The structure of the periodic table
Periodicity is the repeating pattern in chemical and physical properties.

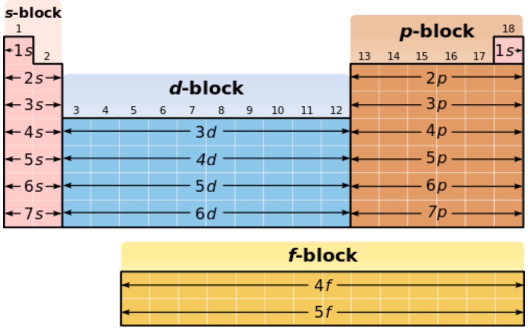
The periodic table is arranged by increasing atomic (proton) number rather than increasing atomic mass because it wouldn't show properties or trends in groups and periods and it wouldn't be in order of increasing electron configuration.

The periods show repeating trends in chemical and physical properties. This is because elements across a period have the same number of subshells e.g. elements in period 2 have 2 sub shells.

Groups have similar chemical properties. Elements down a group have the same number of outer electrons which are used in chemical reactions and the same type of orbitals e.g. all elements in group 2 have 2 outer electron and an s^2

An orbital is a region within an atom that can hold up to two electrons with opposite spins.

Periodic trend in electronic configuration and ionisation energy
Similarities in electron configuration reflect similarities of chemical reactions. For example, atoms of group 2 elements lose one electron to from 2+ ions with a noble gas configuration.



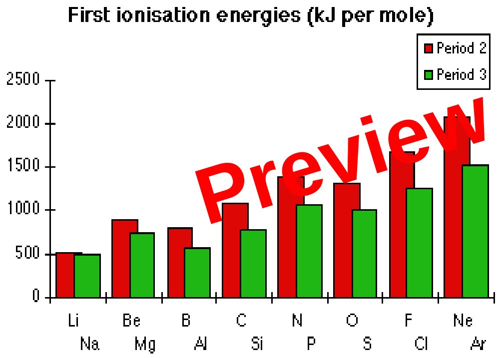
Be: $Be \rightarrow Be^{2+} + e^-$
[He] $2s^2 \rightarrow$ [Ne]
Mg: $Mg \rightarrow Mg^{2+} + e^-$
[Ne] $3s^2 \rightarrow$ [Ar]
Ca: $Ca \rightarrow Ca^{2+} + e^-$
[Ar] $4s^2 \rightarrow$ [Kr]
Sr: $Sr \rightarrow Sr^{2+} + e^-$
[Kr] $5s^2 \rightarrow$ [Xe]

Electron configuration can be based on the previous noble gas, so only outer-shell electrons, responsible for reactions, can be shown.

1st ionisation: $M \rightarrow M^+ + e^-$
2nd ionisation: $M^+ \rightarrow M^{2+} + e^-$
3rd ionisation: $M^{2+} \rightarrow M^{3+} + e^-$
(Where M represents element)

Mg: $1s^2 2s^2 2p^6\ 3s^2 \rightarrow$ [Ne] $3s^2$

First ionisation energy is the energy needed to remove 1 mol of electrons from 1 mol of gaseous atoms.

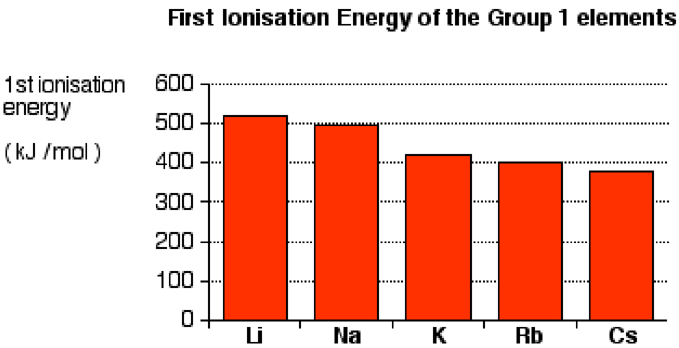


The first ionisation energy increases overall across period 2 and 3 because the number of protons increase so there is greater nuclear charge. The atomic radius decreases because the increased nuclear charge pulls the electrons in towards the nucleus. They have the same number of inner shells so the outermost electrons experience the same shielding. Therefore there is greater nuclear attraction on the electrons and the outer electrons are more strongly attracted to the nucleus.

The first ionisation energy decreases as you move from period 2 to 3 because the atomic radius increases as there is one more inner shell and the electron shielding increases as the number of outer electrons increase. This means the nuclear attraction decreases.

There is a small decrease in first ionisation energy between Be and B because the outer electrons move from an s-orbital to a p-orbital. This increases the atomic radius and the electron shielding as p-orbitals require a little less energy which means there is slightly less nuclear attraction, so it's easier to remove an electron.

There is another small decrease in first ionisation energy between N and O due to p-orbital electron repulsion. It's easier to remove an electron from a singly-occupied orbital so less energy is needed, decreasing the nuclear attraction.



The first ionisation energy decreases down a group because the number of inner shells increases, so the atomic radius increases and the electron shielding increases. The number of protons also increases which increases the nuclear charge. This means there is less nuclear attraction and it is easier to remove an electron from the nucleus.

The nuclear attraction from the increase in atomic radius and electron shielding outweighs the nuclear attraction from the increased nuclear charge.