

**Example 1- K:***Filling of electrons in the 3<sup>rd</sup> period*

The 2<sup>nd</sup> period is ended with the noble gas neon. After neon the 3<sup>rd</sup> period is started with the 3s-orbital:

*Na*,  $1s^2 2s^2 2p^6 3s^1$ , one valence electron in the 3s-orbital and 10 *core electron*.

The electron configuration for the 10 core electrons corresponds to the electron configuration for the noble gas in the previous period (in the case of sodium the core electron configuration corresponds to the electron configuration of neon). To ease the work of writing the full electron configuration, only the electron configuration of the valence electrons are written while the electron configuration for the core electrons is replaced by the chemical symbol for the previous noble gas placed in edged brackets. The electron configuration for sodium can thus more simple be written as follows:

*Na*,  $[Ne]3s^1$ , one valence electron in the 3s-orbital. The electron configuration for 10 core electrons correspond the electron configuration of the noble gas neon.

The filling of valence electrons in the 3<sup>rd</sup> period is continued as in the 2<sup>nd</sup> period:

*Mg*,  $[Ne]3s^2$ , 2 electrons in the 3s-orbital.

*Al*,  $[Ne]3s^2 3p^1$ , 2 electrons in the 3s-orbital and one electron in one of the 3p-orbitals.

*Si*,  $[Ne]3s^2 3p^2$ , 2 electrons in the 3s-orbital and 2 unpaired electrons in two of the 3p-orbitals.

*P*,  $[Ne]3s^2 3p^3$ , 2 electrons in the 3s-orbital and 3 unpaired electrons in three of the 3p-orbitals.

*S*,  $[Ne]3s^2 3p^4$ , 2 electrons in the 3s-orbital and 4 electrons in the 3p-orbitals.

*Cl*,  $[Ne]3s^2 3p^5$ , 2 electrons in the 3s-orbital and 5 electrons in the 3p-orbitals.

*Ar*,  $[Ne]3s^2 3p^6 = [Ar]$ , all orbitals of the period are filled and we have reached the end of the 3<sup>rd</sup> period.

**Example 1- L:***Filling of electrons in the 4<sup>th</sup> period*

The filling of valence electrons through the 4<sup>th</sup> period takes place almost as in the 3<sup>rd</sup> period. The exception is that after the filling of the 4s-orbital, the five 3d-orbitals are then filled before the filling of the three 4p-orbitals according to the row given in (1- 5) on page 30 and Figure 1- 7 on page 30. This is due to the fact that in between the energy levels of the 4s-orbital and the 4p-orbitals the energy level of the five 3d-orbitals is located. During the filling of the five 3d-orbitals, Hund's rule is again followed which means that as long as there are empty 3d-orbitals the "next" electron will be placed in an empty orbital and thus remain unpaired. Some examples of electron configuration for elements from the 4<sup>th</sup> period are given here:

*Ti*,  $[Ar]4s^2 3d^2$ , 2 electrons in the 4s-orbital and 2 electrons in the 3d-orbitals.

*Zn*,  $[Ar]4s^2 3d^{10}$ , all five 3d-orbitals are filled.

*Ga*,  $[Ar]4s^2 3d^{10} 4p^1$ , the filling of the three 4p-orbitals has started.

*Kr*,  $[Ar]4s^2 3d^{10} 4p^6 = [Kr]$ , all the orbitals of the period are filled and the period is ended.

In the 4<sup>th</sup> period there are some exceptions for some of the elements where the electron configuration deviates for the conventional principle of "filling". These exceptions are:

*Cr*,  $[Ar]4s^1 3d^5$ , only *one* electron in the 4s-orbital while all five 3d-orbitals each host one unpaired electron. This configuration gives a particular stable configuration for the d-orbitals

*Cu*,  $[Ar]4s^1 3d^{10}$ , only *one* electron in the 4s-orbital while all five 3d-orbitals each host two electrons.

This configuration gives a particular stable configuration for the d-orbitals.

**Example 1- M:***Filling of electrons in the 5<sup>th</sup> period*

The filling of electrons through the 5<sup>th</sup> period takes place exactly as for the 4<sup>th</sup> period. First the 5s-orbital is filled and then the five 4d-orbitals are filled. Finally the three 5p-orbitals are filled according to the row given in (1- 5) on page 30 and Figure 1- 7 on page 30. In the 5<sup>th</sup> period there are also some deviations for the normal filling principles in which the five 4d-orbitals are either half or completely filled with electrons before the 5s-orbital is filled. These deviations are similar to the deviations in the 4<sup>th</sup> period:

*Mo*,  $[Kr]5s^1 4d^5$ , only *one* single electron in the 5s-orbital while all five 4d-orbitals each host one unpaired electron. This gives a particular stable electron configuration for the d-orbitals.

*Pd*,  $[Kr]5s^0 4d^{10}$ , *no* electrons in the 5s-orbital while all five 4d-orbitals each host two electrons. This gives a particular stable electron configuration for the d-orbitals

*Ag*,  $[Kr]5s^1 4d^{10}$ , only *one* single electron in the 5s-orbital while all five 4d-orbitals each host two electrons. This gives a particular stable electron configuration for the d-orbitals.

Please click the advert



# FOSS

## Sharp Minds - Bright Ideas!

Employees at FOSS Analytical A/S are living proof of the company value - First - using new inventions to make dedicated solutions for our customers. With sharp minds and cross functional teamwork, we constantly strive to develop new unique products - Would you like to join our team?

FOSS works diligently with innovation and development as basis for its growth. It is reflected in the fact that more than 200 of the 1200 employees in FOSS work with Research & Development in Scandinavia and USA. Engineers at FOSS work in production, development and marketing, within a wide range of different fields, i.e. Chemistry, Electronics, Mechanics, Software, Optics, Microbiology, Chemometrics.

**We offer**

*A challenging job in an international and innovative company that is leading in its field. You will get the opportunity to work with the most advanced technology together with highly skilled colleagues.*

*Read more about FOSS at [www.foss.dk](http://www.foss.dk) - or go directly to our student site [www.foss.dk/sharpminds](http://www.foss.dk/sharpminds) where you can learn more about your possibilities of working together with us on projects, your thesis etc.*

The Family owned FOSS group is the world leader as supplier of dedicated, high-tech analytical solutions which measure and control the quality and production of agricultural, food, pharmaceutical and chemical products. Main activities are initiated from Denmark, Sweden and USA with headquarters domiciled in Hillerød, DK. The products are marketed globally by 23 sales companies and an extensive net of distributors. In line with the corevalue to be 'First', the company intends to expand its market position.

**Dedicated Analytical Solutions**

FOSS  
Slangerupgade 69  
3400 Hillerød  
Tel. +45 70103370  
[www.foss.dk](http://www.foss.dk)



**Example 1- N:***Filling of electrons in the 6<sup>th</sup> and 7<sup>th</sup> period*

In the 6<sup>th</sup> and 7<sup>th</sup> period the seven f-orbitals are introduced (the 4f-orbitals and the 5f-orbitals respectively) which is also sketched in (1- 5) on page 30 and Figure 1- 7 on page 30. This means that in the 6<sup>th</sup> period the 6s-orbital is filled first and then the seven 4f-orbitals are filled. After that the five 5d-orbitals are filled follow by the filling of the three 6p-orbitals. In the 7<sup>th</sup> period the 7s-orbital is filled first followed by the filling of the seven 5f-orbitals. After that the five 6d-orbitals are filled and then no more elements exist (or at least they are not found or synthesized yet).

Here are some examples of electron configurations for 6<sup>th</sup> and 7<sup>th</sup> period elements. Europium (Eu), gold (Au), lead (Pb) and einsteinium (Es) are used as examples:

*Eu*,  $[Xe]6s^2 4f^7$ , core electron configuration corresponds to the noble gas xenon. Besides that 2 electrons are in the 6s-orbital and 7 unpaired electrons in each of the seven 4f-orbitals.

*Au*,  $[Xe]6s^1 4f^{14} 5d^{10}$ , core electron configuration corresponds to the noble gas xenon. Besides that *one* electron is in the 6s-orbital, 14 electrons in the 4f-orbitals and 10 electrons in the 5d-orbitals. It is seen that the electron configuration of gold deviates from the normal “filling” principle since only *one* electron is in the 6s-orbital the 5d-orbitals are filled. But as described for the 4<sup>th</sup> and 5<sup>th</sup> period this gives a particular stable configuration for the d-orbitals.

*Pb*,  $[Xe]6s^2 4f^{14} 5d^{10} 6p^2$ , core electron configuration corresponds to the noble gas xenon. Besides that 2 electrons in the 6s-orbital, 14 electrons in the 4f-orbitals, 10 electrons in the 5d-orbitals and 2 electrons in the 6p-orbital.

*Es*,  $[Rn]7s^2 5f^{11}$ , core electron configuration corresponds to the noble gas radon. Besides that there are 2 electrons in the 7s-orbital and 11 electrons in the 5f-orbitals.

The elements with 4f-orbital valence electrons are called *lanthanoids* because the first element before the 4f-orbitals is lanthanum (La). The elements with 5f-orbital valence electrons are called *actinoids* since the first element before the 5f-orbitals is actinium (Ac).

### 1.2.3 Categorization of the elements

The elements in the periodic table can be classified at either

Metals

Half-metal or

None-metal

The metals are placed to the left while the none-metals are placed to the right in the periodic table. The half-metals are placed as a wedge between the metals and the none-metals as sketched in Figure 1- 8.

ALKALI METALS (NOT HYDROGEN)										NOBLE GASSES									
↓ ALKALINE EARTH METALS										↓ HALOGENS									
1.	H																	He	
2.	Li	Be										B	C	N	O	F	Ne		
3.	Na	Mg										Al	Si	P	S	Cl	Ar		
4.	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
5.	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
6.	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	
7.	Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg								
LANTHANOIDS			Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu			
ACTINOIDS			Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr			

**Figure 1- 8: Categorization of the elements in the periodic table**

*The elements can be categorized as metals, half-metals or none-metals. Common names for some of the vertical groups are given as well.*

One of the most important differences between metals and none-metals is that metals have very high electrical conductance in all directions. Carbon for example in the form of graphite only conducts electricity in two dimensions inside the layered structure and is thus characterized as a none-metal. Metals and none-metals also behave very different in association with chemical reactions and the formation of chemical bonds. When a metal reacts with a none-metal, the metal will normally deliver electrons to the none-metal which makes the metal into a cation. The none-metal is thus made into an anion and the chemical bond will thus be ionic. The half-metals are placed as a wedge between the metals and none-metals. The half-metals exhibit both metal characteristics and none-metal characteristics.

The vertical rows in the periodic table are called groups. The elements with the “last attached” electron in a d-orbital are called transition metals while the lanthanoids and actinoids have their “last attached” electron in the 4f-orbitals and 5f-orbitals respectively. The other groups are normally called *main groups*. Some of these groups have common names which are shown in Figure 1- 8. The transition metals in the “middle” (closest to manganese (Mn), Technetium (Tc) and Rhenium (Re)) are generally characterized by the ability to appear in almost all oxidation states, whereas the main group elements in general only are able to appear in one or two different oxidation states besides the oxidation state of zero. This is exemplified in the following example:

**Example 1- O:**

*Oxidation state (transition metal and main group element)*

Transition metals are among other aspects characterized by the ability to appear in many oxidation states. The transition metal osmium can be used as an example:

*Os, Possible oxidation states of osmium:* 0, +2, +3, +4, +6, +8

It is thus seen that osmium can attain six different oxidation states. On the other hand main group elements are generally not able to attain that many different oxidation states. Tin for example can only attain two different oxidation states (besides from zero):

*Sn, Possible oxidation states of tin:* 0, +2, +4

The general trend is that transition metals can attain many different oxidation states which are contradictory to main group elements that in general only can attain a few different oxidation states.

Please click the advert

**Design your  
own future at  
MAN Diesel**

[www.mandiesel.com](http://www.mandiesel.com)

**MAN Diesel** – Powering the World



Download free books at BookBooN.com

### 1.2.4 Periodic tendencies

Different tendencies for the elements are present for the periods (horizontal rows) in the periodic system and different tendencies are present in the groups (vertical rows) in the periodic table. In this section we will look more at the periodic tendencies for the following three terms:

Atomic radius

Ionization energy

Electron affinity

The general trends will be described in this section and only some of the exceptions that appear in the periodic table will be discussed. For more details about all exceptions you should seek the knowledge in more detailed educational textbooks.

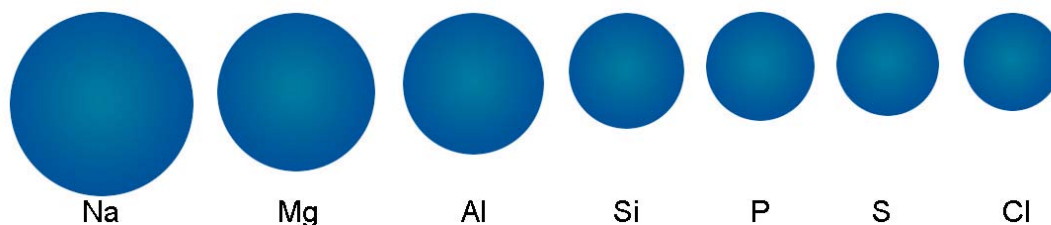
The radius of an atom decreases when you go from the left to the right through a period. This is because when one moves one position to the right (for example when going from lithium to beryllium) one more proton is “added” to the atomic nucleus. Also one electron is “added”. This extra electron will just be hosted in one of the existing orbitals of the period and will not lead to an increased volume. However the “extra” proton in the nucleus will increase the total positive charge of the nucleus by “+1”. This means that the increased positive charge will drag the electrons closer to the nucleus and the total volume and atomic radius will thus decrease.

When you move down a group (vertical row) in the periodic table the atomic radius will of course increase since the atom just below has more electrons and a set of orbital more (*s*- and *p*-orbitals and eventually *d*- and *f*-orbitals). When you move one position down, the primary quantum number *n* will increase by 1 and this means that the orbitals for that particular quantum number are larger which results in a larger atomic radius since the valence electrons are then placed further away from the nucleus. We will look more at atomic radius in the following example:



**Example 1- P:***Atomic radius*

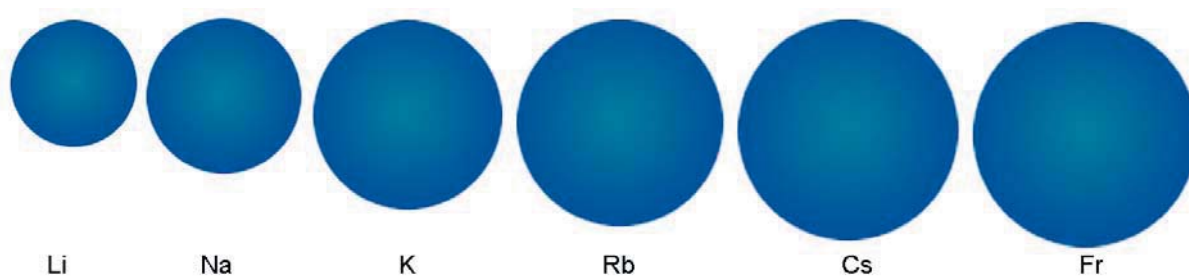
When you move through a period from the right to the left the atomic radius will decrease. This is sketched for the 3<sup>rd</sup> period in Figure 1- 9.



**Figure 1- 9: Relative atomic radius for 3<sup>rd</sup> period elements**

*Relative atomic radius for the elements in the 3<sup>rd</sup> period of the periodic table.*

It is seen that the atomic radius for sodium is almost the double of that of chlorine. When you move down a group (vertically down) the atomic radius increases which is sketched in Figure 1- 10 for the elements in the 1<sup>st</sup> main group.



**Figure 1- 10: Relative atomic radius for 1<sup>st</sup> main group elements**

*Relative atomic radius for the elements in the 1<sup>st</sup> main group.*

In Figure 1- 10 it is seen that the relative atomic radius for the elements in the 1<sup>st</sup> main group (alkali metals) increases which is the case for all vertically groups.

Now we are going to look at ionization energy. When we talk about ionization energy it is implicit that we are talking about the 1<sup>st</sup> ionization energy. The 1<sup>st</sup> ionization energy is the amount of energy required to remove one single electron away from the atom. When one electron is removed the atom becomes a positively charged ion (a cation). Tin and boron can be used as examples:





Thus 708.2 and 800.6 kilo joules are required to ionize 1 mole of tin and boron atoms respectively. When you move from the left to the right in a period in the periodic table the ionization energy increases. This is because the further you are to the right the more protons are present in the nucleus and the larger a positive charge will then drag the electrons towards the nucleus. Thus more energy is required to remove an electron when the larger positive nucleus charge drags the electrons. When you move down a group in the periodic table the ionization energy decreases. This is because the further down the group you are, the further away from the nucleus the valence electrons are placed. Thus the electrons further away from the nucleus are easier to remove and thus the ionization energy decreases. Furthermore the electrons closer to the nucleus constitute a kind of shielding or screening for the outer electrons. Thus the outer electrons are not that much attracted by the nucleus because of this shielding and they are therefore easier to remove and the ionization energy is lower. In the following example we will look more at ionization energies.

Please click the advert

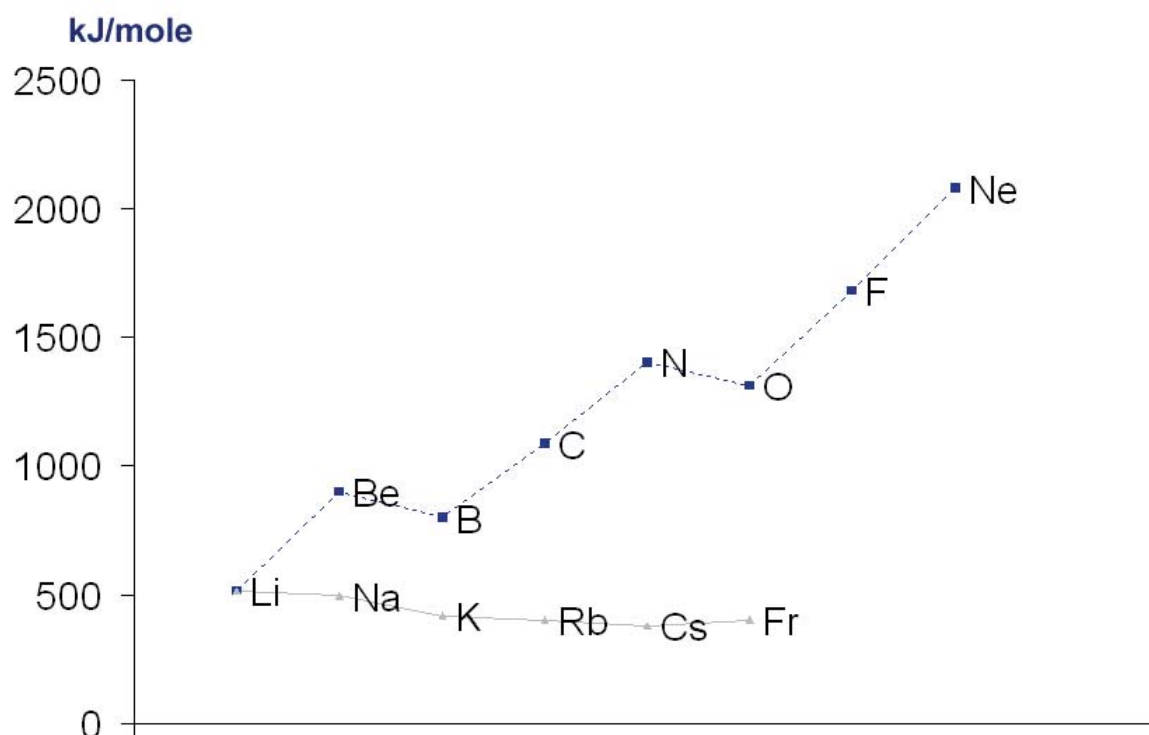
The advertisement features a light blue background with a central graphic. The graphic consists of four square icons arranged horizontally, separated by plus signs and an equals sign. The first icon is dark grey with a white shopping cart icon and the text 'Student Discounts' above it. The second icon is blue with a white calendar icon showing '2009' and the text 'Student Events' above it. The third icon is light grey with a black pound sterling symbol (£) and the text 'Money Saving Advice' above it. The fourth icon is yellow with three white stars and the text 'Happy Days!' above it. Below the icons, the text 'wealthystudent.co.uk' is displayed in a bold, sans-serif font, with 'wealthy' in dark grey and 'student.co.uk' in blue.

**wealthy**student.co.uk

Download free books at BookBooN.com

**Example 1- Q:***Ionization energy*

The ionization energy increases when you move from the left to the right in a period. In Figure 1- 11 the ionization energies for the elements of the 2<sup>nd</sup> period are showed as an example. The ionization energy decreases when you move down a group in the periodic table. This is also shown in Figure 1- 11 for the elements of the 1<sup>st</sup> main group (the alkali metals).



**Figure 1- 11: Ionization energy**

*Ionization energy for the elements in the 2<sup>nd</sup> period (Li, Be, B, C, N, O, F and Ne) and for the elements in the 1<sup>st</sup> main group (Li, Na, K, Rb, Cs and Fr).*

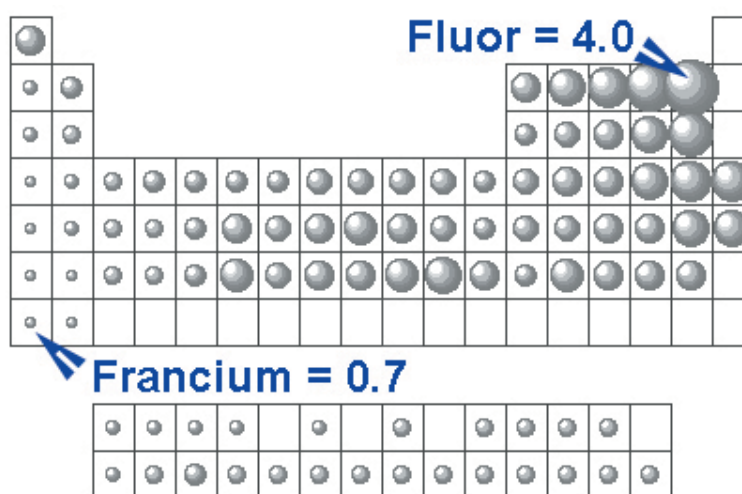
Two exceptions are clear by looking at Figure 1- 11. When you go from beryllium to boron the ionization energy actually decreases. This is because the valence electron of boron in one of the 2p-orbitals is easier to remove than one of the valence electrons of beryllium in the 2s-orbital. The two electrons in the beryllium 2s-orbital constitute a particularly stable electron configuration and the ionization energy is thus relatively large. Nitrogen has three unpaired electrons in each of the degenerated 2p-orbitals which (as described in the section 1.2.2 *Electron configuration*) gives a particularly stable electron configuration. Hence more energy is required to remove one of these unpaired 2p-electrons than the amount of energy required to remove one of the paired 2p-electrons of the oxygen atom. Therefore the ionization energy of oxygen is lower than for nitrogen.

Overall it is seen that the increasing tendency of ionization energy for the periods is much larger than the decreasing tendency down the vertically groups.

Electron affinity is defined as the energy change when an electron is attached to an element. Fluor can be used as an example:



Since the electron affinity is negative, 328 kilo joules are released when 1 mole of electrons are attached to 1 mole of fluor atoms. This relatively large number is caused by the very high electronegativity of fluor. It tells something about the tendency for the atom to accept an extra electron. Electron affinity is thus related to the electronegativity of the elements. In Figure 1- 12 the relative electronegativity for the elements in the periodic table is sketched.



**Figure 1- 12: Electronegativity**

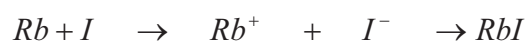
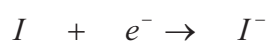
*Electronegativity of the elements in the periodic table. The size of the “bubbles” corresponds to the relative level of the electronegativity. Fluor has the largest electronegativity (4.0) and francium has the lowest electronegativity (0.7).*

The larger electronegativity the more the atom “wants” to adopt an extra electron and the larger a numerical value of the electron affinity. It shall be noticed that all elements have positive electronegativities which means that in principle all elements can benefit from adopting an electron in terms of energy. But here is shall be noted that this extra electron has to be supplied from another atom and that this atom thus has to be lower in electronegativity in order for the total energy to be lowered. In the following example we are going to look more at electronegativities and electron affinities.

**Example 1- R:***Electronegativity*

You have a rubidium atom (Rb) and an iodine atom (I). Both have positive electronegativities which mean that both in principle want to adopt an extra electron. Which of the atoms will become a cation and which will become an anion if they react with each other? Which of the atoms have largest electron?

Rubidium has an electronegativity of 0.8 and iodine has an electronegativity of 2.5. Thus iodine is more likely to adopt an electron than rubidium. Thus rubidium will be forced to deliver an electron to iodine during a chemical reaction.



Because iodine has larger electronegativity than rubidium, the numerical value of the electron affinity for iodine will be larger than for rubidium. This is the reason that during chemical reaction the iodine atom will become the anion (iodide) and rubidium will become the cation (rubidium ion).

Please click the advert

Read The Spanner online on [oxygen.ie](http://oxygen.ie)

Download free books at [BookBooN.com](http://BookBooN.com)