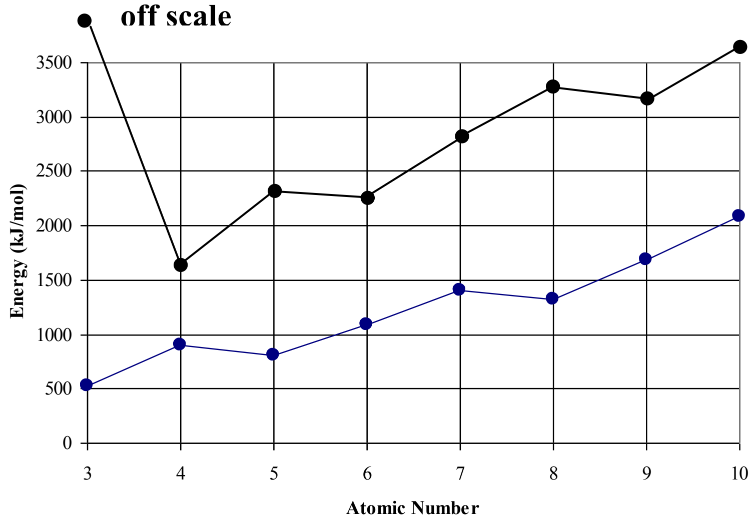
Marks
6

- The graph shows the first ionisation energies for second row elements of the periodic table.



Explain the general trend and both anomalies.

The general trend of an *increase* in ionisation energy across the period is due to the increase in effective nuclear charge (Z_{eff}). The electrons feel a greater pull from the nucleus as Z_{eff} increases. This leads to a decrease in the size of the atom and an increase in the energy required to remove an electron.

First anomaly - the *decrease* in ionisation energy in going from Be (at. no. 4) to B (at. no. 5). Be has an electron configuration of [He] $2s^2$ while B has a configuration of [He] $2s^2\,2p^1$. Due to shielding in a multi-electron atom, the $2p$ orbital is higher in energy than the $2s$ orbital and thus any electron in the $2p$ orbital is held less tightly than those in the $2s$ orbital. B therefore has a lower ionisation energy than Be, despite having a higher nuclear charge.

Second anomaly - another (slight) drop in ionisation energy going form N (at. no. 7) to O (at. no. 8). There are only three *p* orbitals, so the next electron to go into one of the *p* orbitals must pair up. Paired electrons in the same orbital suffer higher electron – electron repulsion so O has a lower ionisation energy than N, despite having a higher nuclear charge.

On the above graph, plot your estimates of the second ionisation energies for the second row elements. Make sure your graph clearly shows the general trends.

See figure above.

The second ionisation of Li is > 7000 kJ mol^{-1} as a core electron is ionised. The second ionisations of the other elements follow the same trends as the first ionisations (for exactly the same reasons), but displaced one atomic number to the right and at a slightly higher energy (as Z_{eff} is greater).