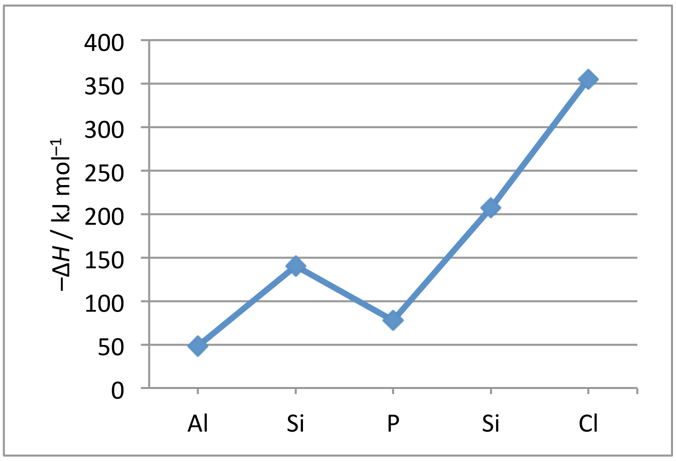
Marks
7

- Electron affinity is the enthalpy change for the reaction $A(g) + e \rightarrow A^-(g)$.
  The graph below shows the trend in electron affinities for a sequence of elements in the third row of the Periodic Table.



Give the electron configurations of the following atoms and singly-charged anions. Use [Ne] to represent core electrons.

Atom	Electron configuration	Ion	Electron configuration
Si	**[Ne] $(3s)^2$ $(3p)^2$**	Si^-	**[Ne] $(3s)^2$ $(3p)^3$**
P	**[Ne] $(3s)^2$ $(3p)^3$**	P^-	**[Ne] $(3s)^2$ $(3p)^4$**
S	**[Ne] $(3s)^2$ $(3p)^4$**	S^-	**[Ne] $(3s)^2$ $(3p)^5$**

Explain why the value for the electron affinity of phosphorus is anomalous.

The general trend across a row is for the electron affinity to increase, as the number of protons in the nucleus increases.

However, in order to form P^-, the extra electron must pair up with an existing electron in one of the *p*-orbitals. The extra repulsion involved leads to the electron affinity being lower for P than for Si despite the higher nuclear charge.

What trend would you expect for the electron affinities for Si^-, P^- and S^-? Explain your answer.

The electron affinities of these anions will be much lower than those of the parent atoms, as adding an electron to an already negatively charged species is much less favourable.

The nuclear charge increase along the series so the electron affinities will increases: $Si^- < P^- < S^-$. The electron affinity of Si^- will also be further decreased because addition of an electron requires pairing again. However, this will not affect the order as Si^- is already has the lowest electron affinity.