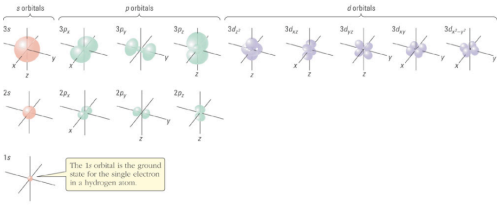
Orbital Shapes



Quantum Numbers

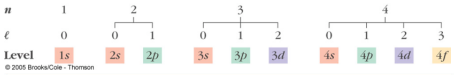


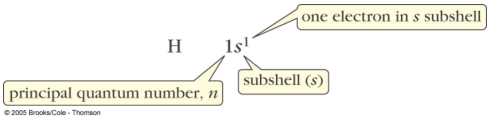
TABLE 7.3 Relationships Among n, ℓ, and m_ℓ for the First Four Principal Energy Levels

n Value	ℓ Value	Subshell Designation	m_ℓ Values	Number of Orbitals in Subshell, $2\ell + 1$	Total Number of Orbitals in Shell, n^2
1	0	1*s*	0	1	1
2	0	2*s*	0	1	
2	1	2*p*	1, 0, −1	3	4
3	0	3*s*	0	1	
3	1	3*p*	1, 0, −1	3	
3	2	3*d*	2, 1, 0, −1, −2	5	9
4	0	4*s*	0	1	
4	1	4*p*	1, 0, −1	3	
4	2	4*d*	2, 1, 0, −1, −2	5	
4	3	4*f*	3, 2, 1, 0, −1, −2, −3	7	
					16

Atom Electron Configurations

Electron Configuration – complete description of orbitals occupied by all the electrons in an atom

- for atoms in ground state – electrons occupy energy shells, subshells and orbitals that give the lowest energy for the atom
- start with 1*s* orbital and work up



Electron Configurations

TABLE 7.5 Electron Configurations of the First Ten Elements

	Electron Configurations		Orbital Box Diagrams				
	Condensed	*Expanded*	*1s*	*2s*	*2p*		
H	$1s^1$		↑				
He	$1s^2$		↑↓				
Li	$1s^2\,2s^1$		↑↓	↑			
Be	$1s^2\,2s^2$		↑↓	↑↓			
B	$1s^2\,2s^2\,2p^1$		↑↓	↑↓	↑		
C	$1s^2\,2s^2\,2p^2$	$1s^2\,2s^2\,2p^1\,2p^1$	↑↓	↑↓	↑	↑	
N	$1s^2\,2s^2\,2p^3$	$1s^2\,2s^2\,2p^1\,2p^1\,2p^1$	↑↓	↑↓	↑	↑	↑
O	$1s^2\,2s^2\,2p^4$	$1s^2\,2s^2\,2p_x^2\,2p^1\,2p^1$	↑↓	↑↓	↑↓	↑	↑
F	$1s^2\,2s^2\,2p^5$	$1s^2\,2s^2\,2p^2\,2p^2\,2p^1$	↑↓	↑↓	↑↓	↑↓	↑
Ne	$1s^2\,2s^2\,2p^6$	$1s^2\,2s^2\,2p^2\,2p^2\,2p^2$	↑↓	↑↓	↑↓	↑↓	↑↓

Electron Configurations

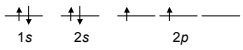
for orbitals with same energy (degenerate) such as the three 2*p* orbitals

- half-fill each orbital with first 3 electrons
- go back and pair electrons with 4, 5 and 6

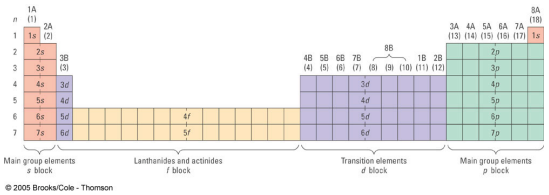
Hund's Rule – electrons pair only after each orbital in a subshell is occupied by a single electron

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

or: [He] $2s^2\,2p^2$



Electron Configurations and the Periodic Table



Valence Electrons

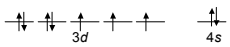
Chemically similar behavior occurs among elements within a group in the periodic table.

Valence electrons – electrons held in outer shell

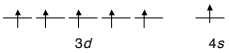
Core electrons – electrons held in filled inner shells

Electron Configurations of Transition Elements

- *d* subshell being filled
- valence electrons - *s* and *p* electrons in outermost shell plus electrons in incompletely filled (*n* - 1)*d* subshell
- (*n* - 1)*d* orbitals are filled after *ns* orbitals and before filling *np* orbitals

Co: [Ar] 

- half-filled *d* subshells are favored if possible (Hund's Rule)

Cr: [Ar] 

Example 5

What family of elements is characterized by having an ns^2np^3 valence-electron configuration?

Group 5A

Ion Electron Configurations

Metal atoms lose electrons to form cations with a positive charge equal to the group number.

Nonmetal atoms gain electrons to form anions with a negative charge equal to the A group number minus eight.

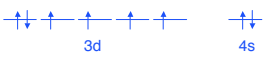
Ca (Group 2) → Ca^{2+}

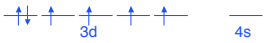
O (Group 6A) → O^{2-}

Transition Metal Ions

The *n*s electrons are at a higher energy than the (*n*-1)*d* electrons so they are always removed before the (*n*-1)*d* electrons when TMs form cations.

Fe → Fe^{2+} + 2e-

Fe^0 [Ar] 

Fe^{2+} [Ar] 

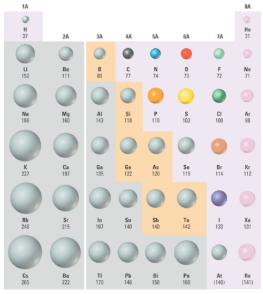
Magnetic Properties of Compounds

- Diamagnetic – atoms and ions with all electrons paired (will be repelled by magnetic field)
- Paramagnetic – atoms and ions with unpaired electrons (will be attracted to magnetic field)

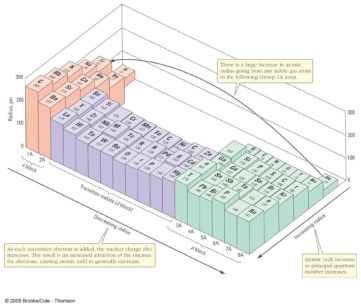
Periodic Trends: Atomic Radii

- move down periodic table – increasing *n*
- as *n* gets larger that shell's radius increases
- **atomic size (radius) increases as you move down the periodic table**
- **atomic size (radius) decreases as you move right in the periodic table**

Periodic Trends: Atomic Radii



Periodic Trends: Atomic Radii

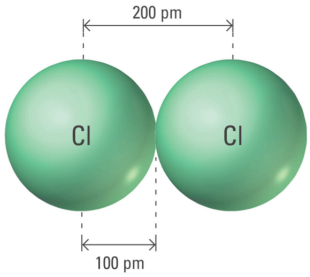


Example 6

Which will be closer to the nucleus, the $n = 3$ electron shell in Ar or the $n = 3$ electron shell in Kr?

Kr
because # protons is higher

Bond Length



Example 7

Predict which will be greater, the P—Br bond length in PBr_3 or the As—Br bond length in $AsBr_3$?

As—Br
atomic radius of As > P

Example 8

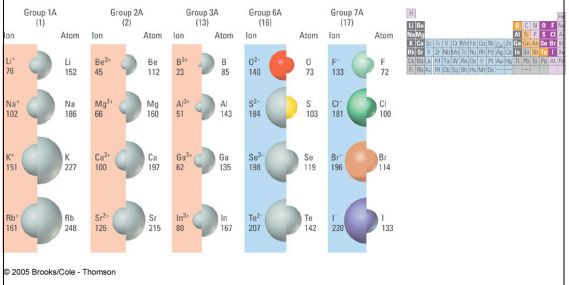
Arrange the following atoms in order of increasing atomic radius: K, Mg, Ca

Mg < Ca < K

Periodic Trends: Ionic Radii

- radii of ions of elements in same group increase in size going down the group
- radius of cation always smaller than that of neutral atom from which derived
- radius of anion always larger than that of neutral atom from which derived

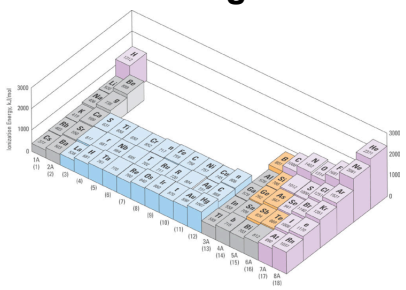
Periodic Trends: Ionic Radii



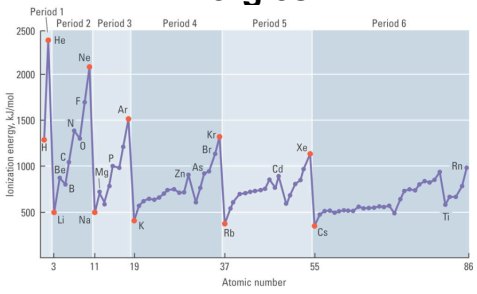
Periodic Trends: Ionization Energies

- ionization energy – energy needed to remove one electron from atom
- *s*- and *p*-block elements – first ionization energies decrease down a group and across a period

Periodic Trends: Ionization Energies



Periodic Trends: Ionization Energies



Example 9

Predict which of the following atoms – C, Si, N or P – has the lowest first ionization energy.

Si

Periodic Trends: Electron Affinities

- electron affinity (EA) – energy change when an electron is added to a gaseous atom to form 1- ion
- measure of attraction an atom has for an additional electron
- large negative value indicates a ready acceptance of an electron

$F(g) + e^- \rightarrow F^-(g)$

Periodic Trends: Electron Affinities

- electron affinity (EA) – energy change when an electron is added to a gaseous atom to form 1- ion
- measure of attraction an atom has for an additional electron
- large negative value indicates a ready acceptance of an electron

$F(g) + e^- \rightarrow F^-(g)$ $\quad \Delta E = EA = -328$ kJ/mol

$Ne + e^- \rightarrow Ne^-$ $\quad EA > 0$ kJ/mol

Periodic Trends: Electron Affinities

TABLE 7.10 Electron Affinities (kJ/mol)

1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
H −73							He >0
Li −60	Be >0	B −27	C −122	N >0	O −141	F −328	Ne >0
Na −53	Mg >0	Al −43	Si −134	P −72	S −200	Cl −349	Ar >0
K −48	Ca −2	Ga −30	Ge −119	As −78	Se −195	Br −325	Kr >0
Rb −47	Sr −5	In −30	Sn −107	Sb −103	Te −190	I −295	Xe >0