화학 General Chemistry
034.020-005

2018 Spring Semester

Tue/Thr 9:30~10:45
Building 028-302

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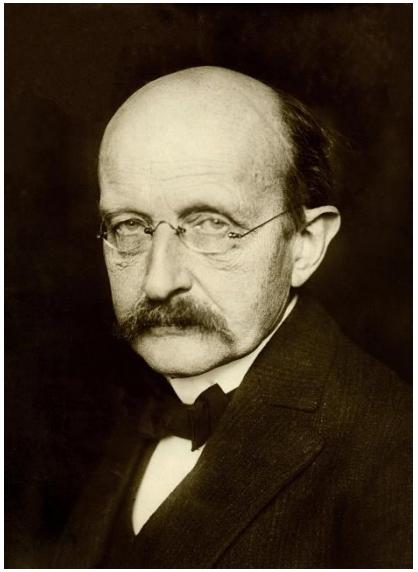
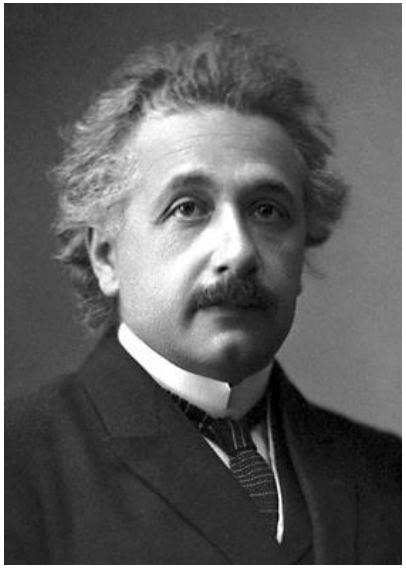


Max Planck
(1858 –1947)

Nobel Prize in Physics in 1918

Quantum theory



Albert Einstein
(1879 -1955)

Nobel Prize in Physics in 1921

Photoelectric effect

Niels Bohr
(1885 –1962)

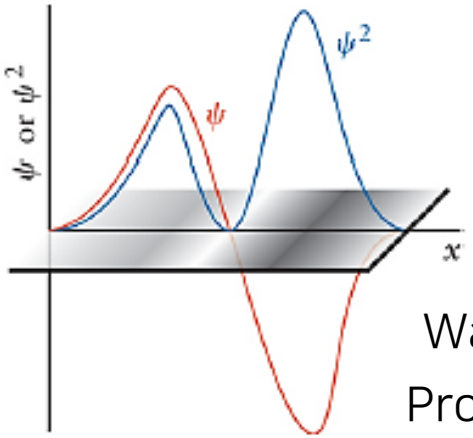
Nobel Prize in Physics in 1922

Erwin Schrödinger
(1887–1961)

Nobel Prize in Physics in 1933

1.7. Wave functions and Energy Levels

A particle has wave-like property; wavefunction



Wave function (ψ)

Probability density Function (ψ^2)

Max Born (1882-1970) Nobel Prize in Physics (1954)

Schrödinger equation

$$H\psi = E\psi$$

$$-\frac{\hbar^2}{2m}\frac{d^2\psi}{dx^2} + V(x)\psi = E\psi$$

Born interpretation

: statistical interpretation of the wave function, probability density function

m : mass of the particle

Quantized energy levels: E

Hamiltonian: H

(Matter) Wave function: $\Psi_n(x)$

$$-\frac{\hbar^2}{2m}\frac{d^2}{dx^2}\Psi(x)+V(x)\Psi(x)=E\Psi(x)$$

$\Psi_n(x,y,z)$

Matter wave associated with the particle.

$$P(x,y,z)=\left|\Psi_n(x,y,z)\right|^2=\Psi_n^*(x,y,z)\times\Psi_n(x,y,z)$$

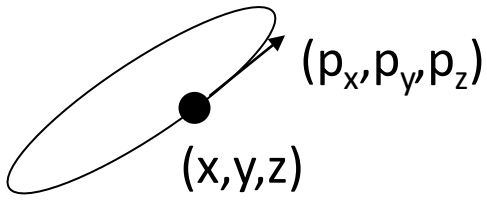
$|\Psi|^2\,dxdydz$
= Probability to find a particle at (x,y,z)

Wave function (ψ)

Probability density (ψ^2)

Classical Mechanics

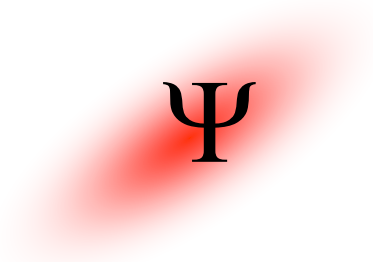
Position and momentum
of a particle can be determined



Energy is continuous

Quantum Mechanics

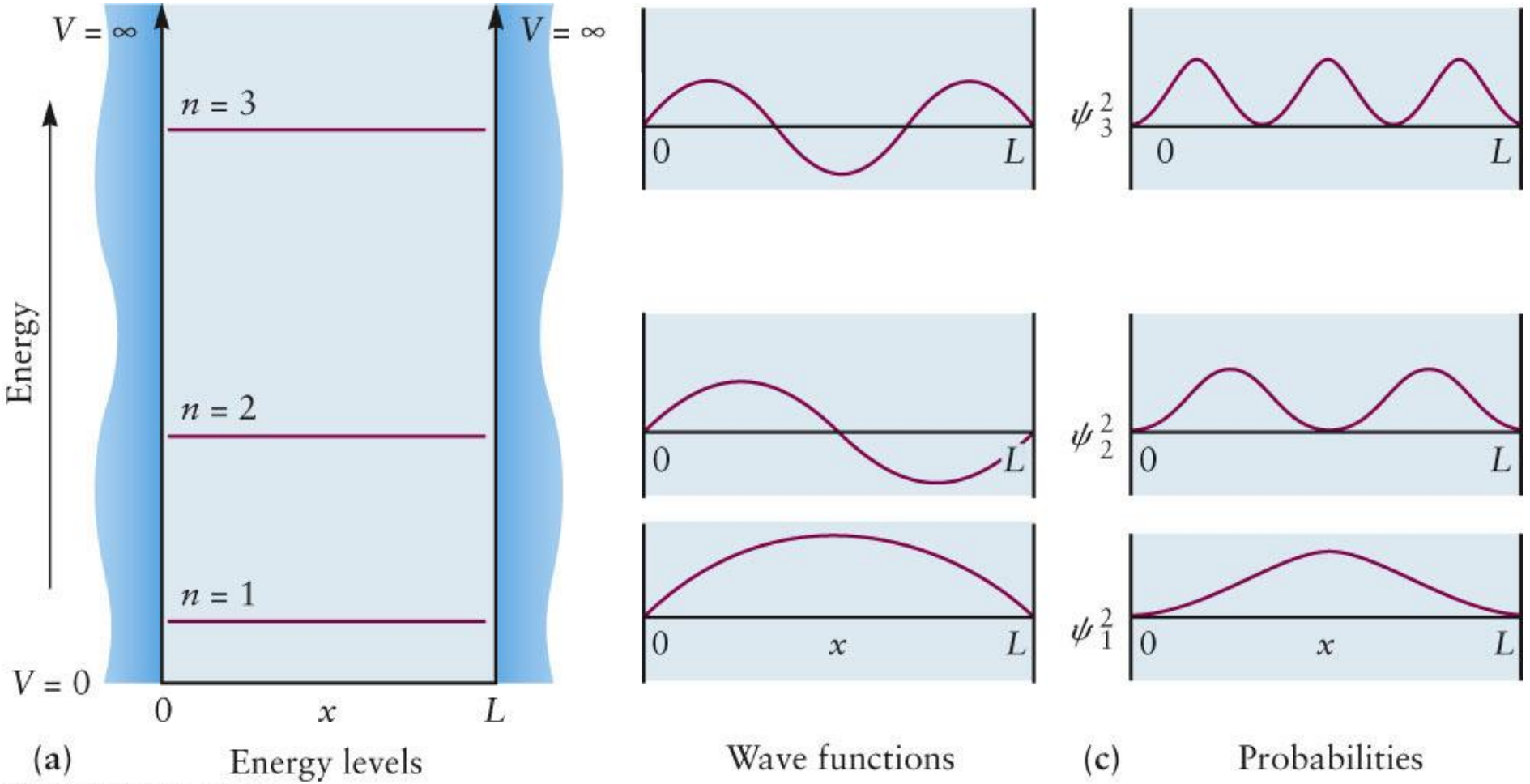
Position and momentum
of a particle cannot be determined
simultaneously



Energy is discrete (quantized)

One-dimensional boxes:
Zero potential energy inside the box
And infinite potential energy outside

$|\Psi|^2$



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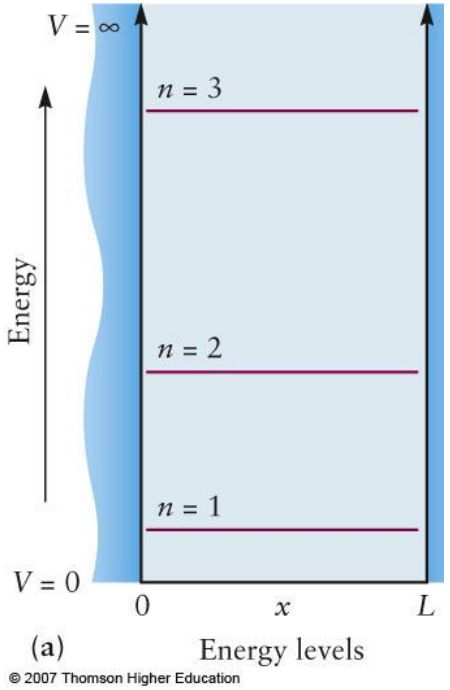
Energy is quantized (determined by n)!

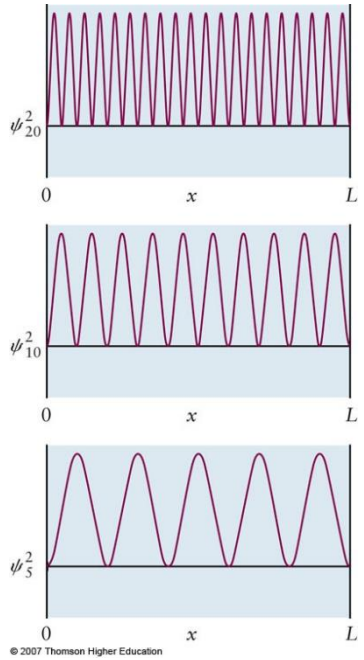
$$E_n = \frac{h^2 n^2}{8mL^2}, \Psi(x) = \sqrt{\frac{2}{L}} \sin\left(\frac{n\pi}{L} x\right)$$

$$E_n = \frac{h^2 n^2}{8mL^2}, \Psi(x) = \sqrt{\frac{2}{L}} \sin\left(\frac{n\pi}{L} x\right)$$

$$E_{n+1} - E_n = \frac{h^2 (n+1)^2}{8mL^2} - \frac{h^2 n^2}{8mL^2} = (2n+1)\frac{h^2}{8mL^2}$$



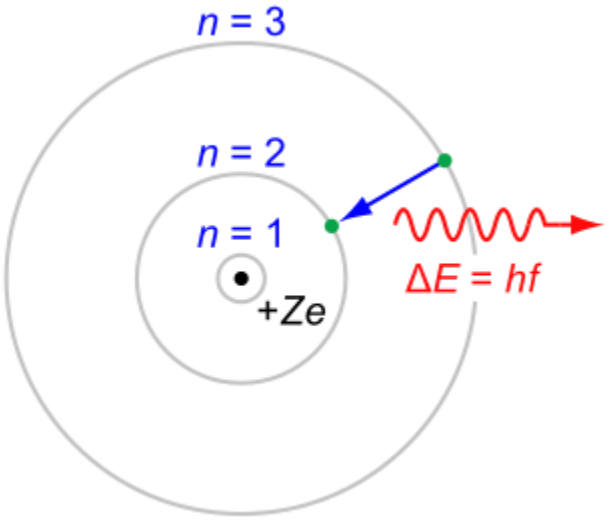
1. Energy level is quantized.
2. As the L or m increases, the separation between neighboring energy levels decreases: The quantized energy level is obvious in an electron in a hydrogen atom.
3. It is impossible to have no energy (E≠0). The lowest energy state is when n=1: E_1 = zero-point energy.
4. There is always some degrees of kinetic energy = always moving (linear momentum)= uncertainty in the position.
5. With large quantum number, the quantum mechanics should converge to classical mechanics: Correspondence Principle

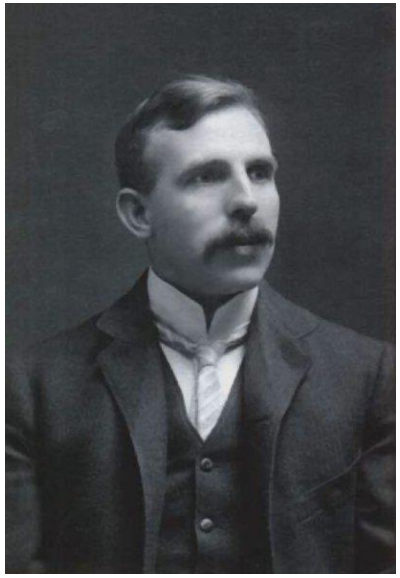


Energy: Principle Quantum Number (n)

Wave function:

4 Quantum Numbers:





The Rutherford–Bohr model of the hydrogen atom

1.8. The Principal Quantum Number

Similarly to a particle in a box, an electron in an atom

For Hydrogen Atom, V(r) = Coulomb potential energy

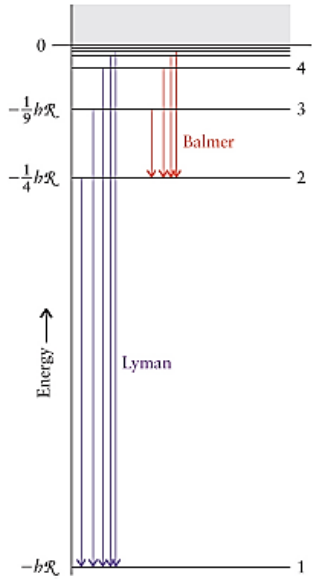
$$V(r) = \frac{(-e)(+e)}{4\pi\varepsilon_0 r} = -\frac{e^2}{4\pi\varepsilon_0 r}$$

where, ε_0 = vacuum permittivity

Solutions of the Schrödinger equation, $-\frac{\hbar^2}{2m}\frac{d^2\psi}{dx^2} + V(x)\psi = E\psi$

$$E_n = -\frac{h\mathcal{R}}{n^2} \qquad \mathcal{R} = \frac{m_e e^4}{8h^3\varepsilon_0^2} \qquad n = 1, 2, \ldots$$

where, $R = 3.29 \times 10^{15}$ Hz ⟶ Rydberg constant!



For other one-electron ions, such as He^+, Li^{2+} and even C^{5+}

$$E_n = -\frac{Z^2 h\mathcal{R}}{n^2} \qquad n = 1, 2, \ldots$$

Proportional to Z^2
= charge effect+ distance effect

Z = atomic number; n = principal quantum number

1.9. Atomic Orbitals

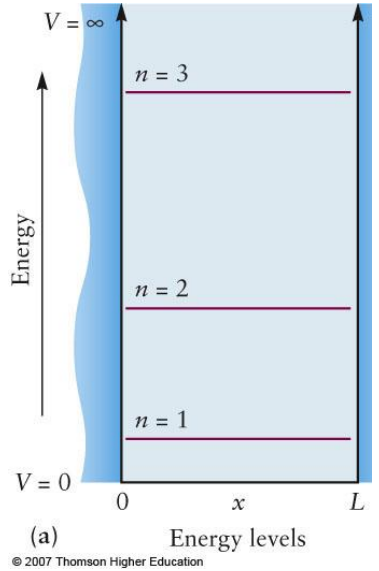
- Atomic Orbitals: the wave functions of electrons in atoms

$$-\frac{\hbar^2}{2m}\nabla^2\psi + V(x,y,z)\,\psi = E\,\psi$$

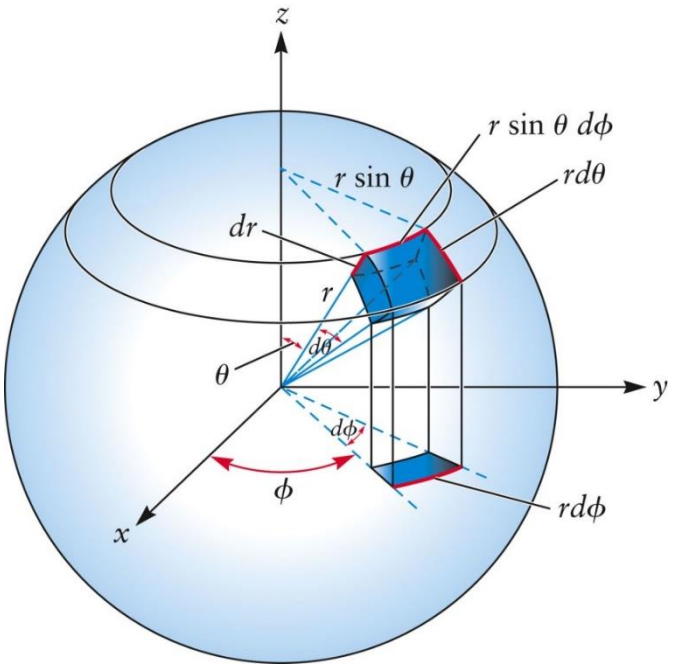
Kinetic energy | Potential energy | Total energy

$$\psi(r,\theta,\phi) = R(r)\,Y(\theta,\phi)$$

Radial wavefunction | Angular wavefunction



$$E_n = \frac{h^2 n^2}{8mL^2}, \Psi(x) = \sqrt{\frac{2}{L}}\sin\left(\frac{n\pi}{L}x\right)$$



$x = r\cos\theta\sin\phi$
$y = r\sin\theta\sin\phi$
$z = r\cos\theta$
$x^2+y^2+z^2 = r^2$

Spherical Coordinate:
$(x, y, z) \rightarrow (r, \theta, \phi)$

→ n, l, m_l

TABLE 1.2 Hydrogen Wavefunctions (Atomic Orbitals), $\psi = RY$

(a) Radial wavefunctions, $R_{nl}(r)$

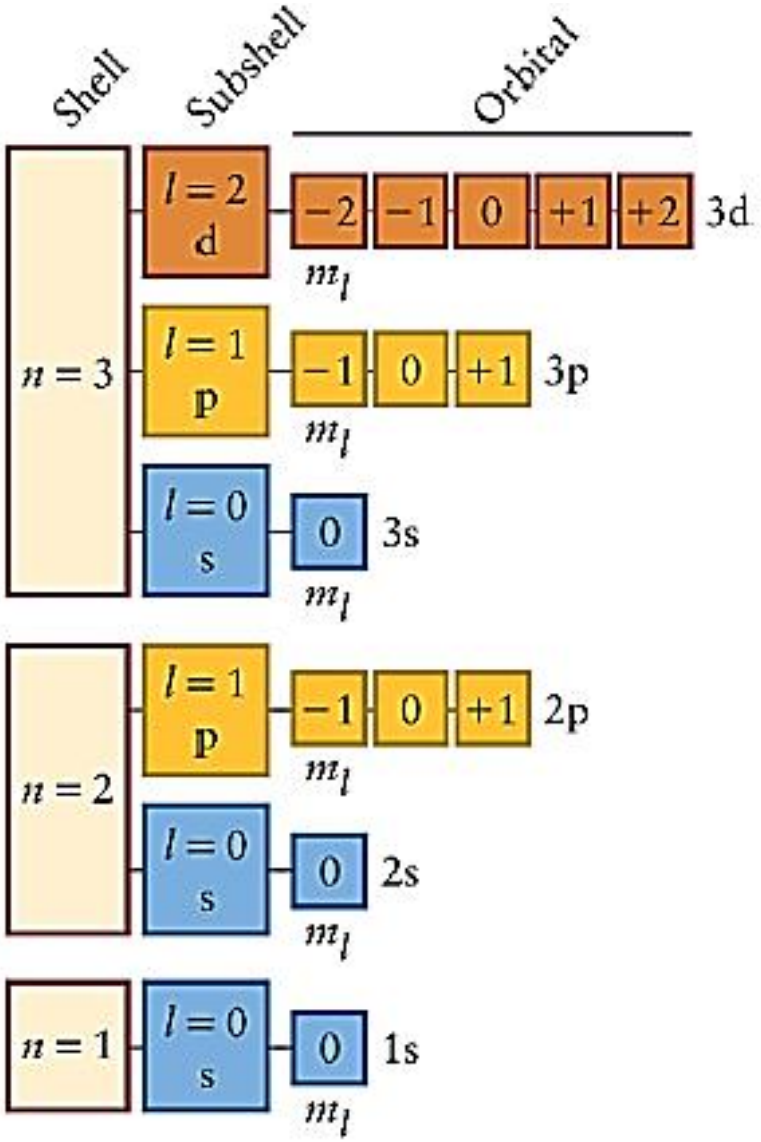
n	l	$R_{nl}(r)$
1	0	$2\left(\frac{Z}{a_0}\right)^{3/2} e^{-Zr/a_0}$
2	0	$\frac{1}{2\sqrt{2}}\left(\frac{Z}{a_0}\right)^{3/2}\left(2-\frac{Zr}{a_0}\right)e^{-Zr/2a_0}$
	1	$\frac{1}{2\sqrt{6}}\left(\frac{Z}{a_0}\right)^{3/2}\left(\frac{Zr}{a_0}\right)e^{-Zr/2a_0}$
3	0	$\frac{1}{9\sqrt{3}}\left(\frac{Z}{a_0}\right)^{3/2}\left(3-\frac{2Zr}{a_0}+\frac{2Z^2r^2}{9{a_0}^2}\right)e^{-Zr/3a_0}$
	1	$\frac{2}{27\sqrt{6}}\left(\frac{Z}{a_0}\right)^{3/2}\left(2-\frac{Zr}{3a_0}\right)e^{-Zr/3a_0}$
	2	$\frac{4}{81\sqrt{30}}\left(\frac{Z}{a_0}\right)^{3/2}\left(\frac{Zr}{a_0}\right)^2 e^{-Zr/3a_0}$

n = principal quantum number 1, 2, ⋯
l = orbital angular momentum quantum number
0, 1, 2, ...n-1
m_l = magnetic quantum number l, l-1,⋯-l

(b) Angular wavefunctions, $Y_{lm_l}(\theta,\phi)$

l	"m_l"*	$Y_{lm_l}(\theta,\phi)$
0	0	$\left(\frac{1}{4\pi}\right)^{1/2}$
1	x	$\left(\frac{3}{4\pi}\right)^{1/2}\sin\theta\cos\phi$
	y	$\left(\frac{3}{4\pi}\right)^{1/2}\sin\theta\sin\phi$
	z	$\left(\frac{3}{4\pi}\right)^{1/2}\cos\theta$
2	xy	$\left(\frac{15}{16\pi}\right)^{1/2}\sin^2\theta\cos 2\phi$
	yz	$\left(\frac{15}{4\pi}\right)^{1/2}\cos\theta\sin\theta\sin\phi$
	zx	$\left(\frac{15}{4\pi}\right)^{1/2}\cos\theta\sin\theta\cos\phi$
	x^2-y^2	$\left(\frac{15}{16\pi}\right)^{1/2}\sin^2\theta\sin 2\phi$
	z^2	$\left(\frac{5}{16\pi}\right)^{1/2}(3\cos^2\theta-1)$

Note: In each case, $a_0 = 4\pi\varepsilon_0{}^2/m_e e^2$, or close to 52.9 pm; for hydrogen itself, $Z = 1$.
*In all cases except $m_l = 0$, the orbitals are sums and differences of orbitals with specific values of m_l.



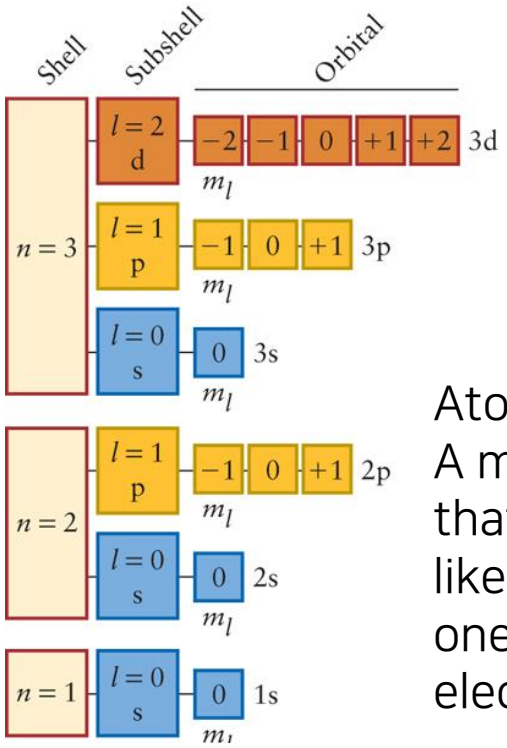
n: principal quantum number
→ related to the size and energy of the orbital; 1, 2, ⋯
→ radius of orbit

l: orbital angular momentum quantum number (*l* = 0, 1, 2, ..., *n*-1)
→ related to the orbital angular momentum of the electron;
→ s, p, d, f orbital

m_l: magnetic quantum number (m_l = *l*, *l*-1, *l*-2, ⋯, -*l*)
→ related to the orientation of the orbital motion of the electron

Energy is dependent only on n; Total number of orbital = n^2

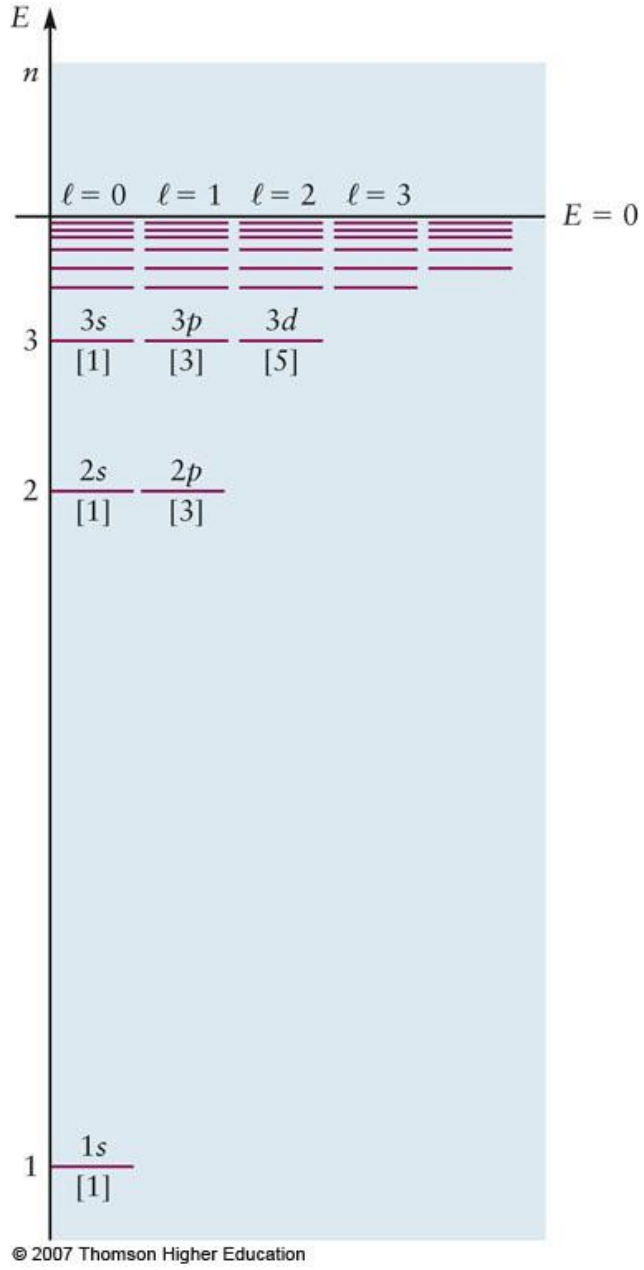
Representation of energy levels



Atomic Orbital =
A mathematical function that describes the wave-like behavior of either one electron or a pair of electrons in an atom.

TABLE 7.2 Quantum Numbers for the First Four Levels of Orbitals in the Hydrogen Atom

n	ℓ	Orbital Designation	m_ℓ	Number of Orbitals
1	0	1*s*	0	1
2	0	2*s*	0	1
	1	2*p*	−1, 0, +1	3
3	0	3*s*	0	1
	1	3*p*	−1, 0, 1	3
	2	3*d*	−2, −1, 0, 1, 2	5
4	0	4*s*	0	1
	1	4*p*	−1, 0, 1	3
	2	4*d*	−2, −1, 0, 1, 2	5
	3	4*f*	−3, −2, −1, 0, 1, 2, 3	7



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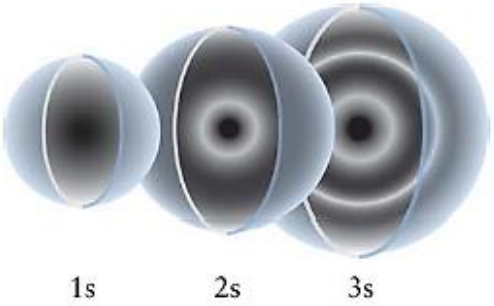
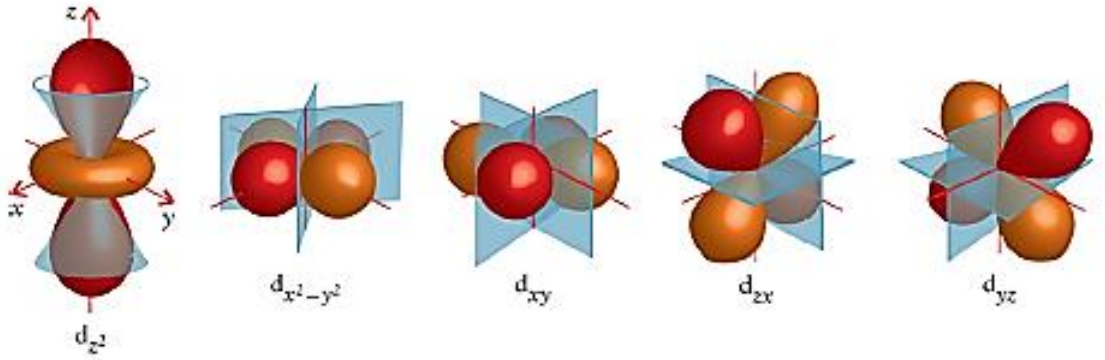
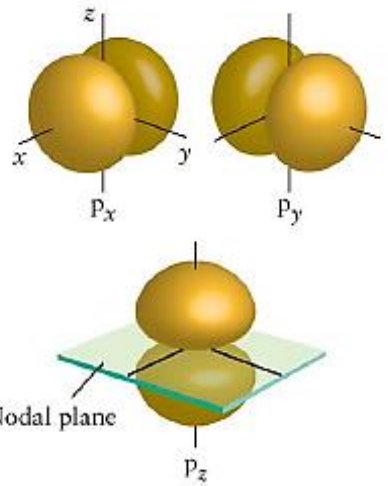
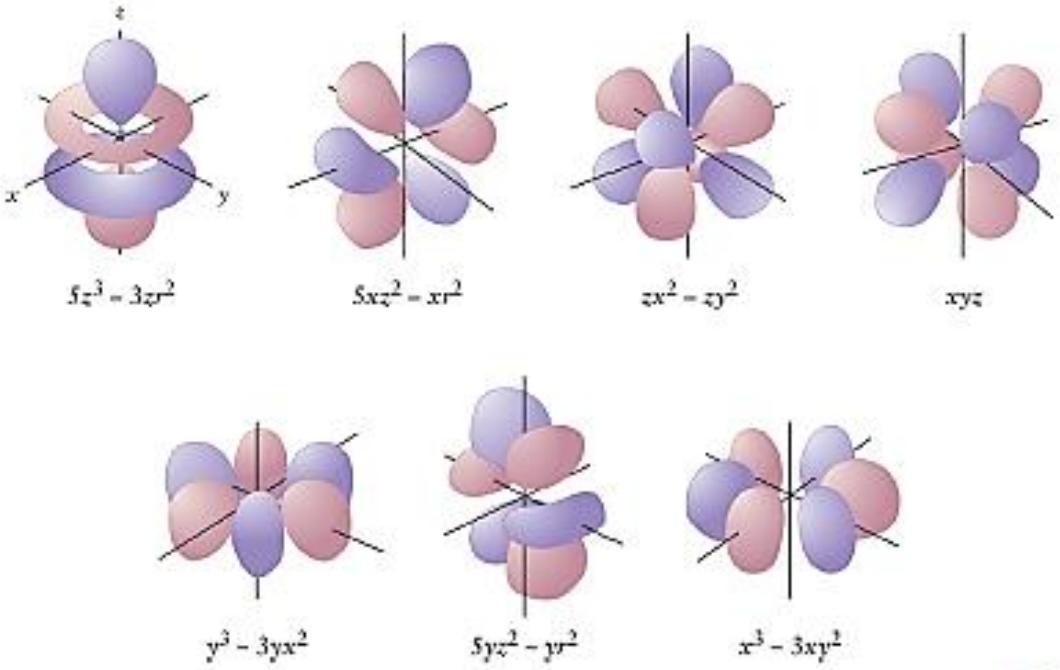
Shape of orbitals



Spherically symmetric.
Number of nodes increase with n







p_x, p_y and p_z have the same energy:
(degenerate states)
p-orbitals have nodal plane along
xy, yz, and zx planes

Size and Shape of orbitals

l = 0: s-orbitals:

Spherically symmetric.

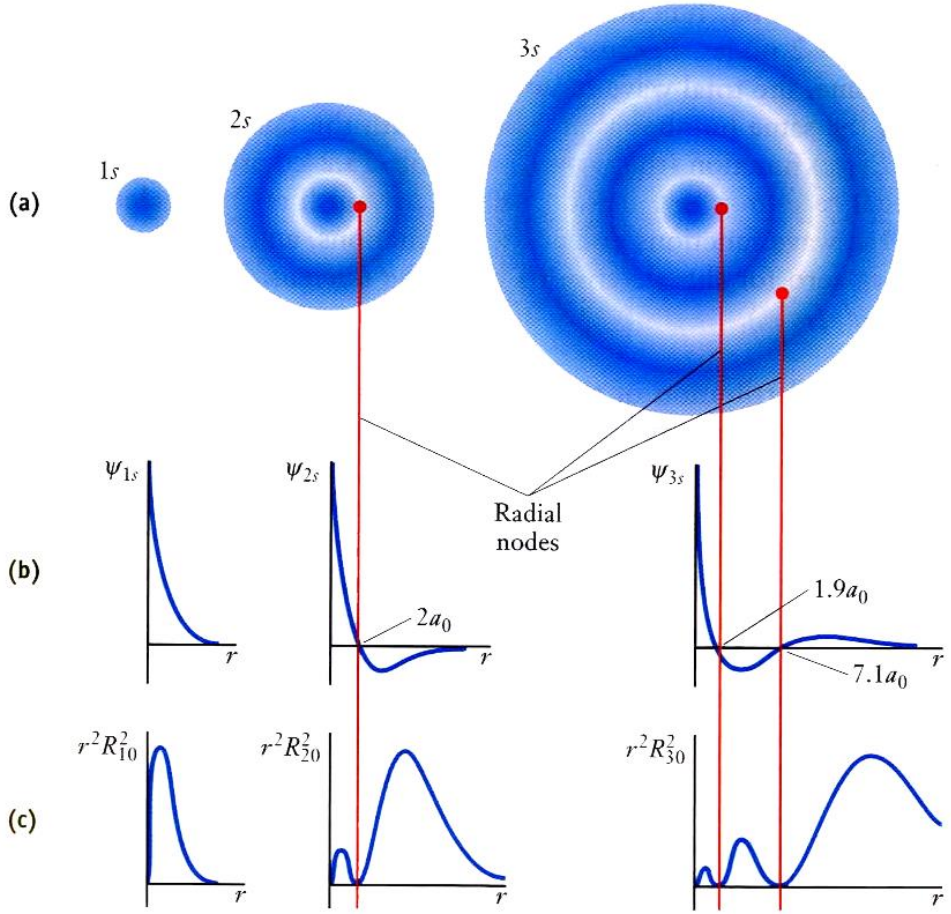
Number of nodes increase with n

(a) Radial wavefunctions, $R_{nl}(r)$

n	l	$R_{nl}(r)$
1	0	$2\left(\frac{Z}{a_0}\right)^{3/2} e^{-Zr/a_0}$
2	0	$\frac{1}{2\sqrt{2}}\left(\frac{Z}{a_0}\right)^{3/2}\left(2-\frac{Zr}{a_0}\right)e^{-Zr/2a_0}$
	1	$\frac{1}{2\sqrt{6}}\left(\frac{Z}{a_0}\right)^{3/2}\left(\frac{Zr}{a_0}\right)e^{-Zr/2a_0}$
3	0	$\frac{1}{9\sqrt{3}}\left(\frac{Z}{a_0}\right)^{3/2}\left(3-\frac{2Zr}{a_0}+\frac{2Z^2r^2}{9a_0^2}\right)e^{-Zr/3a_0}$

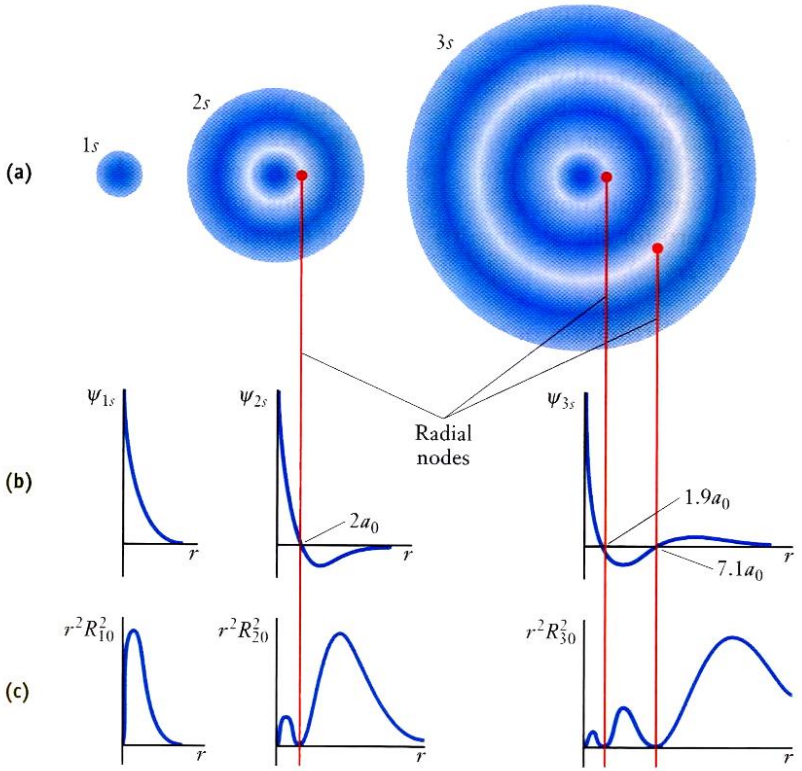
(b) Angular wavefunctions, $Y_{lm_l}(\theta,\phi)$

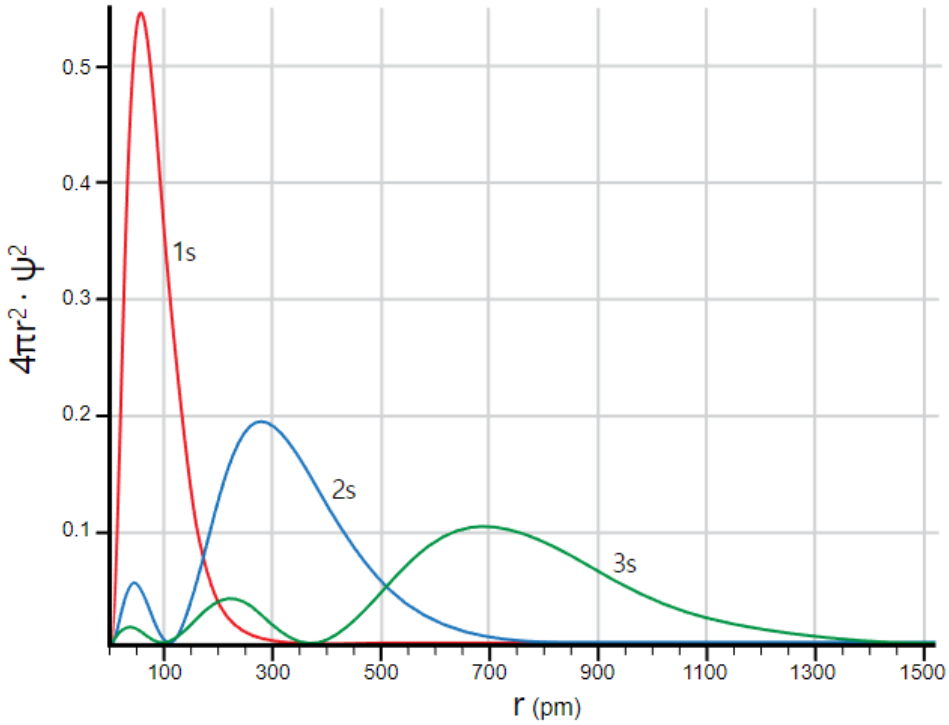
l	"m_l"*	$Y_{lm_l}(\theta,\phi)$
0	0	$\left(\frac{1}{4\pi}\right)^{1/2}$
1	x	$\left(\frac{3}{4\pi}\right)^{1/2}\sin\theta\cos\phi$
	y	$\left(\frac{3}{4\pi}\right)^{1/2}\sin\theta\sin\phi$
	z	$\left(\frac{3}{4\pi}\right)^{1/2}\cos\theta$



$a_0 = 4\pi\varepsilon_0{}^2/m_e e^2$, or close to 52.9 pm; Bohr radius

Hydrogen Wavefunctions





2*p*-orbital: p_x, p_y and p_z

- p_x, p_y and p_z have the same energy = degenerate states
- p-orbitals have nodal plane along xy, yz, and zx planes

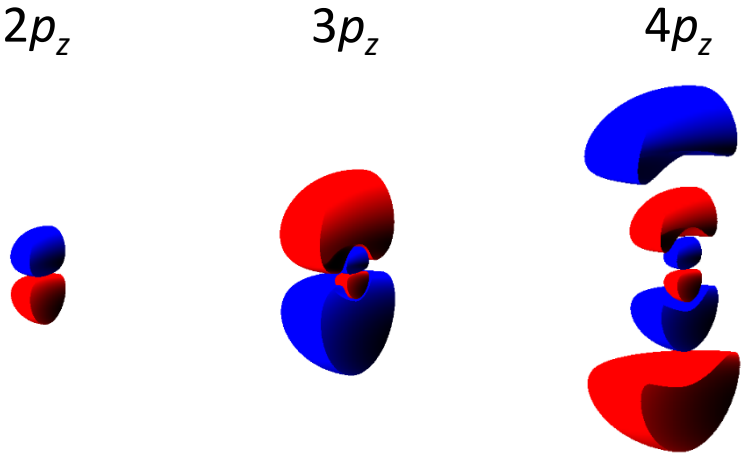
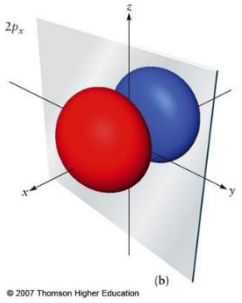


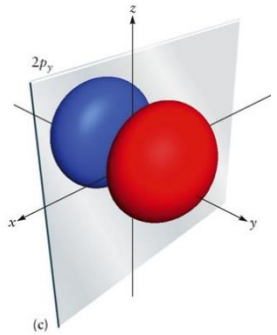
TABLE 1.2 Hydrogen Wavefunctions (Atomic Orbitals), $\psi = RY$

(a) Radial wavefunctions, $R_{nl}(r)$			(b) Angular wavefunctions, $Y_{lm_l}(\theta,\phi)$		
n	l	$R_{nl}(r)$	l	"m_l"*	$Y_{lm_l}(\theta,\phi)$
1	0	$2\left(\frac{Z}{a_0}\right)^{3/2} e^{-Zr/a_0}$	0	0	$\left(\frac{1}{4\pi}\right)^{1/2}$
2	0	$\frac{1}{2\sqrt{2}}\left(\frac{Z}{a_0}\right)^{3/2}\left(2-\frac{Zr}{a_0}\right)e^{-Zr/2a_0}$	1	x	$\left(\frac{3}{4\pi}\right)^{1/2}\sin\theta\cos\phi$
	1	$\frac{1}{2\sqrt{6}}\left(\frac{Z}{a_0}\right)^{3/2}\left(\frac{Zr}{a_0}\right)e^{-Zr/2a_0}$		y	$\left(\frac{3}{4\pi}\right)^{1/2}\sin\theta\sin\phi$
3	0	$\frac{1}{9\sqrt{3}}\left(\frac{Z}{a_0}\right)^{3/2}\left(3-\frac{2Zr}{a_0}+\frac{2Z^2r^2}{9a_0^2}\right)e^{-Zr/3a_0}$		z	$\left(\frac{3}{4\pi}\right)^{1/2}\cos\theta$
	1	$\frac{2}{27\sqrt{6}}\left(\frac{Z}{a_0}\right)^{3/2}\left(2-\frac{Zr}{3a_0}\right)e^{-Zr/3a_0}$	2	xy	$\left(\frac{15}{16\pi}\right)^{1/2}\sin^2\theta\cos 2\phi$
	2	$\frac{4}{81\sqrt{30}}\left(\frac{Z}{a_0}\right)^{3/2}\left(\frac{Zr}{a_0}\right)^2 e^{-Zr/3a_0}$		yz	$\left(\frac{15}{4\pi}\right)^{1/2}\cos\theta\sin\theta\sin\phi$

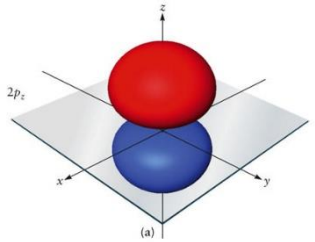
$2p_x$



$2p_y$



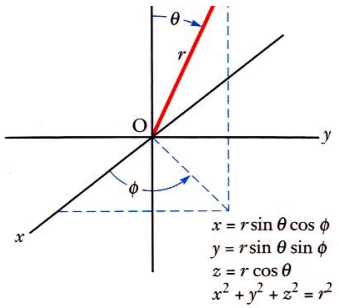
$2p_z$



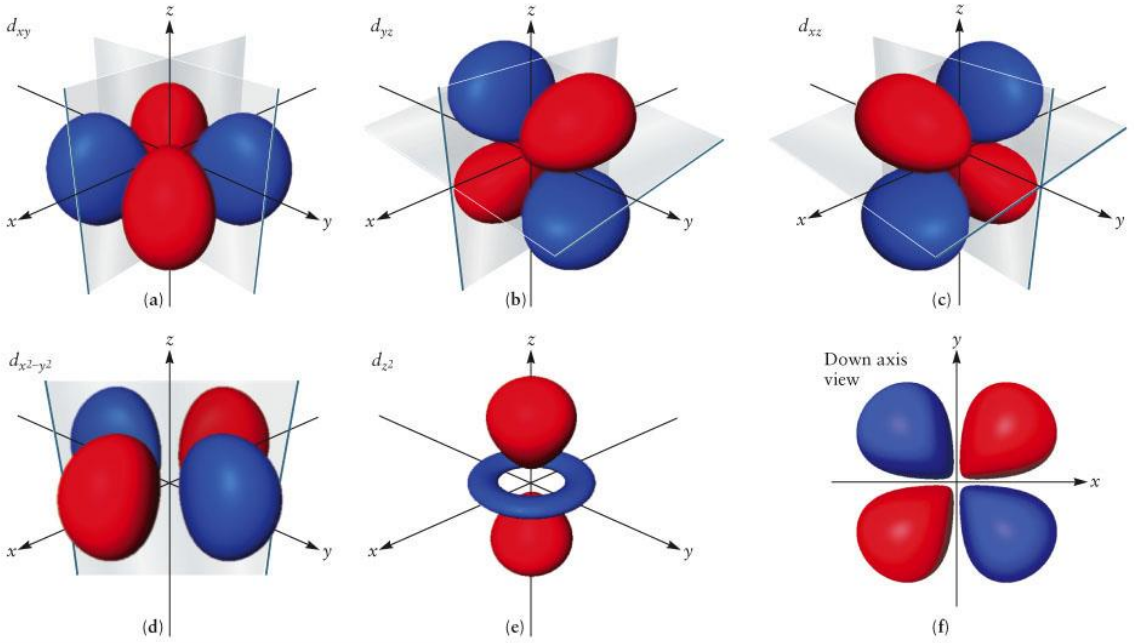
Node: yz plane: Φ = 90

Node: xz plane: Φ = 0

Node: xy plane: θ = 90



3*d*-orbital



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1	0	$2\left(\frac{Z}{a_0}\right)^{3/2} e^{-Zr/a_0}$
2	0	$\frac{1}{2\sqrt{2}}\left(\frac{Z}{a_0}\right)^{3/2}\left(2-\frac{Zr}{a_0}\right)e^{-Zr/2a_0}$
	1	$\frac{1}{2\sqrt{6}}\left(\frac{Z}{a_0}\right)^{3/2}\left(\frac{Zr}{a_0}\right)e^{-Zr/2a_0}$
3	0	$\frac{1}{9\sqrt{3}}\left(\frac{Z}{a_0}\right)^{3/2}\left(3-\frac{2Zr}{a_0}+\frac{2Z^2r^2}{9a_0^2}\right)e^{-Zr/3a_0}$
	1	$\frac{2}{27\sqrt{6}}\left(\frac{Z}{a_0}\right)^{3/2}\left(2-\frac{Zr}{3a_0}\right)e^{-Zr/3a_0}$
	2	$\frac{4}{81\sqrt{30}}\left(\frac{Z}{a_0}\right)^{3/2}\left(\frac{Zr}{a_0}\right)^2 e^{-Zr/3a_0}$

(b) Angular wavefunctions, $Y_{lm_l}(\theta,\phi)$

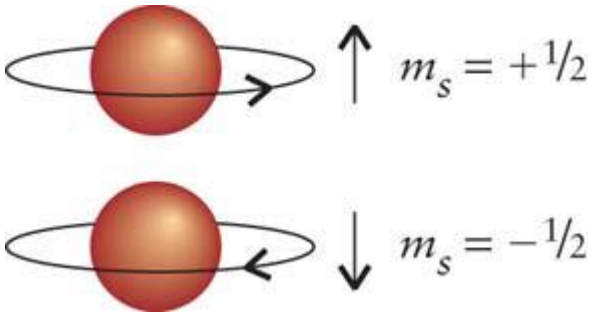
l	"m_l"*	$Y_{lm_l}(\theta,\phi)$
0	0	$\left(\frac{1}{4\pi}\right)^{1/2}$
1	x	$\left(\frac{3}{4\pi}\right)^{1/2}\sin\theta\cos\phi$
	y	$\left(\frac{3}{4\pi}\right)^{1/2}\sin\theta\sin\phi$
	z	$\left(\frac{3}{4\pi}\right)^{1/2}\cos\theta$
2	xy	$\left(\frac{15}{16\pi}\right)^{1/2}\sin^2\theta\cos 2\phi$
	yz	$\left(\frac{15}{4\pi}\right)^{1/2}\cos\theta\sin\theta\sin\phi$
	zx	$\left(\frac{15}{4\pi}\right)^{1/2}\cos\theta\sin\theta\cos\phi$
	x^2-y^2	$\left(\frac{15}{16\pi}\right)^{1/2}\sin^2\theta\sin 2\phi$
	z^2	$\left(\frac{5}{16\pi}\right)^{1/2}(3\cos^2\theta-1)$

1.10. Electron Spin

m_s: spin magnetic quantum number

(m_s = +1/2, -1/2)

→ related to the spin of electron



Name	Symbol	Values	Specifies	Indicates
principal	n	1, 2, . . .	shell	size
orbital angular momentum*	l	0, 1, . . ., $n-1$	subshell: l = 0, 1, 2, 3, 4, . . . s, p, d, f, g, . . .	shape
magnetic	m_l	$l, l-1, \ldots, -l$	orbitals of subshell	orientation
spin magnetic	m_s	$+\frac{1}{2}, -\frac{1}{2}$	spin state	spin direction

- Initially, Schrodinger's theory did not exactly match the observed spectral lines because he did not account for "spin".

1.11. The Electronic Structure of Hydrogen Atom

In the ground state,

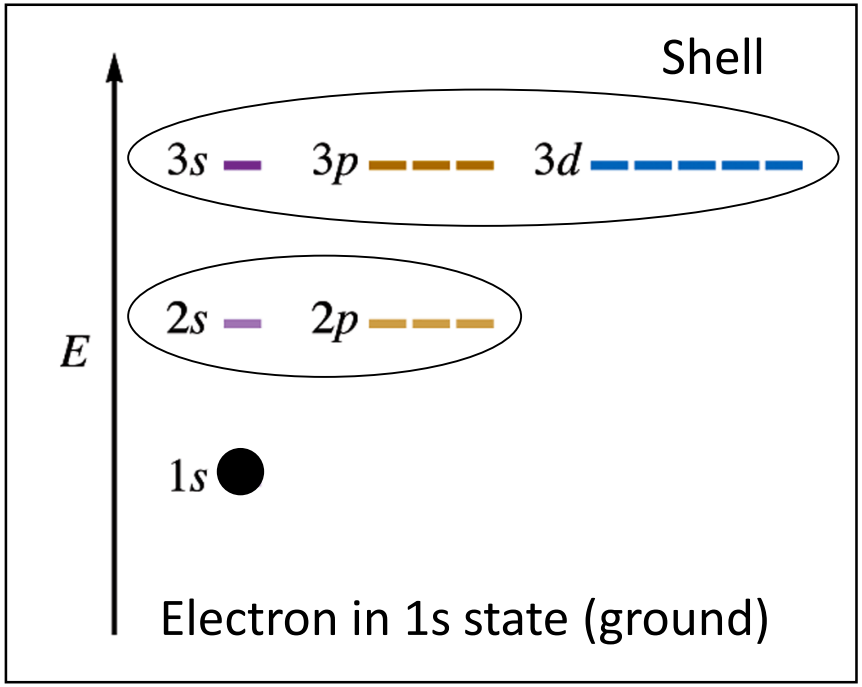
→ n = 1, l = 0, m_l = 0, m_s = +1/2 or -1/2

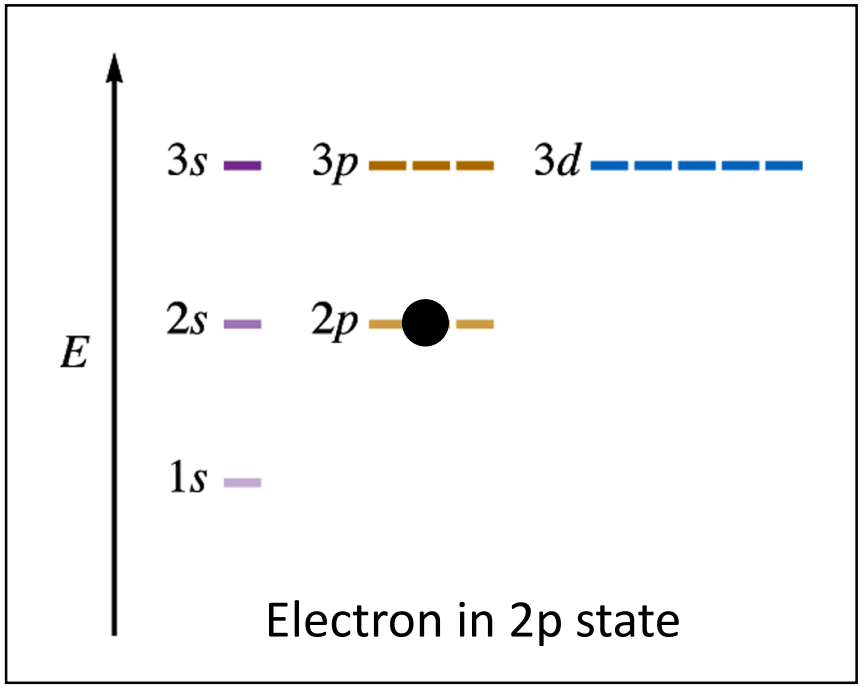
In the excited state

→ n = 2 (four orbitals), n = 3 (nine orbitals), n = k (k^2 orbitals)

n = infinite (ionization)

Degeneracy: same energy for all orbitals in the same n

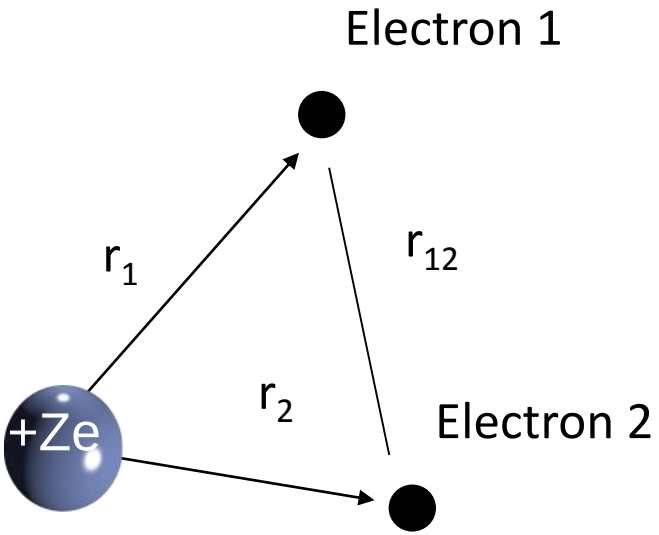




1.12. Orbital Energies

Coulomb potential for many-electron atoms

For poly-electronic atoms, electron-electron repulsion: Slightly change the energy level structure:



We can't obtain analytical solution to this:
→ Approximate solution exists.

1.12. Orbital Energies

$$V(r) = \frac{(-e)(+e)}{4\pi\varepsilon_0 r} = -\frac{e^2}{4\pi\varepsilon_0 r}$$

For one-electron atom

$$E_n = -\frac{h\mathcal{R}}{n^2} \qquad \mathcal{R} = \frac{m_e e^4}{8h^3\varepsilon_0^2}$$

$$E_n = -\frac{Z^2 h\mathcal{R}}{n^2}$$

Coulomb potential for many-electron atoms

$$V = \overbrace{-\frac{2e^2}{4\pi\varepsilon_0 r_1}}^{\text{Attraction of electron 1 to the nucleus}} \overbrace{-\frac{2e^2}{4\pi\varepsilon_0 r_2}}^{\text{Attraction electron 2 to the nucleus}} \overbrace{+\frac{e^2}{4\pi\varepsilon_0 r_{12}}}^{\text{Repulsion between the two electrons}}$$

For example, in a helium atom, r_1: e_1~nucleus; r_2: e_2~nucleus; r_{12}: e_1~e_2

Z = +2

Shielding and Penetration effect

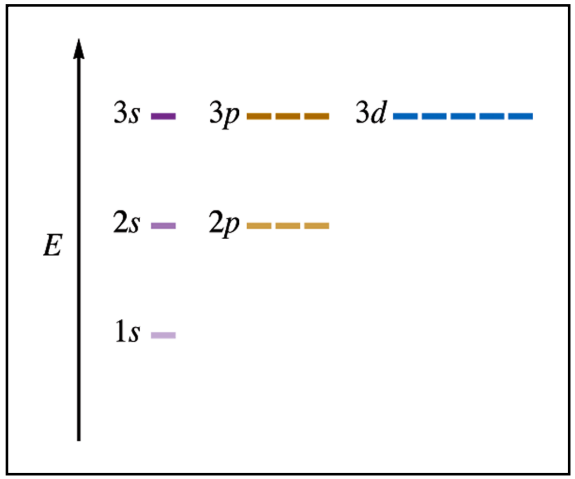
: Each electron is shielded by the other (e^-)s.

$$E_n = -\frac{Z_{\text{eff}}^2 h\mathcal{R}}{n^2}$$

Z_{eff}: effective nuclear charge → energy split

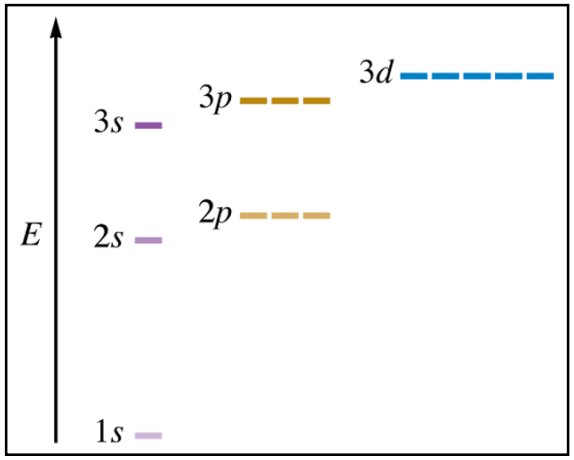
- (e^-)s in p orbital penetrate less than (e-)s in s orbital → $E_s < E_p$

Hydrogenic atom (no e-e repulsion)

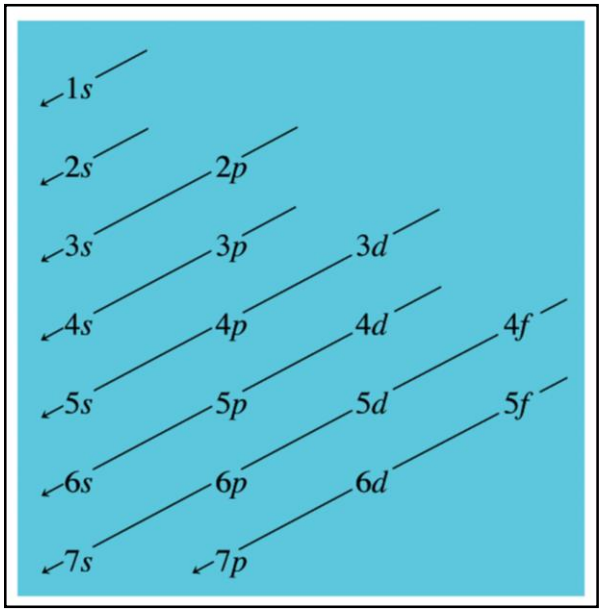


Polyelectronic atom

$$E_{ns} < E_{np} < E_{nd} < E_{nf}$$



(Usual) Energy Ordering

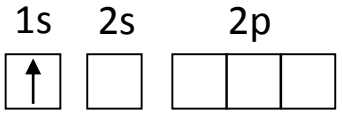


1.13. The Building-Up Principles

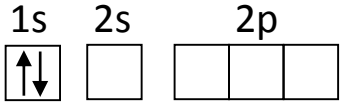
→ no more than 2 (e^-)s in any given orbital: no two electrons in an atom can have the same set of 4 quantum numbers (n, l, m_l, m_s)

→ parallel spin rather than paired

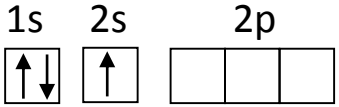
H: $1s^1$



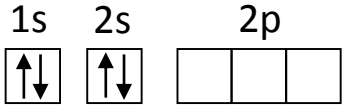
He: $1s^2$



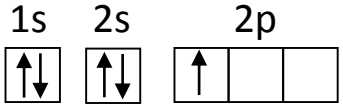
Li: $1s^2 2s^1$



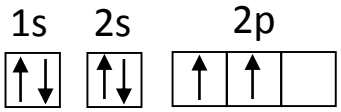
Be: $1s^2 2s^2$



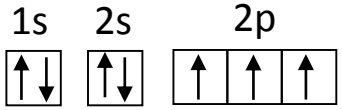
B: $1s^2 2s^2 2p^1$



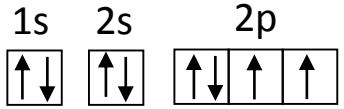
C: $1s^2 2s^2 2p^2$



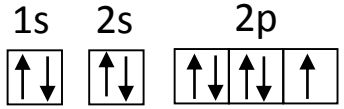
N: $1s^2 2s^2 2p^3$



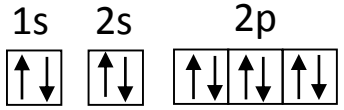
O: $1s^2 2s^2 2p^4$



F: $1s^2 2s^2 2p^5$



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



Shell: all of the orbitals with the same *n*.

Na $1s^2 2s^2 2p^6 3s^1$ [Ne]$3s^1$

Mg $1s^2 2s^2 2p^6 3s^2$ [Ne]$3s^2$

Al $1s^2 2s^2 2p^6 3s^2 3p^1$ [Ne]$3s^2 3p^1$

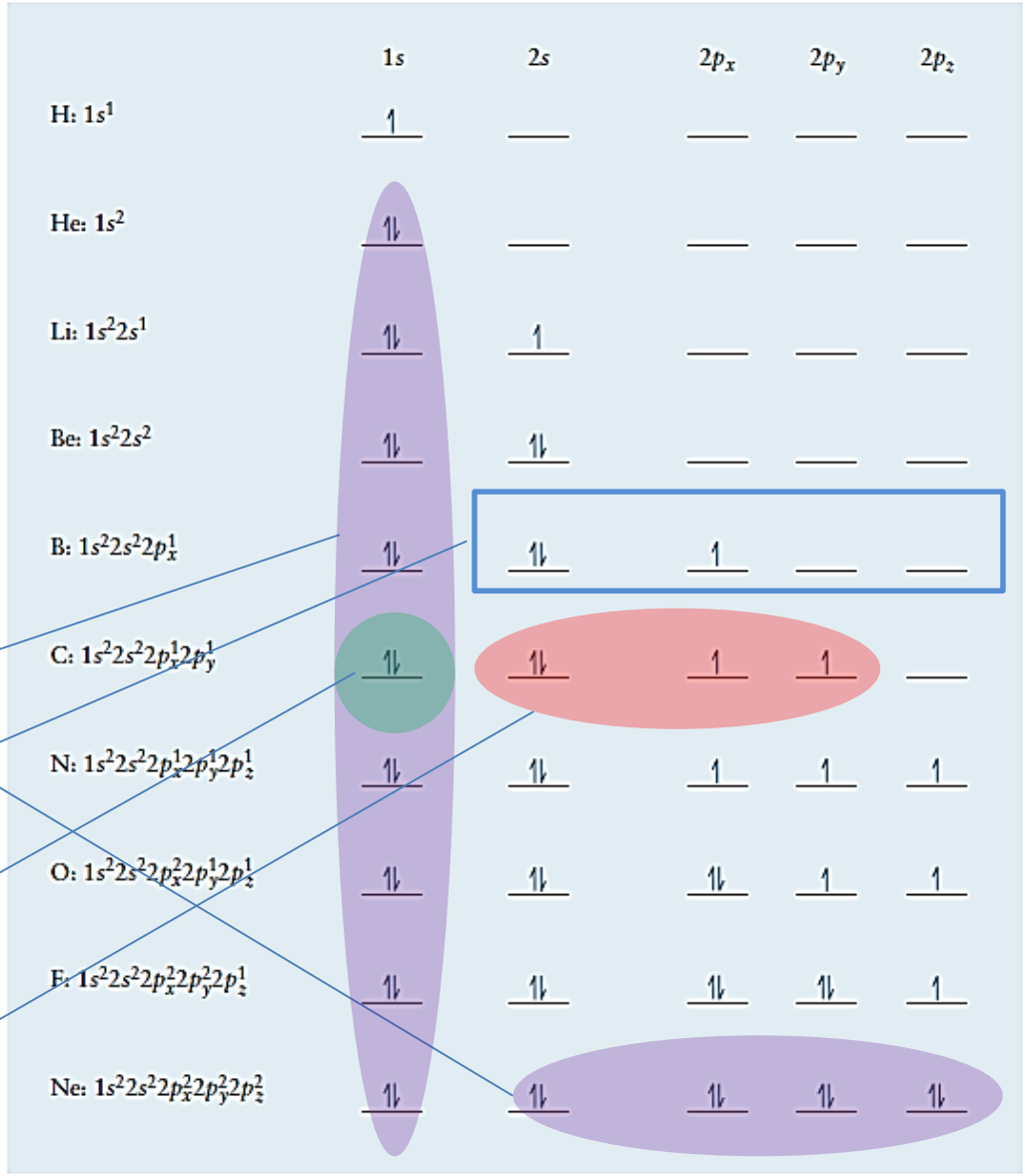
.

.

.

Ar $1s^2 2s^2 2p^6 3s^2 3p^6$ [Ne]$3s^2 3p^6$

Core electrons Valence electrons



closed shell

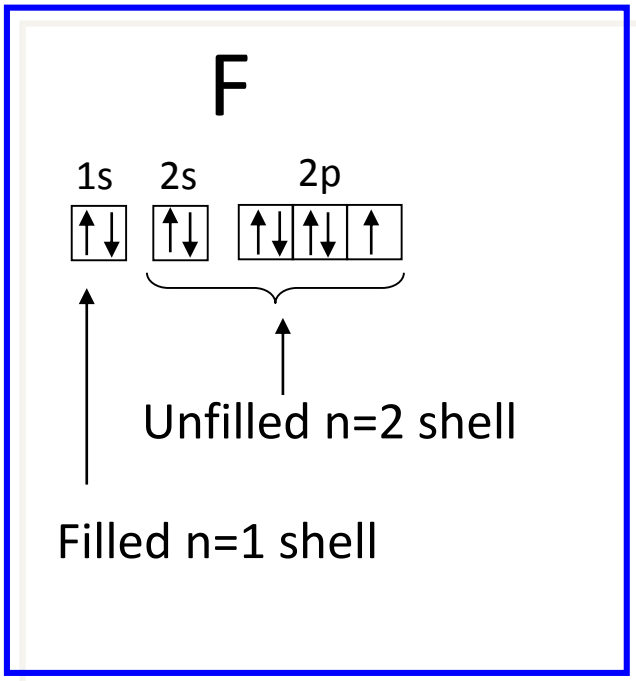
valence shell

core electron

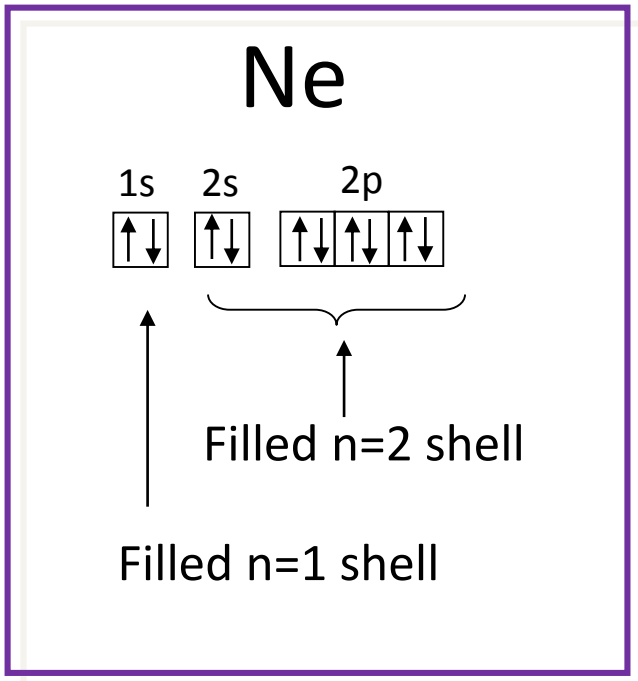
valence electron

1.14. Periodic trends in atomic properties

Atoms want to have “filled electronic shell”.

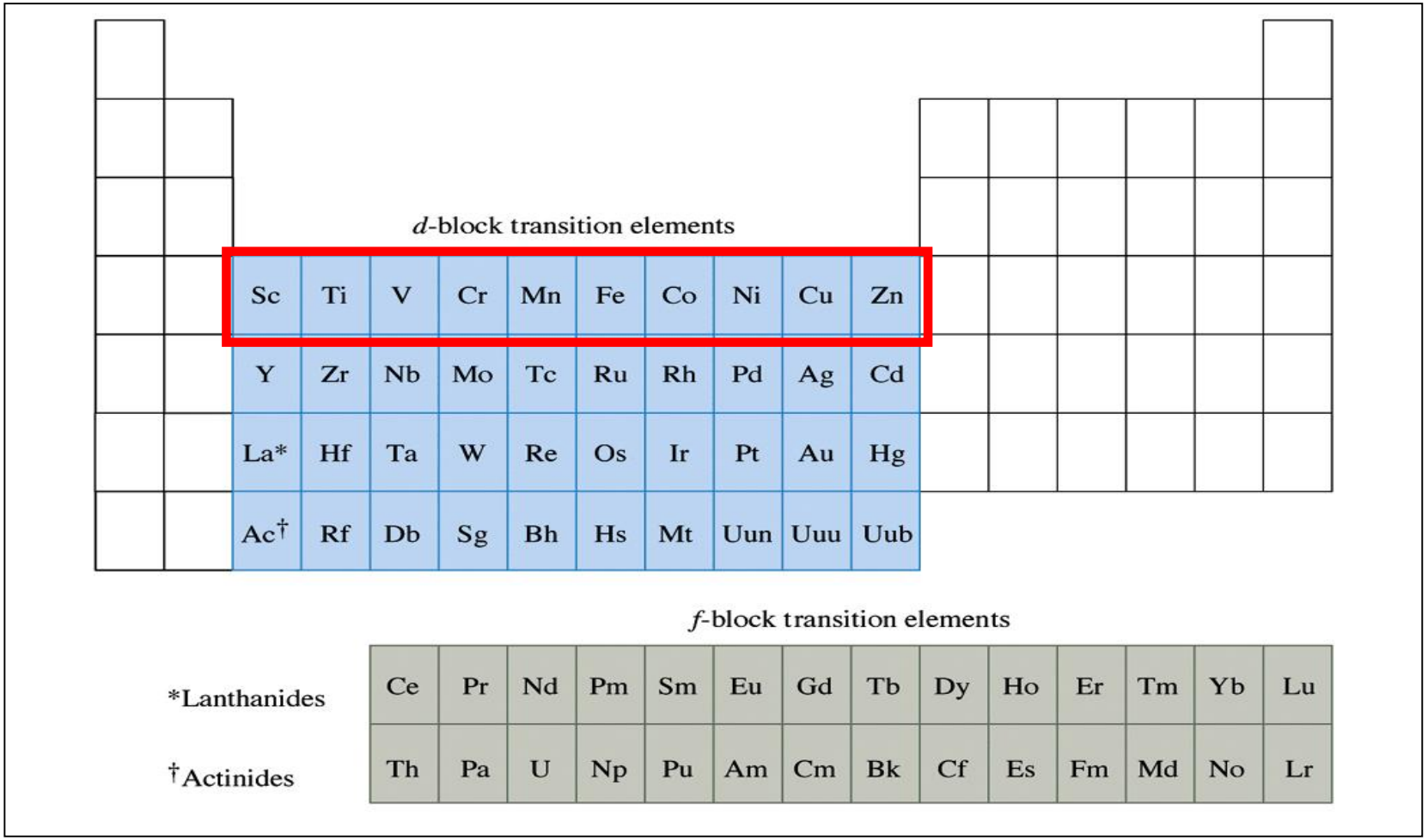


Chemically unstable,
Highly reactive



Chemically Stable,
does not react

Electron Configurations of the first-low transition metals



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

Ca: [Ar]$4s^2$

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

Ti: [Ar]$4s^2 3d^2$

V: [Ar]$4s^2 3d^3$

Cr: [Ar]$4s^2 3d^4$ ➡ actually [Ar]$4s^1 3d^5$

Mn: [Ar]$4s^2 3d^5$

Fe: [Ar]$4s^2 3d^6$

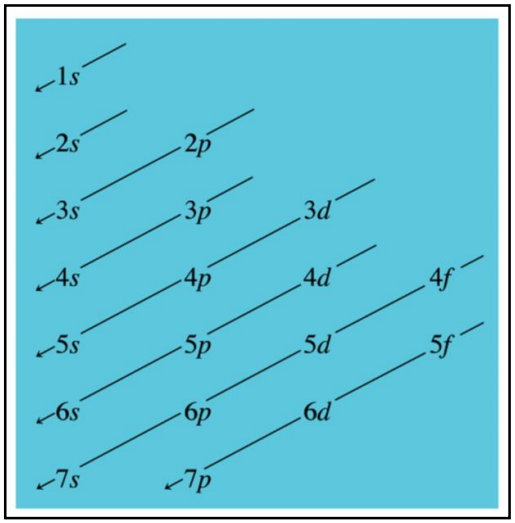
Transition elements favors half-filled and filled orbital!

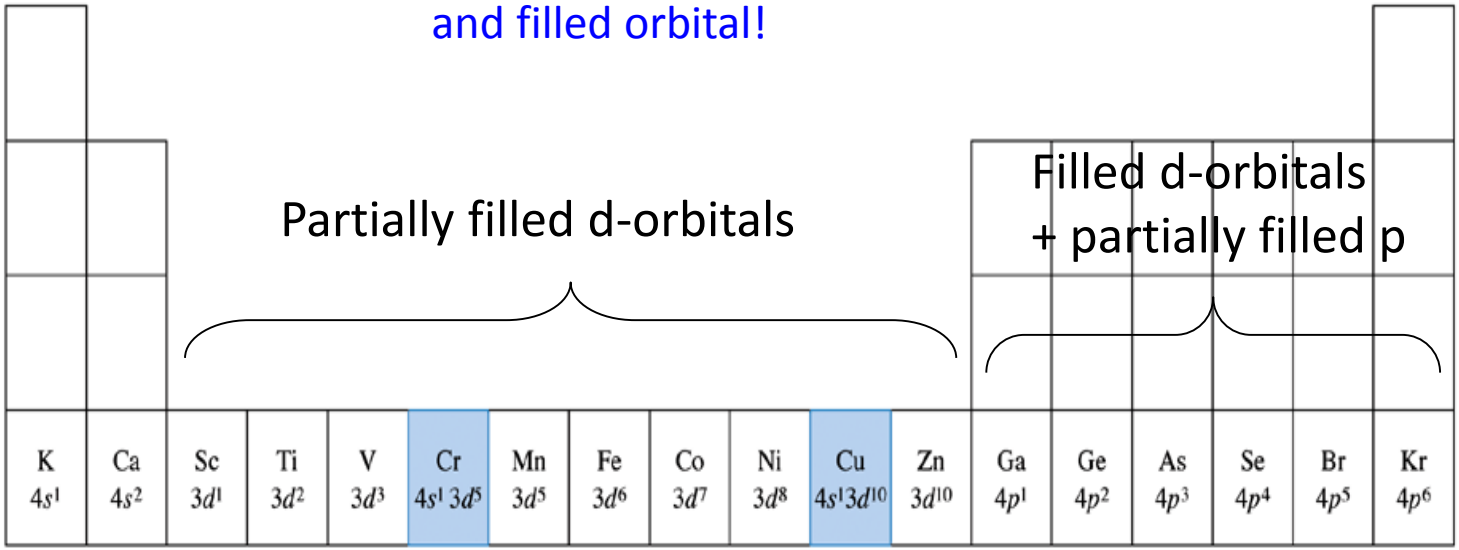
Co: [Ar]$4s^2 3d^7$

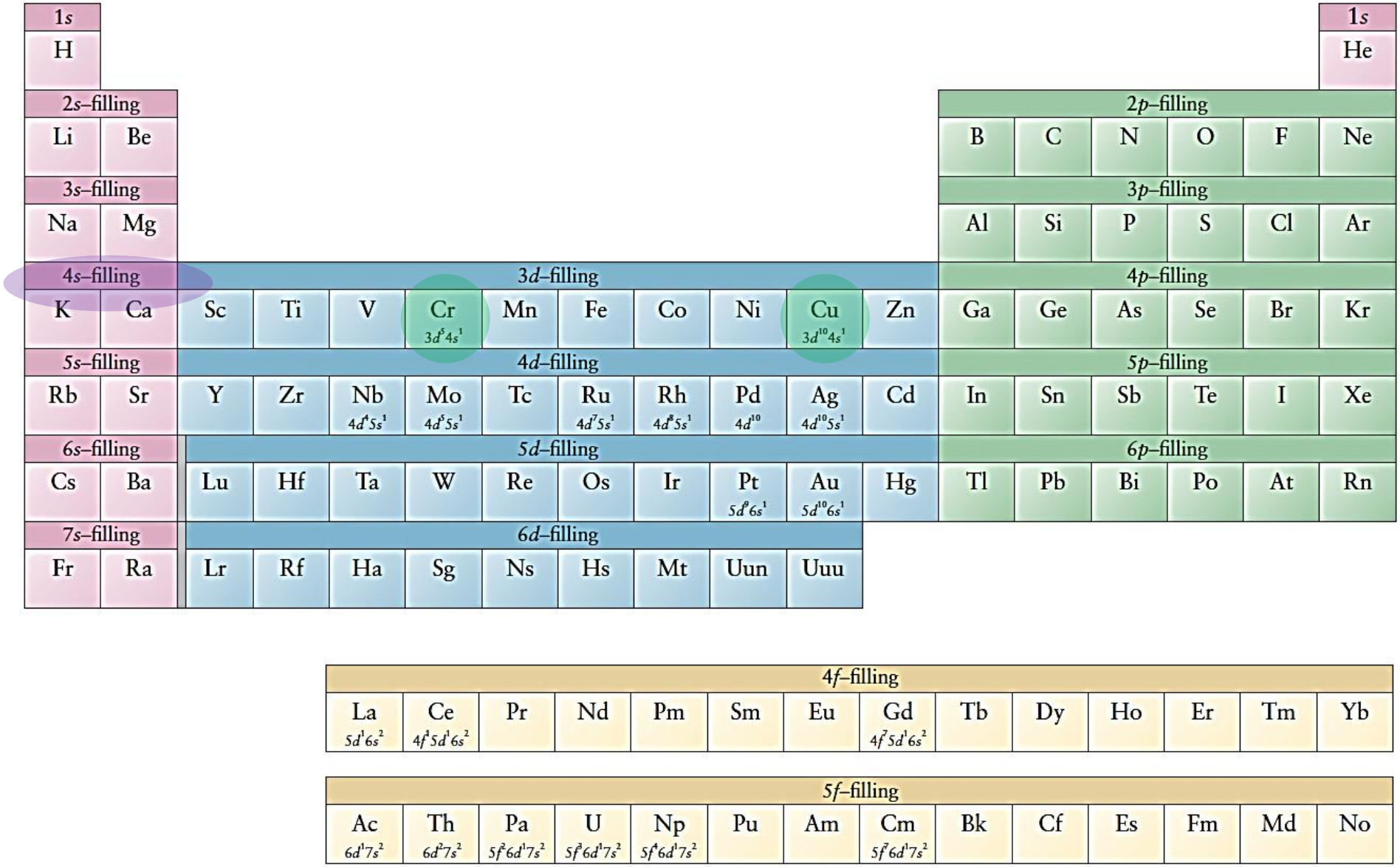
Ni: [Ar]$4s^2 3d^8$

Cu: [Ar]$4s^2 3d^9$ ➡ actually [Ar]$4s^1 3d^{10}$

Zn: [Ar]$4s^2 3d^{10}$

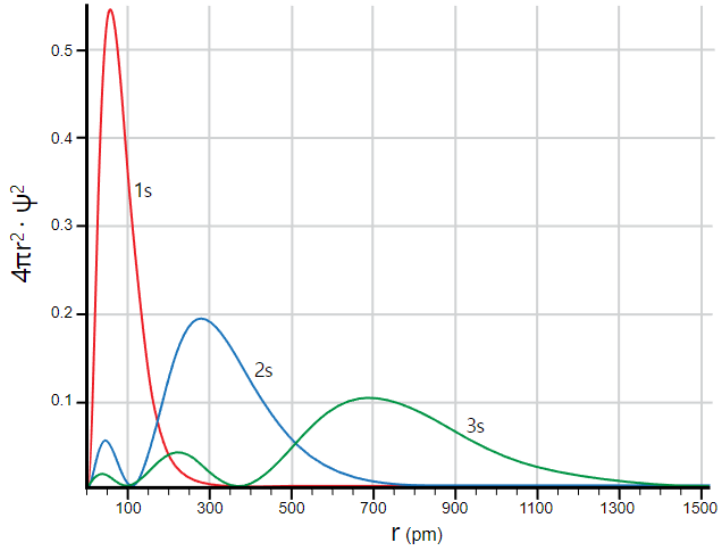


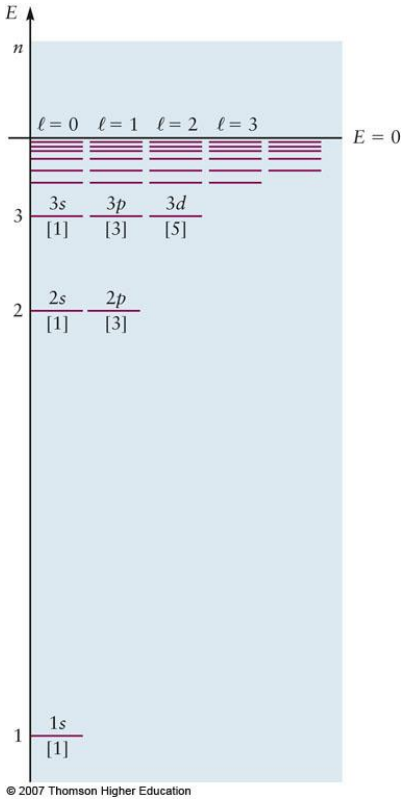



1s
H
1s
He
2s–filling
Li
Be
2p–filling
B
C
N
O
F
Ne
3s–filling
Na
Mg
3p–filling
Al
Si
P
S
Cl
Ar
4s–filling
3d–filling
4p–filling
K
Ca
Sc
Ti
V
Cr
3d5 4s1
Mn
Fe
Co
Ni
Cu
3d10 4s1
Zn
Ga
Ge
As
Se
Br
Kr
5s–filling
4d–filling
5p–filling
Rb
Sr
Y
Zr
Nb
4d4 5s1
Mo
4d5 5s1
Tc
Ru
4d7 5s1
Rh
4d8 5s1
Pd
4d10
Ag
4d10 5s1
Cd
In
Sn
Sb
Te
I
Xe
6s–filling
5d–filling
6p–filling
Cs
Ba
Lu
Hf
Ta
W
Re
Os
Ir
Pt
5d9 6s1
Au
5d10 6s1
Hg
Tl
Pb
Bi
Po
At
Rn
7s–filling
6d–filling
Fr
Ra
Lr
Rf
Ha
Sg
Ns
Hs
Mt
Uun
Uuu
4f–filling
La
5d1 6s2
Ce
4f1 5d1 6s2
Pr
Nd
Pm
Sm
Eu
Gd
4f7 5d1 6s2
Tb
Dy
Ho
Er
Tm
Yb
5f–filling
Ac
6d1 7s2
Th
6d2 7s2
Pa
5f2 6d1 7s2
U
5f3 6d1 7s2
Np
5f4 6d1 7s2
Pu
Am
Cm
5f7 6d1 7s2
Bk
Cf
Es
Fm
Md
No

Why is this happening?

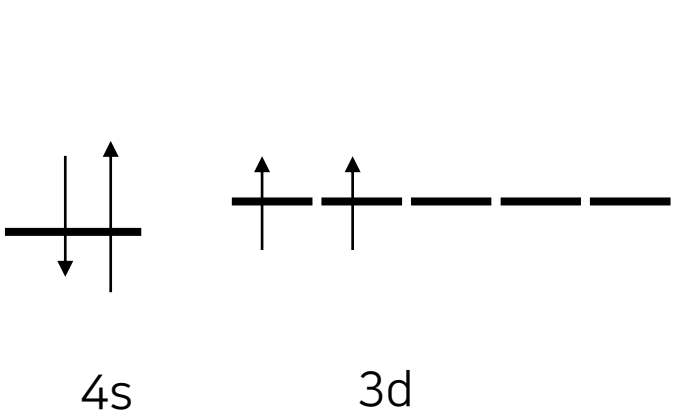
- 4s and 3d orbitals are close in energy.
- In fact, 'orbital picture' of the electrons is an approximation
- Electron-electron repulsion changes the energy level ordering.

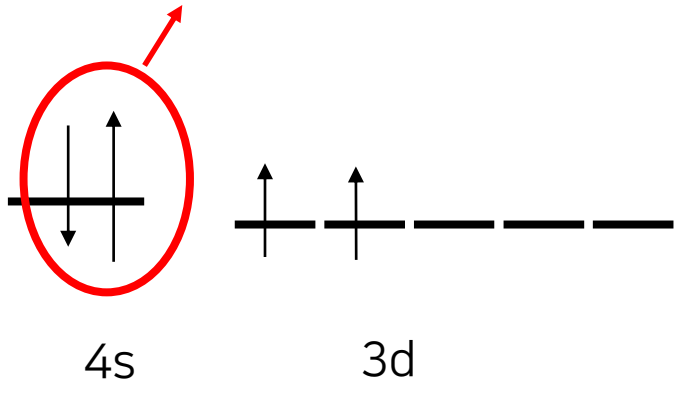




For transition metal,

- Ti: [Ar] $(4s)^2(3d)^2$ Ti^{3+} : $(4s)^0(3d)^1$





When adding electrons:

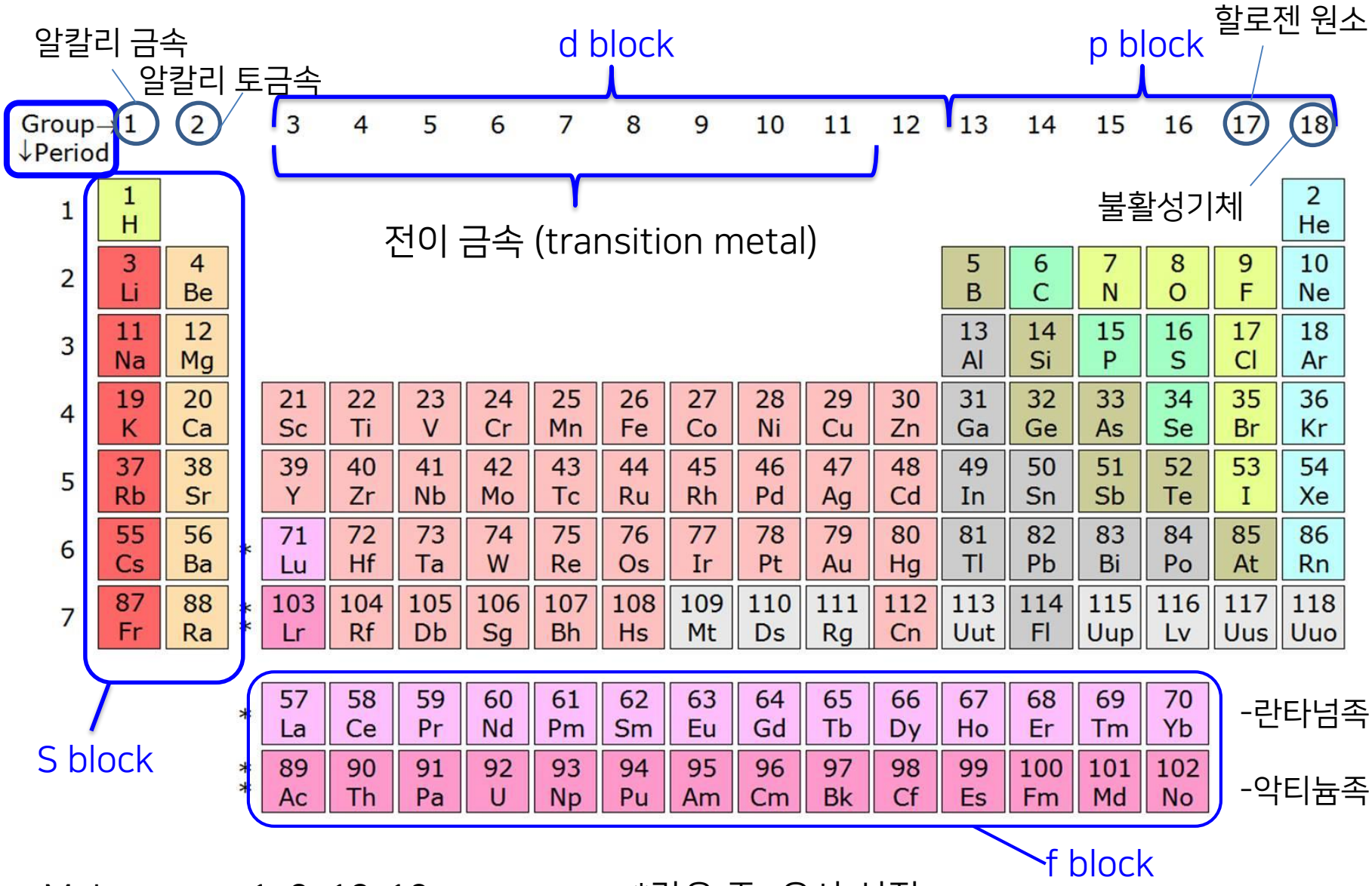
s-orbital fills first
d-orbital fills later

When losing electrons:

s-orbital first
d-orbital later

Electronic Configuration of Ions.

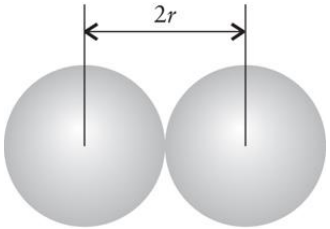
- Sc^{2+}: $[Ar](4s)^0(3d)^1$
- Ti^{2+} : $[Ar]\ (4s)^0(3d)^2$ $\quad Ti^{3+} : (4s)^0(3d)^1$
- V^{2+} : $[Ar](4s)^0(3d)^3$
- Cr^{2+} : $[Ar](4s)^0(3d)^4$
- Mn^{2+} : $[Ar](4s)^0(3d)^5$
- Fe^{2+} : $[Ar](4s)^0(3d)^6$
- Co^{2+} : $[Ar](4s)^0(3d)^7$
- Ni^{2+} : $[Ar](4s)^0(3d)^8$
- Cu^{2+} : $[Ar](4s)^0(3d)^9$
- Zn^{2+} : $[Ar](4s)^0(3d)^{10}$

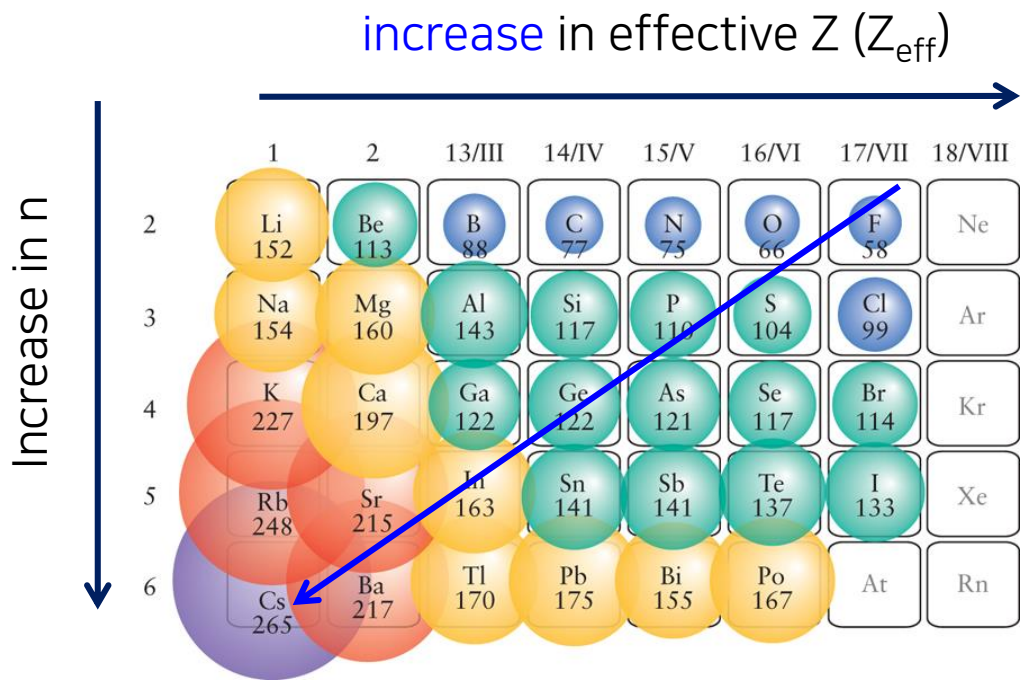


Main group: 1, 2, 13-18 group

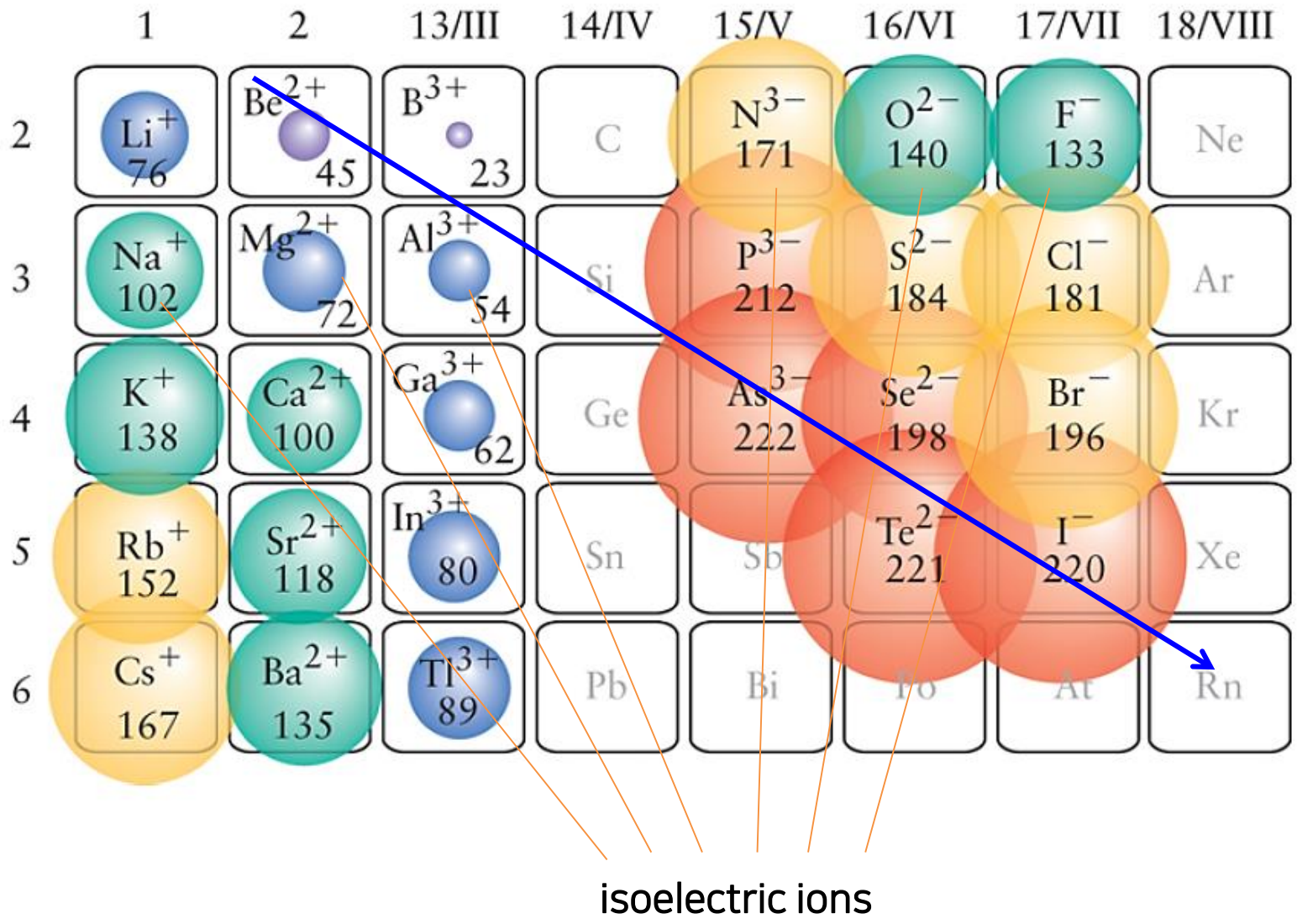
*같은 족=유사 성질

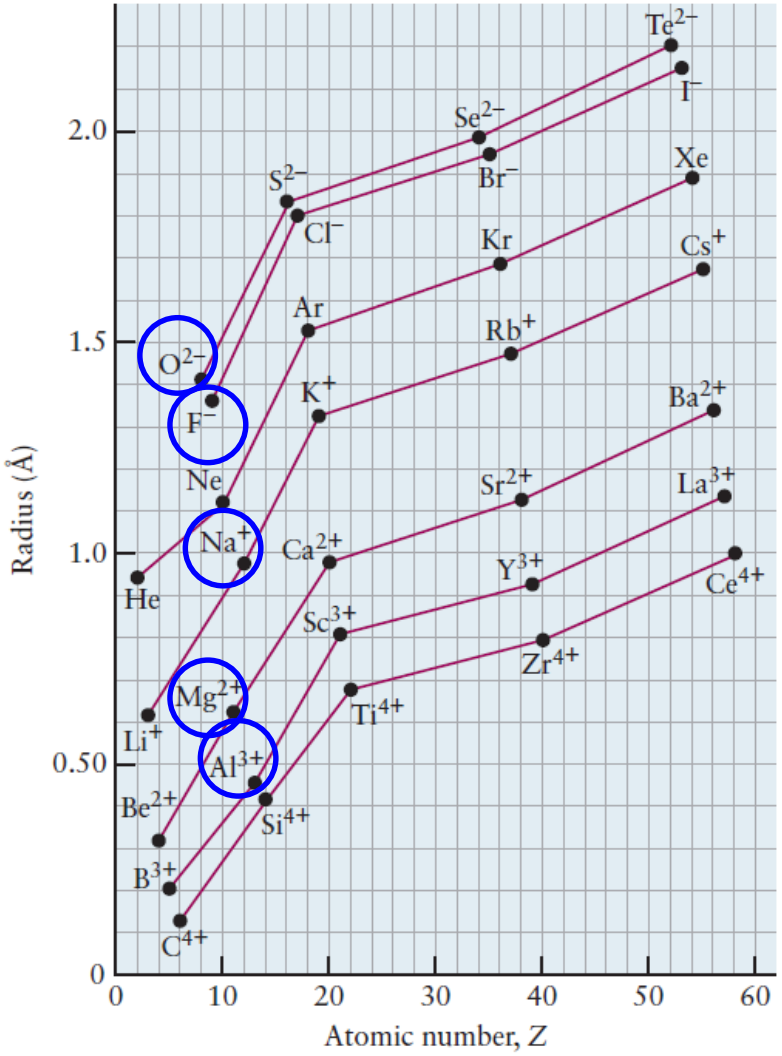
1.15. Atomic Radius





1.16. Ionic Radius





Sizes of atoms and their ions in pm

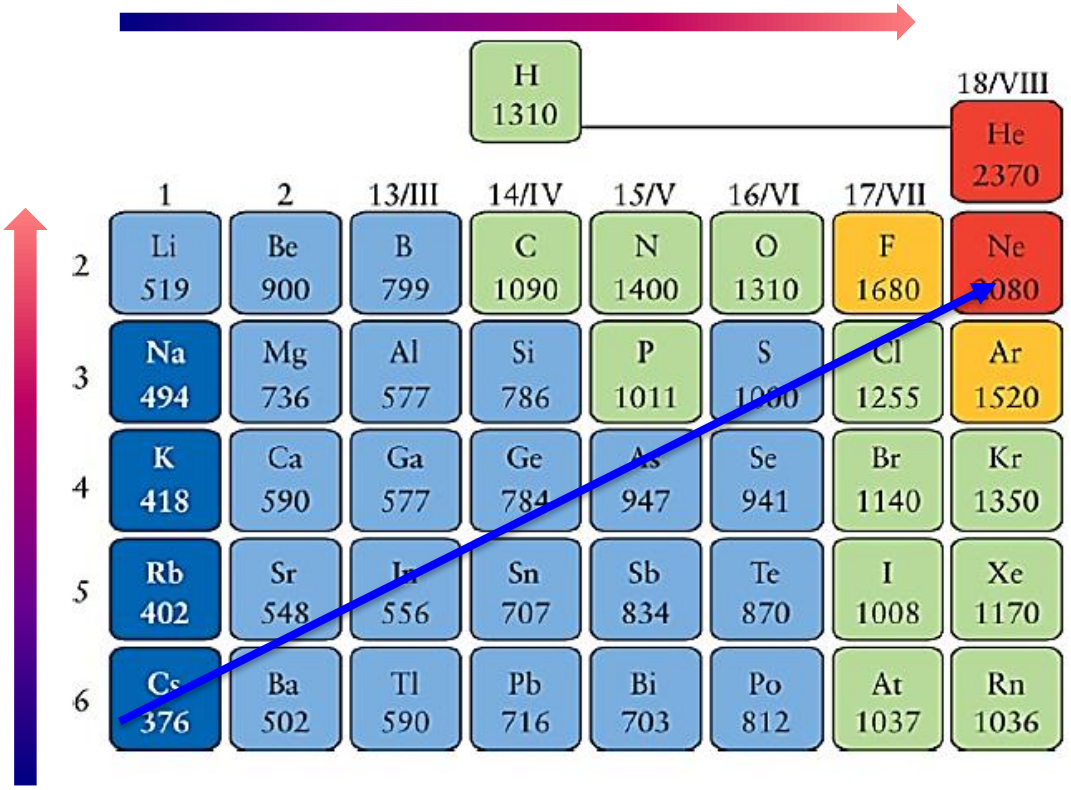
Group 1	Group 2	Group 13	Group 16	Group 17
Li^+ 90 / Li 134	Be^{2+} 59 / Be 90	B^{3+} 41 / B 82	O 73 / O^{2-} 126	F 71 / F^- 119
Na^+ 116 / Na 154	Mg^{2+} 86 / Mg 130	Al^{3+} 68 / Al 118	S 102 / S^{2-} 170	Cl 99 / Cl^- 167
K^+ 152 / K 196	Ca^{2+} 114 / Ca 174	Ga^{3+} 76 / Ga 126	Se 116 / Se^{2-} 184	Br 114 / Br^- 182
Rb^+ 166 / Rb 211	Sr^{2+} 132 / Sr 192	In^{3+} 94 / In 144	Te 135 / Te^{2-} 207	I 133 / I^- 206

1.17. Ionization Energy

The minimum E need to remove an electron from an atom in the gas phase.

$$X\,(g) \rightarrow X^{+}\,(g) + e^{-}\,(g) \qquad I = E(X^{+}) - E(X)$$

Typically, I_1 increases across a period



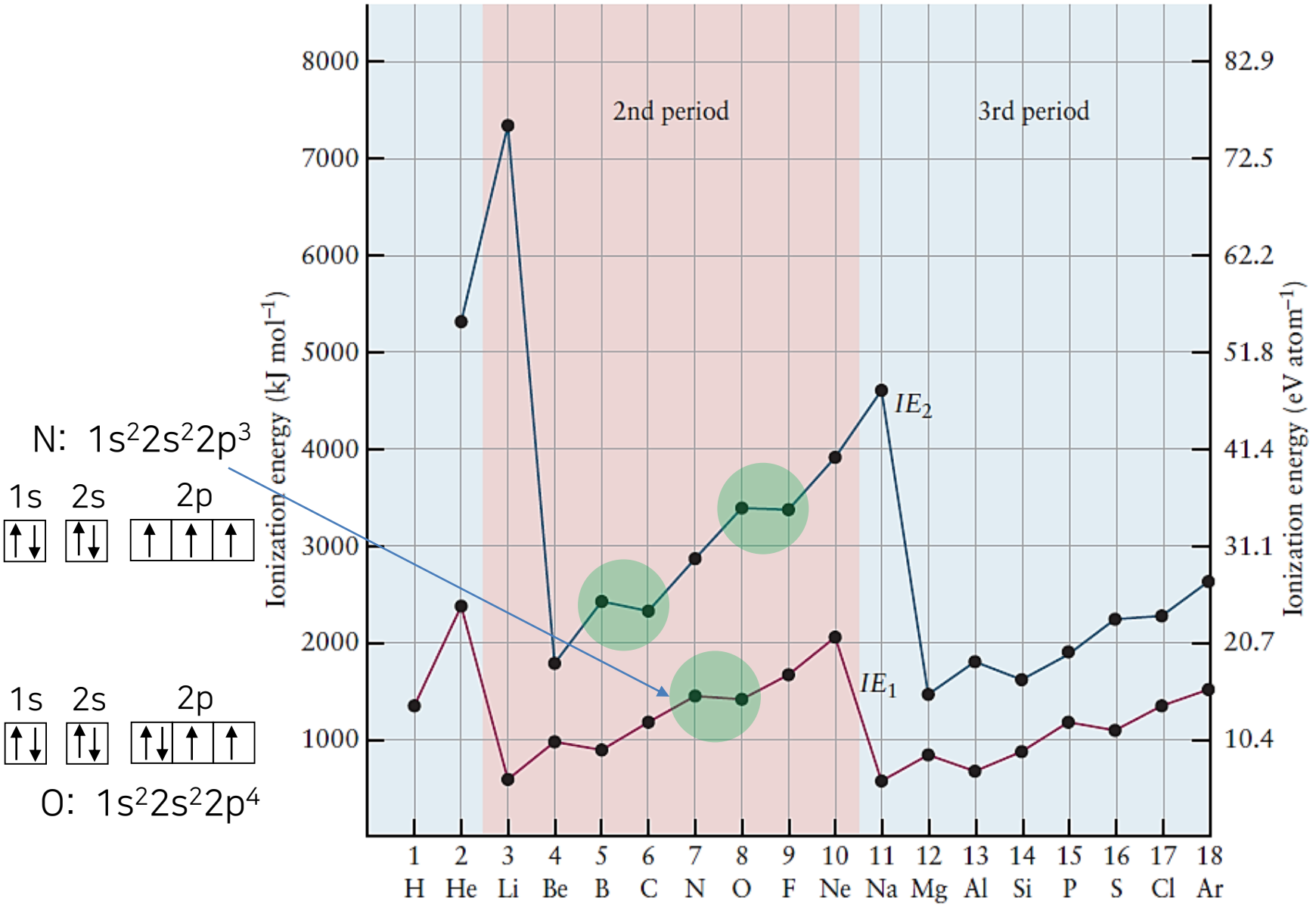
	1	2	13/III	14/IV	15/V	16/VI	17/VII	18/VIII
				H 1310				He 2370
2	Li 519	Be 900	B 799	C 1090	N 1400	O 1310	F 1680	Ne [illegible]
3	Na 494	Mg 736	Al 577	Si 786	P 1011	S [illegible]	Cl 1255	Ar 1520
4	K 418	Ca 590	Ga 577	Ge 784	As 947	Se 941	Br 1140	Kr 1350
5	Rb 402	Sr 548	In 556	Sn 707	Sb 834	Te 870	I 1008	Xe 1170
6	Cs 376	Ba 502	Tl 590	Pb 716	Bi 703	Po 812	At 1037	Rn 1036

Typically,

I_1 decreases down a group

Table: the 1st I (kJ mol^{-1})

Table: the 1st and 2nd I (kJ mol^{-1})



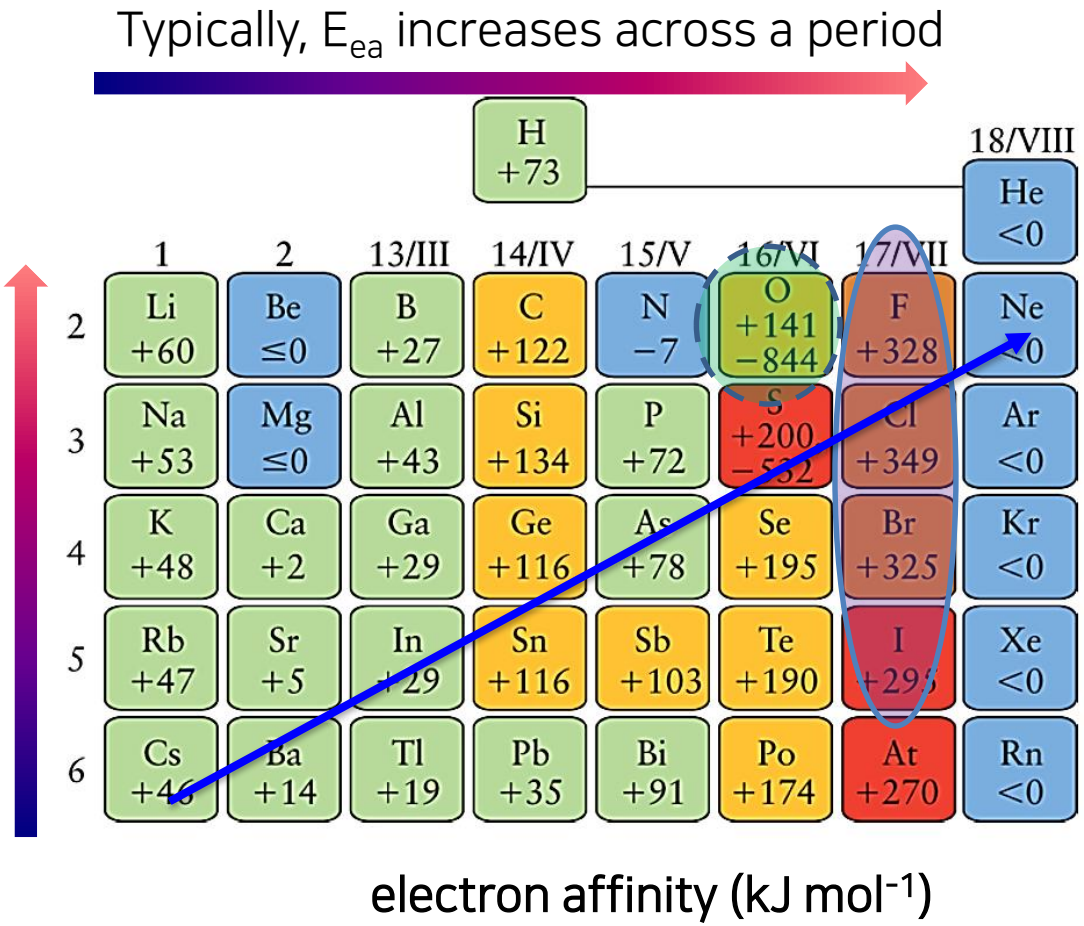
1.18. Electron Affinity

The E released when an electron is added to an atom in the gas phase.

$X(g) + e^-(g) \rightarrow X^-(g)$ $\qquad E_{ea}(X) = E(X) - E(X^-)$

H, Ne··· → Do not want to have one more electron

O, F, Cl,··· → Want to have more electrons

Typically,

E_{ea} decreases down a group