Electron Configuration

Review

How do the quantum numbers play a part in the periodic table arrangement?

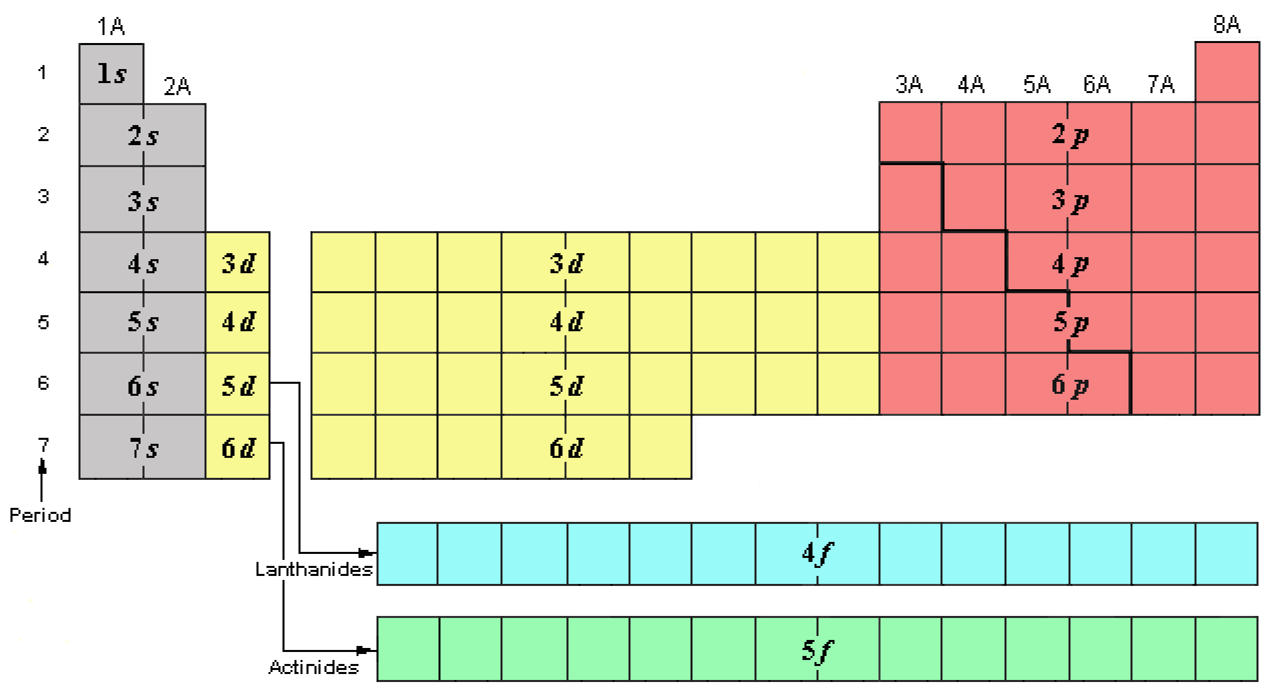
Review: the periodic table is arranged based on its atomic number and thus the number of electrons on its most outer shell.

Principal Quantum number – tells you the primary energy shell where the electrons are found (includes all the sublevels)
Sublevel – tells you where the electrons are by the orbitals that are found in the sublevel. Ex. 2nd energy level would have a "s" and 3 "p" orbitals. So a total of 4 orbitals.
Orbitals – the individual electron clouds that are found based on the probability (s, $p_{x,y,z}$, $d_{xz,\ yz,\ xy,\ x^2\text{-}y^2,\ z^2}$, f (has 7)) each p_x, p_y and p_z are individual orbitals

Arrangement

As you see below, the sublevels (s, p, d, f) are all found in different sections.
The energy level (distance away from the nucleus) increases from s < p < d < f. So the f sublevel has the highest energy or the greatest distance from the nucleus.

As a result, the s sublevels is filled first, then p, and so on.



Each orbital holds a maximum of 2 electrons.
"s" orbital holds – 2 electrons
Each p_x, p_y, p_z orbital holds – 2 electrons

But recall… calculating the maximum number of electrons is $2(n^2)$.
The energy level 2 (n=2) will give you – 8 electrons… how?
2 electrons from s and 6 electrons from p (p_x, p_y, p_z)

Energy level 4 (n=4) would have $2(4^2) = 32$ electrons… how?
2 from s, 6 electrons from p (p_x, p_y, p_z), 10 from d and 14 from f orbitals.

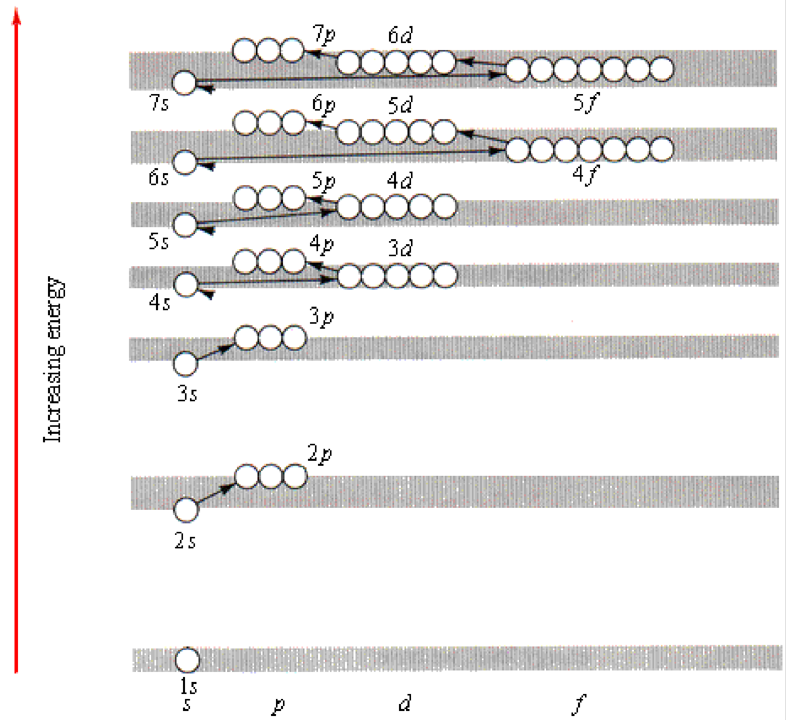
Writing electron configuration

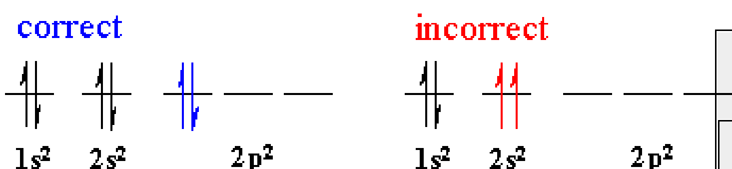
Aufbau Principle – German for "build up", the principle was proposed by Neils Bohr. Electrons are added starting at the lowest energy level and build up to the higher levels.

The "1s" orbital is filled first, and then 2s and then 2p and so forth. *Look at the periodic table above*

Notice 4s orbital has a lower energy than 3d orbital so the 4s is filled before the 3d orbital.

Pauli Exclusion Principle – Only a maximum of two electrons can occupy an orbital if they have opposite spins.





Here's a guide for you to write your electron configuration			
1s			
2s			2p
3s			3p
4s		3d	4p
5s		4d	5p
6s	4f	5d	6p
7s	5f	6d	7p

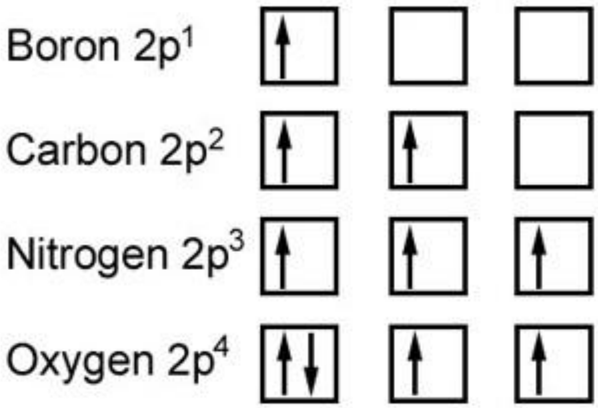
Examples:

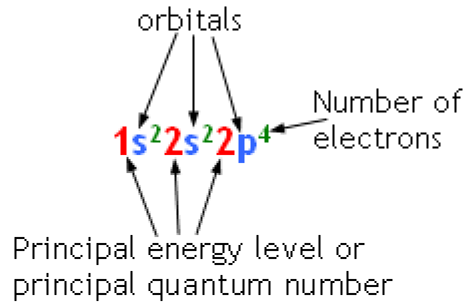
He (Atomic Number 2) - $1s^2$

B (Atomic Number 5) –

C (A.N 6) –

Hund's Rule – Proposed that as the electrons fill up the orbitals, they tend to maximize unpaired electrons in filling all the orbitals before completely filling one. In other words, before filling the first p-orbital with two electrons, an electron is placed into the p_x orbital, then another electron into the p_y then p_z before filling the p_x orbital.





Sample electron configuration

Short hand rule

Instead of writing all the sublevels for very large atoms, using the noble gas in the period before the atom cause save a lot of time.

Example:

6) cobalt ______**[Ar] $4s^2 3d^7$**______

7) silver ______**[Kr] $5s^2 4d^9$**______

8) tellurium ____________

9) radium ____________

10) lawrencium ____________

Exceptions

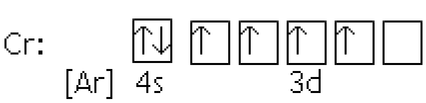
Chromium, **copper** and their families undergo what is called **electron promotion**.

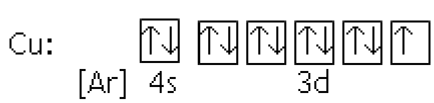
If we follow the aufbau principle, the electron configurations for chromium and copper should be

Cr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$ or $[Ar]4s^2 3d^4$

Cu $1s^2 2s^2 2p^6 3s^2 23p^6 4s^2 3d^9$ or $[Ar]4s^2 3d^9$

If we look at the orbital diagrams for the electrons outside the kernel:





The increased stability of full or half-filled orbitals causes the electron promotion.

The actual configurations are

Cr $1s^2 2s^2 2p^6 3s^2 3p^6$**$4s^1 3d^5$** or [Ar]**$4s^1 3d^5$**

Cu $1s^22s^22p^63s^223p^6$**$4s^13d^{10}$** or [Ar]**$4s^13d^{10}$**

The orbital diagrams illustrate why this configuration is an advantage.

Cr: [Ar] 4s ↑ 3d ↑ ↑ ↑ ↑ ↑ Cu: [Ar] 4s ↑ 3d ↑↓ ↑↓ ↑↓ ↑↓ ↑↓

In chromium, promotion of a 4s electron to the empty 3d orbital creates half-filled orbitals.

In copper, using a 4s electron to fill the 3d orbital is a more stable configuration.

At higher energy levels, the exceptions to the aufbau principle increase because the energy differences between sublevels become smaller. For example, there are many more exceptions when filling the 5d and 4f orbitals.

Electron Configuration Examples:

1) Sodium ____

2) iron ____

3) bromine ____

4) barium ____

5) neptunium ____

Determine what elements are denoted by the following electron configurations:

11) $1s^22s^22p^63s^23p^4$ ____________________

12) $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^1$ ____________________

13) [Kr] $5s^24d^{10}5p^3$ ____________________

14) [Xe] $6s^24f^{14}5d^6$ ____________________

15) [Rn] $7s^25f^{11}$ ____________________

1) sodium **$1s^22s^22p^63s^1$**

2) iron **$1s^22s^22p^63s^23p^64s^23d^6$**

3) bromine **$1s^22s^22p^63s^23p^64s^23d^{10}4p^5$**

4) barium **$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^2$**

5) neptunium **$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^67s^25f^5$**

11) $1s^22s^22p^63s^23p^4$ **sulfur**

12) $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^1$ **rubidium**

13) [Kr] $5s^24d^{10}5p^3$ **antimony**

14) [Xe] $6s^24f^{14}5d^6$ **osmium**

15) [Rn] $7s^25f^{11}$ **einsteinium**

Electron Configurations of Ions

When writing the valence configurations of ions you must consider the charge of the ion. The charge of the ion will determine the number of electrons. A positive ion will have fewer electrons and a negative ion will have more electrons.

Example 3. Write the complete electron configuration for the chloride ion, Cl^-.

Step 1. Determine the number of electrons.

Atomic number = 17

A negative one ion, so one more electron than the neutral atom.

Number of electrons = 17 + 1 = 18

$1s^22s^22p^63s^23p^6$

Example 4. Write the complete electron configuration for the calcium ion, Ca^{2+}.

Step 1. Determine the number of electrons.

Atomic number = 20

A positive two ion, so 2 less electrons than the neutral atom.

Number of electrons = 20 − 2 = 18

$1s^22s^22p^63s^23p^6$

Notice that the calcium ion and the chloride ion have the same electron configuration. They both have the same electron configuration as the noble gas argon. We say the chloride ion, calcium ion and argon atom are **isoelectronic**.

Example 5. Write the complete electron configuration for the iron (II) ion, Fe^{2+}.

Step 1. Determine the number of electrons.

Atomic number = 26

A positive two ion, so 2 less electrons than the neutral atom.

Number of electrons = 26 − 2 = 24

$1s^22s^22p^63s^23p^63d^6$

Electrons removed from metal ions to create ions are always removed from the orbital(s) with the highest principal quantum number.

Example 6. Write the complete electron configuration for the iron (III) ion, Fe3+.

Step 1. Determine the number of electrons.

Atomic number = 26

A positive three ion, so 3 less electrons than the neutral atom.

Number of electrons = 26 − 3 = 23

$1s^22s^22p^63s^23p^63d^5$

Iron can form both a 2+ and a 3+ ion because of its electron configuration. The 2+ ion removes the valence electrons. The 3+ ion forms a very stable structure with each d-orbital containing one electron, or being half-filled.

Other metals are able to form more than one ion charge for a similar reason. Removing the valence shell for one ion charge and then removing electrons from the d-orbital for another.

Valence Configurations

he valence electrons are the electrons found in the outer-most or highest energy level.

Example 4. Write the valence electron configuration for fluorine.

Solution.

Step 1. Write the complete electron configuration.

F = 9 electrons

$1s^22s^22p^5$

Step 2. Choose the electrons in the highest energy level.

The highest energy level is the second, so the electrons in the second energy level are the valence electrons.

$1s^2\boxed{2s^22p^5}$

The valence configuration is $2s^22p^5$

Example 5. Write the valence electron configuration for germanium.

Solution.

Step 1. Write the complete electron configuration.

From a previous example

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$

Step 2. Choose the electrons in the highest energy level.

The highest energy level is the fourth, so the electrons in the fourth energy level are the valence

$1s^2 2s^2 2p^6 3s^2 3p^6 \boxed{4s^2} 3d^{10} \boxed{4p^2}$

Even though the 3d orbital fills after the 4s orbital, the 4s orbital is in the highest or outer-most energy level.

The valence configuration is $4s^2 4p^2$.

Exercise

1. How many electrons in an atom can have the designation

   a. 1s
   b. 2p
   c. $3p_x$
   d. 6f
   e. $3d_{xy}$
   f. n=2
   g. 4p
   h. n=5

2. Write complete electronic configurations for the following atoms and ions:

   a. P
   b. Ca
   c. Cu
   d. Rh
   e. Sb^{3+}
   f. Ni^{2+}
   g. Fe^{2+}
   h. Ni^{4+}
   i. Zn^{2+}
   j. Br^{-}
   k. Sn^{2+}
   l. Co^{3+}

3. How many unpaired electrons are there in each of the following:

   a. Mn
   b. As
   c. Sr
   d. Tl^{+}
   e. Cu^{2+}
   f. V^{3+}
   g. Sn
   h. Lu
   i. Te

4. Write the electronic configurations for the <u>valence</u> electrons of each of the following.

   a. Mg
   b. P
   c. Se
   d. Pb^{2+}
   e. Br
   f. S^{2-}
   g. Ni
   h. Ag^{+}
   i. N^{3-}
   j. Fe^{3+}

Answer Key

1. How many electrons in an atom can have the designation

a. 2
b. 6
c. 2 (individual p-orbital)
d. 14
e. 2 (individual d-orbital)
f. $2n^2 = 8$
g. 6
h. $2n^2 = 50$

2. Write complete electronic configurations for the following atoms and ions:

a. P = $1s^2 2s^2 2p^6 3s^2 3p^3$
b. Ca = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
c. Cu = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
d. Rh = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^7$
e. Sb^{3+} = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$
f. Ni^{2+} = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^8$
g. Fe^{2+} = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^6$
h. Ni^{4+} = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^6$
i. Zn^{2+} = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^{10}$
j. Br^- =$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{104p6}$
k. Sn^{2+} = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$
l. Co^{3+} = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^0 3d^6$

3. How many unpaired electrons are there in each of the following:

a. Mn = 5
b. As = 3
c. Sr = 0
d. Tl^+ = 1
e. Cu^{2+} = 4
f. V^{3+} = 2 (remove 3s electrons first)
g. Sn = 2
h. Lu = 1
i. Te = 2

4. Write the electronic configurations for the valence electrons of each of the following.

a. Mg = $3s^2$
b. P = $3s^2 3p^3$
c. Se = $4s^2 4p^4$
d. Pb^{2+} = $6s^2$
e. Br = $4s^2 4p^5$
f. S^{2-} = $3s^2 3p^6$
g. Ni = $4s^2$
h. Ag^+ = $4s^1$
i. N^{3-} = $2s^2 2p^6$
j. Fe^{3+} = $3s^2 3p^6 3d^5$

Answer the following questions.

1. Give the complete electron configuration for each of the following. (6 marks)

a. potassium
b. argon
c. chromium
d. P^{2-}
e. Se^{2+}
f. Ag^{+}

2. Write the valence electron configurations for each of the following.

   (2 marks)

   a. sulfur
   b. lead
   c. zinc
   d. bromine

3. How many unpaired electrons are present in each of the atoms in #2?

   (2 marks)

4. Identify the following elements: (2 marks)
   a. An excited state of this element has the electronic configuration $1s^2 2s^2 2p^5 3s^1$.
   b. The ground state electron configurations is $[Ne]3s^2 3p^4$.
   c. An excited state of this element has the electronic configuration $[Kr]5s^2 4d^6 5p^2 6s^1$.
   d. The ground state electron configuration contains three unpaired 6p electrons.