Unit 8

On completion of the unit you should be able to:

1. state the mass, charge and location of protons, electrons and neutrons.
2. given the atomic number of an element, draw diagrams showing the protons in the nucleus and the electrons in shells (or energy levels) around the nucleus.
3. write the electron dot formulae for the elements in groups 1 to 8 in the periodic table.
4. define isotopes, mass number and atomic mass.
5. given the relative abundance and isotopic masses for an element, calculate the atomic mass of that element.
6. given the atomic number and mass number of an element, calculate the number of electrons.

ATOMS AND THE PERIODIC TABLE

8.1 **Atom**

- **Subatomic particles**

8.2 **Isotope**

Reading: Hebden – page 144-146

8.3 **Arrangement of electrons**

- **Bohr model of the atom**
  Reading: Hebden – page 142
- **Lewis electron-dot formulas of elements**
  Reading: Hebden – page 167
- **Quantum mechanical model of the atom**

8.4 **Electron configurations of atoms**

8.5 **Atomic mass**

- **Tabulated atomic mass**

Reading: Hebden – page 150

8.6 **Periodic table of the elements**

- **Metals Nonmetals and Semimetals**
  - **Properties of metals**
  - **Properties of nonmetals**
  - **Properties of semimetals**

Today's focus.

Quantum Mechanical model and writing electron configuration of the atom.

Quantum Mechanical Model

The **Quantum Mechanical Model** of the atom presents a more accurate model of the atom. It is a more sophisticated model based on complex mathematical calculations and interpretations. We will take a look at this model and summarize the results based on these mathematical calculations without carrying them out ourselves.

The Quantum Mechanical Model introduces the concept of

- sublevels *(s, p, d, f)*
- atomic orbitals

The following table summarizes the number of orbitals in each sublevel.

Type of Atomic Orbital	Number of orbitals
s	There is 1 *s*-type orbital.
p	There are 3 *p*-type orbitals.
d	There are 5 *d*-type orbitals.
f	There are 7 *f*-type orbitals.*

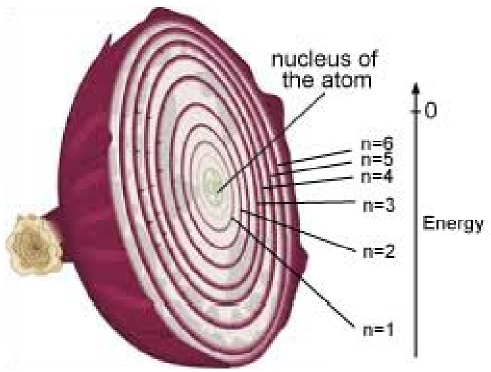
Within a principle energy level, this is the order they appear

Atomic Orbitals

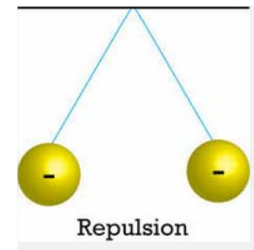
These *atomic orbitals* in the sublevels represent the probability of where in the space surrounding the nucleus the electron in each type of atomic sublevel could be found.

Let's look at these how electrons occupy these atomic orbitals!

Rules for Writing *Electron Configurations*:



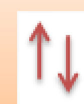
For each element, we assign electrons to the orbitals to give the most *stable* state of the element (i.e. electrons occupy the lowest energy possible)



Based on the "Aufbau" Principle (German for "building-up"):

1. Determine the number of electrons for the element. For a nucleus of +*z* charge, it must be surrounded by *z* electrons, since elements are neutral charge.

2. There may not be more than 2 electrons in each atomic orbital. In addition, the 2 electrons must have *opposite* spins. This is the Pauli Exclusion Principle.



Spin up; Spin down

3. When electrons fill orbitals with the same energy, the order of filling is such that as many electrons remain unpaired as possible with their spins parallel. This is Hund's Rule.

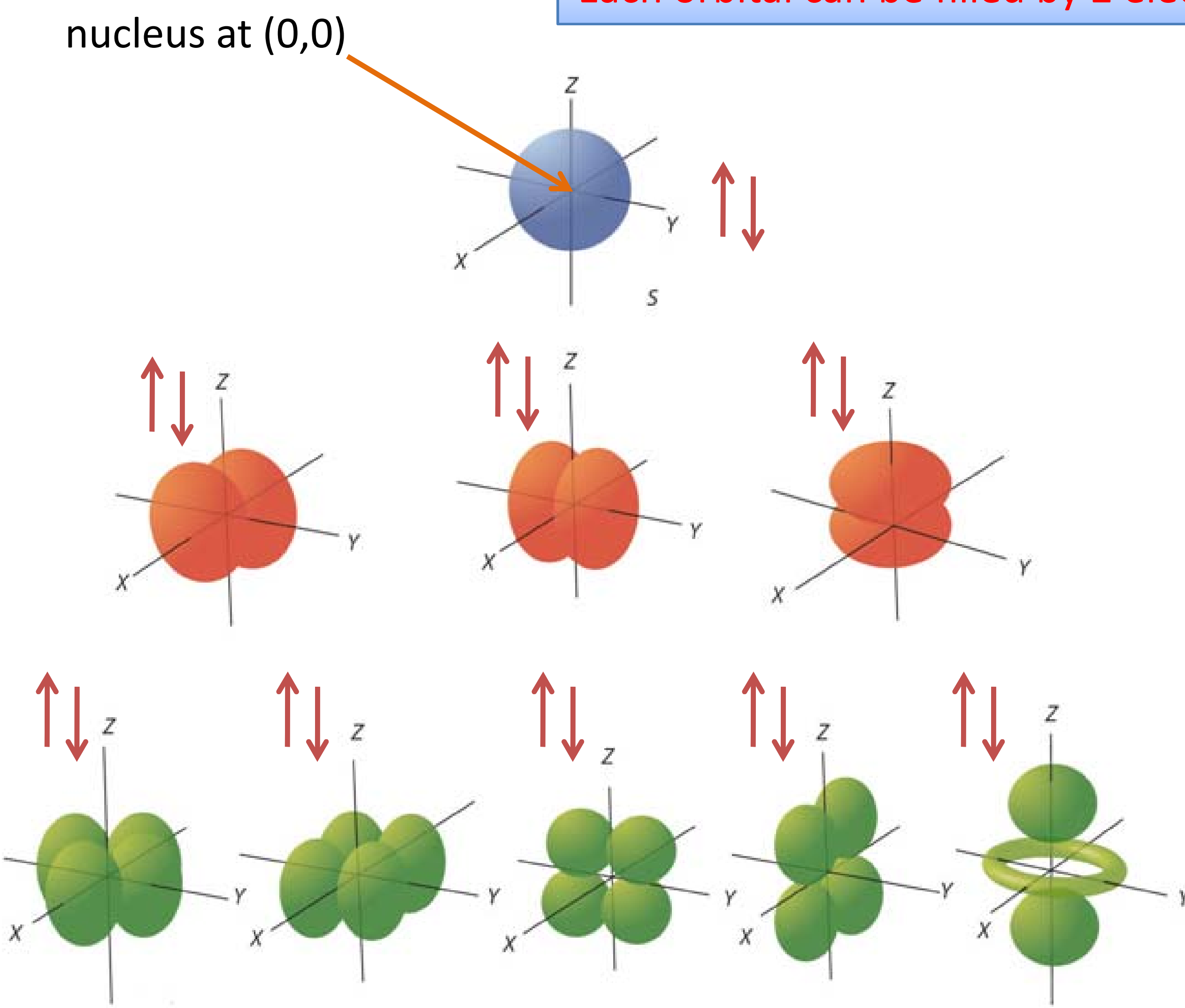
4. When electrons fill the orbitals, they fill them in the order from the lowest energy to the higher energy orbitals. This electron configuration gives the lowest energy to the atom (the most *stable* state), and it is called the ground state of the atom.

Quantum Mechanical Model

These *atomic orbitals* in the sublevels represent the probability of where in the space surrounding the nucleus the electron in each type of atomic sublevel could be found.

Each orbital can be filled by 2 electrons.

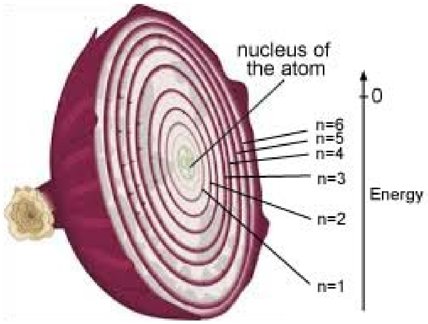
↑↓ Spin up; Spin down



Atomic Orbital	Maximum # of electrons
s-orbital	2
p-orbitals	6
d-orbitals	10
f-orbitals *Google the shapes	14 (there are 7 *f*-orbitals)

Image credit: http://2012books.lardbucket.org/books/beginning-chemistry/s12-02-quantum-numbers-for-electrons.html

Bohr model / Quantum Mechanics Model of the atom



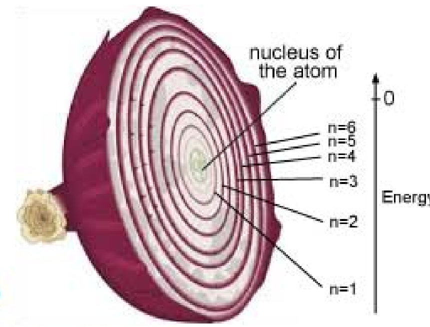
Maximum number of electrons in an energy level $= 2n^2$

Quantum Mechanics Model tells us more …

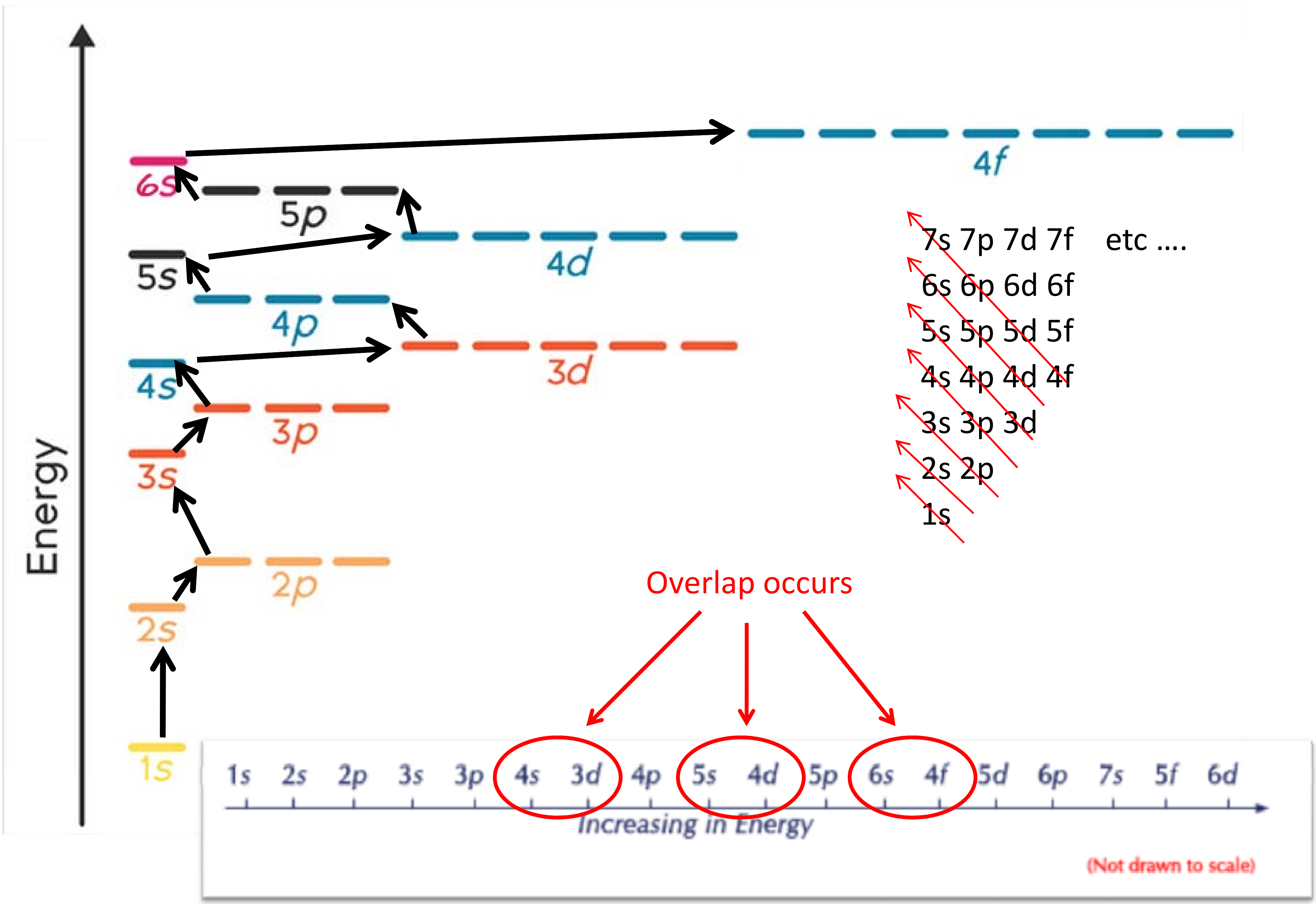
Energy Level	*# of Sublevels*	*Types of atomic orbital*	*When the orbital is fully occupied, # of electrons is*	$2n^2$	Max # of electrons
n=1	1	*s*	2	$2(1)^2 =$	2 electrons
n=2	2	*s* *p*	*2* *6*	$2(2)^2 =$	8 electrons
n=3	3	*s* *p* *d*	*2* *6* *10*	$2(3)^2 =$	18 electrons
n=4	4	*s* *p* *d* *f*	*2* *6* *10* *14*	$2(4)^2 =$	32 electrons

This is the structure of the atom. Let's fill this structure with electrons!

The Energy Level Diagram for an Atom

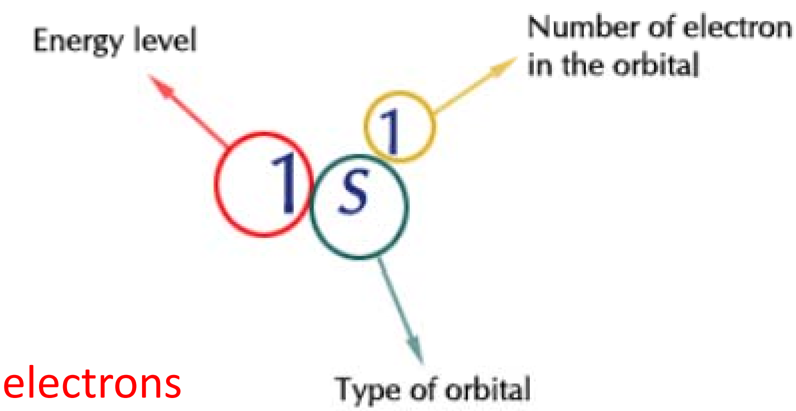


Electrons fill orbitals from low to high energy

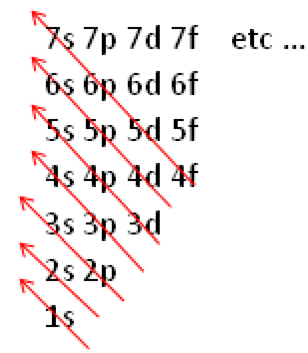


Write the Ground State Electron Configuration for the Elements

Hydrogen: 1 electron



Valence electrons are bolded.



Period 1 elements

Period 1: n=1
1*s*-orbital completely filled

Helium: 2 electrons → $\mathbf{1s^2}$

Period 2 elements

Lithium: 3 electrons → $1s^2$ $\mathbf{2s^1}$

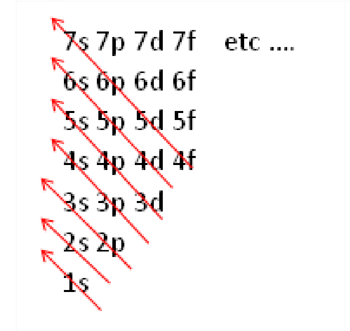
Beryllium: 4 electrons → $1s^2$ $\mathbf{2s^2}$

2*s*-orbital completely filled

Boron: 5 electrons → $1s^2$ $\mathbf{2s^2\ 2p^1}$

Carbon: 6 electrons → $1s^2$ $\mathbf{2s^2\ 2p^2}$

Write the Ground State Electron Configuration for the Elements



Nitrogen: 7 electrons → $1s^2$ **$2s^2$ $2p^3$**

Oxygen: 8 electrons → $1s^2$ **$2s^2$ $2p^4$**

Fluorine: 9 electrons → $1s^2$ **$2s^2$ $2p^5$**

Neon: 10 electrons → $1s^2$ **$2s^2$ $2p^6$**

Period 2: n=2
***2p*-orbitals completely filled**

Sodium: 11 electrons → $1s^2$ $2s^2$ $2p^6$ **$3s^1$**

Period 3 elements

Magnesium: 12 electrons → $1s^2$ $2s^2$ $2p^6$ **$3s^2$**

3*s*-orbital completely filled

Aluminum: 13 electrons → $1s^2$ $2s^2$ $2p^6$ **$3s^2$ $3p^1$**

Silicon: 14 electrons → $1s^2$ $2s^2$ $2p^6$ **$3s^2$ $3p^2$**

Phosphorus: 15 electrons → $1s^2$ $2s^2$ $2p^6$ **$3s^2$ $3p^3$**

Sulfur: 16 electrons → $1s^2$ $2s^2$ $2p^6$ **$3s^2$ $3p^4$**

Chlorine: 17 electrons → $1s^2$ $2s^2$ $2p^6$ **$3s^2$ $3p^5$**

Argon: 18 electrons → $1s^2$ $2s^2$ $2p^6$ **$3s^2$ $3p^6$**

Period 3: n=3
3*p*-orbitals completely filled

H: $1s^1$

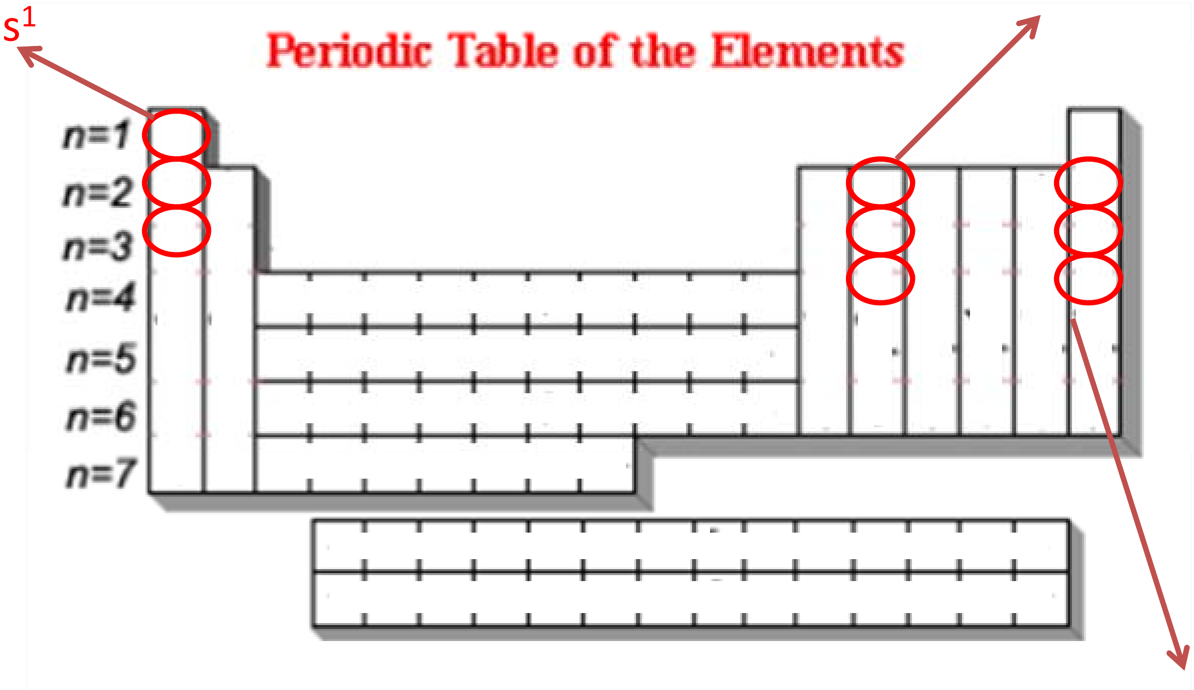
Li: $1s^2 2s^1$

Na: $1s^2 2s^2 2p^6 3s^1$

C: $1s^2 2s^2 2p^2$

Si: $1s^2 2s^2 2p^6 3s^2 3p^2$

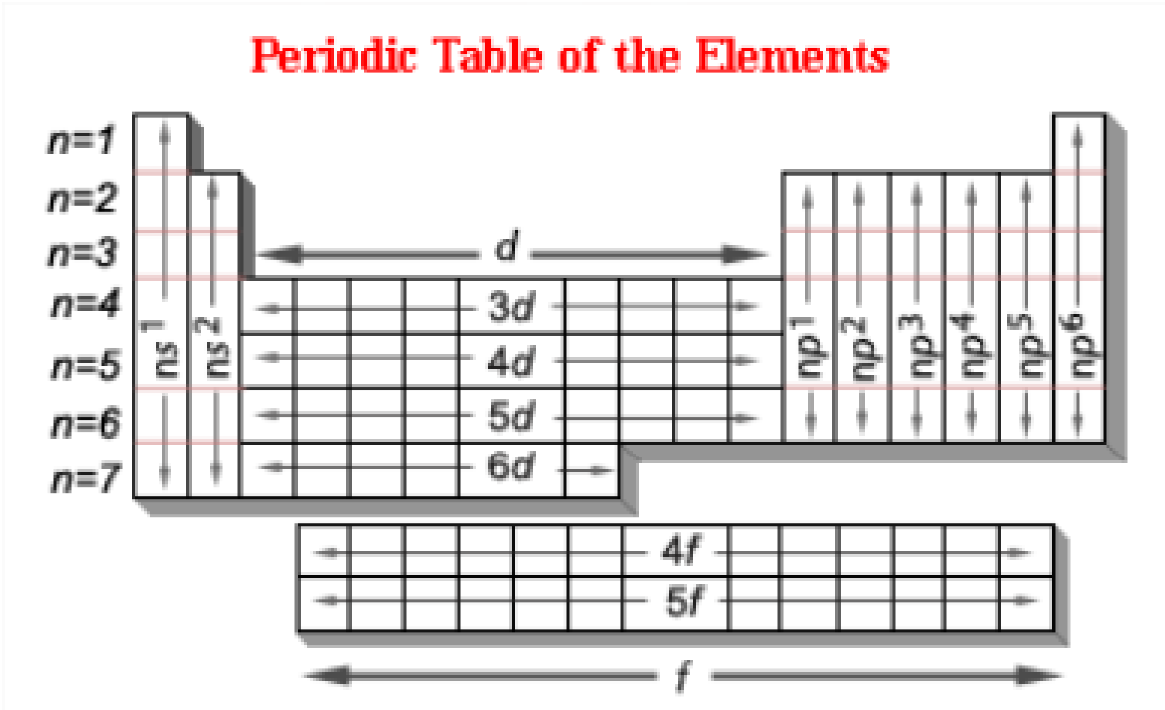
Ge: $1s^2 2s^2 2p^6$.......$4p^2$



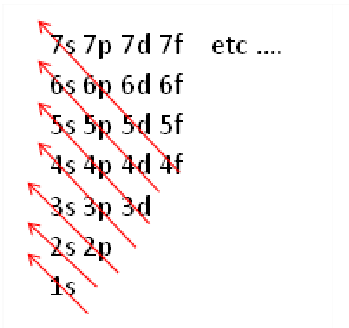
Ne: $1s^2 2s^2 2p^6$

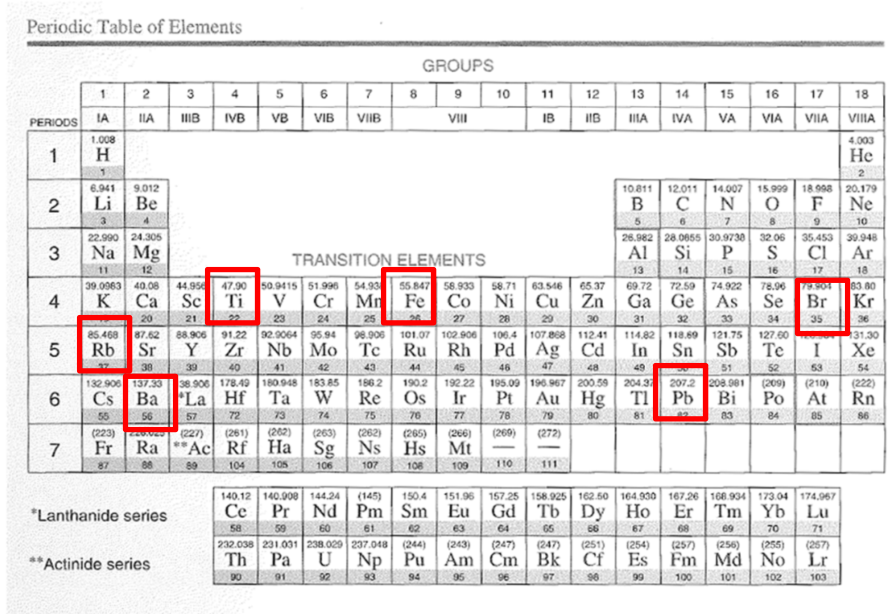
Ar: $1s^2 2s^2 2p^6 3s^2 3p^6$

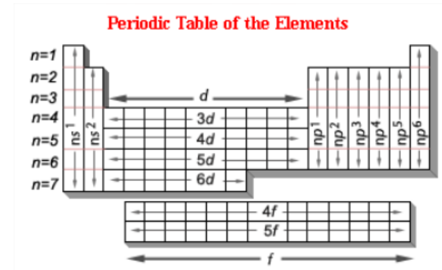
Kr: $1s^2 2s^2 2p^6$.........$4p^6$


Periodic Table of the Elements
n=1
n=2
n=3
n=4
n=5
n=6
n=7
ns 1
ns 2
d
3d
4d
5d
6d
np 1
np 2
np 3
np 4
np 5
np 6
4f
5f
f

Write the Ground State Electron Configuration for the following elements:







Element	# of e^-	End with:	Write the Electron Configuration
Iron, Fe	26	$3d^6$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
Bromine, Br	35	$4p^6$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
Barium, Ba	56	$6s^2$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$
Titanium, Ti	22	$3d^2$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
Rubidium, Rb	37	$5s^1$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$
Lead, Pb	82	$6p^2$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^2$

The sum of the superscripts = # of electrons