Announcements

Wednesday, November 18, 2009

Quiz 3 will be next Wed, Nov 25.

MasteringChemistry due dates (all at 11:59 pm)

- Ch 7: Wed, Nov 25
- Ch 8: Wed, Dec 2
- Ch 9: Fri, Dec 4

Chapter 8: Periodic properties of the elements

Electron configuration: shows which orbitals are occupied in an atom, and how many electrons they contain

Ground state: lowest energy, most stable state for an atom - has all electrons in the lowest energy possible orbitals

The energy of an electron in an H atom depends only on the principal quantum number, *n*, so in its ground state, the electron in an H atom occupies a 1s orbital.

Electron configuration: $1s^{1}$ ← #electrons in subshell

Orbital diagram: figure that organizes electrons into their orbitals. Orbital = box, electron = arrow in box

H: [↑]

Electron spin: direction of an electron's inherent angular momentum - this creates a magnetic field around the electron that either points **up** or **down**.

Spin quantum number, m_s, defines electron spin

- $m_s = +\frac{1}{2}$: up spin
- $m_s = -\frac{1}{2}$: down spin

Multi-electron atoms

The first 3 quantum numbers define an orbital

All 4 quantum numbers define **one electron** in an atom

- ***n*:** size of orbital
- ***ℓ*:** shape of orbital
- ***m_ℓ*:** orientation of orbital
- ***m_s***: spin of electron in that orbital

Pauli exclusion principle: no two electrons in an atom can have the same four quantum numbers

This means each orbital can hold no more than 2 electrons

He, 2 electrons, ground state electron configuration:

Configuration: $1s^2$

Orbital diagram: 1s [↑↓]

	n	ℓ	m_ℓ	m_s
↑:	1	0	0	$+\frac{1}{2}$
↓:	1	0	0	$-\frac{1}{2}$

Be, 4 electrons, configuration: $1s^2\ 2s^2$

Orbital diagram:

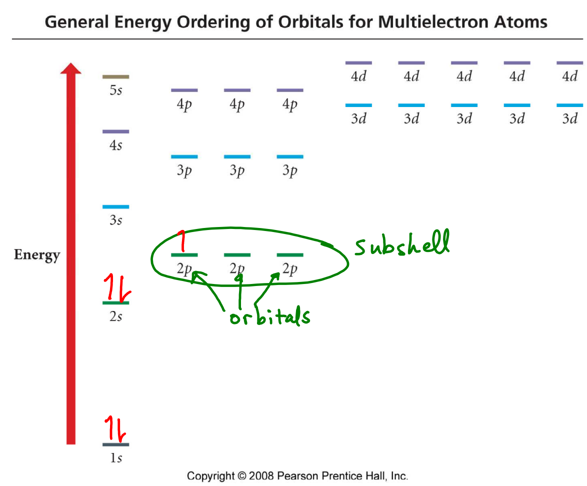
[↑↓] 1s [↑↓] 2s

Energy of electrons in multi-electron atoms

In hydrogen, *n* is the only quantum number necessary to calculate the energy of an orbital

In multi-electron atoms, both *n* and *ℓ* influence the energy of an orbital

Energies of *ℓ*: (lowest) **s < p < d < f** (highest)

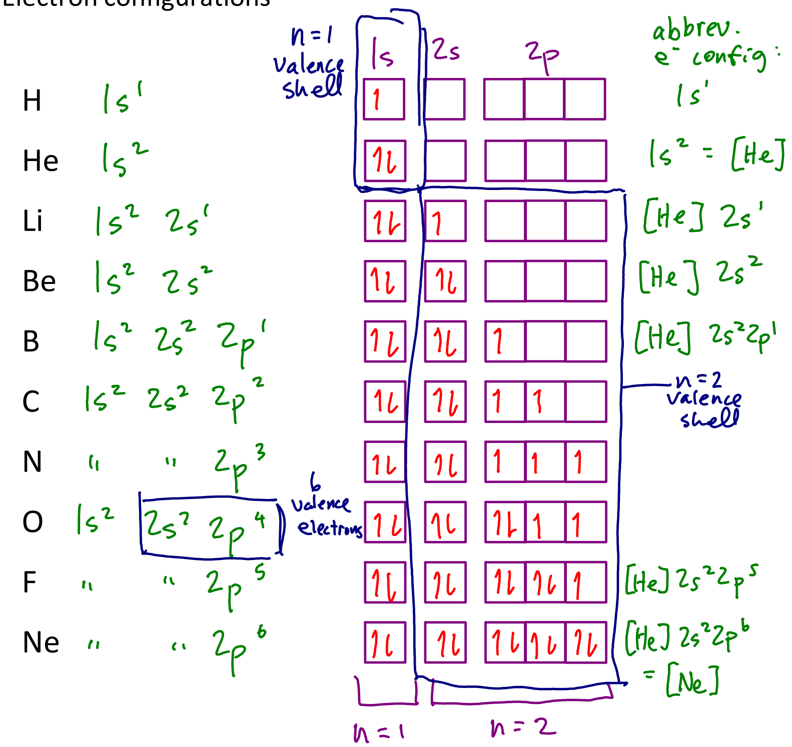


B (5 electrons) - configuration: $1s^2\ 2s^2\ 2p^1$

Orbital diagram: 1s [↑↓] 2s [↑↓] 2p [↑][ ][ ]

Also, since energies get closer together as *n* increases, a **4s** orbital is lower in energy than a **3d** orbital

Electron configurations



Hund's rule: Electrons fill orbitals with equal energy **singly** first with parallel spins.

Inner electrons: A full noble gas electron configuration inside an atom - represented with noble gas symbol in brackets to make **abbreviated electron configuration.**

Periodic table and valence electrons

Outer Electron Configurations of Elements 1–18

valence electrons

1A	2A	3A	4A	5A	6A	7A	8A
1 H $1s^1$							2 He ($1s^2$)
3 Li $2s^1$	4 Be $2s^2$	5 B $2s^2 2p^1$	6 C $2s^2 2p^2$	7 N $2s^2 2p^3$	8 O $2s^2 2p^4$	9 F $2s^2 2p^5$	10 Ne $2s^2 2p^6$
11 Na $3s^1$	12 Mg $3s^2$	13 Al $3s^2 3p^1$	14 Si $3s^2 3p^2$	15 P $3s^2 3p^3$	16 S $3s^2 3p^4$	17 Cl $3s^2 3p^5$	18 Ar $3s^2 3p^6$

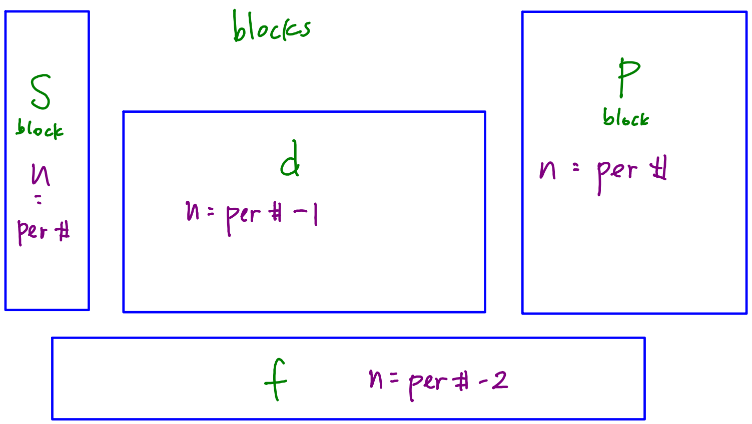
8

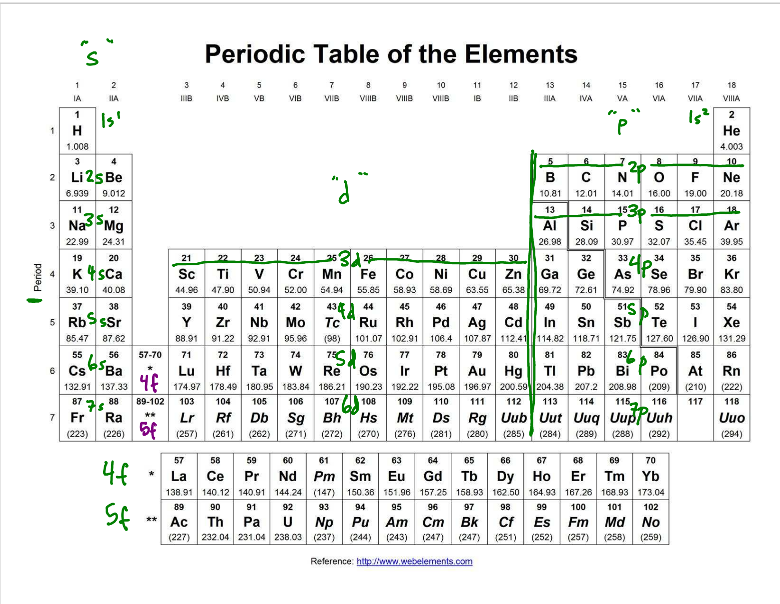
Valence electrons: number of electrons in outermost principal energy level (*n*) (plus outermost d electrons for transition elements)

Core electrons: inner electrons plus filled d or f sublevels

Periodic table and filling order

The ***reason*** the periodic table has its shape is because of the orbitals occupied in those elements.





Electron configurations

Ca: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$: $[Ar]\ 4s^2$

Sc: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^1$: $[Ar]\ 4s^2 3d^1$

rearranged: $3p^6\ 3d^1\ 4s^2$

Sometimes, electron configurations are rearranged into order of increasing *n* (to group valence electrons better)

periodic table (filling) order

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

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^4$ ← valence electrons

<u>Abbreviated configurations:</u>

Po: filling:

rearranged:

Bh: filling:

rearranged: