**Atoms - Electronic Configuration**

Electronic configuration is about the arrangement of electrons within the atom.

As you already know, atoms contain protons, neutrons and electrons. The protons and neutrons are in the nucleus of the atom. The electrons orbit the nucleus in energy levels, also called **shells.** The positive charges of the protons in the nucleus hold the electrons in the energy levels.

You also know that each energy level can contain only a certain number of electrons. For your purposes you only need to think about the first three energy levels. The first level can hold a maximum of 2 electrons, the second a maximum of 8 and the third a maximum of 8.

The Bohr Diagrams are a simple representation of electrons in the energy levels (shells).

For example: Here are 4 elements showing the Bohr diagrams and the notation used to indicate the distribution of the atoms in the neutral elements.

|  |  |  |  |
| --- | --- | --- | --- |
| Fluorine | Neon | Sodium | Calcium |
| Structure of a fluorine atom. A black dot represents the nucleus. The small circle around this has two red dots on it, representing the first energy level with two electrons. A larger outer circle has seven red dots on it, representing the second energy level with seven electrons | Structure of a neon atom. A black dot represents the nucleus. The small circle around this has two red dots on it, representing the first energy level with two electrons. A larger outer circle has eight red dots on it, representing the second energy level with eight electrons | Structure of a sodium atom. A black dot represents the nucleus. The small circle around this has two red dots on it, representing the first energy level with two electrons. A larger middle circle has eight red dots, representing the second energy level with eight electrons. A larger outer circle has one red dot on it, representing the third energy level with one electron | Structure of a calcium atom. A black dot represents the nucleus. The small circle around this has two red dots on it, representing the first energy level with two electrons. A larger circle has eight red dots, representing the second energy level with eight electrons. Another larger circle has eight red dots on it, representing the third energy level, with eight electrons. An even larger outer circle has two red dots, representing the fourth energy level with two electrons |
| F 2,7 | Ne 2,8 | Na 2,8,1 | Ca 2,8,8,1 |

The energy of an electron depends mainly on which shell it occupies. Electrons in shells further from the nucleus have a higher energy than electrons in shells closer to the nucleus.

If you were to write out the Bohr diagrams for all of the elements on the Periodic Table you would see a pattern. For the Noble gases (Ne is an example above) the outermost energy level is full. For elements in Group 1 the outermost shell holds just 1 electron (see Na and Ca above) and for the halogens of Group 17 the outermost shell holds 7 electrons (see F above).

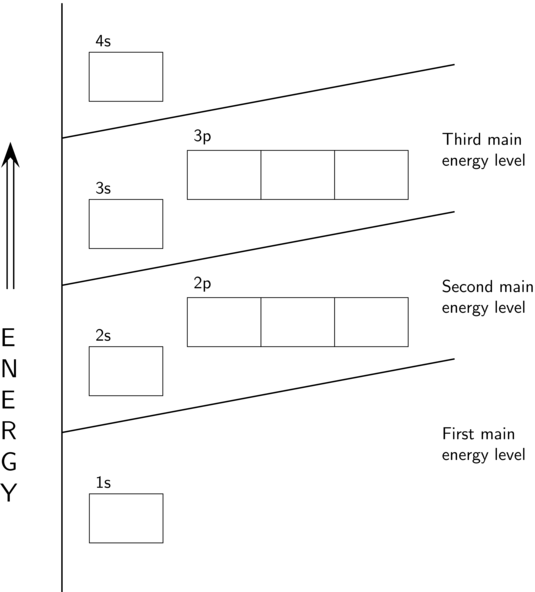
Later electrons were discovered to occupy a three dimensional volume of space with a characteristic shape, called an **orbital**. *Each orbital can hold a maximum of two electrons.*

So the first energy level, shell 1, only 2 electrons can be accommodated and there is one orbital, called the *s orbital*. The *s* orbital has a spherical shape. The s orbital for shell 1 is called 1*s*.

Electrons in an orbital mostly move around inside the orbital, but they can spend a small amount of time completely outside the orbital

The second energy level can hold 8 electrons, so more orbitals are required. Again there is a spherical s orbital. In addition there are 3 *p orbitals,* each of which can hold 2 electrons. The orbitals for the second energy level are called 2*s* and 2*p*.

The third energy level can hold 8 electrons, it has the following orbitals: 3*s* and 3*p*.  
  
In the following diagram the energy levels of the orbitals form 1*s*to 4*s* are shown:



The electrons fill an orbital of lower energy before starting to occupy an orbital of higher energy. So they start at 1*s,* moving to *2s* and so on.

**Representing electronic configuration**

**Several Principles govern how the electrons are located in orbitals:**

1. The **Aufbau principle** electrons orbiting the nucleus of an atom fill the lowest available energy levels before filling higher levels (e.g. 1s is filled with two electrons before 2s can be occupied). This produces the most stable position for the electrons.  
   For the first 20 elements in the Periodic Table the electron filling pattern is therefore: 1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*. The *s* orbitals can each hold 2 electrons and the 3 p orbitals hold 2 electrons each which makes a total of 6 electrons for the full p orbitals of a particular energy level shell).  
     
   Example: For Carbon (C): 1*s*2, 2*s*2
2. The **Pauli exclusion principle** states that two electrons occupying a single orbitals must have opposite spins. This is a quantum physics concept, beyond the understanding of this course. However, just remember that when you represent electrons in a box diagram the arrows point in opposite direction to indicate opposite spin (see below).
3. **Hund's rule**: every orbital in a orbital (subshell) is singly occupied with one electron before any one orbital is doubly occupied, and all electrons in singly occupied orbitals have the same spin (the arrows point in the same direction – see below).

There are two popular methods used to write out the positions of electrons in a particular element.

* **The Aufbau diagram method**  
    
  In this method the orbitals are represented by boxes. Each box can hold 2 electrons. The presence of electrons within a box is shown by an arrow.

If two electrons are present the arrows point in opposite directions (see *Paulis exclusion principle*).  
  
  
Let’s use F (fluorine) as an example. It has 9 electrons.  
  
F:   
 1*s* 2*s* 2*p* 2*p 2p*  
  
There’s something else to remember!

Within a group of p orbitals (remember there are 3) of a single energy level e.g. the three p orbitals of energy level 2 you would draw 3 boxes. When filling the boxes with arrow (to represent electrons) you would also need to obey *Hunds Rule* (see above) and put one electron in each box before starting to double up.  
  
Let’s look at boron (5 electrons), carbon (6 electrons), nitrogen (7 electrons) and oxygen (8e electrons):   
  
B:

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1*s* 2*s* 2*p* 2*p 2p*  
  
C:

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1*s* 2*s* 2*p* 2*p 2p*

N:  
 1*s* 2*s* 2*p* 2*p 2p*

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O:  
  
 1*s* 2*s* 2*p* 2*p 2p*

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* **The spectroscopic electronic configuration**This method uses the orbitals *s* and *p* only in your case, although there are also *d* and *f* orbitals for elements with more than 20 electrons.  
    
  The electrons are denoted in the following forms:  
    
  1*s*2 means that there are 2 electrons in the *s* orbital of energy level (shell) 1 while 2*s*1 means that there is only 1 electron in the *s* orbital of energy level (shell) 2.  
    
  2*p*3 means that there are 3 electrons in the *p* orbital of energy level 2 while 3*p*1 means that there is only 1 electron in the *p* orbital of energy level (shell) 3.  
    
  As you already know, the order of filling (for the first 20 elements) is as follows: 1*s* 2*s* 2*p* 3*s*4*s*

For example: Sodium has 11 protons and therefore 11 electrons to place in orbitals.  
1*s*2 takes up 2 electrons  
2*s*2 takes up 2 more electrons (4 in total)  
2*p*6 takes up 6 more electrons (10 in total)  
3*s*1 takes up 1 more electron (all 11 accounted for)

The full spectroscopic electron configuration for sodium is therefore:  
 1*s*22*s*22*p*63*s*1

Note: Elements with one or two *s* electrons in their outermost shell are called s-block elements. They fall into Groups 1 and 2 of the Periodic Table.

Elements with one to six *p* electrons in their outermost shell are called p-block elements. They fall into Groups 13 to 18 of the Periodic Table.

**Relationship to core and valence electrons**

**Core** electrons are the electrons that occupy filled electron orbitals. They are stable electrons.

**Valence** electrons are electrons that occupy partially filled electron orbitals. These are the electrons that are of particular interest as they are less stable than core electrons They are the ones that get involved in chemical reactions. You will learn more about this later.

For example:

* Na: 1*s*22*s*22*p*63*s*1  
  The 1*s*22*s*22*p*6 orbitals are all full. The core electrons are in these orbitals.   
  3*s*1 is short of one electron so is not full. The single electron present is a valence electron.
* N: 1*s*22*s*22*p*3  
  The 1*s*22*s*2 orbitals are all full. The core electrons are in these orbitals.   
  2*p*3 is short of three electrons so is not full. The three electrons present are valence electrons.

Activity : Answer the following questions:

1. Draw Aufbau diagrams to show the electron configuration of each of the following elements:  
   magnesium, potassium, neon, nitrogen
2. Give the spectroscopic electron configuration for phosphorus (P) and draw an Aufbau diagram.
3. Complete the following table:

|  |  |  |  |
| --- | --- | --- | --- |
| **Element** | **Electron configuration** | **Core electrons** | **Valence electrons** |
| Sulfur (S) |  |  |  |
| Magnesium (Mg) |  |  |  |
| Lithium (Li) |  |  |  |
| Aluminium (Al) |  |  |  |

1. Answer the following multiple choice questions:
   1. Most of the volume of an atom is occupied by   
      (A). electrons.   
      (B). protons.   
      (C). neutrons.   
      (D). empty space
   2. The atomic number of an atom identifies the number of   
      (A). protons.   
      (B). neutrons.   
      (C). quantum orbits.   
      (D). excited states.
   3. What is the electron configuration for potassium?   
      (A). 1s22p63s43p64s2   
      (B). 1s42p63s23p64s2   
      (C). 1s22s22p63s23p64s1   
      (D).1s22s22p63s23p8

Answers :

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1. Mg:  
     
    1*s* 2*s* 2*p* 2*p 2p* 3*s*   
     
   K:   
    1*s* 2*s* 2*p* 2*p 2p* 3*s* 3*p* 3*p* 3*p* 4*s*  
     
   Ne:   
     
    1*s* 2*s* 2*p* 2*p 2p*N: 1*s* 2*s* 2*p* 2*p 2p*

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1. P has 15 electrons that need to be accommodated.  
   Spectroscopic electron configuration: 1*s*22*s*22*p*63*s*23*p*3  
     
   Aufbau diagram:

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1*s* 2*s* 2*p* 2*p 2p* 3*s* 3*p* 3*p* 3*p*

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| --- | --- | --- | --- |
| **Element** | **Electron configuration** | **Core electrons** | **Valence electrons** |
| Sulfur (S) | 1*s*22*s*22*p*63*s* 23*p*4 | 10 (1*s*22*s*22*p*6) | 6 (3*s* 23*p*4) |
| Magnesium (Mg) | 1*s*22*s*22*p*63*s* 2 | 10 (1*s*22*s*22*p*6) | 2 (3*s* 2) |
| Lithium (Li) | 1*s*22*s1* | 2 (1*s*2) | 1 (2*s1*) |
| Aluminium (Al) | 1*s*22*s*22*p*63*s* 23*p1* | 10 (1*s*22*s*22*p*6) | 3 (3*s* 23*p1*) |

1. 1. D
   2. A (and therefore electrons in a neutral atom).
   3. C

**Representing the spectroscopic electronic configuration of the first 20 elements of the Periodic Table**

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| 1A | 2A | 3A | 4A | 5A | 6A | 7A | 8A |
| 1 **H** 1s1 |  | | | | | | 2 **He** 1s2 |
| 3 **Li** 1s2 2s1 | 4 **Be** 1s2 2s2 | 5 **B** 1s2 2s22p1 | 6 **C** 1s2 2s22p2 | 7 **N** 1s2 2s22p3 | 8 **O** 1s2 2s22p4 | 9 **F** 1s2 2s22p5 | 10 **Ne** 1s2 2s22p6 |
| 11 **Na** 1s2 2s22p6  3s1 | 12 **Mg** 1s2 2s22p6 3s2 | 13 **Al** 1s2 2s22p6 3s23p1 | 14 **Si** 1s2 2s22p6 3s23p2 | 15 **P** 1s2 2s22p6  3s23p3 | 16 **S** 1s2 2s22p6 3s23p4 | 17 **Cl** 1s2 2s22p6 3s23p5 | 18 **Ar** 1s2 2s22p6 3s23p6 |

Below is a repeat of the third row in the table above. It shows an abbreviated way to express the spectroscopic electronic configuration of the elements represented. You will see that the first 2 lines of the configuration have been replaced by the symbol [Ne]. This is because the electronic configuration up to that point is exactly the same as that of Neon itself.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| 11 **Na** [Ne] 3s1 | 12 **Mg** [Ne] 3s2 | 13 **Al** [Ne] 3s23p1 | 14 **Si** [Ne] 3s23p2 | 15 **P** [Ne] 3s23p3 | 16 **S** [Ne] 3s23p4 | 17 **Cl** [Ne] 3s23p5 | 18 **Ar** [Ne] 3s23p6 |