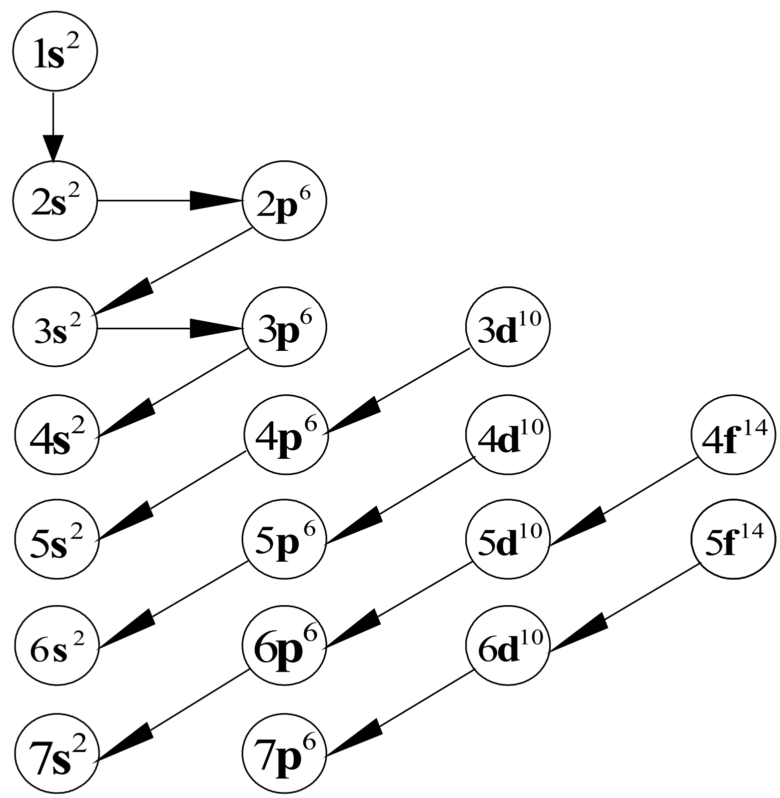
ELECTRON CONFIGURATIONS OF ELEMENTS(GROUND STATE)

When given an atom in chemistry, we often want to know how the electrons are arranged in that atom. The arrangement is called the electron configuration of that atom. The electrons fall into orbits which are different distances from the nucleus of the atom. These orbits can be looked upon like orbits in the solar system. The orbits are called levels. There are 7 energy levels that will be studied: **1(*lowest*), 2, 3, 4, 5, 6, 7**(*highest*). Levels correspond to different energy levels of the electrons of an atom and roughly correspond to the rows of the periodic table. The electrons with higher energy are in the outermost levels of the atom.

Levels are divided into sublevels. There are 4 sublevels: **s, p, d, f**. Sublevels are divided into orbitals which hold two electrons apiece. The **s** sublevel has one orbital(*2 electrons*). The **p** sublevel has three orbitals(*6 electrons*). The **d** sublevel has five orbitals(*10 electrons*), and the **f** sublevel has seven orbitals(*14 electrons*).

To find the electron configuration of a particular atom, look up the atomic number of that atom. Recall that the atomic number is the number of protons in the nucleus of an atom which also equals the number of electrons. Use the chart below and follow the arrows. As you follow the arrows, add the exponents and stop at the first bubble where the sum of the exponents equals or exceeds the atomic number. If you end up at an **s** orbital with no arrow, "jump" to the next orbital as follows: $4\mathbf{s}^2$ jumps to $3\mathbf{d}^{10}$; $5\mathbf{s}^2$ jumps to $4\mathbf{d}^{10}$; $6\mathbf{s}^2$ jumps to $4\mathbf{f}^{14}$; $7\mathbf{s}^2$ jumps to $5\mathbf{f}^{14}$. Write down the sublevels(*the bubbles in the below diagram*) as you go. Make sure that you have the exponents add to the atomic number of the atom. If necessary, reduce the exponent in the last bubble you stopped at.



Example 1: Find the electron configuration for carbon(**C**). Looking in the periodic table, we find that carbon has atomic number **6**. Starting at the $1\mathbf{s}^2$ bubble, adding the exponents, and stopping at the first bubble that equals or exceeds the atomic number **6**, we have $1\mathbf{s}^2 2\mathbf{s}^2 2\mathbf{p}^2$. To make sure that the exponents add up to **6**, we reduced the exponent in $2\mathbf{p}^6$ to **2**.

Example 2: Find the electron configuration for rhodium(**Rh**). Looking in the periodic table, we find that rhodium has atomic number **45**. Starting at the $1\mathbf{s}^2$ bubble, adding the exponents, and stopping at the first bubble that equals or exceeds the atomic number **45**, we have $1\mathbf{s}^2 2\mathbf{s}^2 2\mathbf{p}^6 3\mathbf{s}^2 3\mathbf{p}^6 4\mathbf{s}^2 3\mathbf{d}^{10} 4\mathbf{p}^6 5\mathbf{s}^2 4\mathbf{d}^7$. To make sure that the exponents add up to **45**, we reduced the exponent in $4\mathbf{d}^{10}$ to **7**.

There are exceptions to the above rule as the below examples will show:

Example 3: Find the electron configuration for silver(**Ag**). Looking in the periodic table, we find that silver has atomic number **47**. Starting at the $1\mathbf{s}^2$ bubble, adding the exponents, and stopping at the first bubble that equals or exceeds the atomic number **47**, we expect to have $1\mathbf{s}^2 2\mathbf{s}^2 2\mathbf{p}^6 3\mathbf{s}^2 3\mathbf{p}^6 4\mathbf{s}^2 3\mathbf{d}^{10} 4\mathbf{p}^6 5\mathbf{s}^2 4\mathbf{d}^9$. To make sure that the exponents add up to **47**, we reduced the exponent in $4\mathbf{d}^{10}$ to **9**.

This is not the case because the 4**d** sublevel is one electron shy of being full and *a full sublevel paired with a half-full sublevel is more stable*. Hence, an electron is taken from the 5**s** sublevel making it half-full and put with the 4**d** sublevel making it full. Thus, the electron configuration of silver is $1\mathbf{s}^2 2\mathbf{s}^2 2\mathbf{p}^6 3\mathbf{s}^2 3\mathbf{p}^6 4\mathbf{s}^2 3\mathbf{d}^{10} 4\mathbf{p}^6 5\mathbf{s}^1 4\mathbf{d}^{10}$. Copper(**Cu**) and gold(**Au**), the other two coinage metals, also display this behavior.

Example 4: Find the electron configuration for chromium(**Cr**). Looking in the periodic table, we find that chromium has atomic number **24**. Starting at the $1\mathbf{s}^2$ bubble, adding the exponents, and stopping at the first bubble that equals or exceeds the atomic number **24**, we expect to have $1\mathbf{s}^2 2\mathbf{s}^2 2\mathbf{p}^6 3\mathbf{s}^2 3\mathbf{p}^6 4\mathbf{s}^2 3\mathbf{d}^4$. To make sure that the exponents add up to **24**, we reduced the exponent in $3\mathbf{d}^{10}$ to **4**.

Again, this is not the case because the 3**d** sublevel is one electron shy of being half-full and *a half-full sublevel paired with a half-full sublevel is also more stable*. Hence, an electron is taken from the 4**s** sublevel making it half-full and put with the 3**d** sublevel make it a half-full sublevel. Thus, the electron configuration of chromium is $1\mathbf{s}^2 2\mathbf{s}^2 2\mathbf{p}^6 3\mathbf{s}^2 3\mathbf{p}^6 4\mathbf{s}^1 3\mathbf{d}^5$.