Trishit Dedhia

2020
e-Book
Edition

PERIODICITY CRACKED

PART I

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The author of this document is Trishit Dedhia.

Special thanks to Anupa Subramanian, IBDP chemistry teacher at Don Bosco International School.

If you have any doubts, queries or feedback, you may reach the CrackBox Team using email: crackboxofficial@gmail.com.

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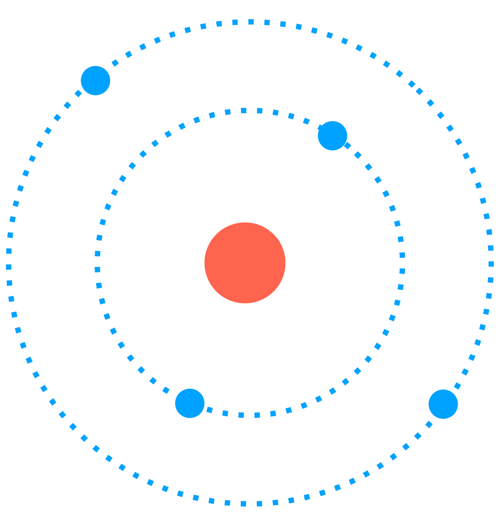

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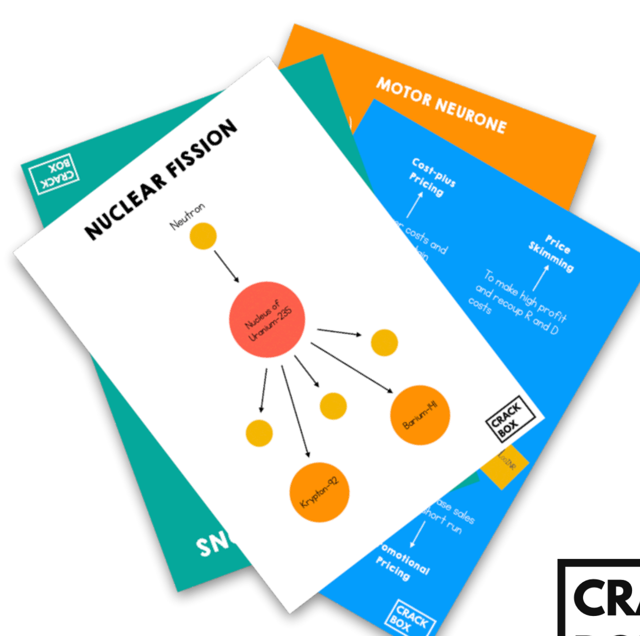

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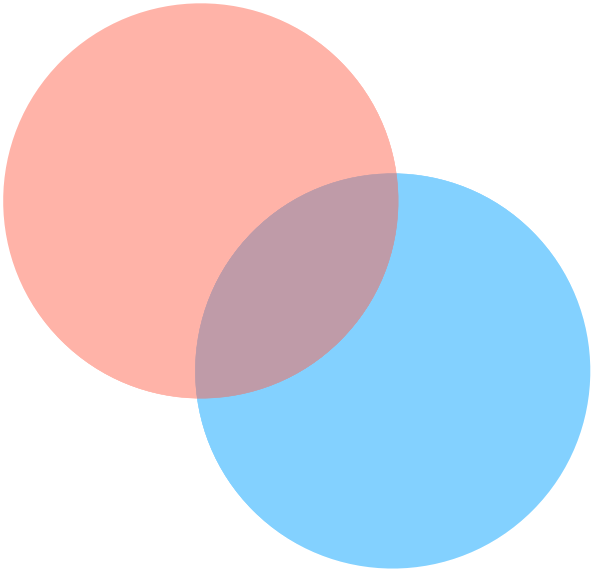

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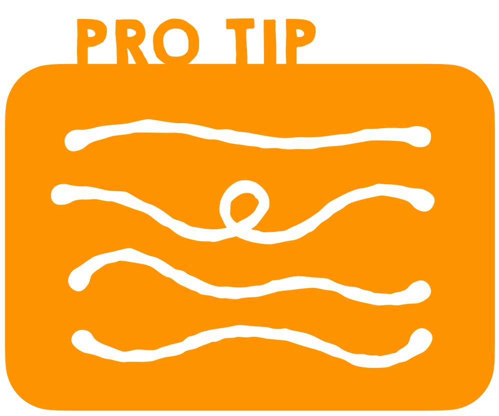

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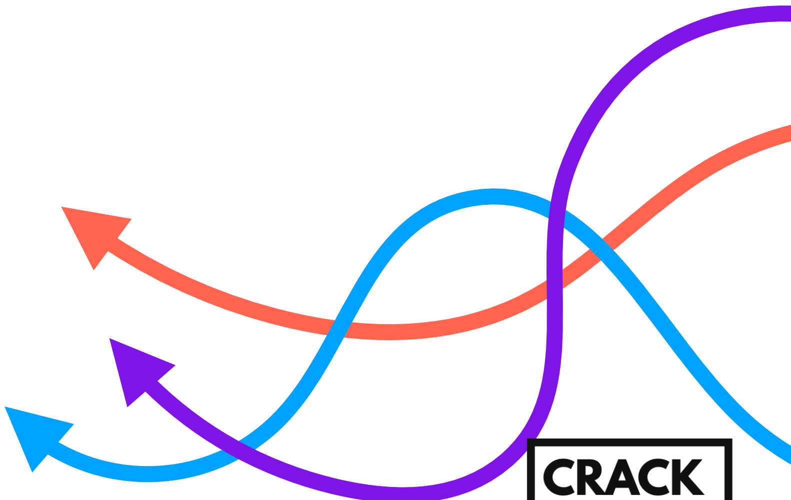

SUBCONTENTS

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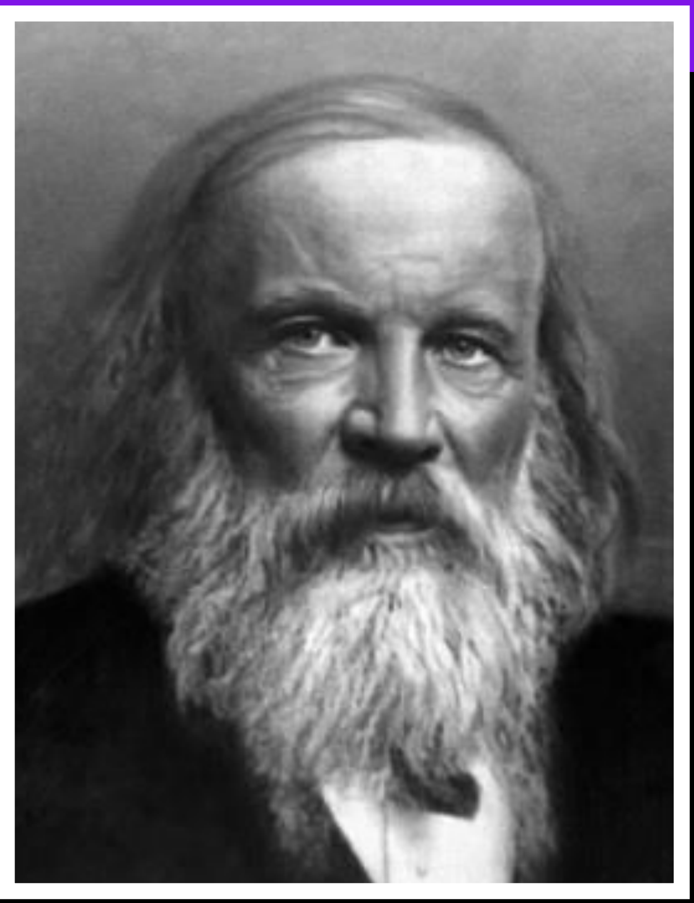
The Periodic Table

The periodic table is indeed a scientific wonder! The periodic table we study today is an adaption to Dimitri Mendeleev's who was the first scientist to arrange the elements in their atomic number while leaving gaps between elements for undiscovered elements. Not just that, before even experimenting of discovering the missing elements, Mendeleev predicted their physical and chemical properties!

A Tinge of History



DMITRI IVANOVICH MENDELEEV

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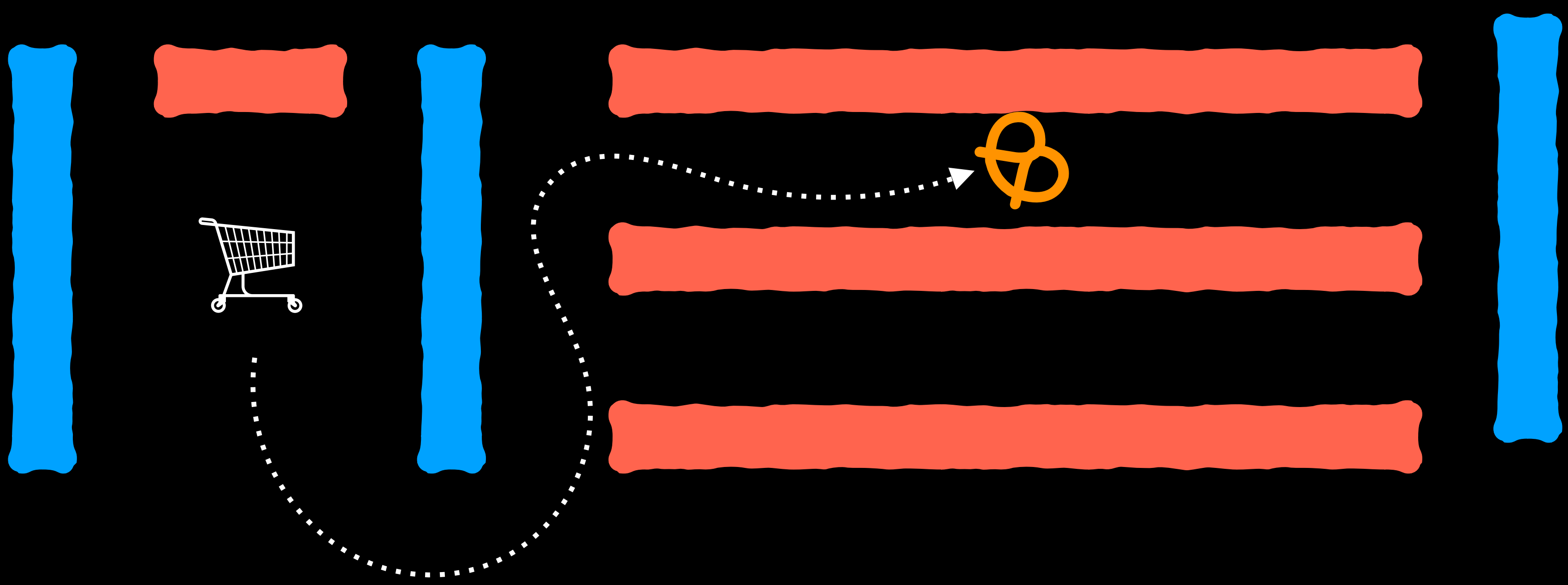
A supermarket has multiple aisles and racks of stuff neatly organised They have sections for food items, with which there are rows for veggies, beverages, etc.



A periodic table is similar with multiple rows and columns called periods and groups respectively.

The elements are placed in order of increasing atomic number(Z). The position of the element is also related to its electronic configuration.

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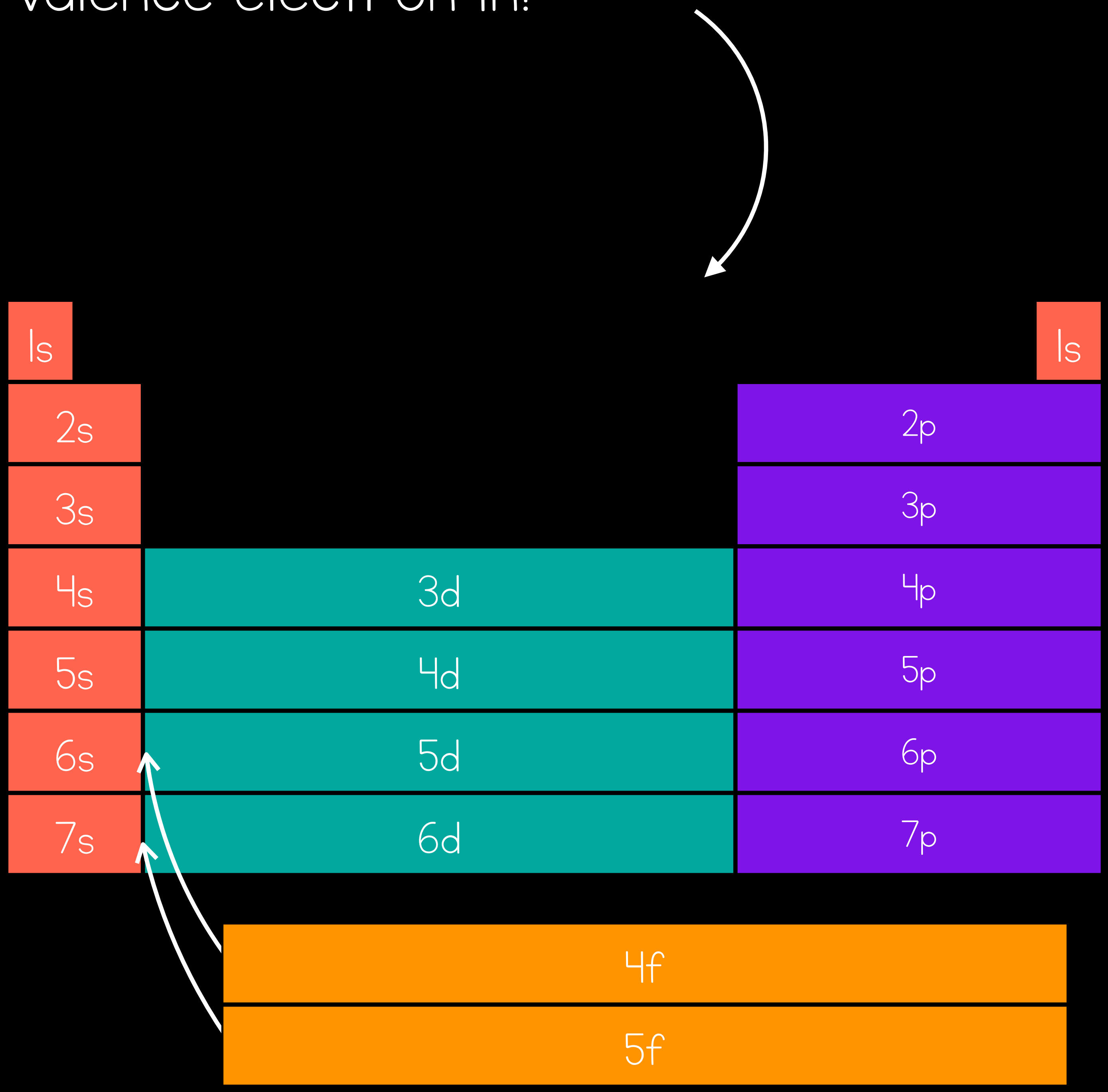
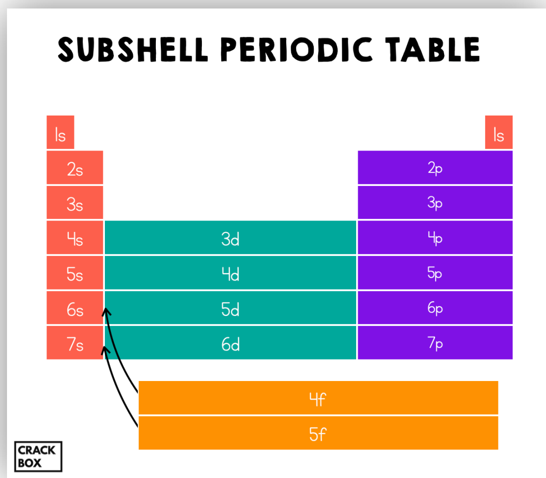
The periodic table is very interestingly also categorised as per which sub shell is the valence electron in!





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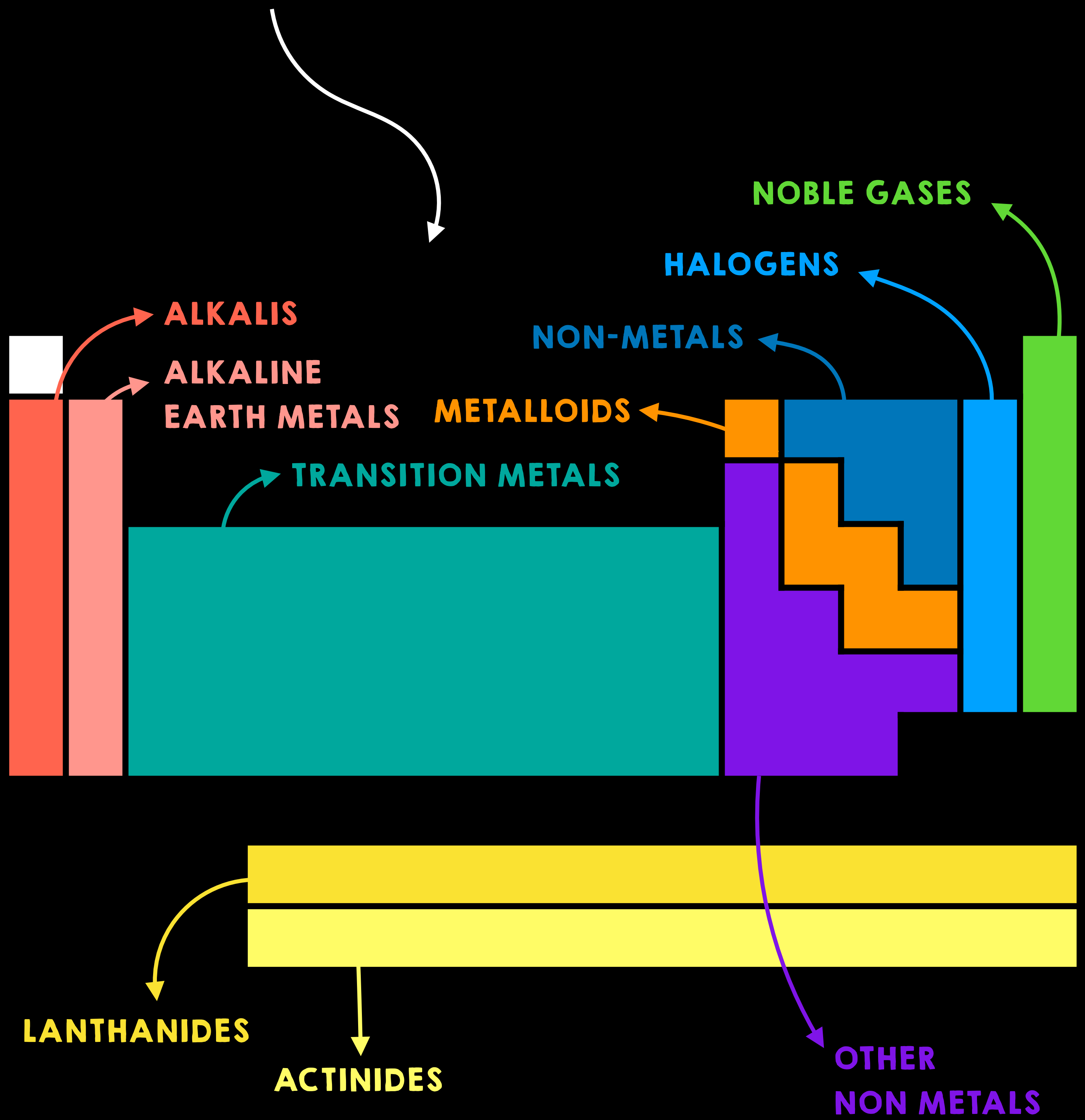
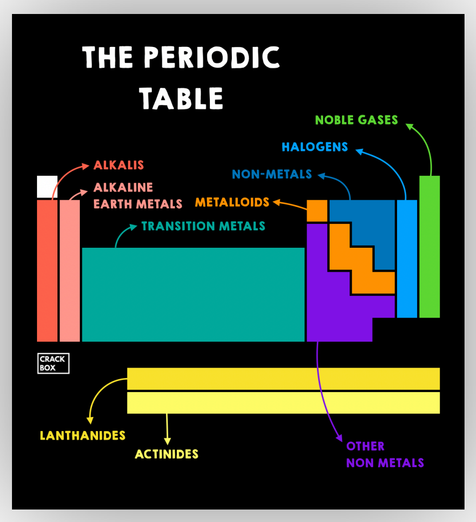
The table is also divided into various parts like metals, non-metals, etc.





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The non-metals can be found on the upper right side of the p-block on the table.

The halogens for example are a reactive group of the non-metals unlike the noble gases, which too are non-metals, but don't react.

The metals are found on the left side of the table in the s-block and the transition metals can be found in the d-block.

The alkali metals is a group of reactive metals found on the first column to the left.

The diagonal staircase separating the metals and non metals are the metalloids. They share the characteristics of both, metals and non-metals. Although, physical properties are usually metallic while chemical properties are mostly non-metallic.

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PHYSICAL TRENDS

Fascinatingly, the period table has all the elements arranged according to their properties as well, elements can decrease or increase in a property periodically along groups and periods! A very important property of an atom is its nuclear charge.

Nuclear charge & Shielding Effect

Just to get things simpler, there are 3 different properties:

- Nuclear Charge
- Shielding Effect
- Effective Nuclear Charge

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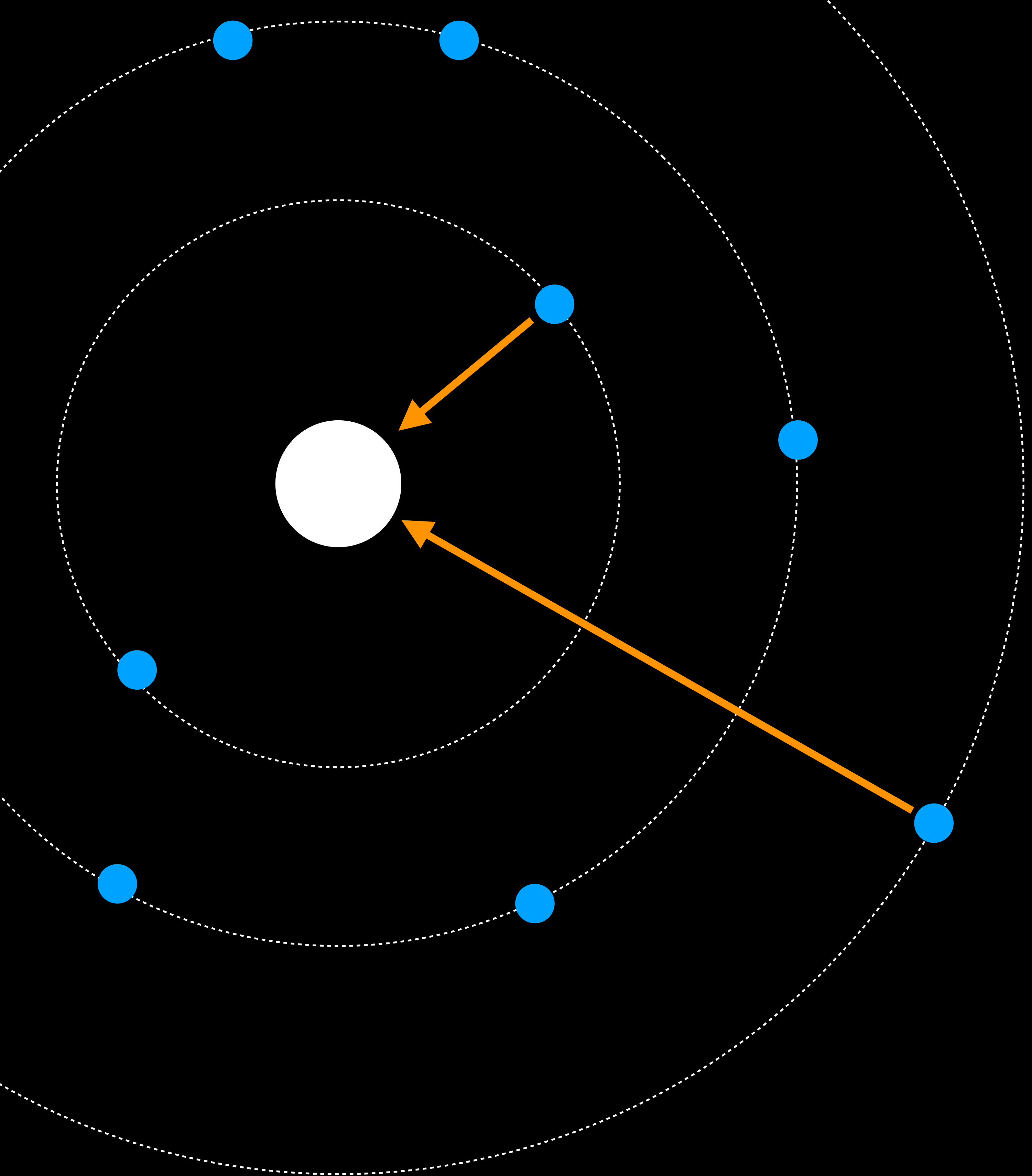
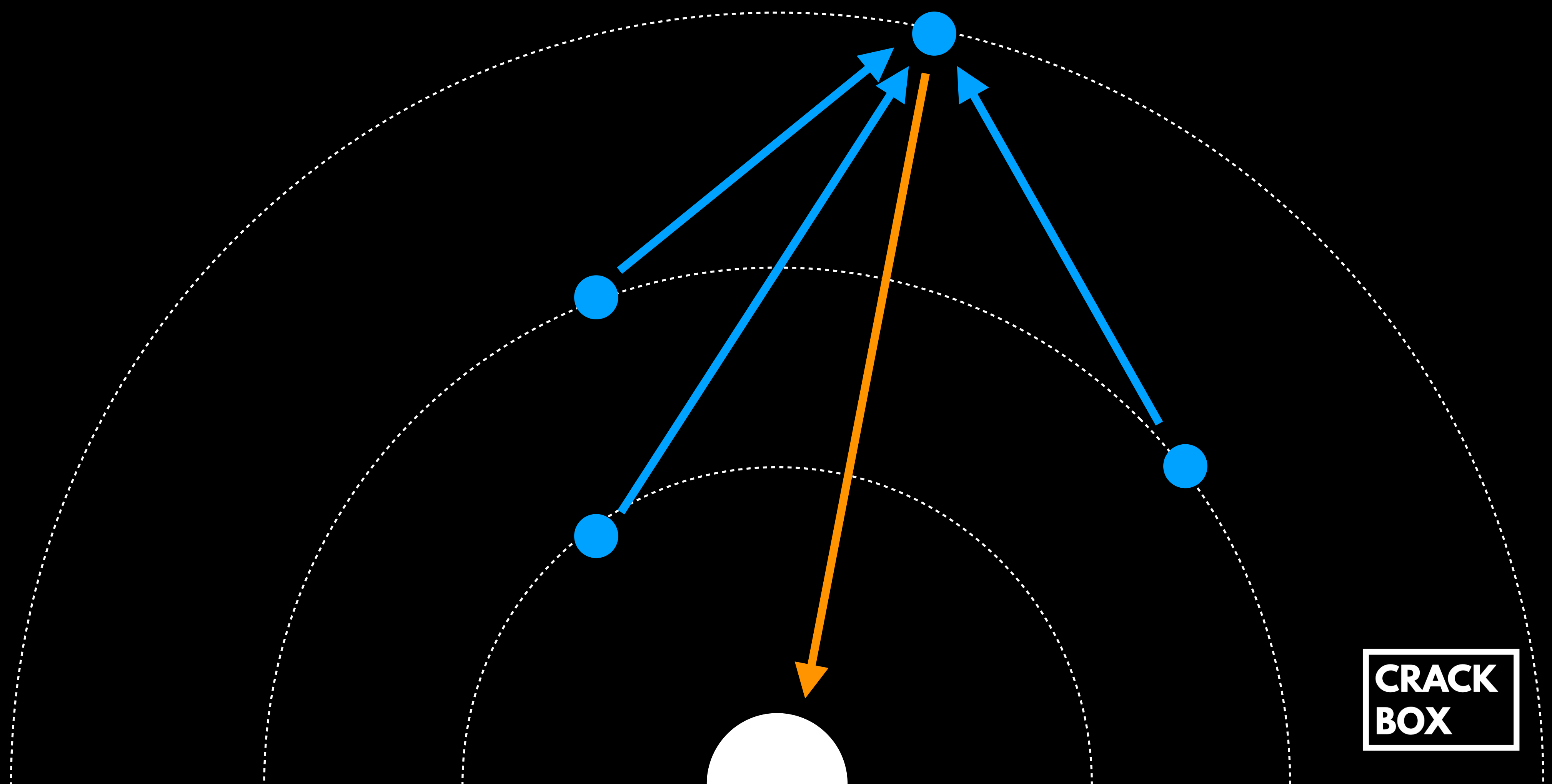
Every atom has a positively charged nucleus that attracts the negatively charged electrons. That charge experienced by an electron is known as Nuclear Charge.



The value of Nuclear Charge is given by the Atomic Number of an element and hence the Nuclear Charge increases as the Atomic Number (Z) increases.

Now, the only force isn't the positively charged nucleus of the atom, the tiny electrons also do have an impact due to the negatively charged electrons repelling other electrons in the atom. The force experienced by an electron from the other inner shell electrons is known as the Shielding Effect.



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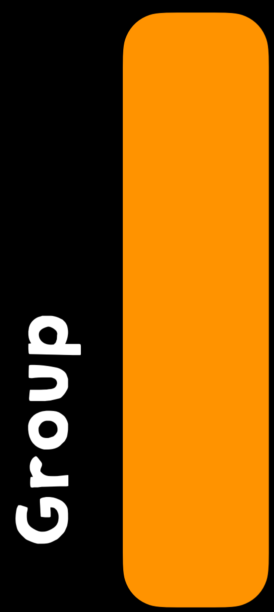
Now, the Shielding Effect acts against the Nuclear Charge, and hence the actual pull experienced by an electron towards the nucleus is known as the Effective Nuclear Charge(Z_{eff}). Where this can be deduced by subtracting the number of inner shell electrons(s) from the atomic number(Z).

$$Z_{eff} = Z - s$$

The Effective Nuclear Charge increases across a period.

Element	Na	Mg	Al	Si
Z_{eff}	+1	+2	+3	+4

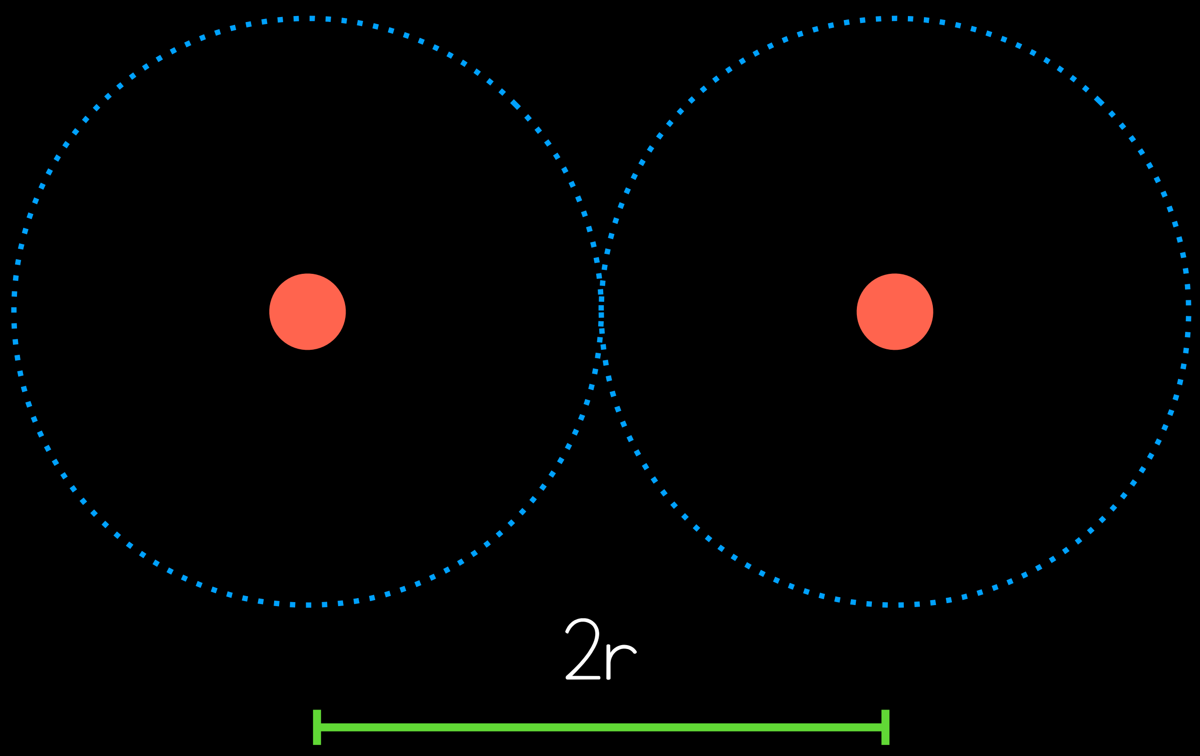
The Effective Nuclear Charge stays the same down the group because the inner electrons increase by 8 and so does the proton number.



Z	Element	Z_{eff}
3	Li	+1
11	Na	+1
19	K	+1

Atomic & Ionic Radius

An atom doesn't have sharp boundaries, we can only estimate the probability of finding an electron; hence, the atomic radius is defined has half distance between neighbouring nuclei and not the distance between the nucleus to the outmost electron.



Ionic radius on the other hand is the same thing as atomic radius but between two neighbouring nuclei of ions.

What affects the size of an ion's radius?

An ion's radius depends on factors such the number of energy levels and the effective nuclear charge.

More energy levels means a larger ionic radius. Hence, the radius increases down the group as more energy levels are added.

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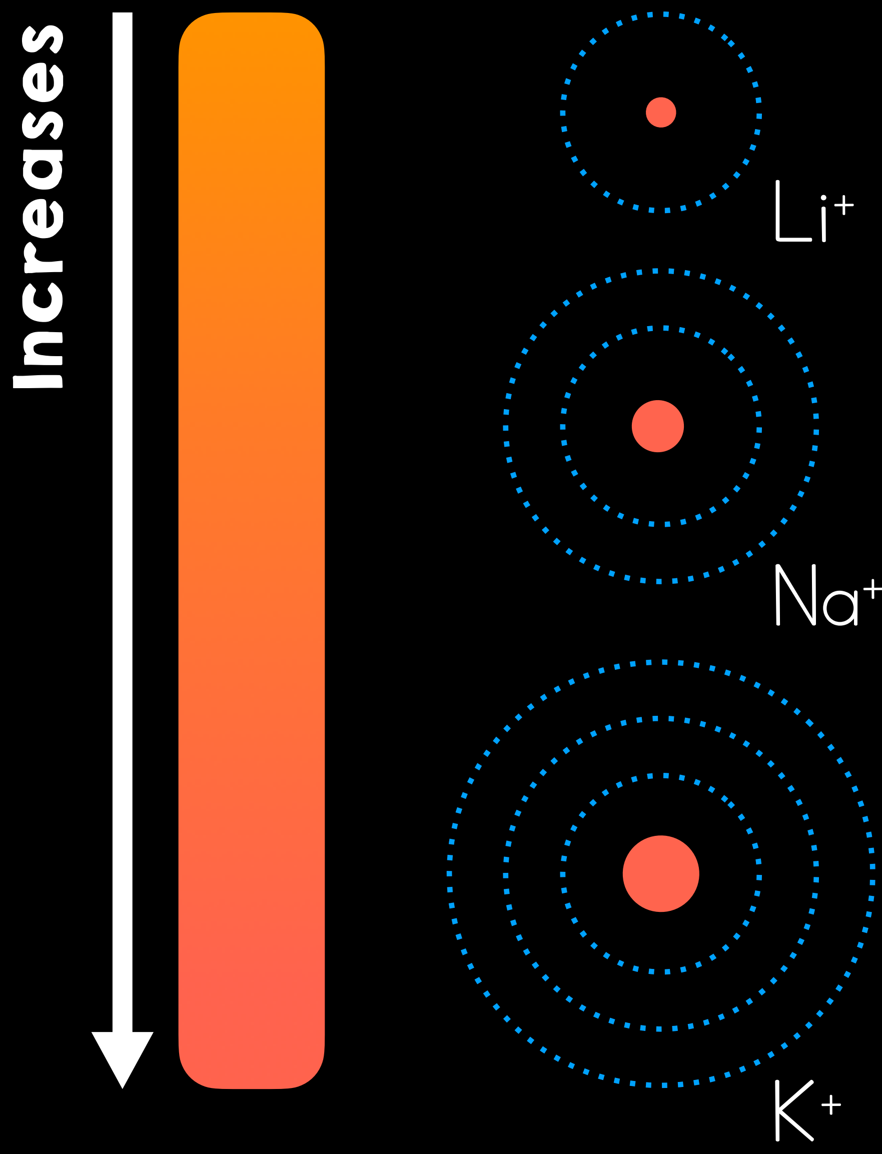

Electronic Config:

$1s^2$

$1s^2\ 2s^2\ 2p^6$

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

When we are looking at the ionic radius across a period, it generally decreases because the ions are isoelectronic as the electron number stays same but the proton number keeps increasing. This causes a larger positive charge that is able to pull the electrons closer. However, there is a small exception at group 4.

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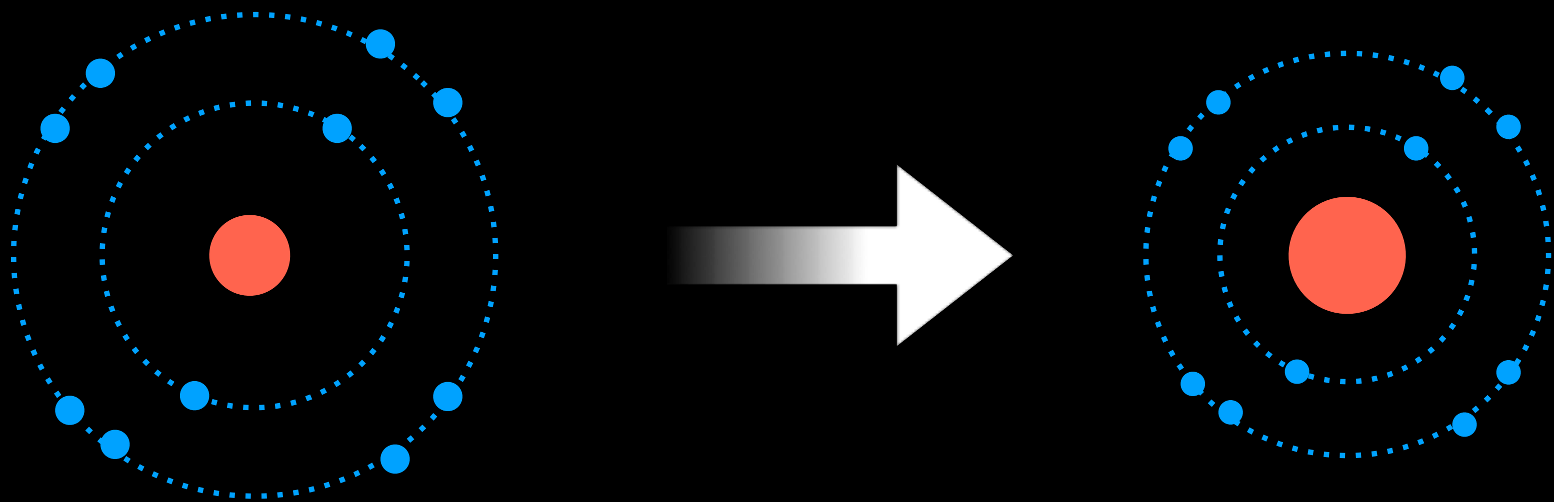
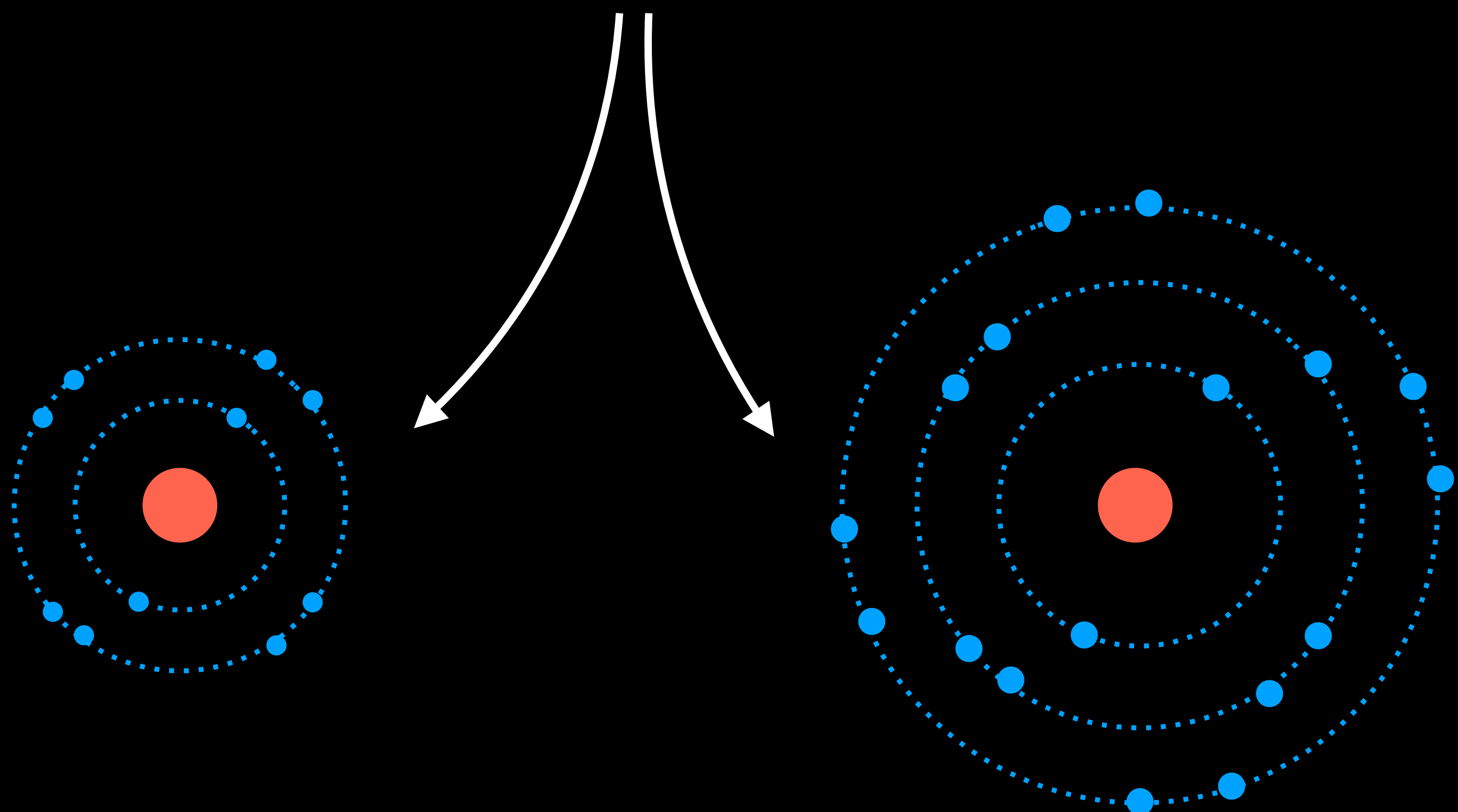


Na^{+1} Mg^{+2}

Once we reach an element from group 4, for example silicon, there are 2 possible ions that we can have:



Si^{+4} Si^{-4}

You can see that the Ionic Radius drastically increases because of Silicon having a whole new energy level. Now if we put all the ionic radius in a row for period 3 elements, we can see that the Ionic Radius in general, decreases; but there is a jump at group 4.

Ionisation Energy

Ionisation Energy is the energy required to knock off the valance electron (The very last electron).

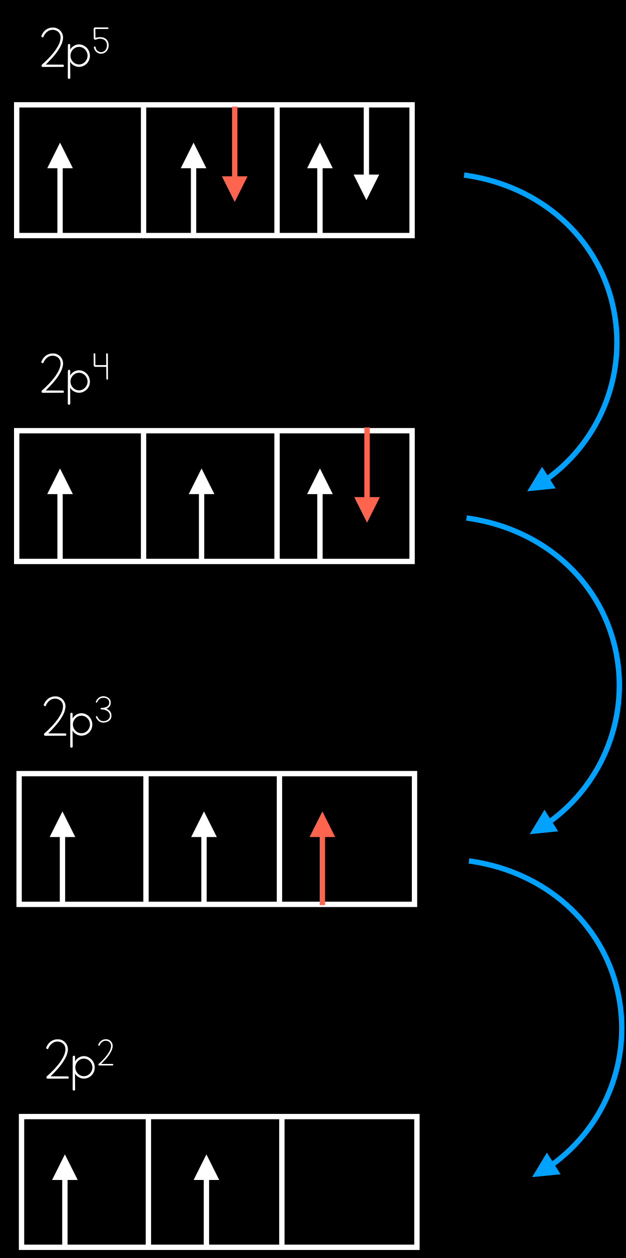
Determining which electron gets knocked next is simple. The electron from the very last sub shell which is preferably unpaired.

Fluorine $1s^2\ 2s^2$

Fluorine $1s^2\ 2s^2$

Fluorine $1s^2\ 2s^2$

Fluorine $1s^2\ 2s^2$



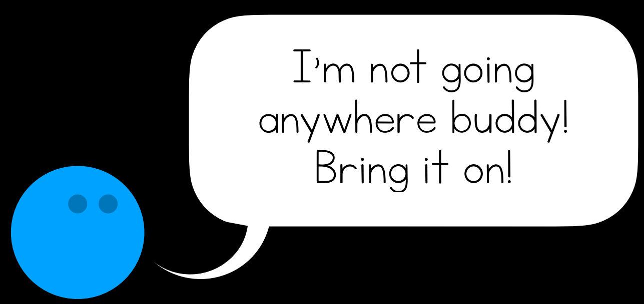
Every time an electron is knocked

We can see that the paired ones are removed first.

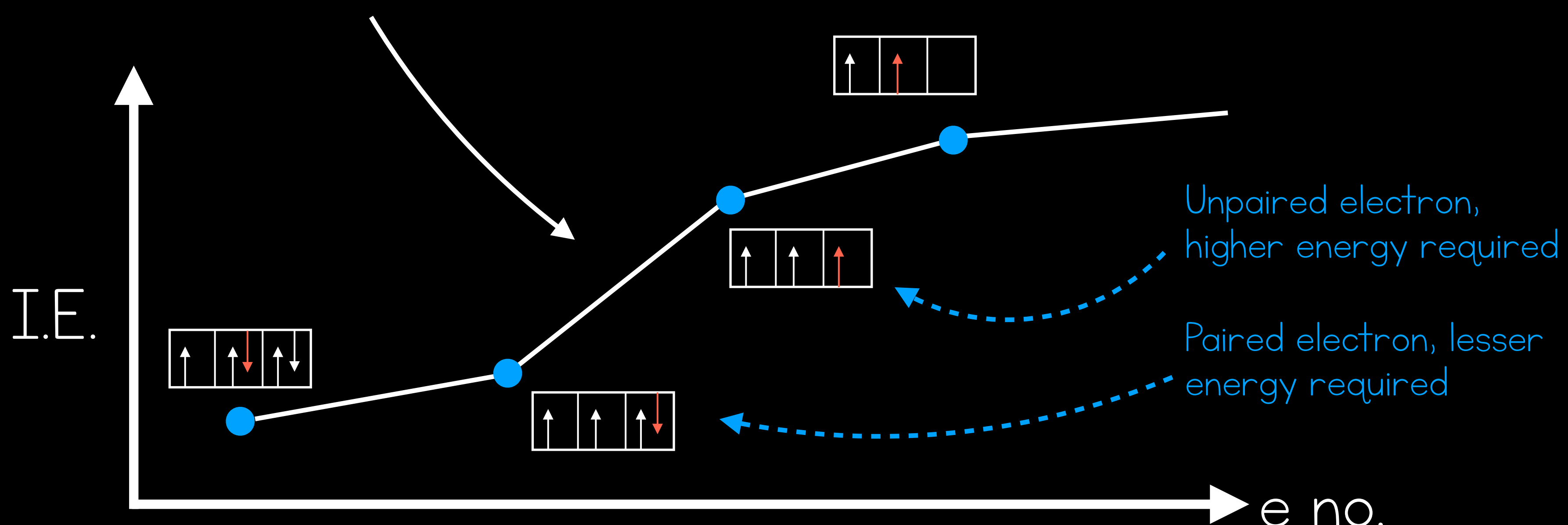
And after 3 successful knocks we can get F^{+3}

When an electron is unpaired, the energy required to remove it is greater compared to an electron that is paired.

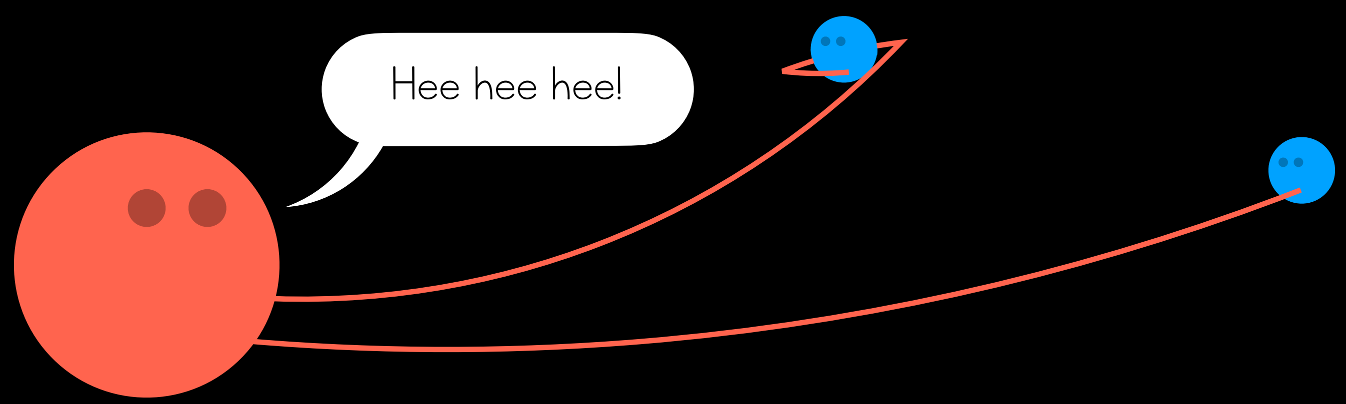
Because a paired electron already has an electron moving in the opposite direction exerting a repulsive force, hence it becomes easier to knock off that electron.



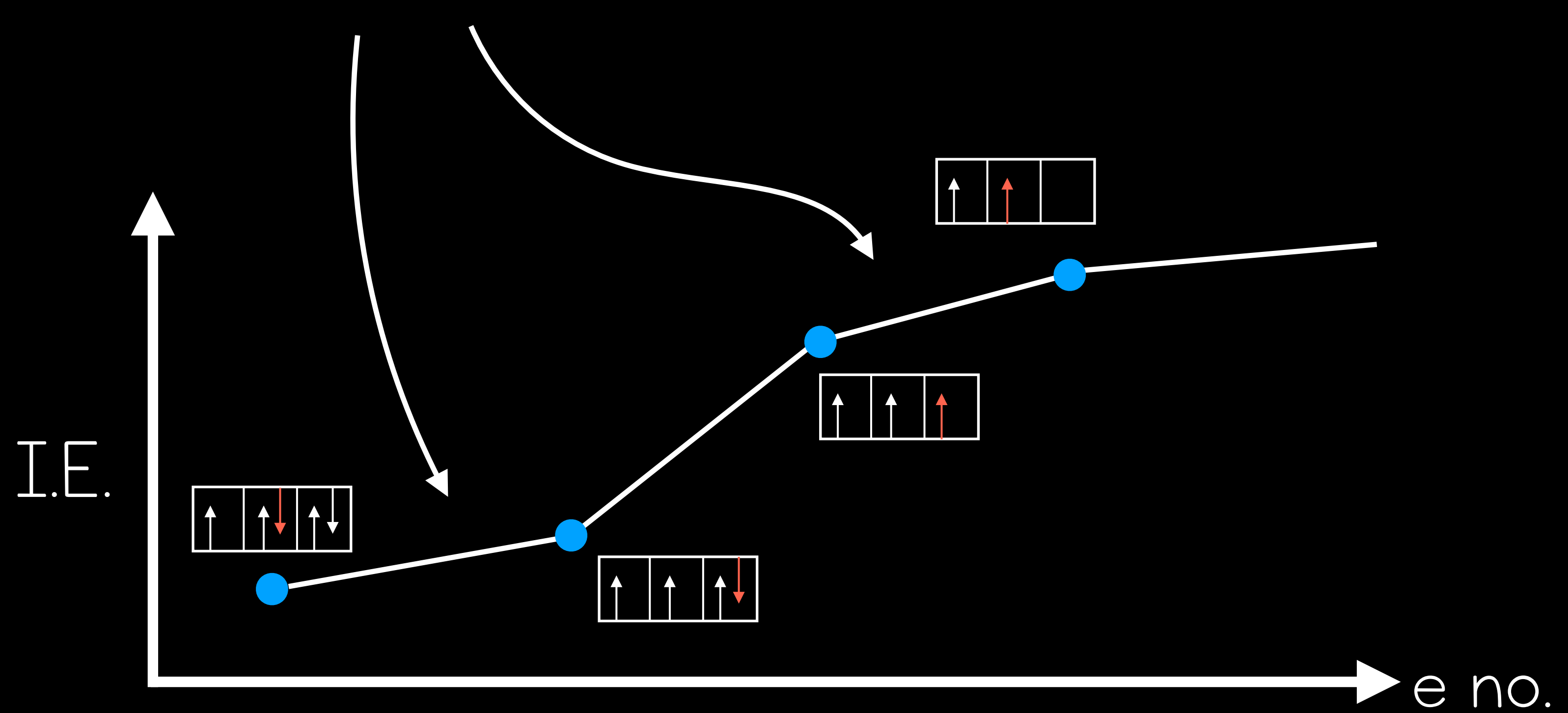
Hence, if we were to make a graph, of I.E. against electrons in fluorine ion, it would have a bump here.



Another factor affecting I.E. is the Effective Nuclear Charge, More the nucleus's pull, more the I.E. required to knock the electron off.



Hence, the I.E. always <u>increases</u> a little bit as an ion loses an electron. Because the Effective Nuclear Charge <u>increases</u>. That is why the graph is never flat.



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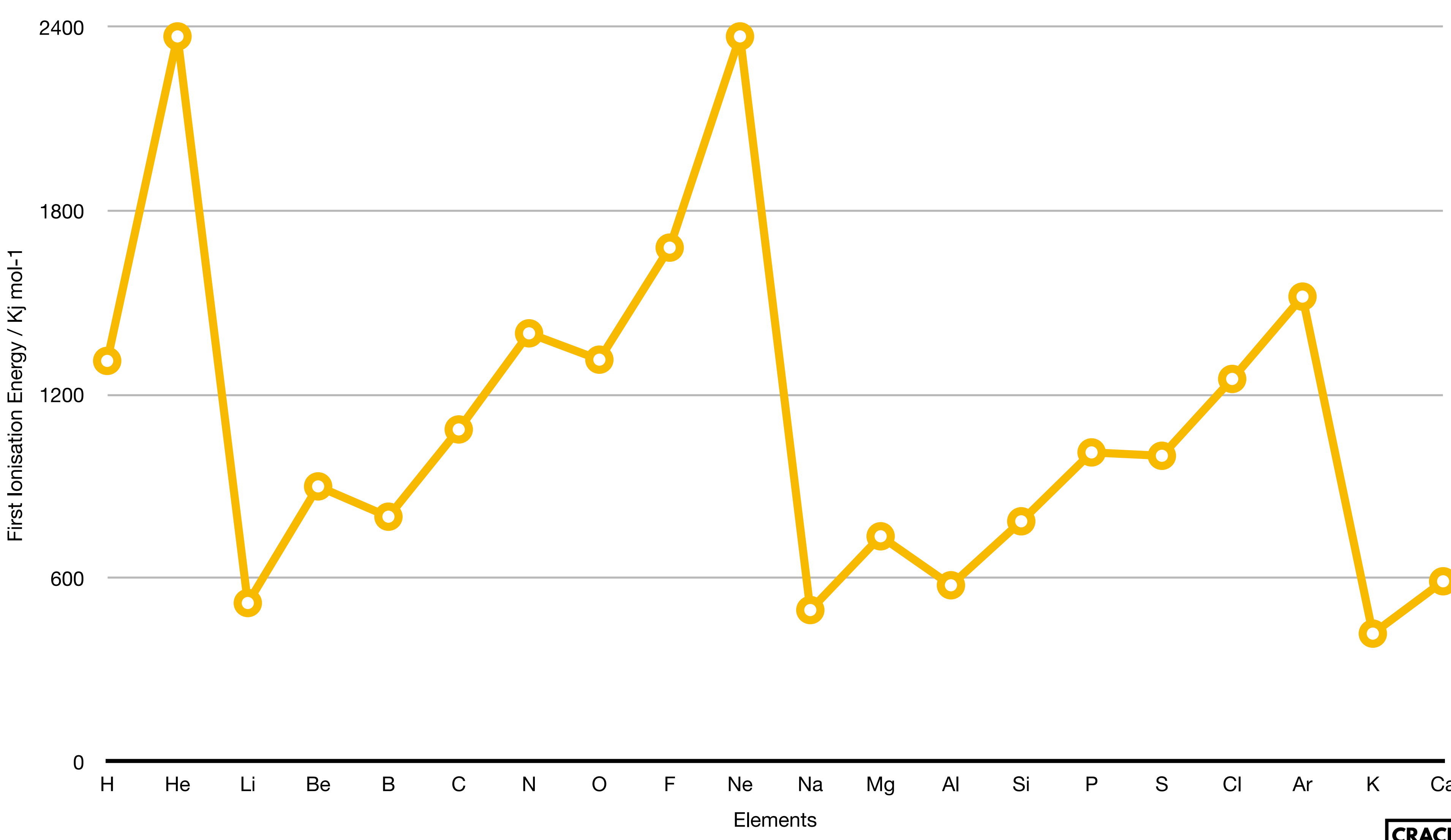
The I.E. required for the first (the furthermost) electron of an atom, is called the First Ionisation Energy.

For example, if ionising Li to produce Li^{+} took 500 kJ/mol^{-1}, this 500 kJ/mol^{-1} is your First Ionisation Energy.

Let's now look at a graph of the First Ionisation Energy of different atoms.



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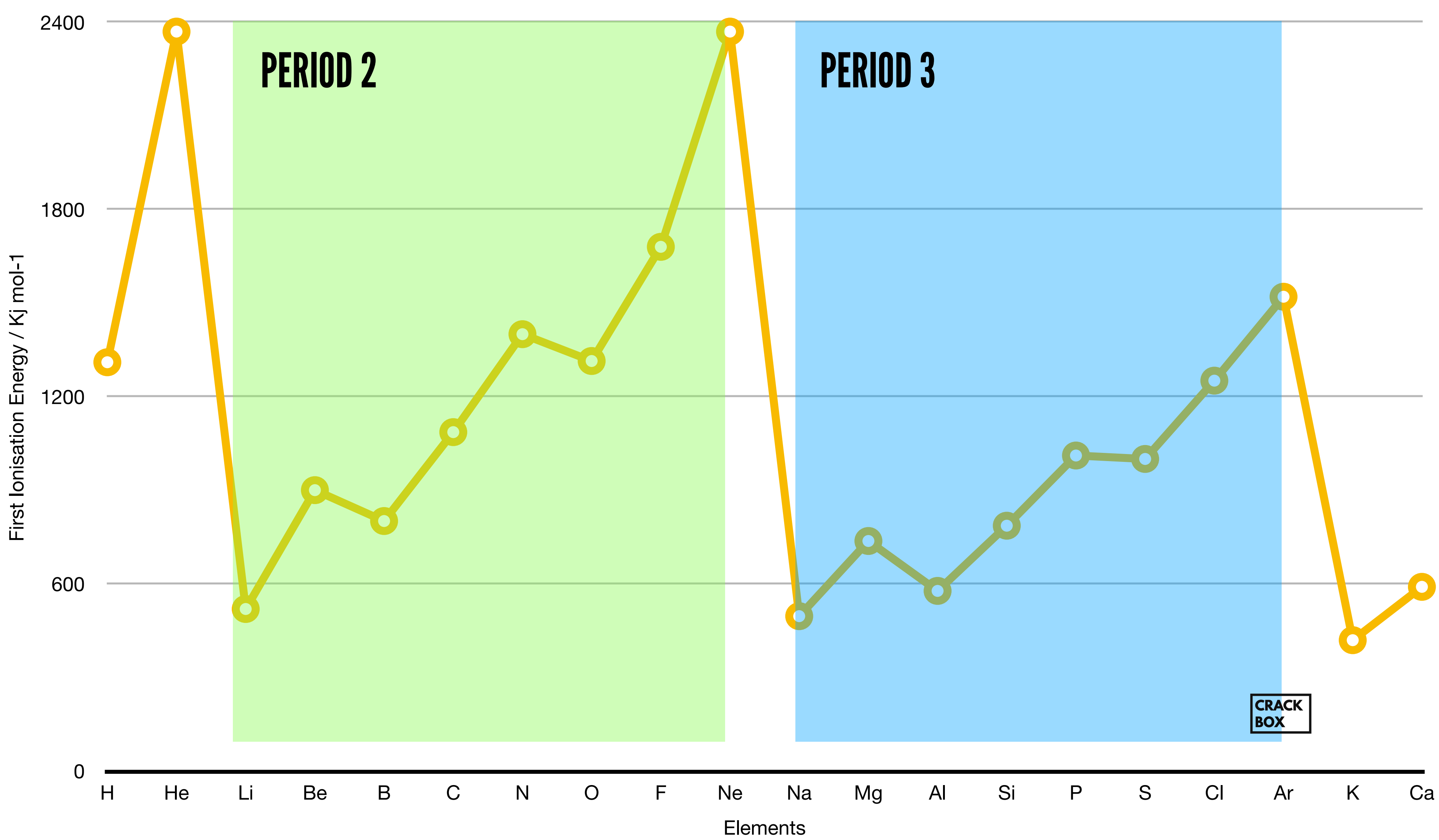
Let's now divide the graph into segments as per periods...



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If you see the general trend is that the First I.E. <u>increases</u> along a period.

But if we look closely, we see that the graph has a few ups and downs along every period. And that is because the valence electrons are paired in some cases and unpaired in others.

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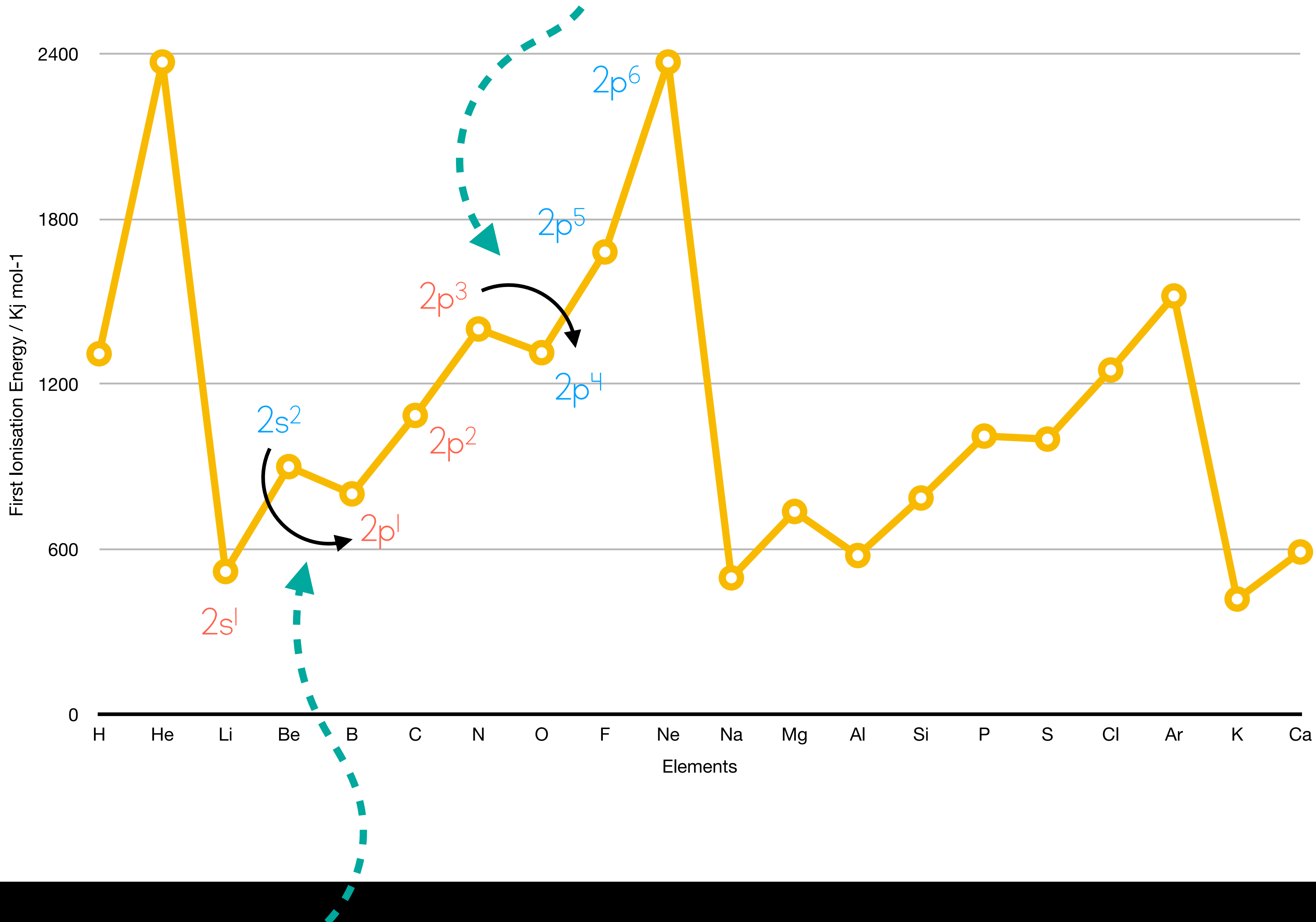
Over here we can see that since Oxygen has a paired electron, the First I.E. is lesser than Nitrogen's.



And here we can see that the First I.E. has reduced despite becoming an unpaired electron. That's because the electron is in the p sub-shell which has a lower energy level than the s sub-shell because it's further away from the nucleus.

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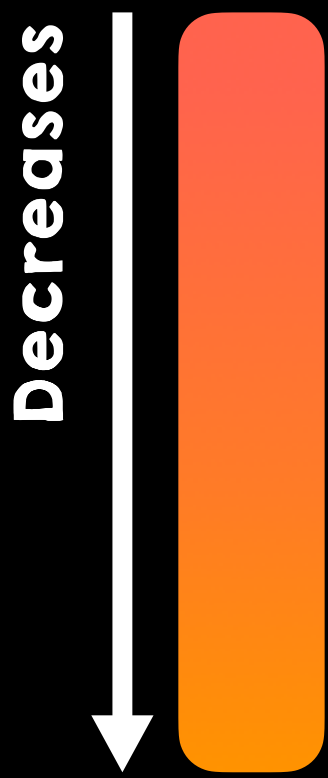
So ignoring the bumps, the First Ionisation Energy in general, increases along a period.

When we look at the First Ionisation Energies down a group, it generally decreases because the distance between the nucleus and the electron becomes greater.

This happens even though the nuclear charge increases because even the shielding effect increases and the effective nuclear charge remains constant. Hence the increased distance between the nucleus and the electron reduces the I.E.



Hence, the First Ionisation Energy <u>reduces</u> as we go down the groups.

PRO TIP

An easy way to remember this trend is to know that the I.E. trend is the <u>reverse</u> of the Atomic Radii trend.

Electron Affinity

The First Electron Affinity of an element is the energy change when one mole of electrons is added to one mole of gaseous atoms to form one mole of gaseous ions.

$$X (g) + e^- \longrightarrow X^- (g)$$

Electron Affinity is like the opposite of Ionisation Energy.

Noble Gases however don't have an EA because they don't give up or take electrons in the first place due to their stable nature. Hence they are excluded.

As the positively charged nucleus of an atom readily attracts anything like electrons, the first EA is usually exothermic as it doesn't require energy to force the electron into the the atom, but in fact releases energy when an electron comes.

$$O (g) + e^- \longrightarrow O^- (g)$$

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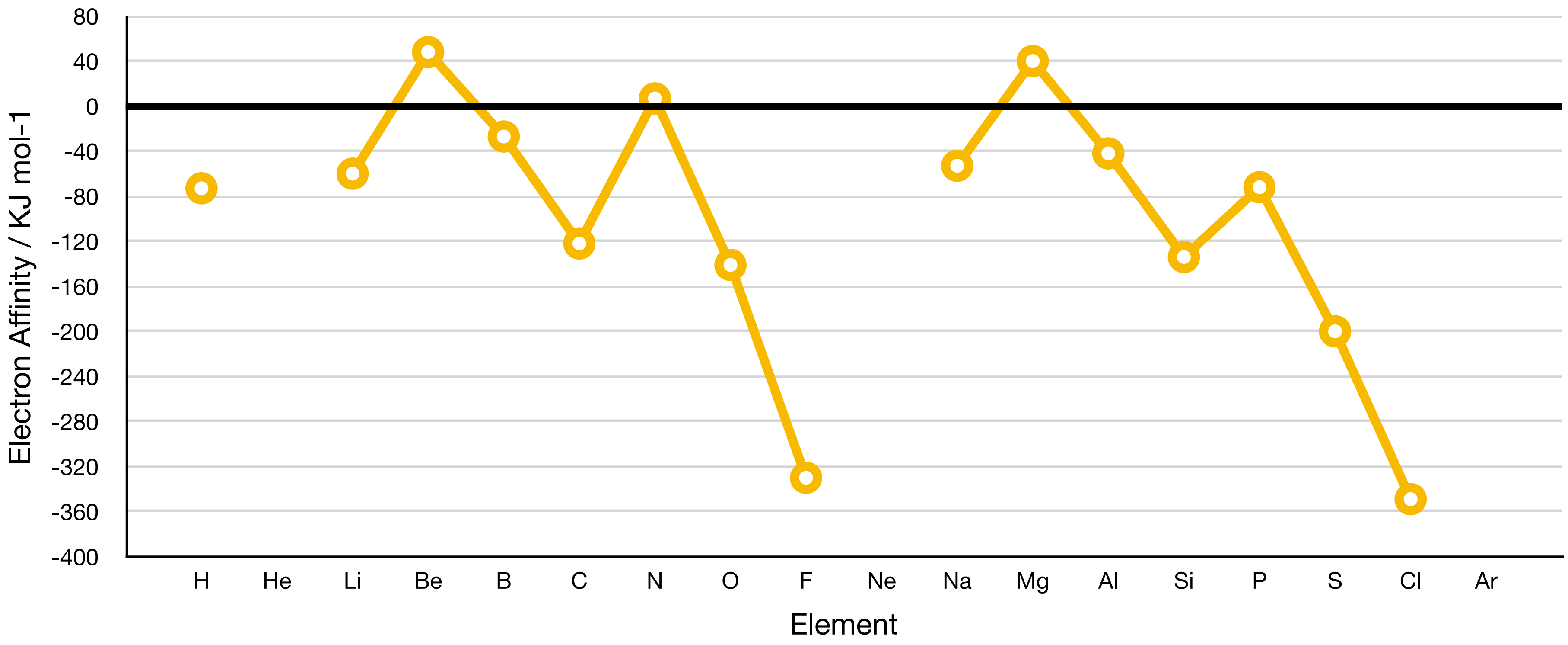
But for the second EA, the atoms are negatively charged ions. Therefore, the ion repels the extra electron. So in order to add another electron, energy is required to force it and the reaction becomes endothermic.

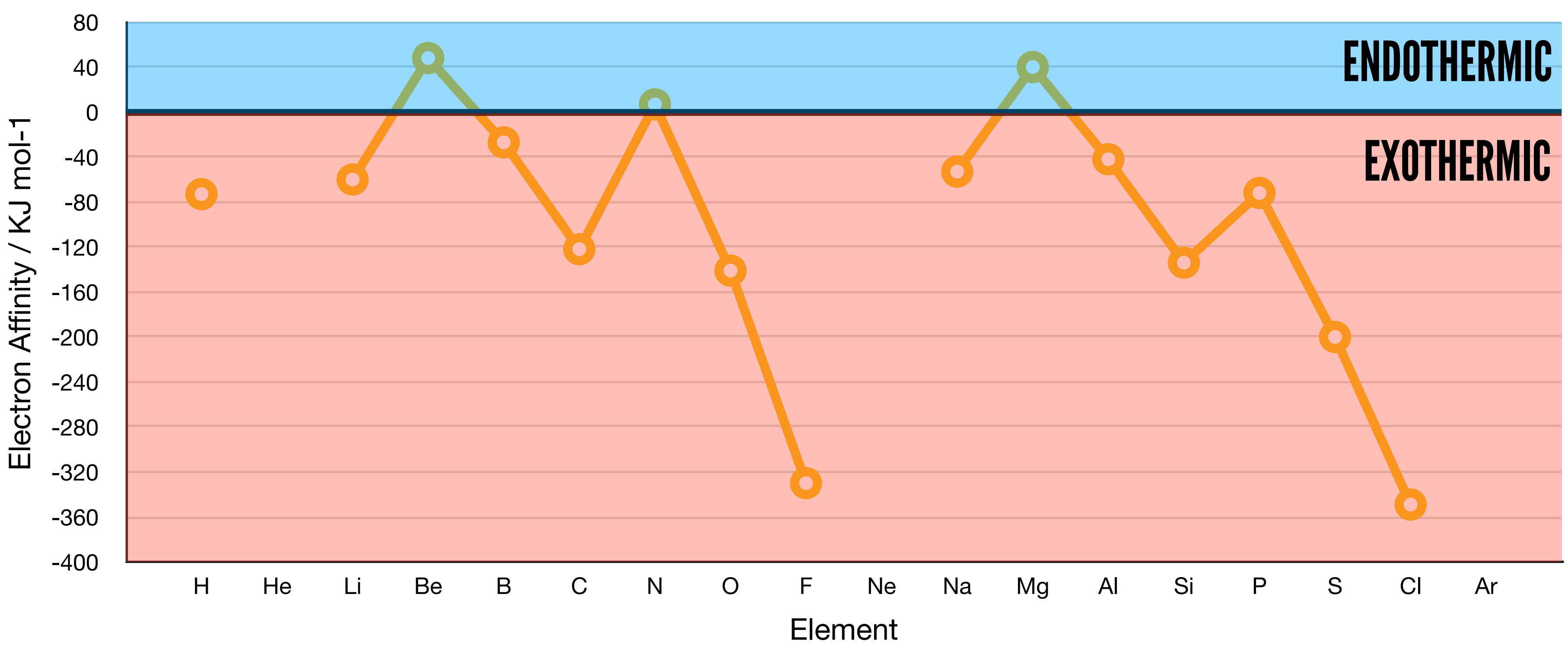
$$O^{-} (g) + e^{-} \longrightarrow O^{-2} (g)$$

Now, the larger the nuclear charges, greater the EA because the nucleus attracts electrons more strongly.

Hence group 17 elements have the greatest EAs due to their nuclear charges of approximately +7.

And group 1 elements have the lowest EAs.





As we can see, the negative energy changes are a result of an exothermic reaction. Most of the elements have a negative Electron Affinity. But group 2 elements (Be, Mg) show an exception.

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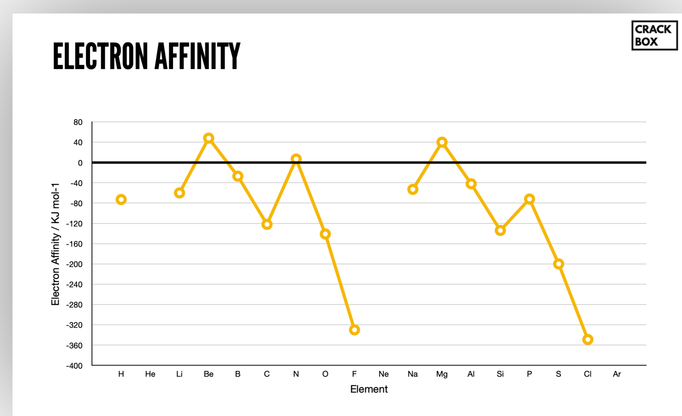

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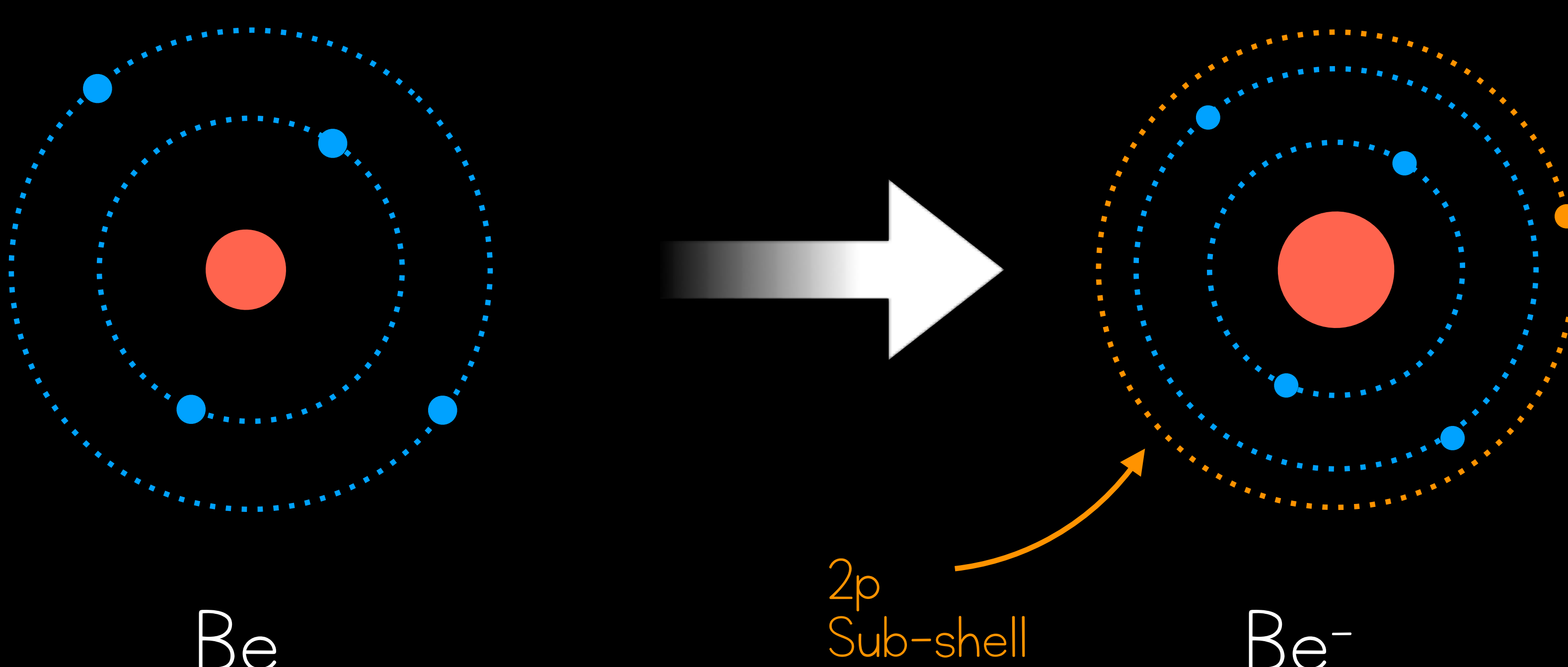
To understand the exception with the group 2 elements (Be, Mg) and the group 5 elements (N, P), we need to recollect their Ionisation Energies.

Let's take Be for example, Be has an electronic configuration of $1s^2$, $2s^2$. If we were to add another electron, it would have to come in the 2p sub-shell.



The 2p sub-shell may not be a whole new shell, but it has a higher energy level than the 2s sub-shell and because of this, it experiences a -4 shielding effect from all the electrons which cancels out the +4 nuclear charge. And so the atom does not readily accept the extra electron and energy needs to be given to force the electron into the atom.

The same concept applies to the other atoms which break the trend like Mg, P, etc.

Looking from the top, the Electron Affinity <u>increases</u> across a period. Note: EA is the change in energy, a larger EA means a larger change which could be positive or negative.

Electronegativity

Electronegativity is a measure of the atom's ability to attract electrons in a covalent bond.

An element with a high electronegativity has a strong electron pulling power.

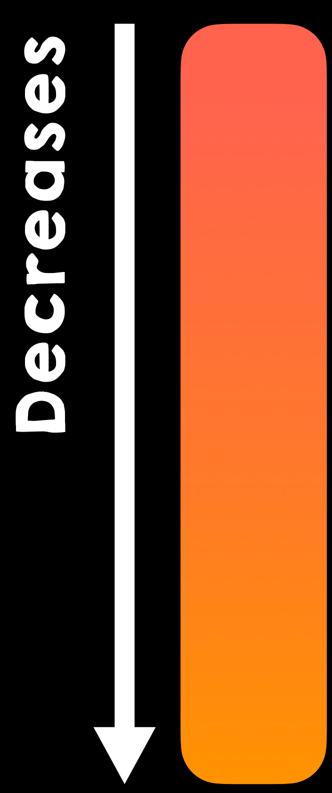
An element with a low electronegativity has a weak electron pulling power.

PRO TIP

The general trends are the same as ionisation energy.

Noble gases are excluded from this trend because they don't attract electrons at all.

Electronegativity <u>increases</u> across a period because of the increased effective nuclear charge.



Electronegativity <u>decreases</u> down the group because the electrons are furtherest from the the nucleus.

Melting Points

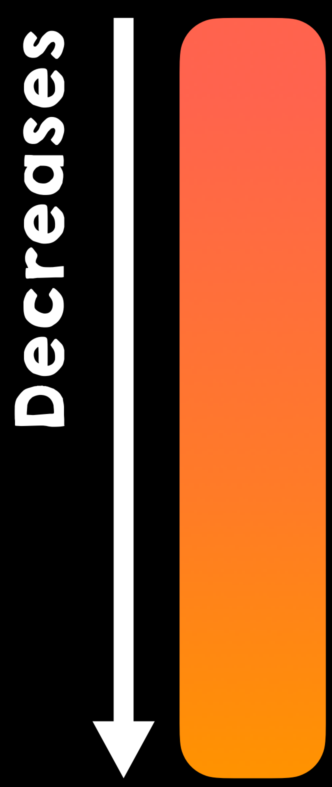
Melting Point trends are different for:

METALS

NON-METALS

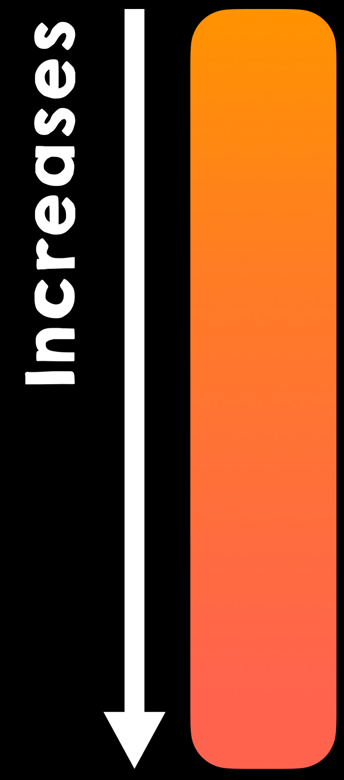
METALS form strong lattice structures. And smaller the atomic radii of a metal atom, the more difficult it becomes to break the lattice structure.

For NON-METALS however, the intermolecular forces like the London forces and Van der Waals forces of attraction keep the molecules close. More the electrons, more these forces and hence larger amounts of energy are required to overcome them.



Hence, in group 1 & 2 metals, the melting & boiling points decrease down the group.

In group 7 & 8 Non-metals, the melting & boiling points increase down the group.



In other areas of the table like the transition metals, the trends become different

For metals, the melting & boiling points <u>increase</u> along the period.

For non-metals, the melting & boiling points <u>decrease</u> along the period.

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the end.