ionisation energy

Ionisation energy is defined as **the amount of energy required to remove one electron completely from a gaseous atom or ion.** The amount of energy required to remove one the first electron from an atom is called the **first ionisation energy**. It is also called the **enthalpy change**.

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

The energy required to remove the second electron is called the **second ionisation energy**.

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

Ionisation energy for the second and subsequent electrons are called successive ionisation energies. It is measured in energy per mole or $kJ\ mol^{-1}$

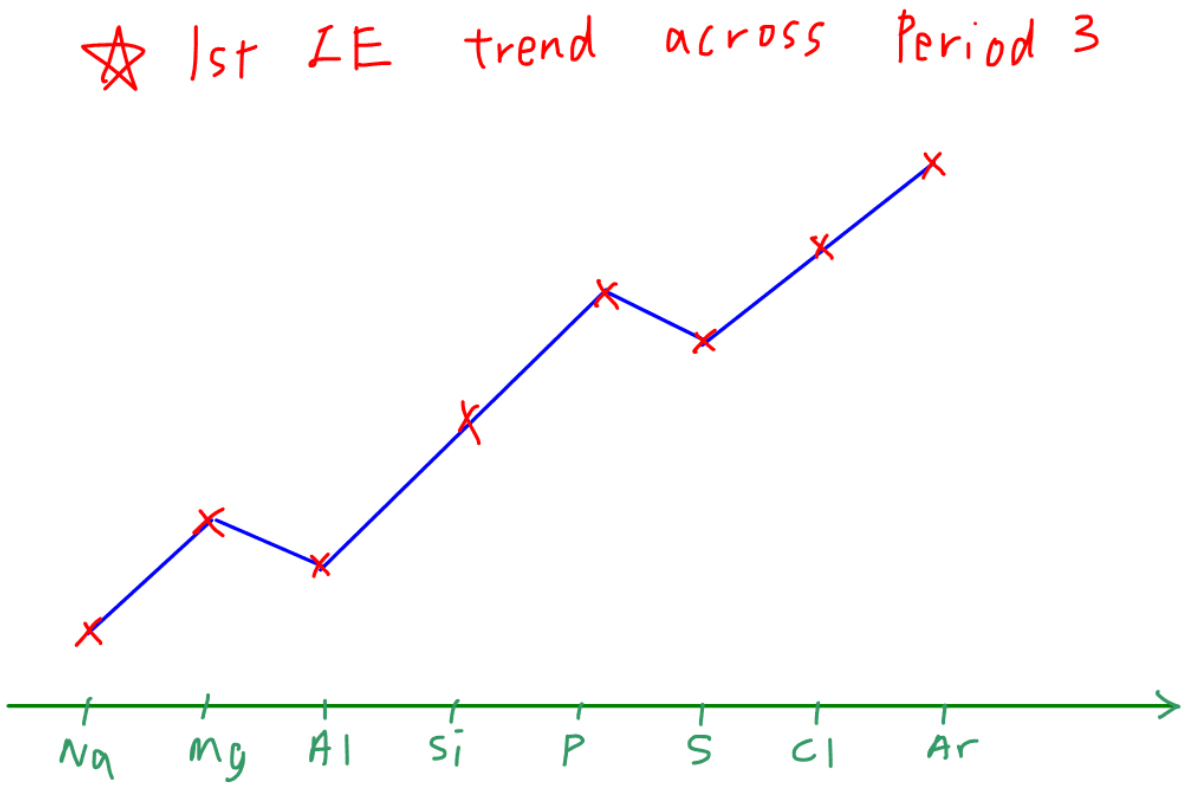
Trends in ionisation energy

Ionisation energies depend on the strength of the attraction between the electrons and the nucleus. Specifically, these are the 4 factors:

- nuclear charge
- atomic or ionic radius
- shielding of the nuclear charge
- spin pair repulsion of electrons

Ionisation energies **decrease down a group**. This is because the atomic radius increase and there is more shielding between the nucleus and the outer most electron.

Ionisation energies **increase from left to right** in the same period (with a few exceptions).



In sodium, the outermost electron fills the 3s shell and it is easy to remove. In magnesium, the outermost electron fills the 3s shell and is now a paired electron + there is one more proton so the nuclear charge is increased.

In aluminium, the outermost electron fills the 3p shell. Since 3p shells have more energy than 3s shells, it is easier to remove, so the ionisation energy dips down. In silicon, there is one more proton so the nuclear charge increases, making it harder to remove. Same for phosphorus.

In sulfur, the outermost electron fills the first pair of the 3p shell, so it is easier to remove the outermost electron than phosphorus. In chlorine and argon, there is more nuclear charge so the IE is higher.

Predicting an element’s group from its ionisation energy

In magnesium, the successive ionisation energies are

$Mg(g) \rightarrow Mg^{+}(g) + e^{-} = +738\ kJ\ mol^{-1}$

$Mg^{+}(g) \rightarrow Mg^{2+}(g) + e^{-} = +1451\ kJ\ mol^{-1}$

$Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^{-} = +7733\ kJ\ mol^{-1}$

Here we can see that there is a big jump in IE from the second to the third IE. This is because the first 2 electrons are from the 3s shell, which is the outermost shell. After the 2 3s electrons are removed, the remaining electrons are now in the 2p shell, which is a lot harder to remove. So, we can predict that magnesium is from the second group of the periodic table.