Section 5

Electron Configuration

Electron Configuration

Introduction

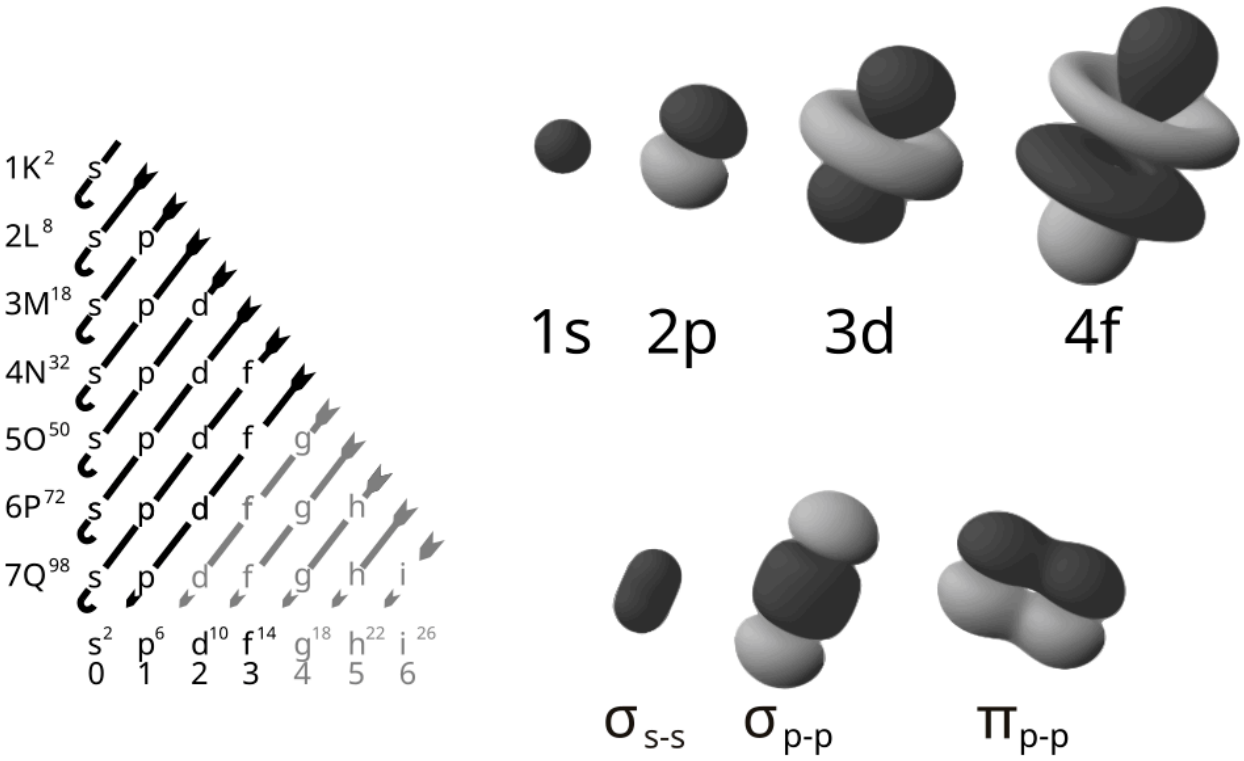
- Know the electronic configurations of the elements; exceptions to know are Cr (both 4s and 3d orbitals are half full) and Cu (4s is half full and 3d is completely full), the rest will be examples that follow the periodic table
- If you learn how to recognize it, this information is displayed on the periodic table. The position of an element on the table corresponds to its highest energy electron. The elements with their highest energy electron in an s type orbital are found in the first 2 left-hand columns corresponding to the alkali and alkaline earth metals; those with the highest energy electron in a p type orbital are found in the six right-hand side columns corresponding to the six columns headed by the elements B through Ne (all of the elements with highest energy electrons in either an s or p type orbital comprise the main group elements); those with this highest energy electron in a d type orbital are found in the ten middle columns corresponding to the transition metal elements; those with the highest energy electron in a f type orbital are found in the two separate rows at the bottom of the table and correspond to the lanthanides (rate earths) and actinides.
- Know the general shape of the orbits: s (spherical) and p type orbitals (bilobal or dumbbell). Be able to recognize a d orbital and don't worry about the f orbitals (if its not an s or p type orbital then, it must be a d type). Know the maximum number of electrons allowed in s, p, d, and f type orbitals (again the periodic table has the answer).
- Be able to correlate an element's position on the periodic table with its electronic structure and with a group, ie. alkali metals, alkaline earths, halogens, etc.

2.5.1 Electron Configuration

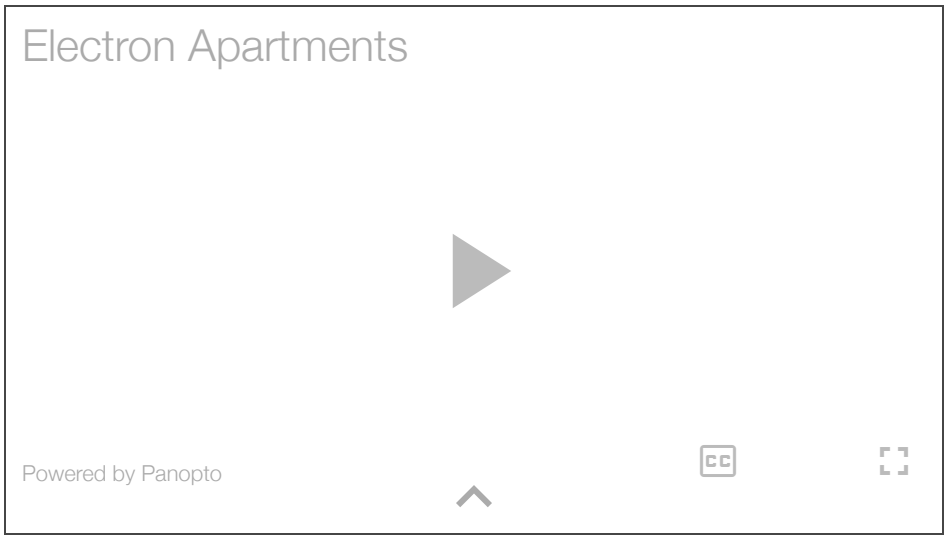
The available electron orbitals or "apartments" always fill from the lowest energy orbital or "ground floor" up in an orderly way. The $n = 1$ level has one s orbital and is occupied first by a spin-up and then by a spin-down electron. In the $n = 2$ level, the s-type orbital is of lower energy than the three p-type orbitals. The s orbital fills with two electrons before the p orbitals begin to fill up. First, each p orbital will fill with one electron, then when all three orbitals are half full, the fourth through sixth electrons will double-fill each of the p orbitals. The fact that all available electron orbitals of the same energy level half-fill before they double-fill is known as Hund's rule. This trend continues as more and more electrons are added. There are, however, some

unexpected exceptions to this order. Most notably, the $n = 4$ s orbital fills before the $n = 3$ d orbitals fill. In addition, the electrons in neutral chromium (Cr) and copper (Cu) atoms adopt more favorable half-filled or fully filled electron arrangements that are overall lower in energy.

The electron occupancy of an element is known as its electron configuration. There are different ways of writing out the electron configuration of an element. Individual orbitals can be represented as circles or boxes labeled with the principal quantum number and orbital type, and the electrons themselves fill the boxes as arrows pointing either up or down to represent spin up or spin down. Orbitals may also be represented by their principal quantum number and type of orbital—s, p, d, or f. When the configuration is given in this format, the number of electrons in any given orbital follows the orbital type as a superscripted number. For example, the element carbon has six electrons. Two of the six electrons will occupy the lowest energy 1s orbital, which is written as $1s^2$. The next lowest energy orbital is the 2s orbital, and two electrons also occupy it. This is written as $2s^2$. Finally, the remaining two electrons occupy the 2p orbitals, written as $2p^2$. Taken together in this format, the electron configuration of carbon is $1s^2\,2s^2\,2p^2$. This notation is often shortened by placing the elemental symbol of the appropriate noble gas element in brackets to indicate the lower energy-filled subshells. Using carbon and calcium as examples, written in this format the electron configuration of carbon is $[\mathrm{He}]\,2s^2\,2p^2$ and calcium is $[\mathrm{Ar}]\,4s^2$.



$$1s^2_2\,2s^2_4\,2p^6_{10}\,3s^2_{12}\,3p^6_{18}\,4s^2_{20}\,3d^{10}_{30}\,4p^6_{36}\,5s^2_{38}\,4d^{10}_{48}\,5p^6_{54}\,6s^2_{56}\,4f^{14}_{70}\,5d^{10}_{80}\,6p^6_{86}\,7s^2_{88}\,5f^{14}_{102}\,6d^{10}_{112}\,7p^6_{118}$$



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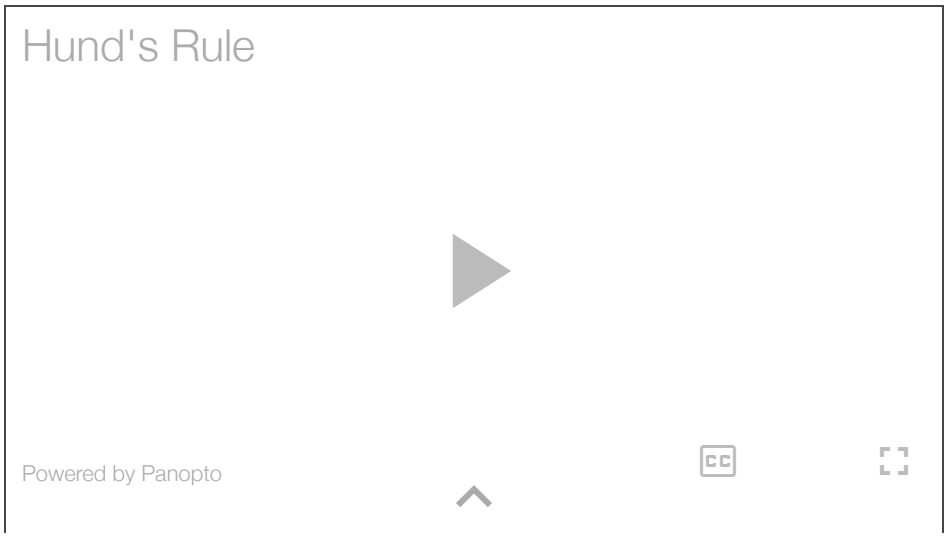

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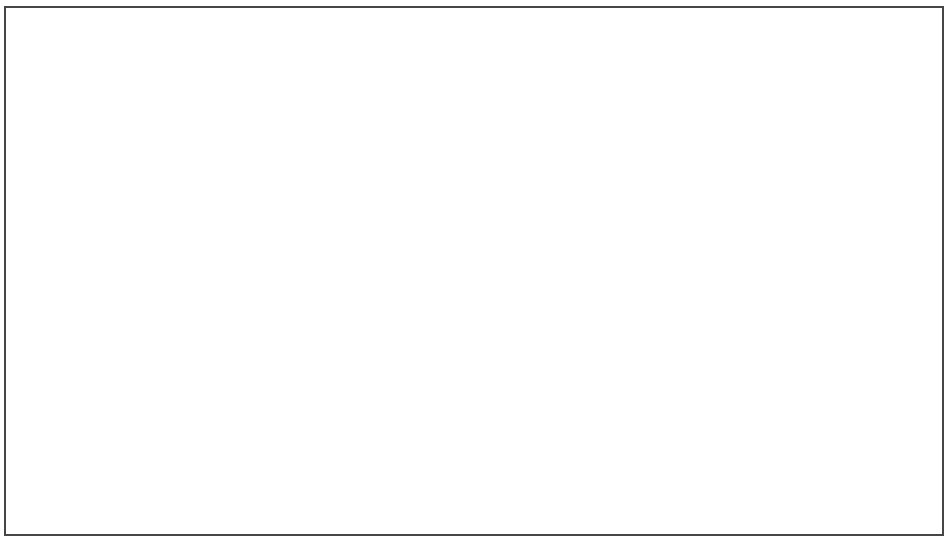

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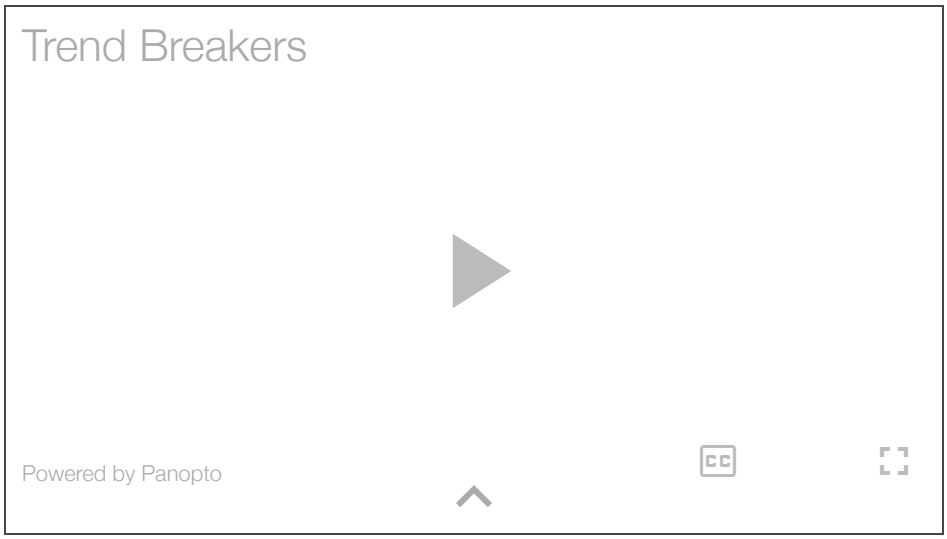

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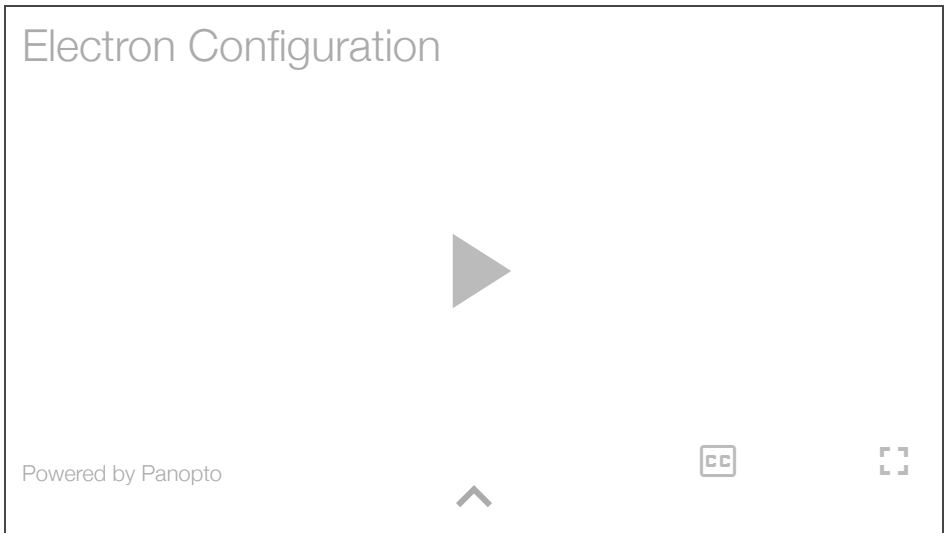

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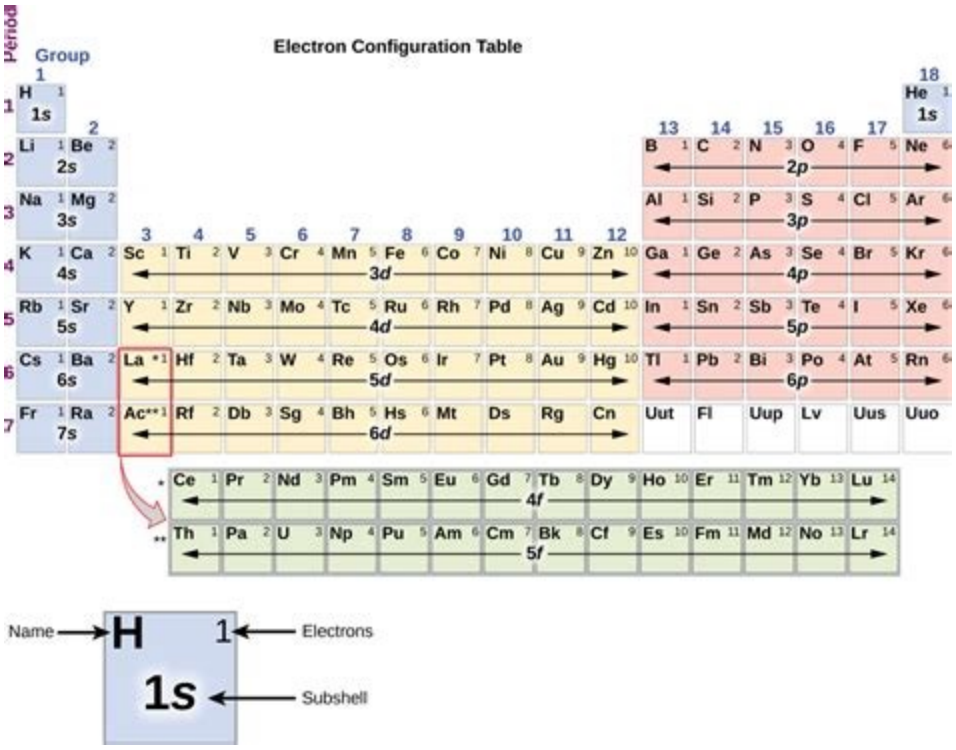
2.5.2 Electron Configuration and the Periodic Table

The electron configuration of each element identifies the orbital location of the highest energy electrons in neutral or uncharged atoms of the element. Consequently, the placement of an element on the periodic table is directly correlated to its electron configuration. For example, the electron configuration of the element sodium, symbol Na, is $[\mathrm{Ne}]\,3s^1$, and the highest energy electron in a neutral sodium atom occupies a $3s$ orbital. Likewise, the highest energy electron in a neutral potassium atom (symbol K), whose electron configuration is $[\mathrm{Ar}]\,4s^1$, occupies a $4s$ orbital. On the periodic table, the element K lies directly below Na. Both elements belong to the alkali metals group. This same positional pattern is true for all of the elements on the periodic table. This is seen clearly in the electronic configuration of the halogens F, Cl, and Br. Their electron configurations are $[\mathrm{He}]\,2s^2\,2p^5$, $[\mathrm{Ne}]\,3s^2\,3p^5$, and $[\mathrm{Ar}]\,3d^{10}\,4s^2\,4p^5$, respectively. The electron configurations of the highest energy electrons for each of these elements, which appear in the same column of the periodic table, is $s^2\,p^5$. They all have the same number of electrons in the same types of orbitals.

Starting with the element H in the upper left corner of the periodic table, each row or period across the table correlates with the principal quantum number increased by one. The elements in the two columns on the far left all have their highest energy

electrons in s type orbitals. Elements in these two columns are often referred to as s-block elements. The six columns on the far right side of the periodic table have their highest energy electrons in p-type orbitals. These are the p-block elements. Similarly, the ten columns between the s-block and **p-block** elements are the d-block elements, and the two rows of fourteen elements at the bottom of the table are the f-block elements. The periodic table is a direct reflection of the electronic configuration of the elements, and as such, the electronic configuration of most all of the elements can be read directly from it.



Electronic structure of the elements is reflected on periodic table and an element's position on the periodic table reflects its chemistry. Therefore, the electronic structure of an element reflects its chemistry. It is, in fact, what determines chemistry of an element. The far two right-hand columns correlate with s type orbitals, the far left-hand six columns correlate with the three p type orbitals, the middle ten columns of the transition metals correlate with the five d type orbitals, and the fourteen columns at the bottom of the table correlate with the seven f type orbitals.

In general, the electronic configuration of a neutral uncharged element can be "read" directly from the periodic chart.

The elements in the same column have similar chemistries and the same number of highest energy electrons in same orbital configuration.

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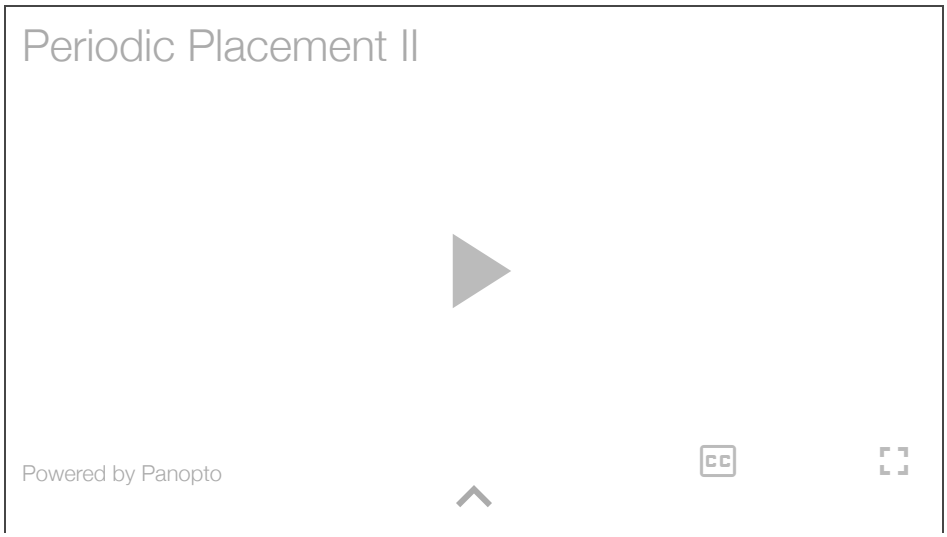

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Electron Configuration Practice

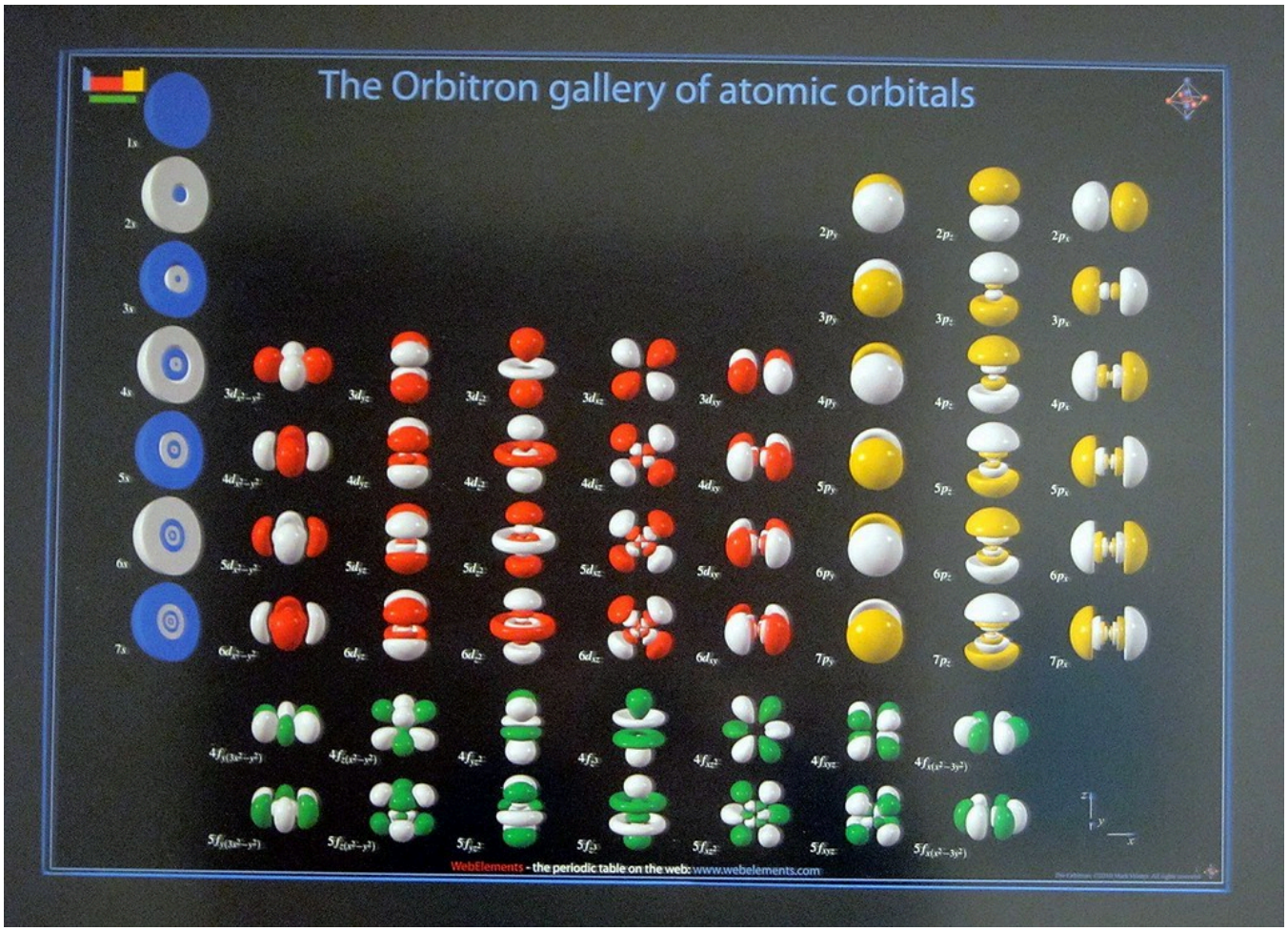
Electron Configurations Concept Check Activity

What is the correct electron configuration of Argon (Ar)?
Hint: The exponents always represent the number of electrons in that orbital

- $1s^2 1p^6 2s^2 2p^6$
- $1s^2$
- $1s^2 1p^6 2s^2 2p^6 3s^2 3p^6 4s^2 4p^6$
- $1s^2 1p6^3 s2^3 p^6$

Check

2.5.2 Configuration and the Periodic Table



What you need to know:

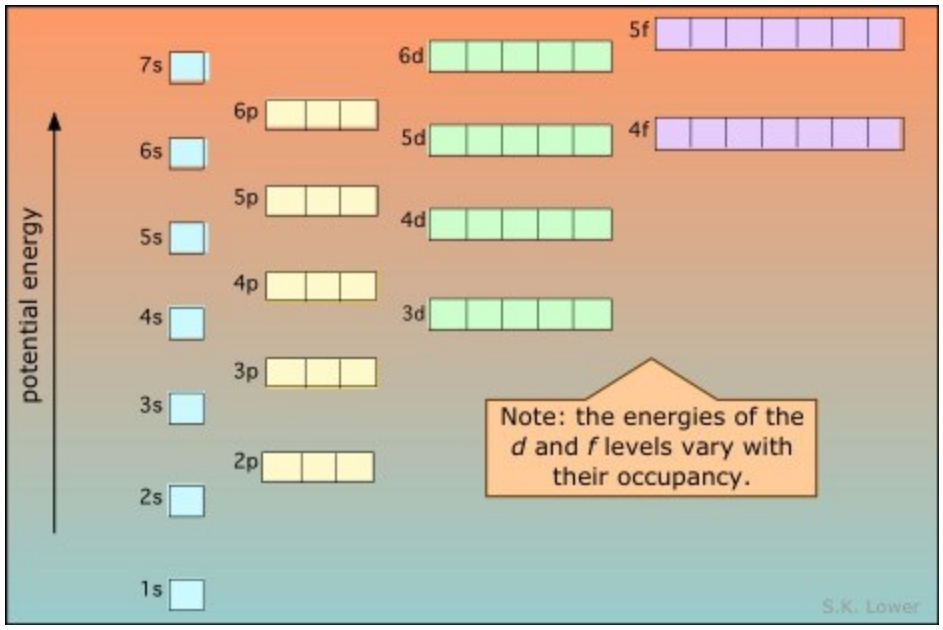
You need to know how determine the electron configuration of an element and to look at the electron configuration and identify the element. You will be provided a periodic table to assist you with determining electron configurations. You will also need to be able to visualize the orbitals occupied by electrons to some degree.

How to learn it:

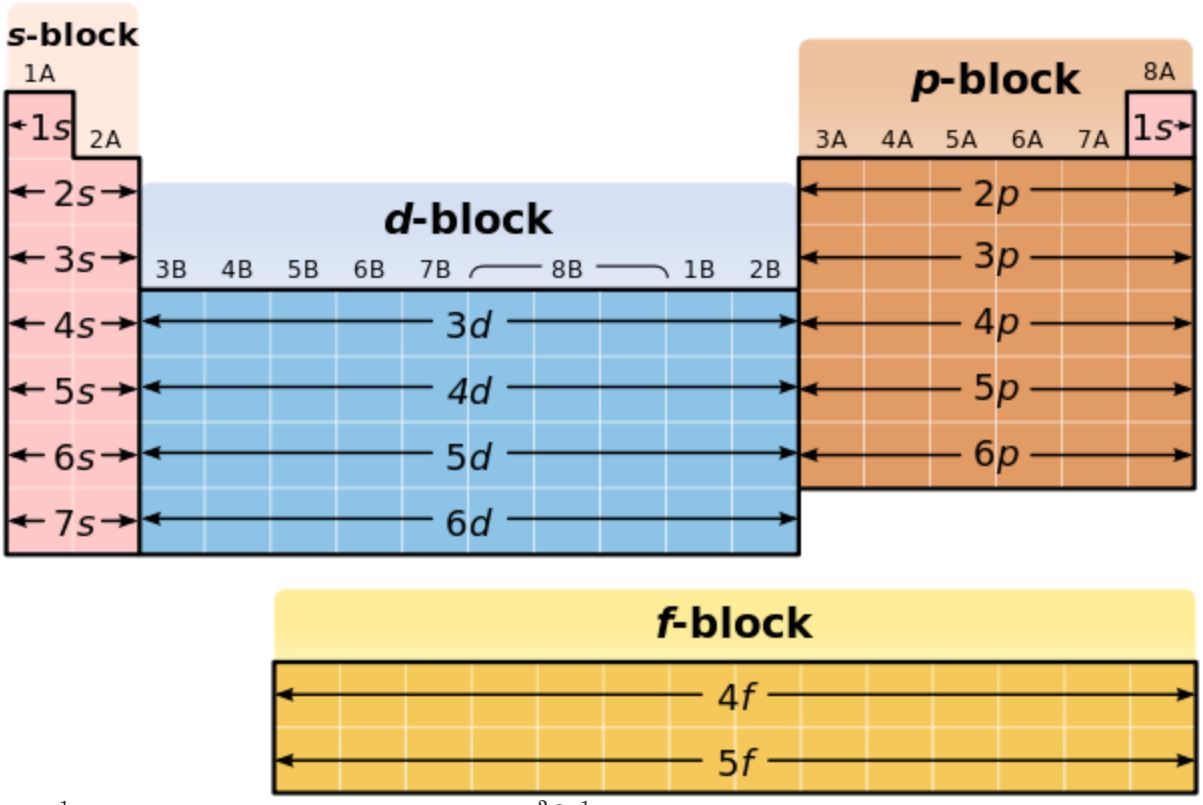
The periodic table is your best friend when you wish to determine electron configuration. It is arranged in rows with increasing atomic number and in columns with similar chemical properties. The chemical properties of an element are a reflection of its electron configuration. Elements in the same group have similar electron configurations of their highest energy electron. Thus, the periodic table reflects an element's electron configuration, because it groups elements with similar chemical properties in columns.

The electron configuration of oxygen is $1s^2\,2s^2\,2p^4$. To read this, take it step by step. Superscripted numbers represent the number of electrons present in an orbital. The letters in an electron configuration represent the type of orbital. The numbers preceding the letters represent the energy level the orbital resides in. $1s^2$ means that there are two electrons in the s orbital ($\ell = 0$) in the first energy level ($n = 1$). $2s^2$ means that there are two electrons in the s orbital ($\ell = 0$) in the second energy level ($n = 2$). $2p^4$ means that there are four electrons in the p orbital ($\ell = 1$) in the second energy level ($n = 2$).

The rows of the periodic table correspond to the primary quantum number, n. The topmost row of the periodic table corresponds to $n = 1$, the next one down is $n = 2$, and so forth. Elements on rows beyond the first row have a complete set of all the electrons in the rows preceding.



The two left-hand columns of the periodic table correspond to the s orbital. The farthest left group, the alkali metals, correspond to elements that have one electron in their respective s orbital. Thus, Hydrogen has an electron configuration of $1s^1$, Lithium has an electron configuration of $1s^2\,2s^1$. The second-to-the-left group, the alkaline earth metals, correspond to elements that have two electrons in their respective s orbitals, and thus have full s orbitals. Thus, Beryllium has an electron configuration of $1s^2\,2s^2$, and Magnesium has an electron configuration of $1s^2\,2s^2\,2p^6\,3s^2$.



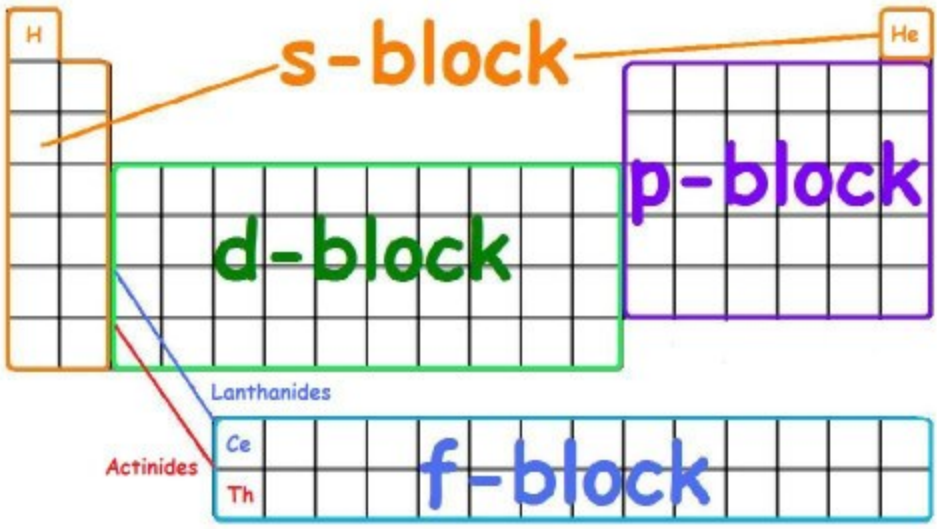
Note: Helium appears to be in the "wrong" place on the periodic table, because it has a $1s^2$ electron configuration, and should thus be in the alkaline earth metals group. However, Helium reacts similarly to the noble gases because it has a full subshell and is placed at the far right-hand side at the head of the noble gas column.

The p orbitals correspond to the block of 6 groups on the right side of the periodic table. The noble gas group has an electron configuration that ends in p^6 (Helium, however, has an electron configuration of $1s^2$).

The d orbitals correspond to the 10 transition metal groups present in the middle of the periodic table. Interestingly, the top row of the d block corresponds to $3d$ even though it's on the $n = 4$ row. This is because $4s$ is filled with electrons first before the $3d$ orbitals because they are actually lower in energy than the $3d$ orbitals.

The f orbitals correspond to the Lanthanides and Actinides, listed below the main section of the periodic table. We will not be using these often in our introductory chemistry course. The top row of the f block corresponds to $4f$, even though it is on the $n = 6$ row. This is due to the discrepancies in energy associated with the f orbitals compared to the s, p, and d orbitals.

When writing an electron configuration, write it either in full or use the noble gas shorthand. The noble gas shorthand substitutes the symbol for the noble gas that corresponds to the core electrons, which are those electrons that occupy completely filled inner shells. For example, Strontium has an extended electron configuration of



$1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6\,5s^2$. The valence electrons are the $5s^2$ electrons, which are those electrons in partially filled shells. The core electrons are $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2\,4p^6$, which is the electron configuration for Krypton. Thus, the noble gas shorthand for Strontium is $[\mathrm{Kr}]\,5s^2$.

There are some exceptions to the general rule with electron configurations. Cr, Mo, Cu, Ag, and Au are the exceptions you should know about. Chromium has an electron configuration of $[\mathrm{Ar}]\,3d^5\,4s^1$. This can also be written as $[\mathrm{Ar}]\,4s^1\,3d^5$ or as $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^5\,4s^1$. This should give you pause, as the $4s$ orbitals are lower in energy than the $3d$ orbitals and thus should be filled first. So by this reasoning, the electron configuration should be $[\mathrm{Ar}]\,4s^2\,3d^4$. However, when the $3d$ orbital is half-full or full, it's lower in energy than the $4s$ orbital. Consequently, one of the $4s$ electrons occupies a $3d$ position such that the $3d$ sub-shell is half (5 electrons) or fully filled (10 electrons). Thus, $[\mathrm{Ar}]\,4s^1\,3d^5$ is lower in energy than $[\mathrm{Ar}]\,4s^2\,3d^4$, and therefore the true electron configuration is $[\mathrm{Ar}]\,4s^1\,3d^5$, because nature always preferentially forms the lowest energy configuration. This same pattern holds true for the five exceptions: Cr, Mo, Cu, Ag, Au. W is not an exception; its electron configuration is $[\mathrm{Xe}]\,4f^{14}\,5d^4\,6s^2$.

Why it matters:

Reactivity and periodic trends are almost solely due to electron configurations of elements and ions. Understanding electron configurations will help you understand how an element may react with another element, and how many bonds it can possibly form. We will discuss this further.

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