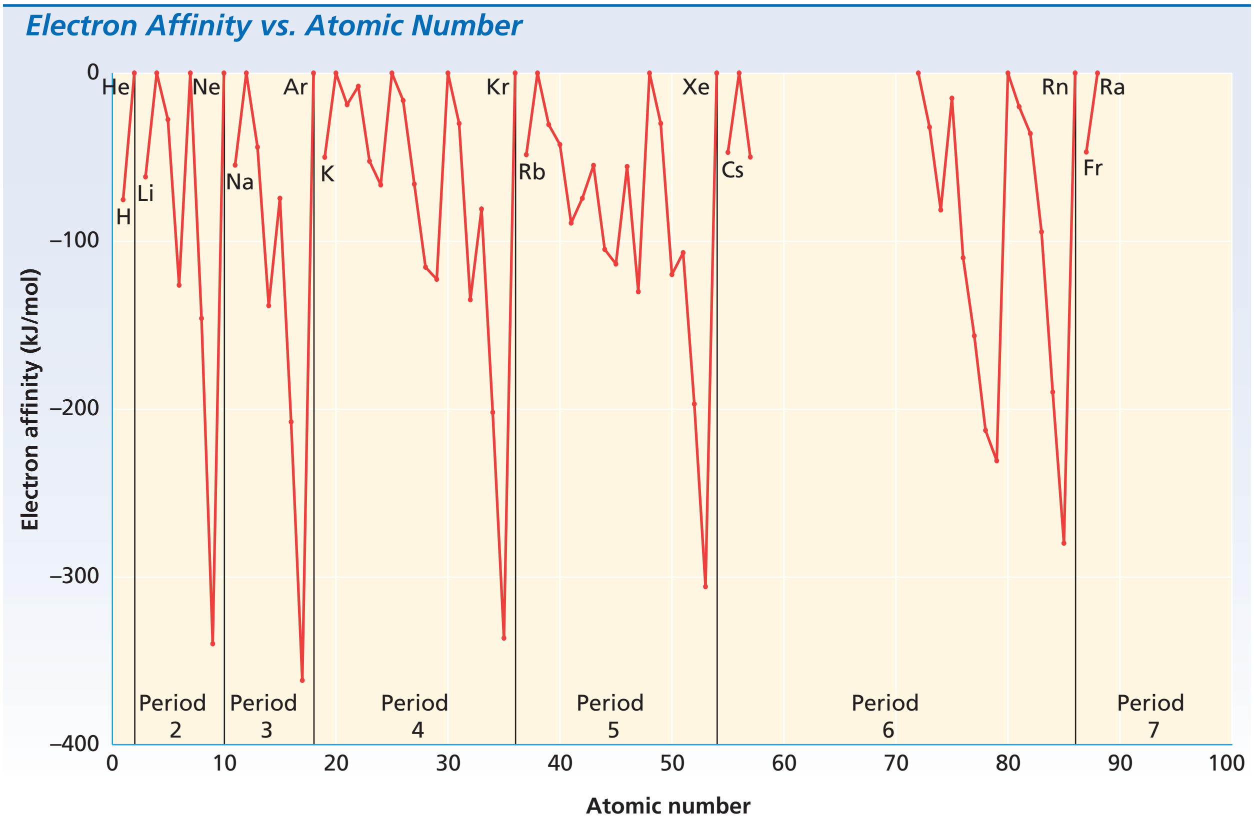

FIGURE 18 The plot of electron affinity versus atomic number shows that most atoms release energy when they acquire an electron, as indicated by negative values.

than forcing an electron to pair with another electron in an orbital of the already half-filled *p* sublevel of a nitrogen atom.

Group Trends

Trends for electron affinities within groups are not as regular as trends for ionization energies. As a general rule, electrons add with greater difficulty down a group. This pattern is a result of two competing factors. The first is a slight increase in effective nuclear charge down a group, which increases electron affinities. The second is an increase in atomic radius down a group, which decreases electron affinities. In general, the size effect predominates. But there are exceptions, especially among the heavy transition metals, which tend to be the same size or even decrease in radius down a group.

Adding Electrons to Negative Ions

For an isolated ion in the gas phase, it is always more difficult to add a second electron to an already negatively charged ion. Therefore, second electron affinities are all positive. Certain *p*-block nonmetals tend to form negative ions that have noble gas configurations. The halogens do so by adding one electron. For example, chlorine has the configuration [Ne]$3s^23p^5$. An atom of chlorine achieves the configuration of the noble gas argon by adding an electron to form the ion Cl^- ([Ne]$3s^23p^6$). Adding another electron is so difficult that Cl^{2-} never occurs. Atoms of