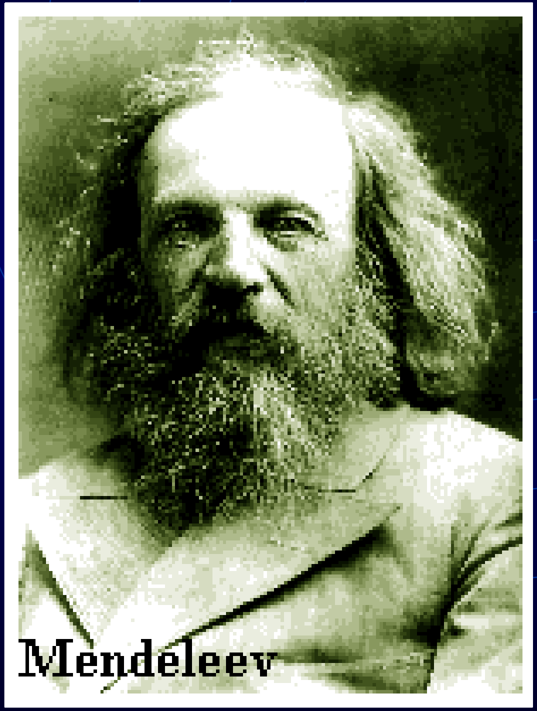
Periodic Trends

Elemental Properties and Patterns

Who created it?



- The quest for a systematic arrangement of the elements started with the discovery of individual elements.
- By 1860 about 60 elements were known and a method was needed for organization.
- In 1869, Russian chemist Dimitri Mendeleev proposed arranging elements by **atomic weights** and properties.
- The table contained gaps but Mendeleev predicted the discovery of new elements.

The Periodic Law

- Mendeleev understood the ‘Periodic Law’ which states:
- When arranged by increasing atomic number, the chemical elements display a regular and repeating pattern of chemical and physical properties.

The Periodic Law

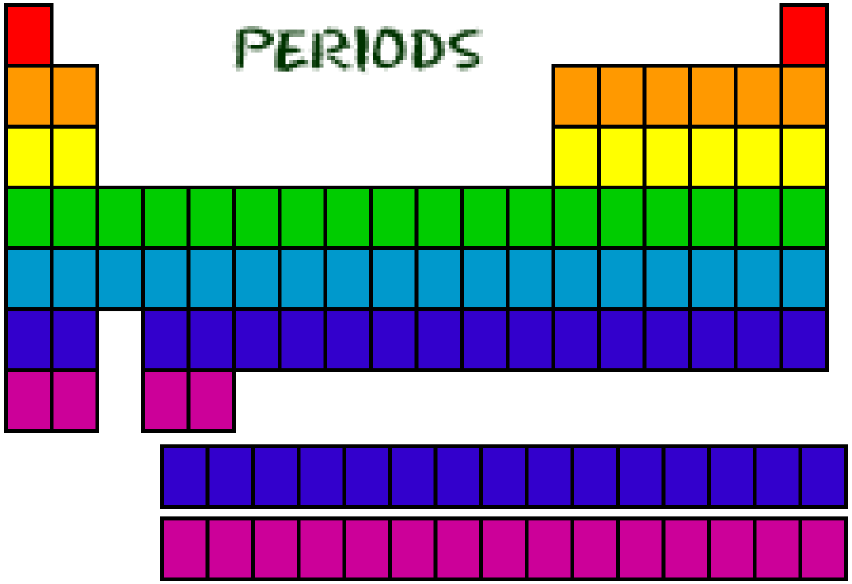
- Atoms with similar properties appear in groups or families (vertical columns) on the periodic table.
- They are similar because they all have the same number of valence (outer shell) electrons, which governs their chemical behavior.

So how is it arranged?

- The genius of the periodic table "is that it is organized like a big grid. The **elements** are placed in specific places because of the way they look and act. If you have ever looked at a grid, you know that there are rows (left to right) and columns (up and down). The periodic table has rows and columns, too, and they each mean something different."

- quoted from http://www.chem4kids.com/files/elem_pertable.html

You've got Your Periods...



- Even though they skip some squares in between, all of the rows go left to right. When you look at a periodic table, each of the rows is considered to be a different **period** (Get it? Like PERIODic table.)

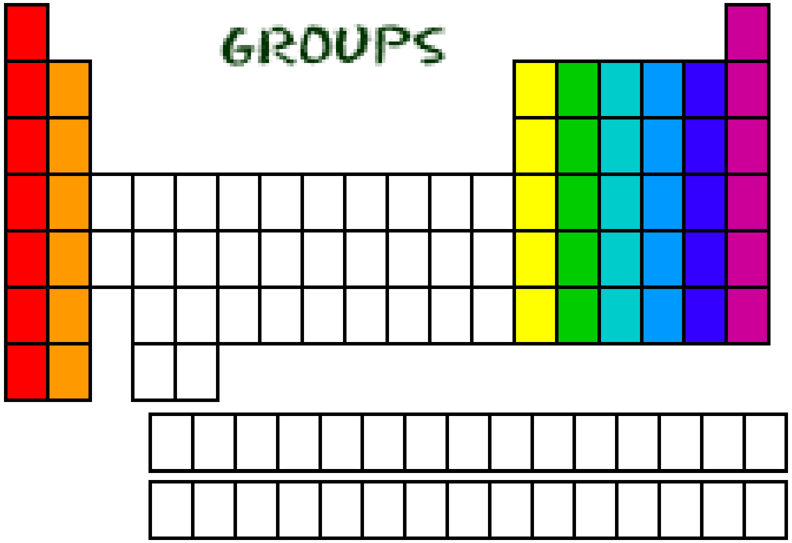
- quoted from http://www.chem4kids.com/files/elem_pertable.html

Periods = Rows

- In the periodic table, elements have something in common if they are in the same row.
- All of the elements in a period have the same number of atomic orbitals.
- Every element in the top row (the first period) has one orbital for its electrons. All of the elements in the second row (the second period) have two orbitals for their electrons. It goes down the periodic table like that.

- quoted from http://www.chem4kids.com/files/elem_pertable.html

And you got your groups…



- The periodic table has a special name for its columns, too. When a column goes from top to bottom, it's called a **group**.

- quoted from http://www.chem4kids.com/files/elem_pertable.html

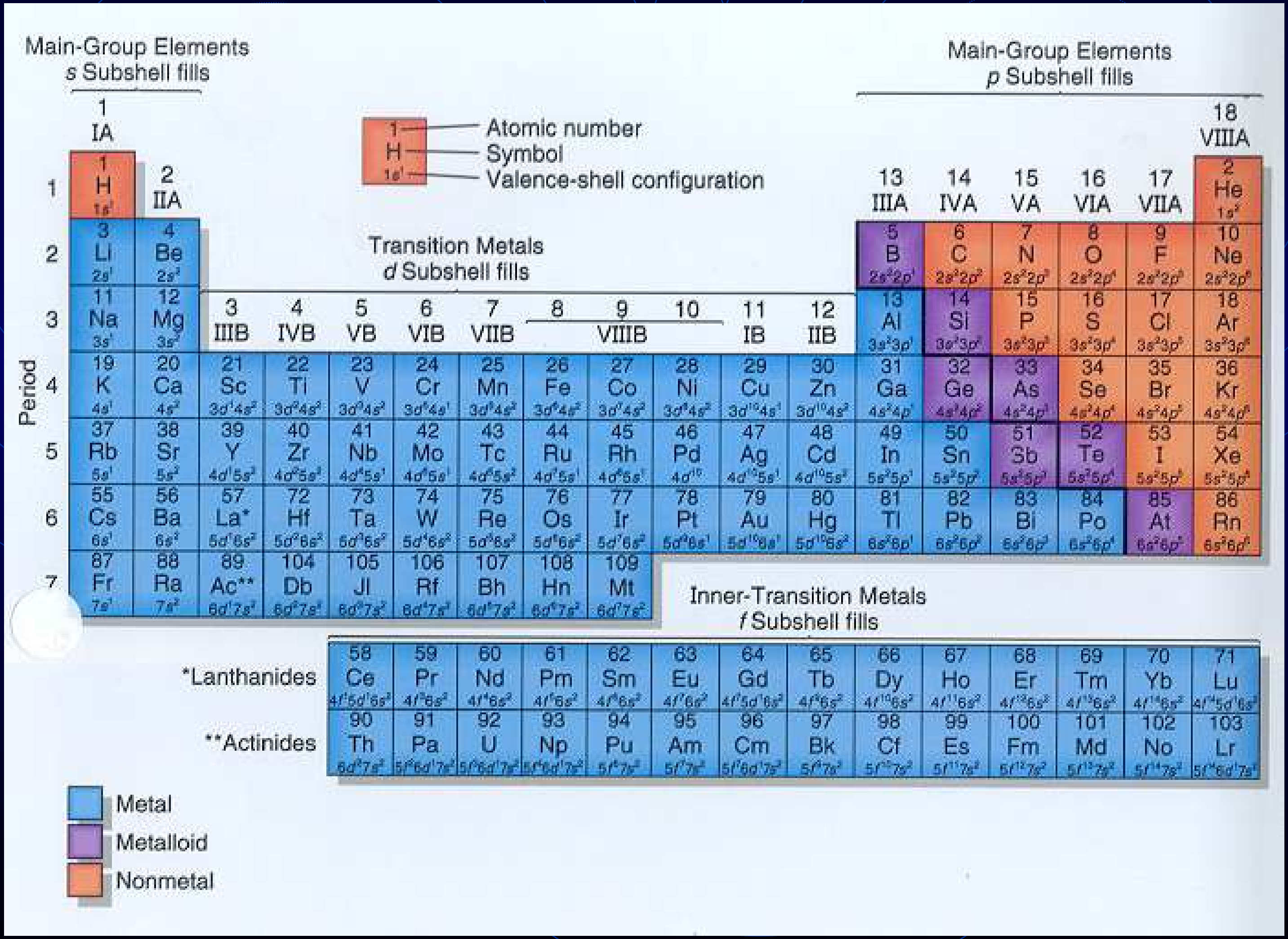
Groups = Columns

- The elements in a group have the same number of electrons in their outer orbital.
- Every element in the first column (group one) has one electron in its outer shell. Every element on the second column (group two) has two electrons in the outer shell. As you keep counting the columns, you'll know how many electrons are in the outer shell.
- There are some exceptions to the order when you look at the transition elements, but you get the general idea.

A Different Type of Grouping

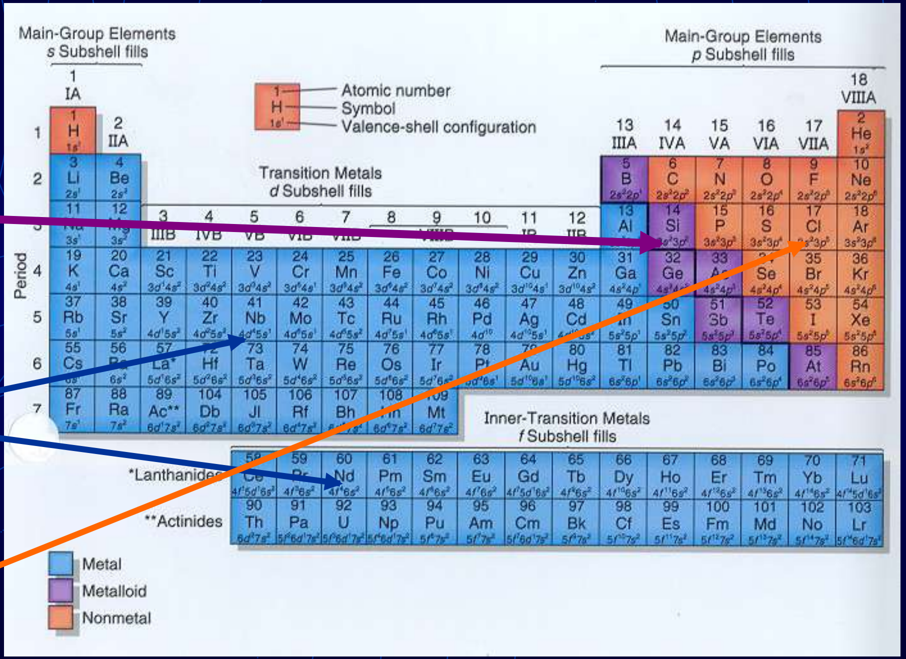
- Besides the 4 blocks of the table, there is another way of classifying element:
- Metals
- Nonmetals
- Metalloids or Semi-metals.
- The following slide shows where each group is found.

Metals, Nonmetals, Metalloids



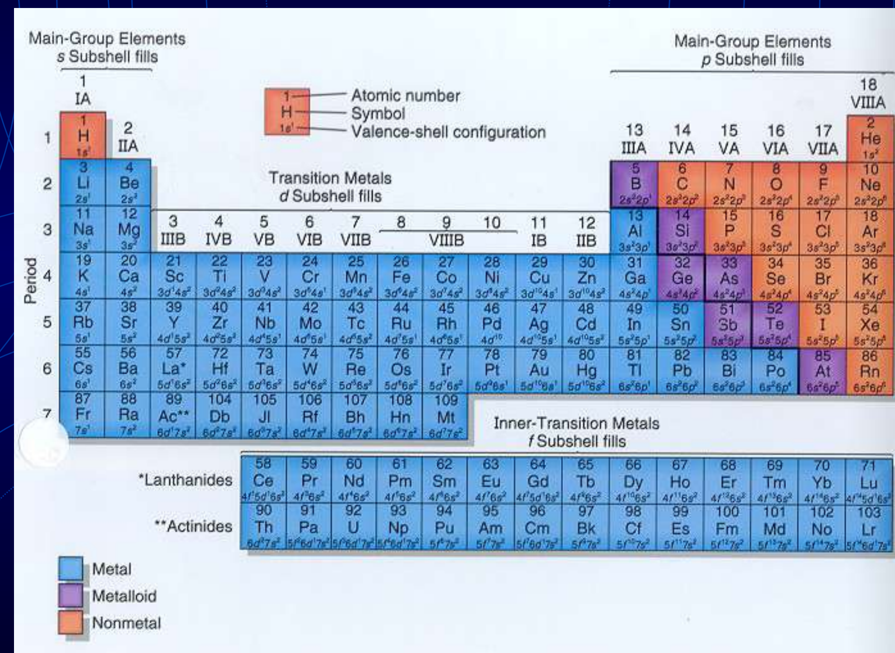
Metals, Nonmetals, Metalloids

- There is a zig-zag or staircase line that divides the table.
- Metals are on the left of the line, in blue.
- Nonmetals are on the right of the line, in orange.



Metals, Nonmetals, Metalloids

- Elements that border the stair case, shown in purple are the metalloids or semi-metals.
- There is one important exception.
- Aluminum is more metallic than not.



Metals

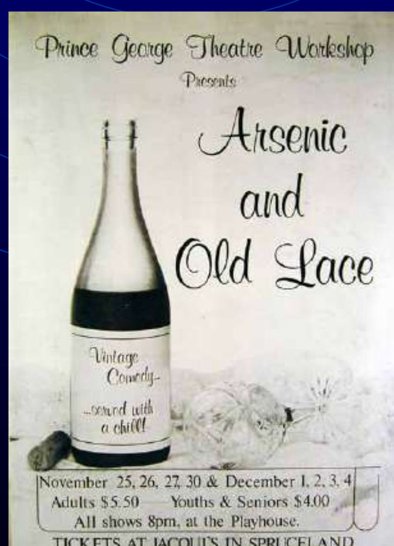
- Metals are lustrous (shiny), malleable, ductile, and are good conductors of heat and electricity.
- They are mostly solids at room temp.
- What is one exception?

Nonmetals

- Nonmetals are the opposite.
- They are dull, brittle, nonconductors (insulators).
- Some are solid, but many are gases, and Bromine is a liquid.

Metalloids

- Metalloids, aka semi-metals are just that.
- They have characteristics of both metals and nonmetals.
- They are shiny but brittle.
- And they are semiconductors.
- What is our most important semiconductor?



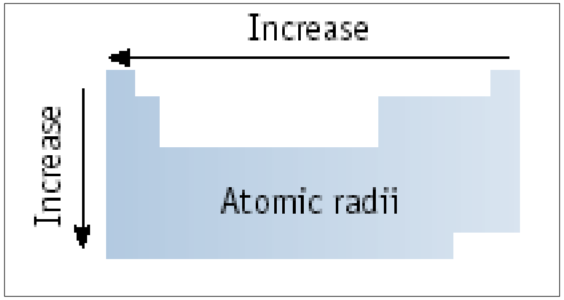
Periodic Trends

- There are several important atomic characteristics that show predictable trends that you should know.
- The first and most important is atomic radius.
- Radius is the distance from the center of the nucleus to the "edge" of the electron cloud.

Atomic Radius

- Since a cloud's edge is difficult to define, scientists use define covalent radius, or half the distance between the nuclei of 2 bonded atoms.
- Atomic radii are usually measured in picometers (pm) or angstroms (Å). An angstrom is 1×10^{-10} m.

Atomic Size



- **Size goes UP** on going down a group.
- Because electrons are added farther from the nucleus, there is less attraction.
- **Size goes DOWN** on going across a period.

Atomic Radii

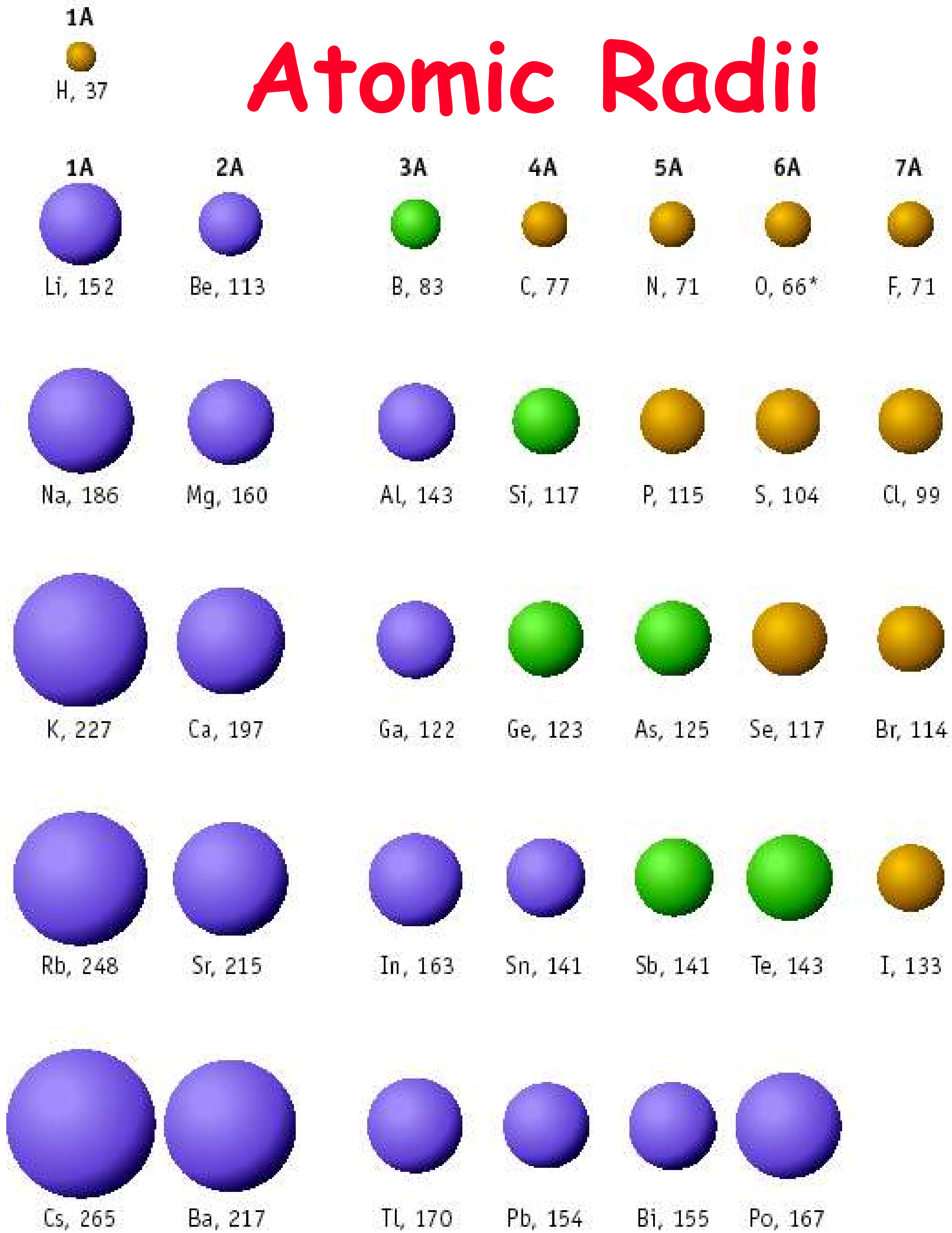
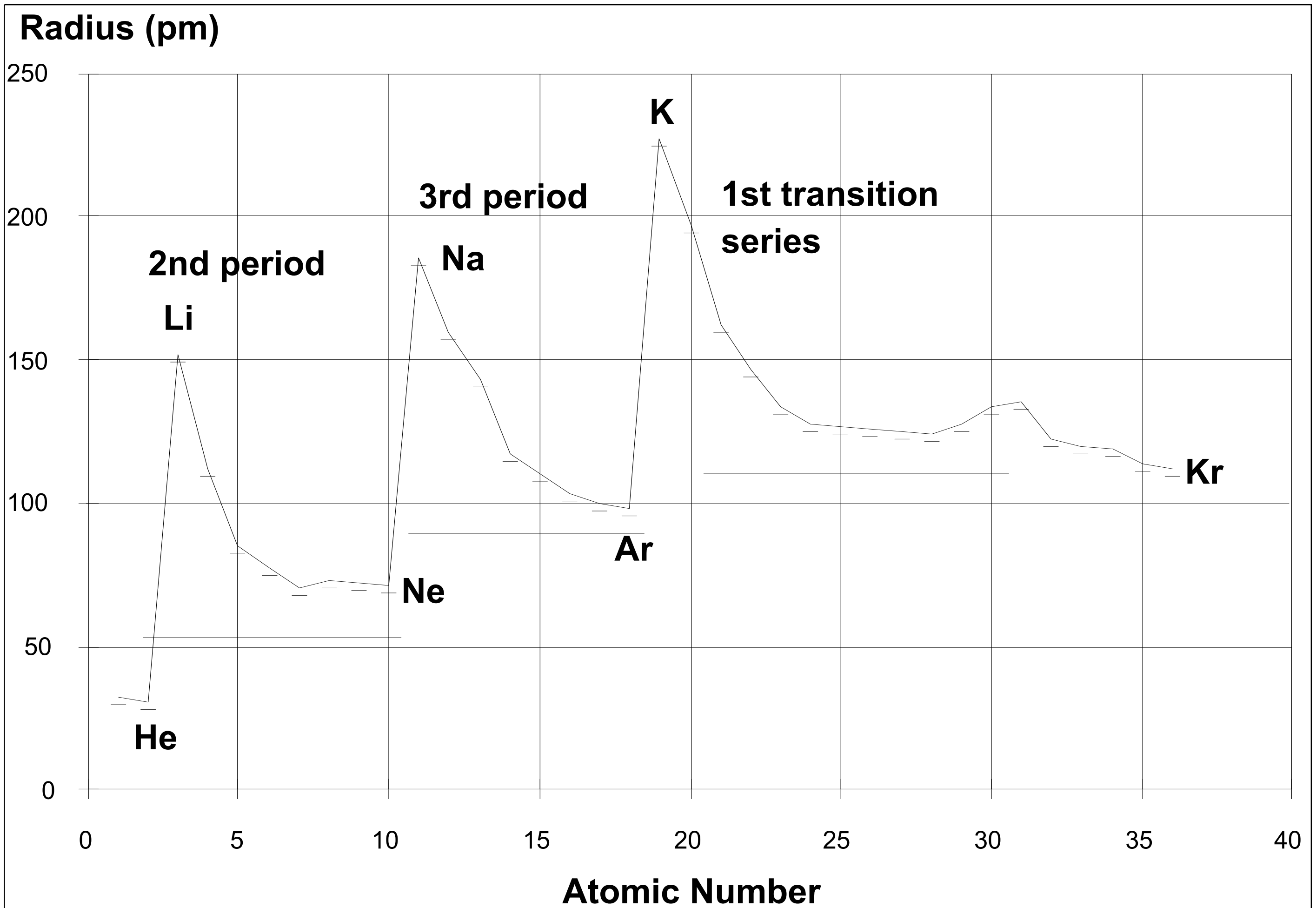


Figure 8.9

Trends in Atomic Size

See Figures 8.9 & 8.10



Atomic Radius

- The trend for atomic radius in a vertical column is to go from smaller at the top to larger at the bottom of the family.
- Why?
- With each step down the family, we add an entirely new EL to the electron cloud, making the atoms larger with each step.

Atomic Radius

- The trend across a horizontal period is less obvious.
- What happens to atomic structure as we step from left to right?
- Each step adds a proton and an electron (and 1 or 2 neutrons).
- Electrons are added to existing ELs or sublevels.

Atomic Radius

- The effect is that the more positive nucleus has a greater pull on the electron cloud.
- The nucleus is more positive and the electron cloud is more negative.
- The **increased attraction pulls the cloud in**, making atoms smaller as we move from left to right across a period.

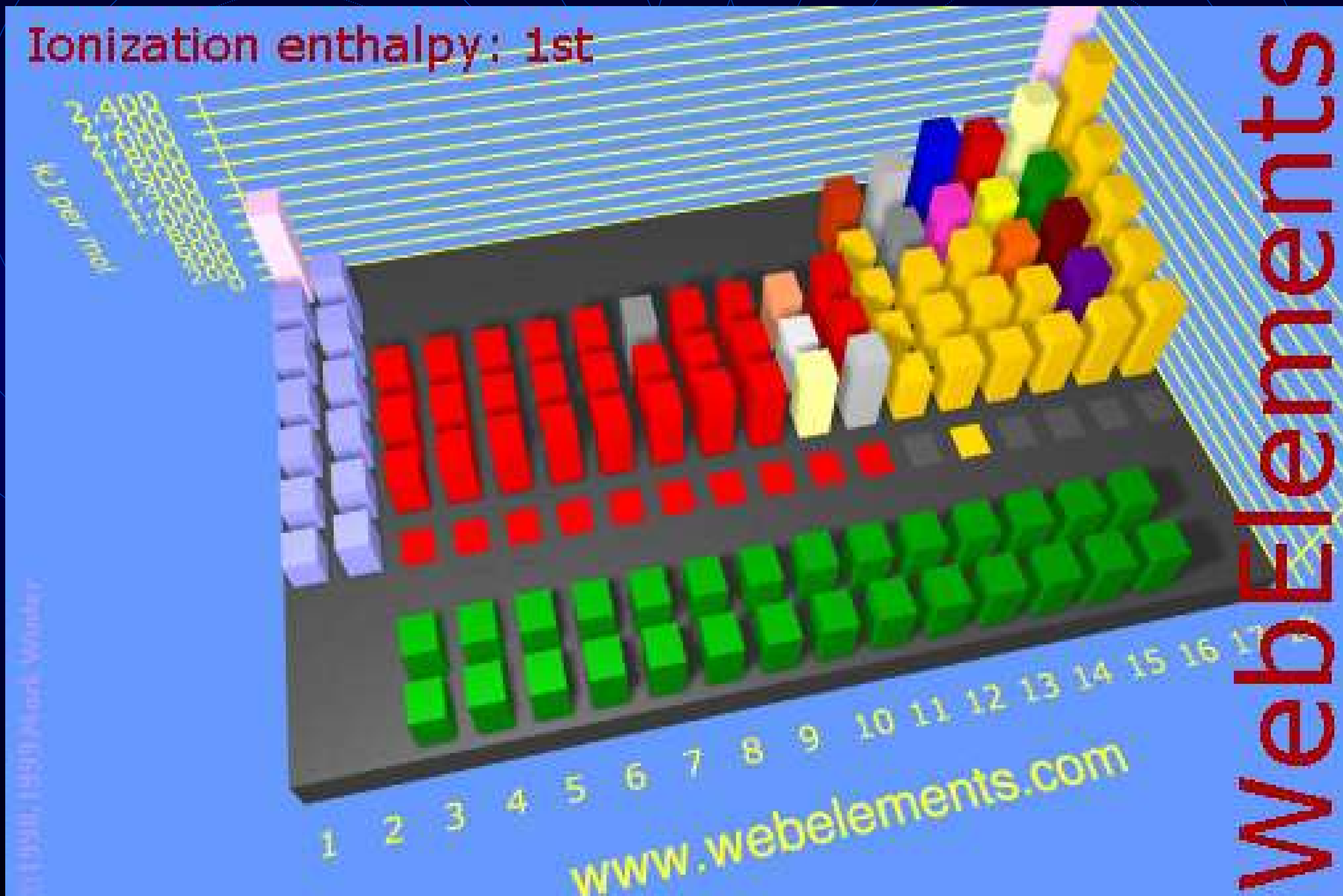
Ionization Energy

- This is the second important periodic trend.
- If an electron is given enough energy (in the form of a photon) to overcome the effective nuclear charge holding the electron in the cloud, it can leave the atom completely.
- The atom has been "ionized" or charged.
- The number of protons and electrons is no longer equal.

Ionization Energy

- The energy required to remove an electron from an atom is ionization energy. (measured in kilojoules, kJ)
- The larger the atom is, the easier its electrons are to remove.
- **Ionization energy and atomic radius are inversely proportional.**
- Ionization energy is always endothermic, that is energy is added to the atom to remove the electron.

Ionization Energy



Electron Affinity

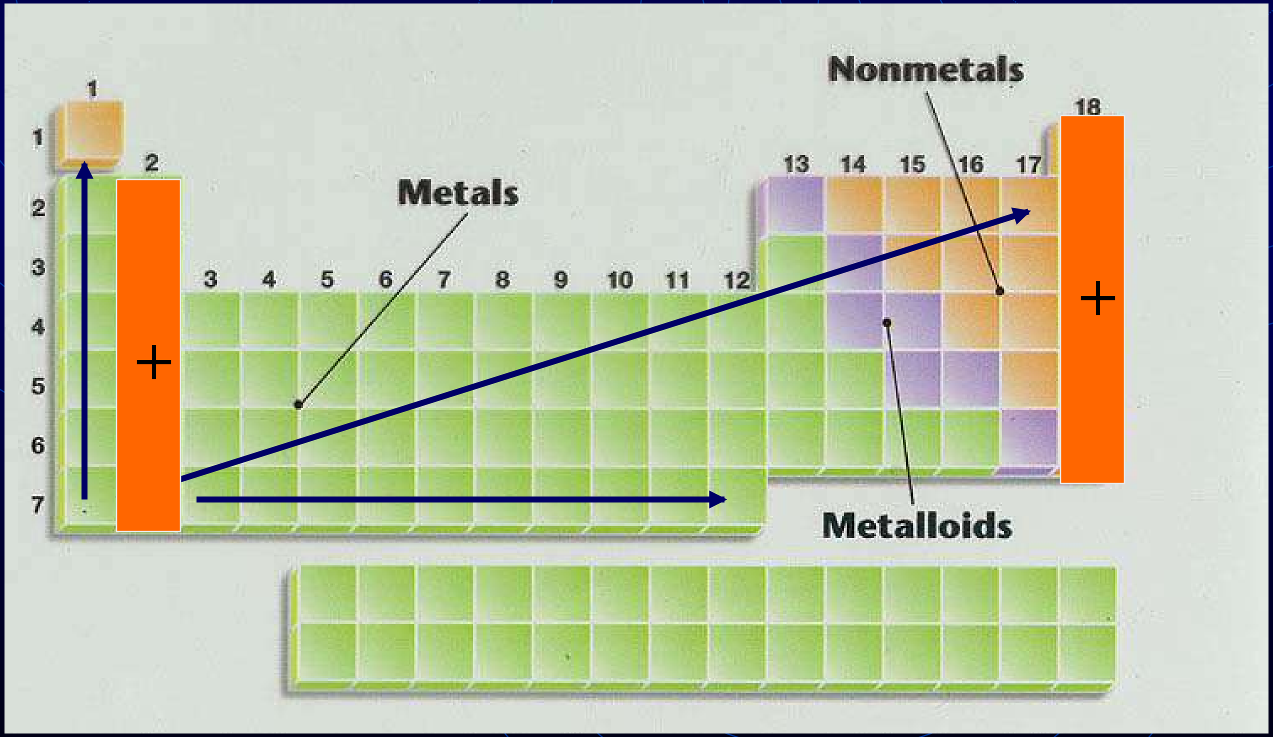
- What does the word ‘affinity’ mean?
- Electron affinity is the energy **change** that occurs when an atom **gains an electron** (also measured in kJ).
- Where ionization energy is always endothermic, electron affinity is usually exothermic, but not always.

Electron Affinity

- Electron affinity is exothermic if there is an empty or partially empty orbital for an electron to occupy.
- If there are no empty spaces, a new orbital or PEL must be created, making the process endothermic.
- This is true for the alkaline earth metals and the noble gases.

Electron Affinity

- Your help sheet should look like this:

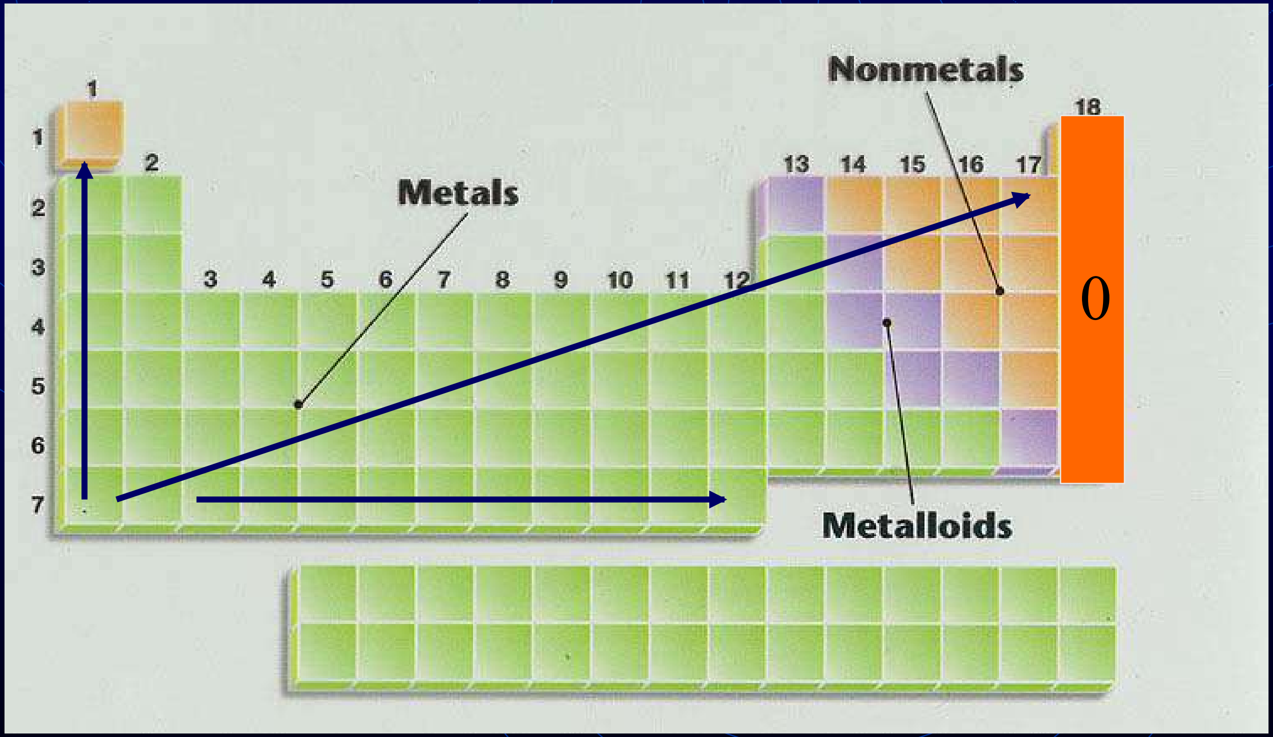


Electronegativity

- Electronegativity is a measure of an atom's attraction for another atom's electrons.
- It is an arbitrary scale that ranges from 0 to 4.
- The units of electronegativity are Paulings.
- Generally, metals are electron givers and have low electronegativities.
- Nonmetals are are electron takers and have high electronegativities.
- What about the noble gases?

Electronegativity

- Your help sheet should look like this:



Overall Reactivity

- This ties all the previous trends together in one package.
- However, we must treat metals and nonmetals separately.
- The most reactive metals are the largest since they are the best electron givers.
- The most reactive nonmetals are the smallest ones, the best electron takers.

The Octet Rule

- The “goal” of most atoms (except H, Li and Be) is to have an octet or group of 8 electrons in their valence energy level.
- They may accomplish this by either giving electrons away or taking them.
- Metals generally give electrons, nonmetals take them from other atoms.
- Atoms that have gained or lost electrons are called ions.

Ions

- When an atom gains an electron, it becomes negatively charged (more electrons than protons) and is called an **anion.**
- In the same way that nonmetal atoms can gain electrons, metal atoms can lose electrons.
- They become positively charged **cations**.

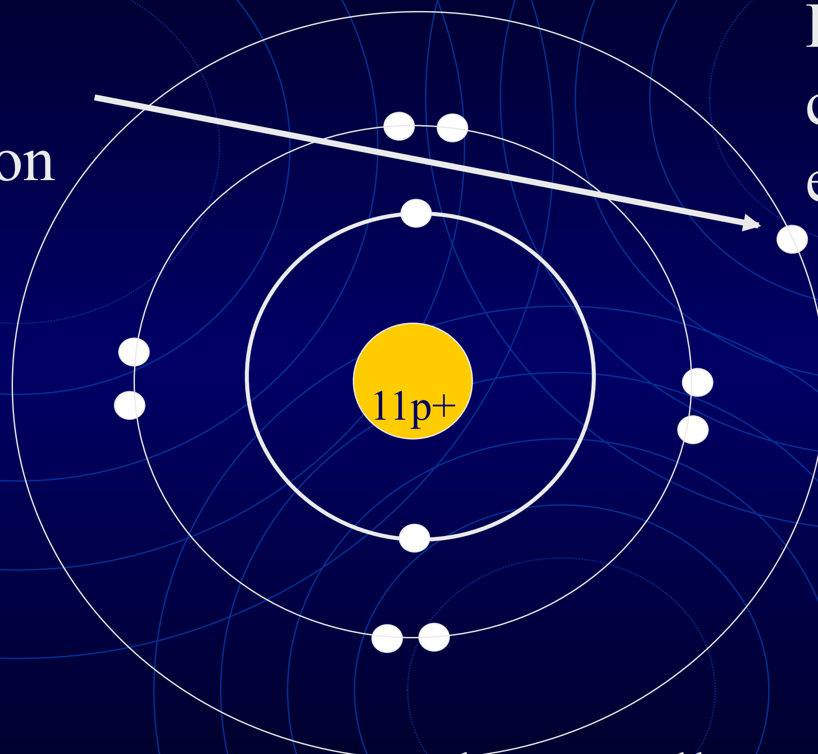
Ionic Radius

- Cations are always smaller than the original atom.
- The entire outer PEL is removed during ionization.
- Conversely, anions are always larger than the original atom.
- Electrons are added to the outer PEL.

Cation Formation

Na atom

1 valence electron



Effective nuclear charge on remaining electrons increases.

Remaining e- are pulled in closer to the nucleus. Ionic size decreases.

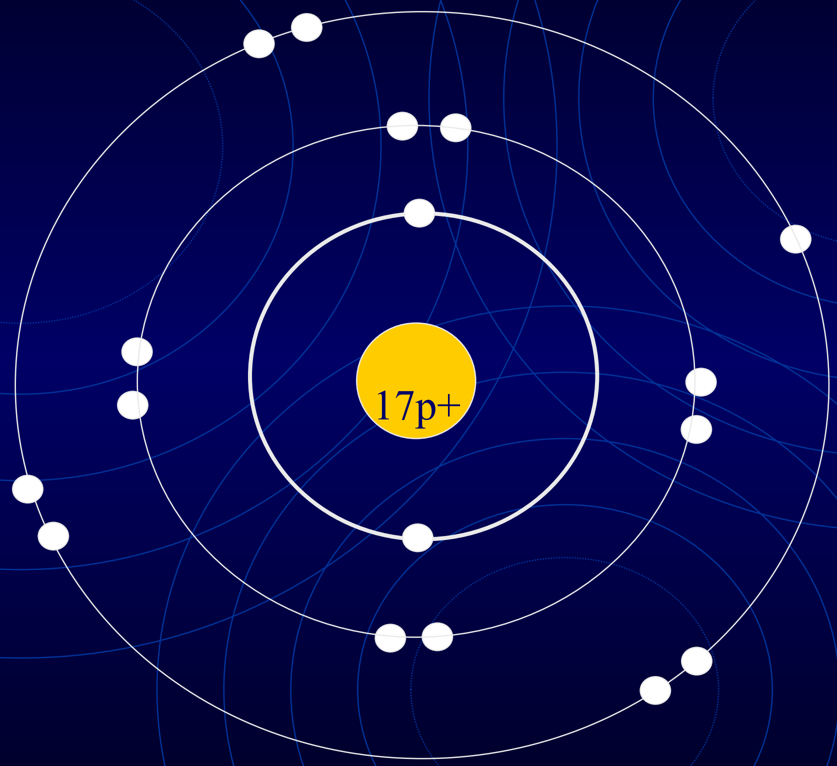
Valence e- lost in ion formation

Result: a smaller sodium cation, Na^+

Anion Formation

A chloride ion is produced. It is larger than the original atom.

Chlorine atom with 7 valence e-

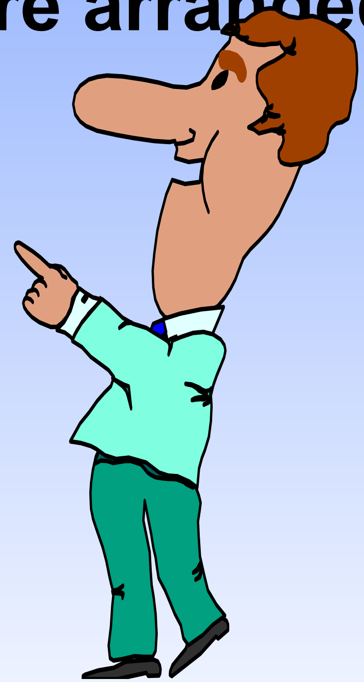


One e- is added to the outer shell.

Effective nuclear charge is reduced and the e- cloud expands

Arrangement of Electrons in Atoms

Electrons in atoms are arranged as



SHELLS (n)

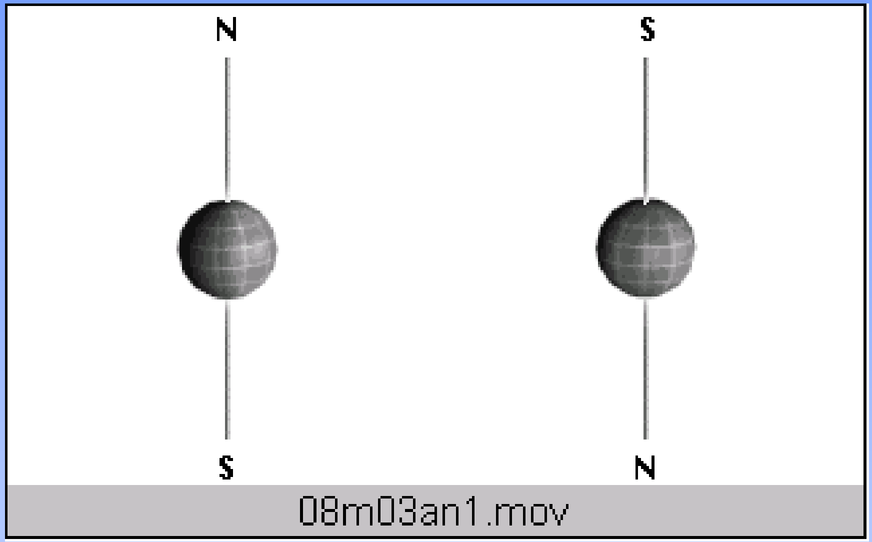
SUBSHELLS (l)

ORBITALS (m_l)

Arrangement of Electrons in Atoms

Each orbital can be assigned no more than 2 electrons!

This is tied to the existence of a 4th quantum number, the electron spin quantum number, m_s.



Electron Spin Quantum Number, m_s

Can be proved experimentally that electron has a spin. Two spin directions are given by m_s where m_s = +1/2 and -1/2.

QUANTUM NUMBERS

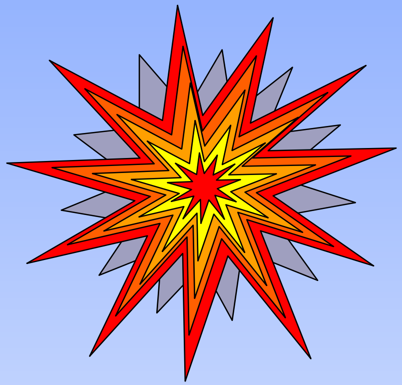
n ---> shell 1, 2, 3, 4, ...

l ---> subshell 0, 1, 2, ... n - 1

m_l ---> orbital -l ... 0 ... +l

m_s ---> electron spin +1/2 and -1/2

Pauli Exclusion Principle



- No two electrons in the same atom can have the same set of 4 quantum numbers.
- That is, each electron in an atom has a unique address of quantum numbers.

Electrons in Atoms

When n = 1, then l = 0

this shell has a single orbital (1s) to which 2e- can be assigned.

When n = 2, then l = 0, 1

2s orbital 2e-

three 2p orbitals6e-

TOTAL = 8e-

Electrons in Atoms

When n = 3, then l = 0, 1, 2

3s orbital 2e-

three 3p orbitals6e-

five 3d orbitals10e-

TOTAL = 18e-

Electrons in Atoms

When n = 4, then l = 0, 1, 2, 3

4s orbital 2e-

three 4p orbitals 6e-

five 4d orbitals 10e-

seven 4f orbitals 14e-

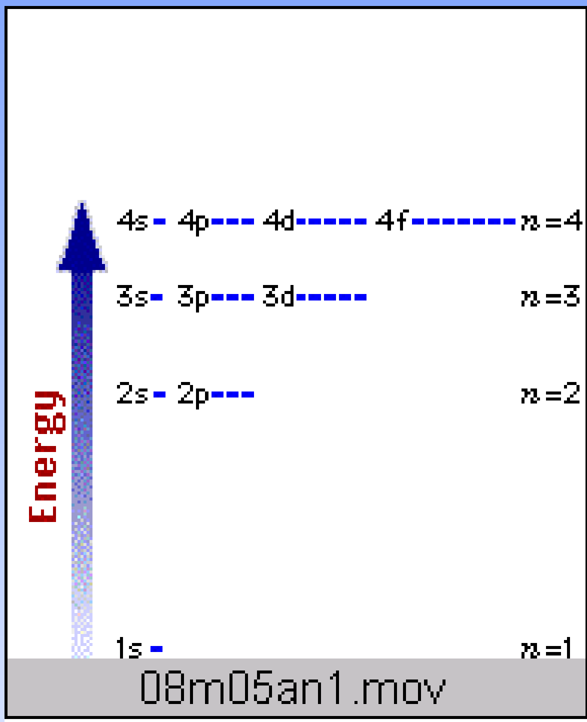
TOTAL = 32e-

Table 8.1 • Number of Electrons Accommodated in Electron Shells and Subshells With $n = 1$ to 6

Electron Shell (n)	Subshells Available	Orbitals Available ($2\ell + 1$)	Number of Electrons Possible in Subshell [$2(2\ell + 1)$]	Maximum Electrons Possible for nth Shell ($2n^2$)
1	*s*	1	2	2
2	*s*	1	2	8
	p	3	6	
3	*s*	1	2	18
	p	3	6	
	d	5	10	
4	*s*	1	2	32
	p	3	6	
	d	5	10	
	f	7	14	
5	*s*	1	2	50
	p	3	6	
	d	5	10	
	f	7	14	
	*g**	9	18	
6	*s*	1	2	72
	p	3	6	
	d	5	10	
	*f**	7	14	
	*g**	9	18	
	*h**	11	22	

*These orbitals are not used in the ground state of any known element.

Assigning Electrons to Subshells



- In H atom all subshells of same n have same energy.
- In many-electron atom:

a) subshells increase in energy as value of (n + l) increases.

b) for subshells of same (n + l), the subshell with lower n is lower in energy.

Electron Filling Order

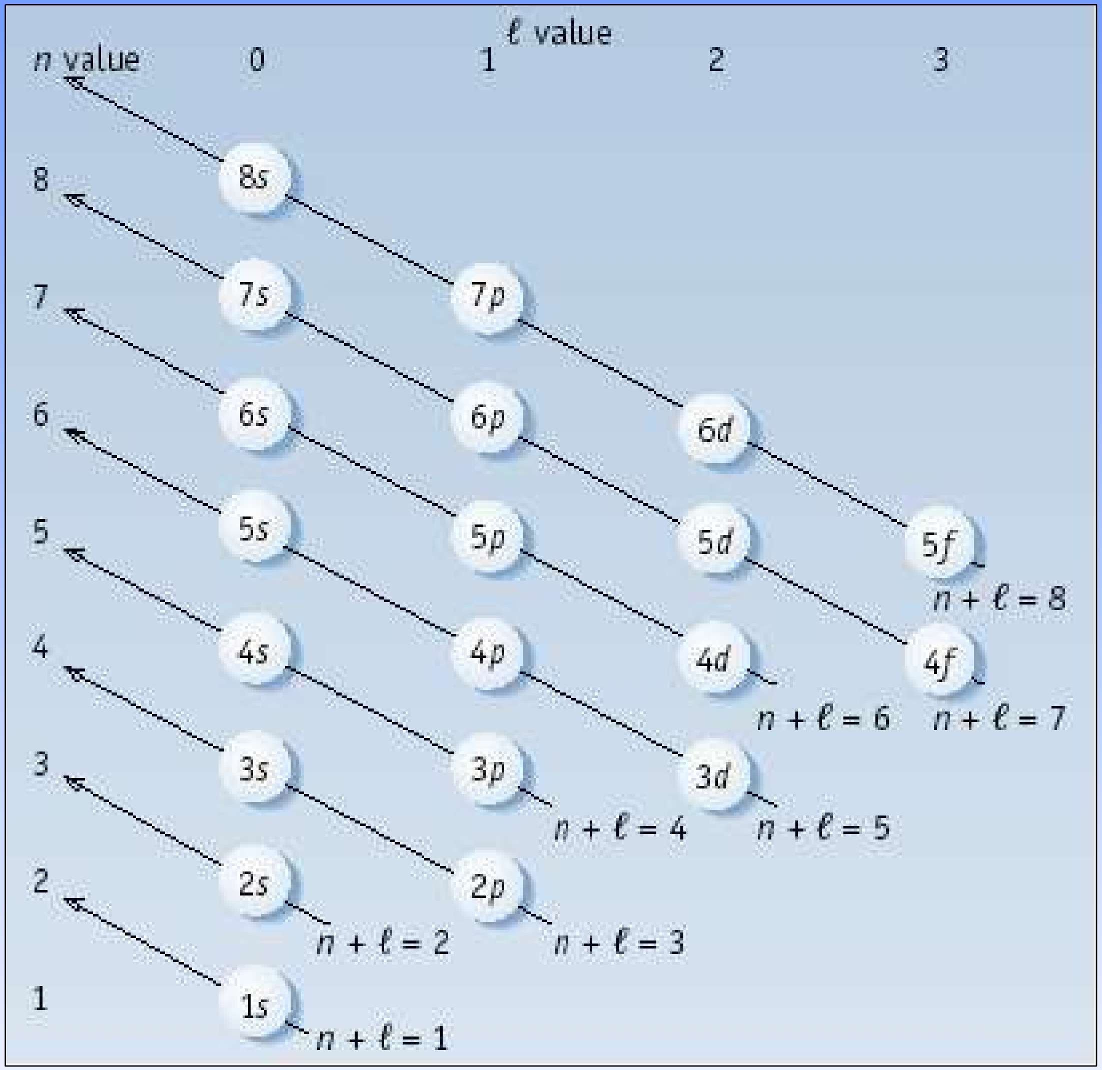


Figure 8.5

Writing Atomic Electron Configurations

Two ways of writing configs. One is called the **spdf notation.**

spdf notation

for H, atomic number = 1

$$1s^1$$

no. of electrons

value of l

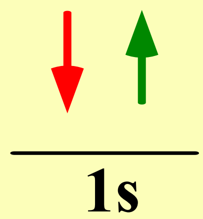
value of n

Writing Atomic Electron Configurations

Two ways of writing configs. Other is called the **orbital box notation.**

ORBITAL BOX NOTATION for He, atomic number = 2

$1s^2$

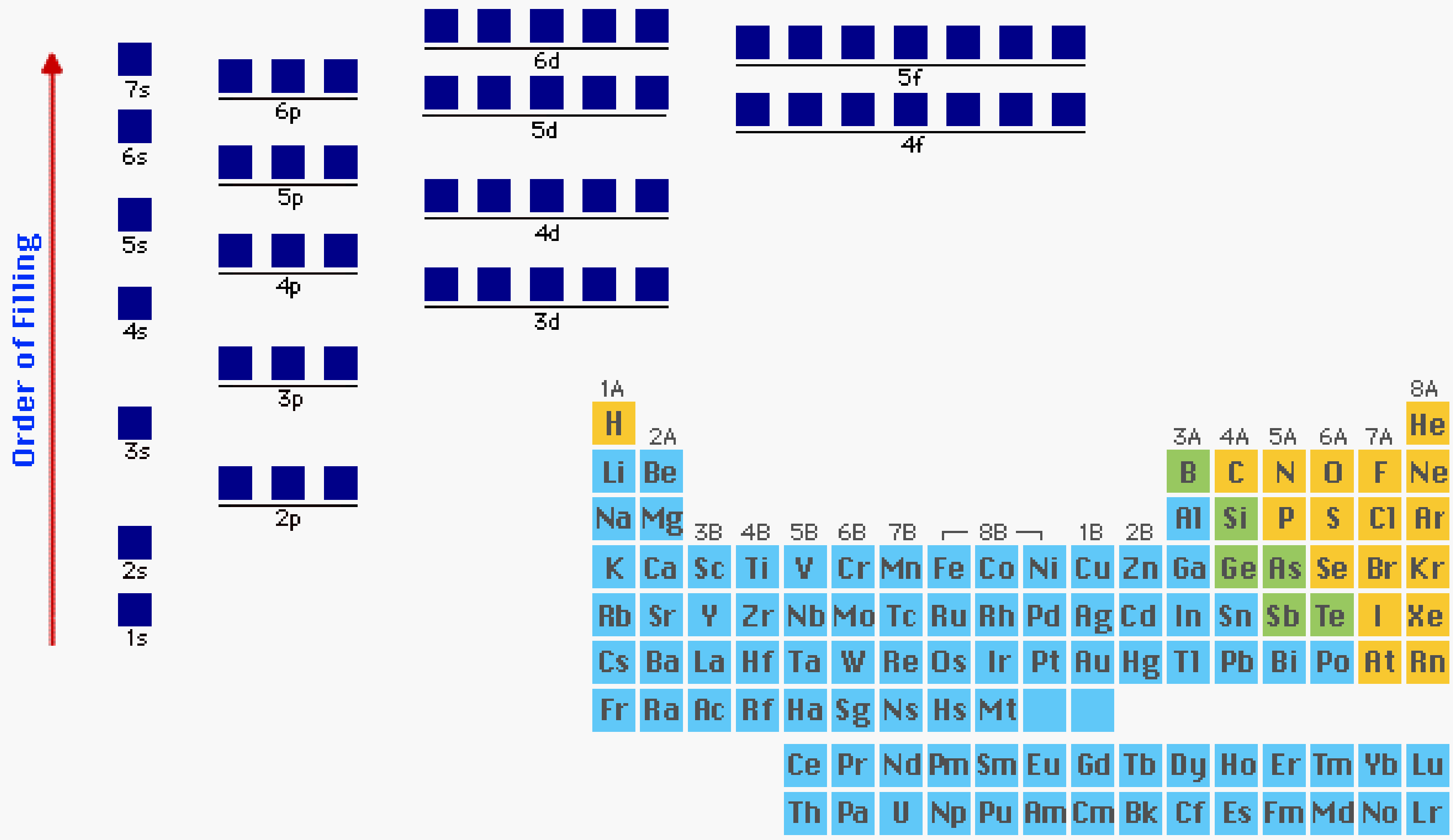


Arrows depict electron spin

One electron has $n = 1$, $l = 0$, $m_l = 0$, $m_s = + 1/2$

Other electron has $n = 1$, $l = 0$, $m_l = 0$, $m_s = - 1/2$

Atomic Electron Configurations



See “Toolbox” for Electron Configuration tool.