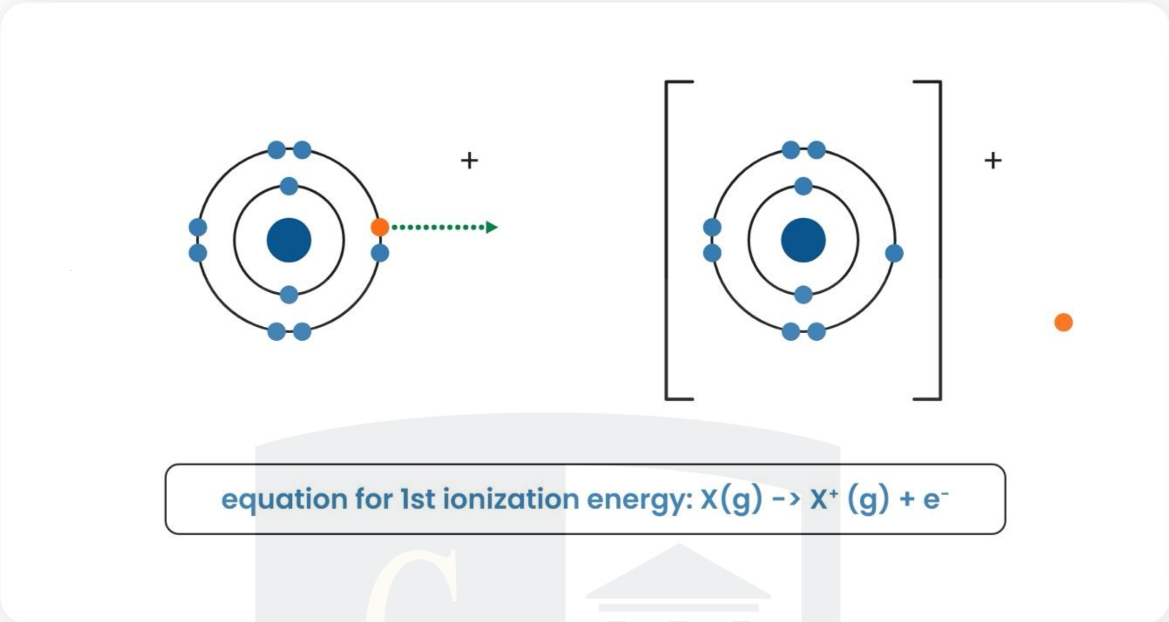


Ionization of an Atom

First ionization energy is the enthalpy change when one mole of gaseous atoms forms one mole of gaseous ions having a single positive charge with a removal of one mole of electrons.

Exam point:

(Learn the first ionization energy definition as it's the exam demand)

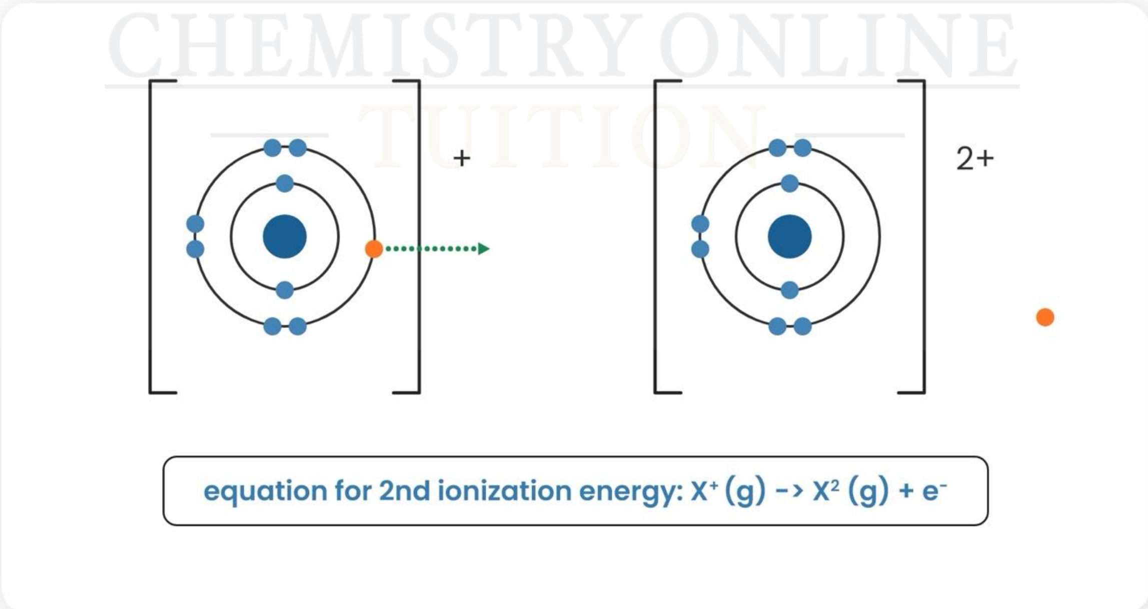


Exam Questions:

Write an equation to show the process that occurs when the first ionisation energy of calcium is measured.

(Please visit the resources section on Chemistryonlinetuition.com for more resources)

Second ionization energy is the enthalpy change when one mole of gaseous ions with a single positive charge forms one mole of gaseous ions with a double positive charge.



Factors that affect Ionization energy

- **Atomic Radius** - When the atomic radius is higher, the ionization energy is lower because the outer electrons are further from the nucleus, reducing the attractive pull from the nucleus.

- **Nuclear Charge** – the **higher** the nuclear charge, the **higher** the ionization energy. This is because the greater the positive charge of the nucleus, the stronger the attraction for the outer electrons.

- **Shielding effect**– An atom's inner shells affect its ionization energy. Inner shells are the electron shells between the outermost shell and the nucleus. For instance, if the outermost electron is in the fourth shell, then three inner shells repel the outer electrons, reducing the attraction towards the nucleus. This phenomenon is known as shielding.

Exam tip:

Learn all these three points affecting ionization as they would be required to answer the questions.

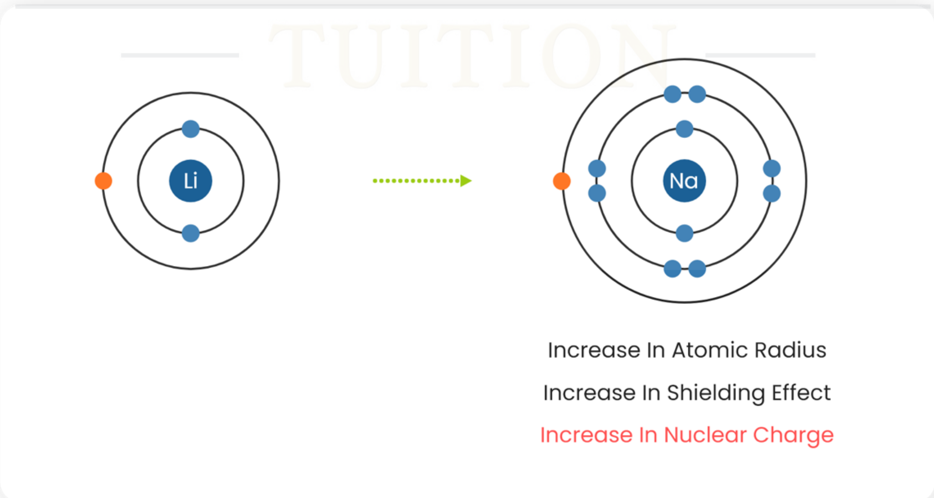
Trend Down a Group

Down a group in the periodic table, the first ionization energy increases, meaning less energy is required to remove an electron from one mole of gaseous atoms. Let's talk about each factor:

- **Atomic radius** - As you move down a group in the periodic table, the atomic radius increases. The reason for this is that an extra electron shell is added each time, which results in a greater distance between the nucleus and the outermost electron. Consequently, the attraction between the outermost electron and nucleus decreases.

- **Nuclear charge** - **increases** down a group, but the two factors above outweigh the increasing nuclear charge.

- **Shielding effect** - As you go down a group in the periodic table, the shielding effect increases. This is because an extra electron shell is added, leading to an increase in the number of inner electron shells. As a result, the attraction between the outermost electron and the nucleus decreases due to more repulsion between the electrons.



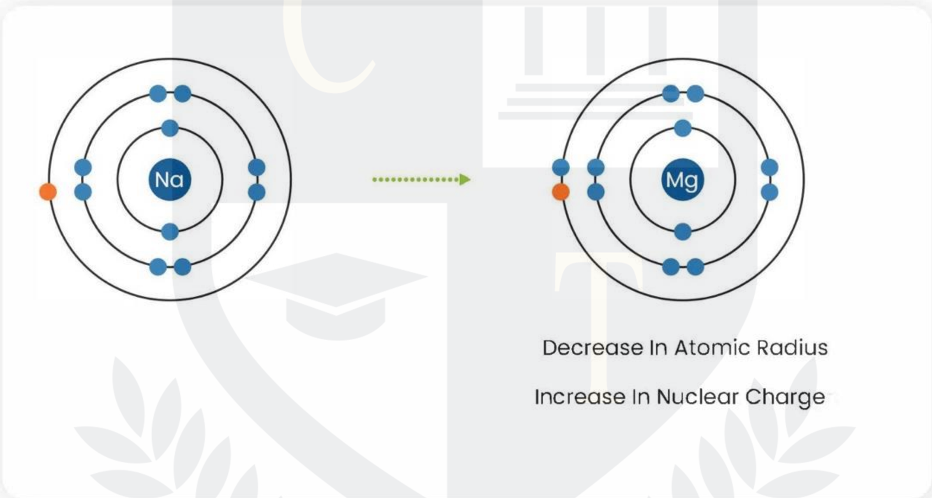
Trend Across a Period

First, ionization energy increases across a period. This is due to two factors:

- **Nuclear charge** – The nuclear charge increases across a period because the atomic number and number of protons in the nucleus increase.

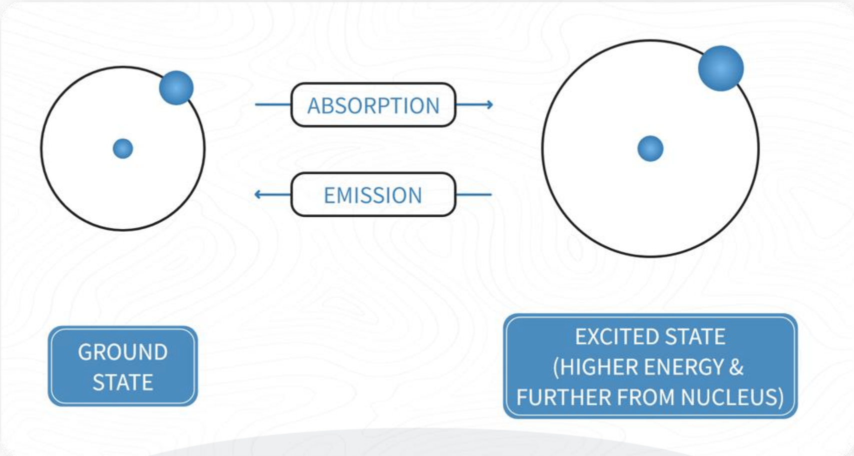
- **Atomic radius** – the atomic radius **decreases slightly** as you go across a period. This is a subtle increase because the nuclear charge of the nucleus increases, which means that the nucleus pulls the outer electrons closer to it, which decreases the atomic radius.

- **Shielding effect** - There is no change in the shielding effect across a period. Electrons are added to the same outer shell as you go across a period.



Evidence for electronic Configuration from emission spectra

- The electrons orbit around the nucleus at a high speed and are arranged in different energy levels called shells.
- If an electron's energy increases, it can move to a higher energy level.
- The electron transition process is reversible, meaning that electrons can move back to their original energy levels.
    - When this occurs, it causes the release of energy.
- The energy frequency remains the same; it is emitted, not absorbed.



The difference between absorption and emission depends on the direction of electron energy level transitions.

- If the electrons jump from a higher to a lower energy level, they emit radiations of a particular frequency.
- The emission of radiation in the visible region produces a line emission spectrum.

Line emission spectra

- Each line represents a specific energy level, indicating electrons have limited energy options.

- These packets of energy are called '**quanta**' (plural **quantum**)
- Notice that the lines on the spectrum become closer as they move towards the blue end
- This set of lines is called convergence, as they converge towards the higher energy end, indicating that the electron has reached its maximum energy.
- The maximum value represents the amount of energy required to remove an electron from an atom, which is also known as the ionization energy.
- Swiss school teacher Johannes Balmer discovered these lines, which are named after him.
- We now know that these lines correspond to electrons transitioning from higher energy levels to the second, or n=2, level.

<u>The ionization energies of an element provide valuable insight into its electronic structure</u>.

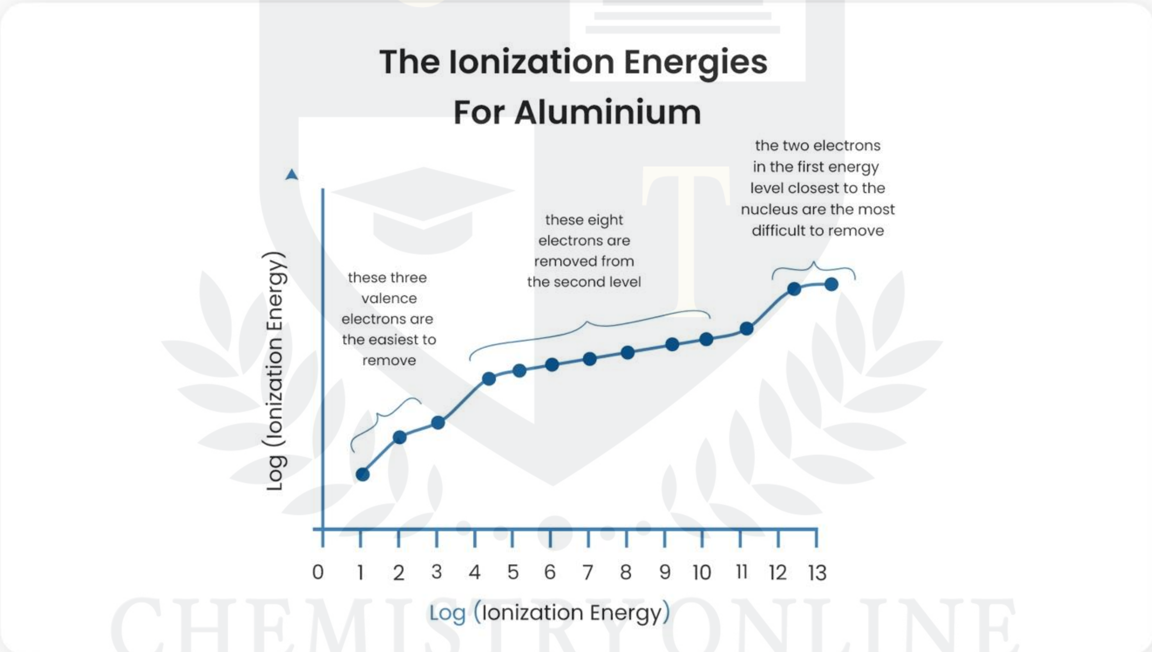
- **Successive ionization energies increase between shells.**
  For instance, in the case of sodium, which has one outer electron, there is a significant increase in ionization from the first to the second electron. This is because the second electron is being removed from a shell closer to the nucleus, which means there is a stronger attraction from the nucleus.

- **Successive ionization energies increase within each shell.**
  Each time an electron is removed, the ionization energy increases because you are removing an electron from an increasingly positive ion. For example, for magnesium, which has two outer electrons, the second ionization energy is greater than the first because it is harder to remove an electron from an Mg^{2+} ion than an Mg+ atom because there is a stronger attraction experienced by the remaining electrons.

Interpreting the ionization energy data for Shell structure

In an atom, each time there is a sudden increase in the ionization energy, it indicates the entry into a new electron shell. Referring to the diagram given below, we can observe that the outermost shell contains 3 electrons, followed by 8 electrons in the next inner shell, and so on.

Moreover, by counting the number of electrons removed before the first sudden rise in ionization energy, we can determine the group of the element in the periodic table. The element belongs to Group 3 in this particular diagram, as it has three outer electrons. Find below the ionization energy graph for Aluminium.



According to this graph for sodium, there are 3 electrons removed before the first sudden rise, then 8 before the next rise, and finally 2 electrons, so the electronic structure for aluminum is:

Shell	Number of electrons
1st	2
2nd	8
3rd	3

Exam Question

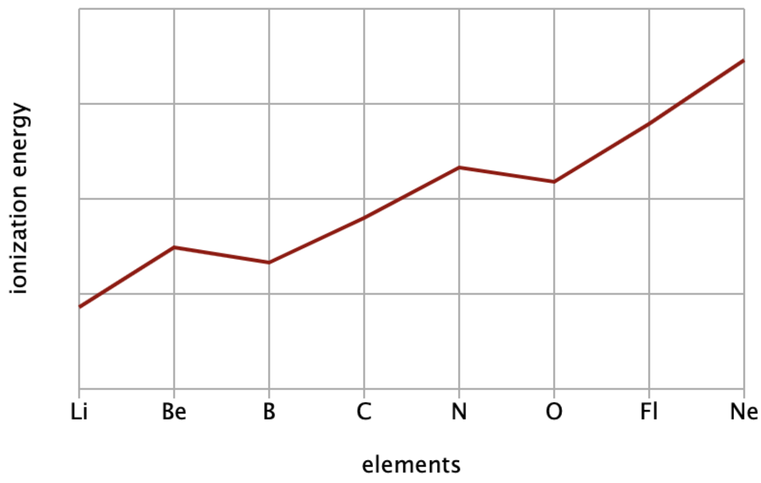
Which group does the element belong to?

	1	2	3	4	5
Ionisation energy kJ mol^{-1}	790	1350	5140	6680	8320

There is a big jump between the 2nd and 3rd ionization energies, which means that this element must be in group 2 of the periodic table as the 3rd electron is removed from an electron.

Exam Points to Remember:

The following graph shows the ionization energy for the second period. As a general trend, there is an Increase in ionization energy, but there are dips between beryllium and boron, as well as between nitrogen and oxygen. Can you explain the reason behind these dips?



Drop between Groups 2 and 3

There is a drop in first ionization energy between Group 2 and 3 elements. For example, Beryllium (Group 2) and Boron (Group 3) exhibit this trends.

Element	Electrons	Configuration
Beryllium	4 Electrons	$1s^2 2s^2$
Boron	5 Electrons	$1s^2 2s^2 2p^1$

We would expect Beryllium's first ionization energy to be higher than Boron's due to increased nuclear charge and decreased atomic radius. Isn't it?

Boron has its outer electron located in a 2p sub-shell, which means it is positioned at a higher energy level compared to the 2s sub-shell. Due to this, the electron is found farther away from the nucleus, resulting in a higher atomic radius. Moreover, it is also subject to more shielding from inner electrons.

Two factors lead to a lower first ionization energy of Boron than Beryllium, despite having more protons in the nucleus.

The drop between Group 2 and 3 provides **evidence** for electron sub-shells. The change in the usual pattern of ionization energies can only be proven by the presence of sub-shells.

Key Exam point:

Explain how Group 2 (Be–Ba) ionization gives evidence for electron configuration in sub-shells and in shells.
''The drop between Group 2 and 3 provides **evidence** for electron sub-shells. The change in the usual pattern of ionization energies can only be proven by the presence of sub-shells'' – In order to obtain full marks, please provide additional explanation about the cause of the drop.

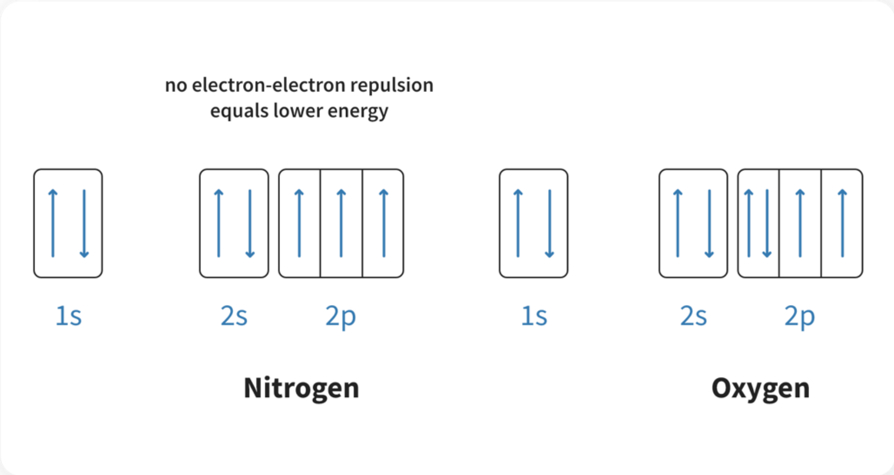
Drop between Groups 5 and 6

There is a drop in the first ionization energy between Group 5 and Group 6 elements. For instance, consider Nitrogen (Group 5) and Oxygen (Group 6):

Element	Electrons	Configuration
Ntrogen	7 Electrons	$1s^2 2s^2 2p^3$
Oxygen	8 Electrons	$1s^2 2s^2 2p^4$

We now understand that the 2p sub-shell consists of three orbitals.:

- Each orbital in nitrogen contains one electron.
- In oxygen, there are two orbitals each with one electron and one with two electrons.



In nitrogen, an electron is removed from an orbital with one electron, while in oxygen, it is removed from an orbital with two electrons.

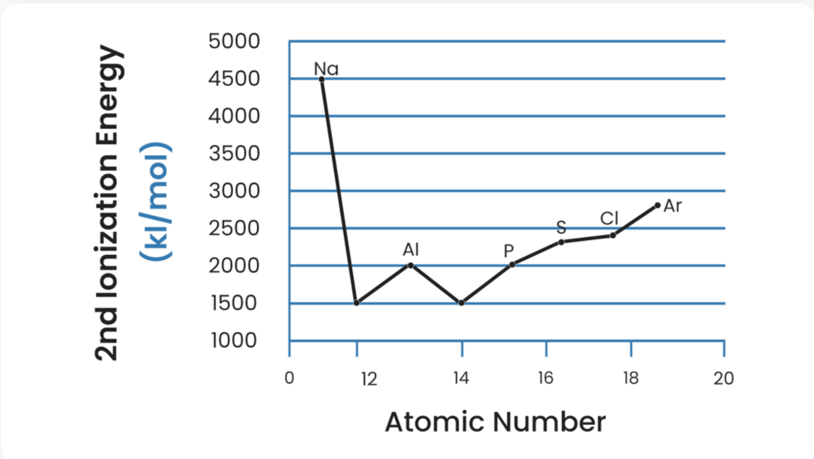
When two electrons are in the same orbital, they repel each other, which makes it easier to remove them. As a result, electrons in fully-filled orbitals are easier to remove than those in half-filled orbitals. This is why oxygen has a lower first ionization energy than nitrogen, meaning less energy is required to remove the first electron from oxygen than nitrogen .

Exam Question:

Why is there a small drop from P to S?

Patterns in the second ionization energy

If the graph of the second ionization or each successive element is plotted, then a similar pattern to the first ionization energy is observed, but all the elements will have shifted one to the left.



In period 2 of the periodic table, Lithium would have the second largest ionization energy among all elements. This is because its second electron would be removed from the first 1s shell, which is closest to the nucleus and has no shielding effect from inner shells. As a result, Lithium requires more energy to remove its second electron than Helium, which has a smaller second ionization energy due to having fewer protons.

