Chemistry, The Central Science, 10th edition

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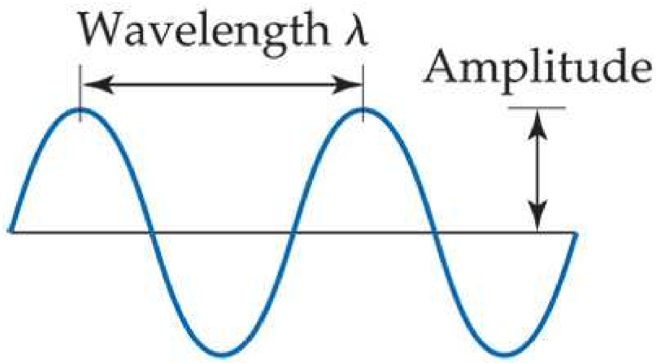
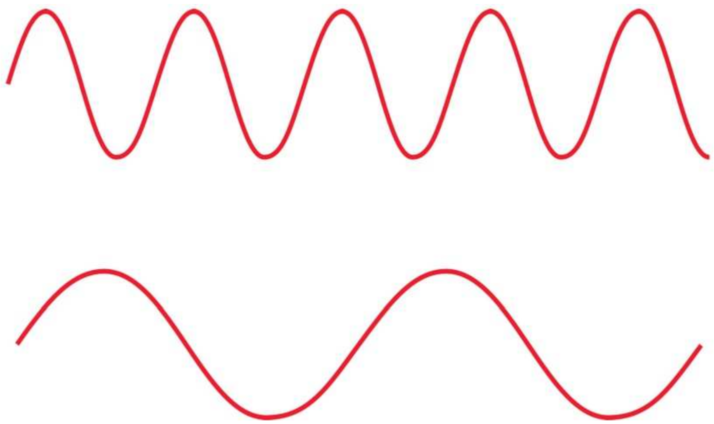
Chapter 6
Electronic Structure of Atoms

Waves



- To understand the electronic structure of atoms, one must understand the nature of electromagnetic radiation.
- The distance between corresponding points on adjacent waves is the wavelength (λ).

Waves



$E \ \alpha \ \nu$

where α means 'is prapotionate to '

Therefore

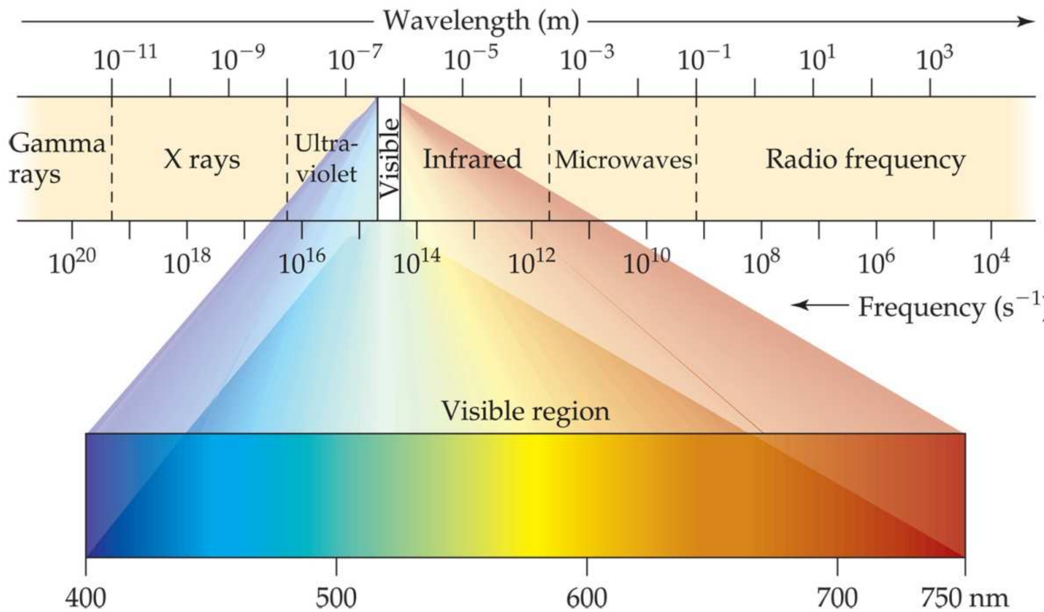
$\nu\lambda = \text{constant}$

this constant is the speed of light

- The number of waves passing a given point per unit of time is the frequency (ν).
- For waves traveling at the same velocity, the longer the wavelength, the smaller the frequency.

Electronic Structure of Atoms

Electromagnetic Radiation



- All electromagnetic radiation travels at the same velocity: the speed of light (c), 3.00 $\times 10^8$ m/s.
- Therefore,

$$c = \lambda\nu$$

- Isaac Newton believed that light consists of beam of particles.
- Thomas Young shoes that light had the ability to diffract which is a wave property.
- Einstein postulated that light had both wave and particulate properties.

The Nature of Energy

- When solids are heated they emit radiation.
- Red hot objects give out red light and at higher temperatures white light is emitted.
- Max Planck explained this phenomenon by making an assumption that in light energy comes in packets called quanta.

The Nature of Energy

- He concluded that energy is proportional to frequency:

$$E \ \alpha \ \nu$$

$$E = h\nu$$

where *h* is Planck's constant,

$$= 6.63 \times 10^{-34} \text{ J-s.}$$

- The units for planks constant:

$$E = h\nu$$

$$h = \frac{E}{\nu} = \frac{J}{/s} = J.s$$

The Nature of Energy

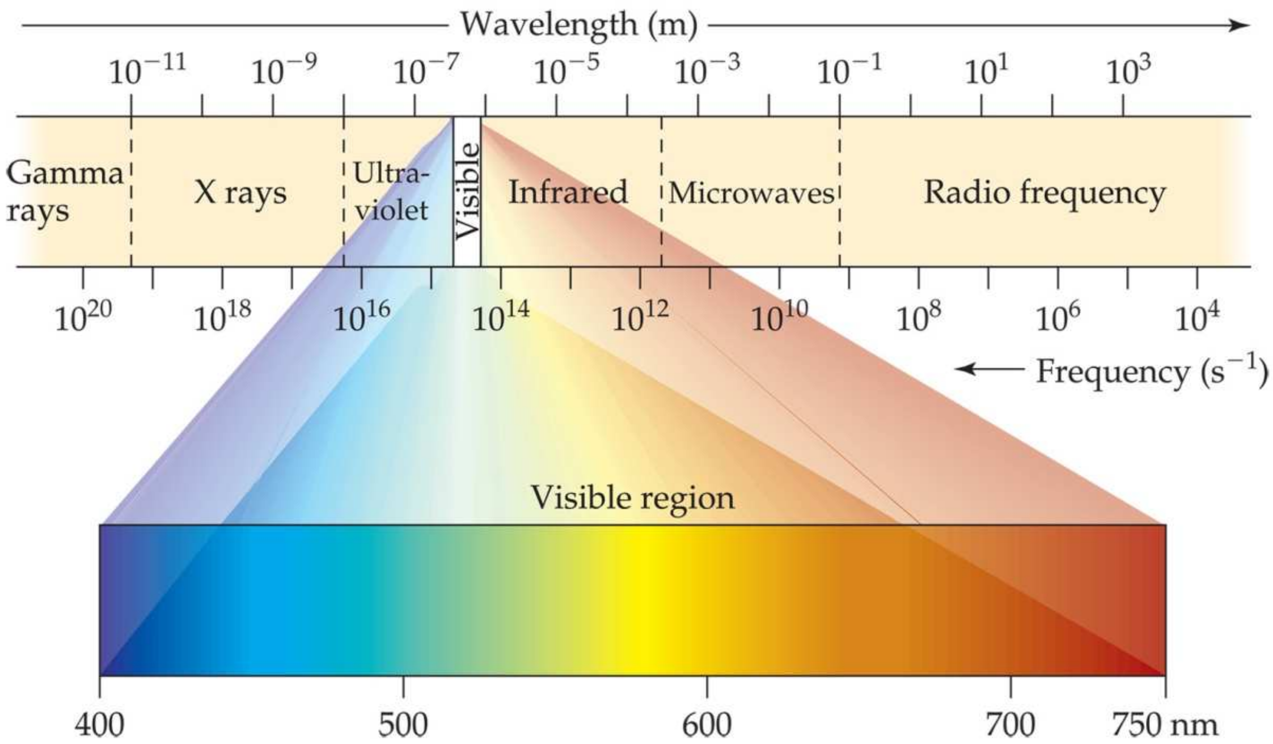
- Therefore, if one knows the wavelength of light, one can calculate the energy in one photon, or packet, of that light:

$$c = \lambda \nu$$

Gives us the frequency

$$E = h\nu$$

Gives us the energy



You can calculate the energy of the photon of every kind of wave as long as you know that

$$h = 6.63 \times 10^{-34} \text{ J-s.and}$$

$$c = 3.0 \times 10^{8} \text{ m/s}$$

Electronic Structure of Atoms

- On Monday we were calculating the energy in photon from the visible light that had the wavelength of 400nm.
- We know that $\lambda \nu = c$

$$\text{So } \nu = c/\lambda = \frac{3\text{x}10^{8}\text{m/s}}{400\text{nm x } \dfrac{1\text{m}}{10^{9}\text{nm}}}$$

$$= 7.5 \text{ x } 10^{14} \text{ /s}$$

$$E = h\nu = 6.63 \times 10^{-34} \text{ J-s x } 7.5 \text{ x } 10^{14} \text{ /s}$$

$$= 4.9725 \text{x} 10^{-19} \text{ J}$$

$$= 5 \text{ x} 10^{-19} \text{ J}$$

Electronic Structure of Atoms

- The Si unit for frequency is Hz or Hertz.
- It is the same as / s.

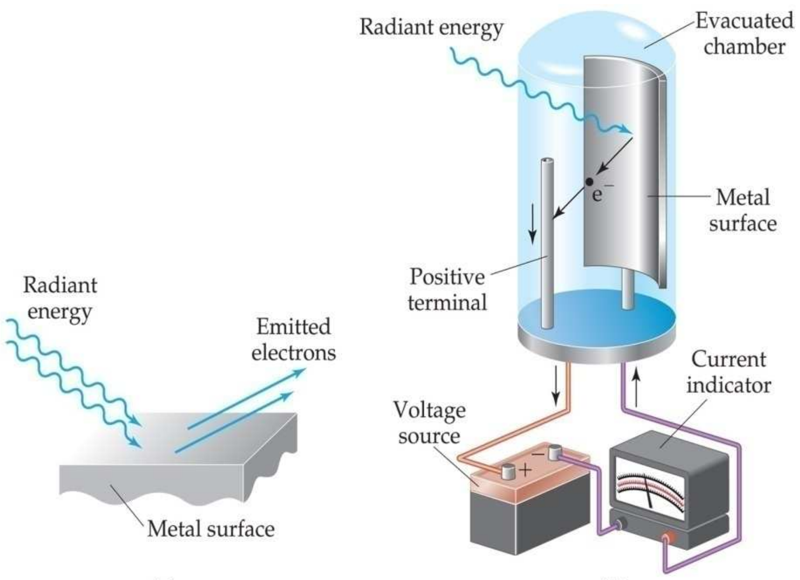
- Max Plank in 1900 postulated that energy can either be released or absorbed by atoms in discreet chunks

  This is related to the frequency by the equation

$$E = h\nu$$

- He also said that the atom can absorb energy in multiples of hv only.
- The smallest value of hv is called a quanta of energy

Photoelectric effect



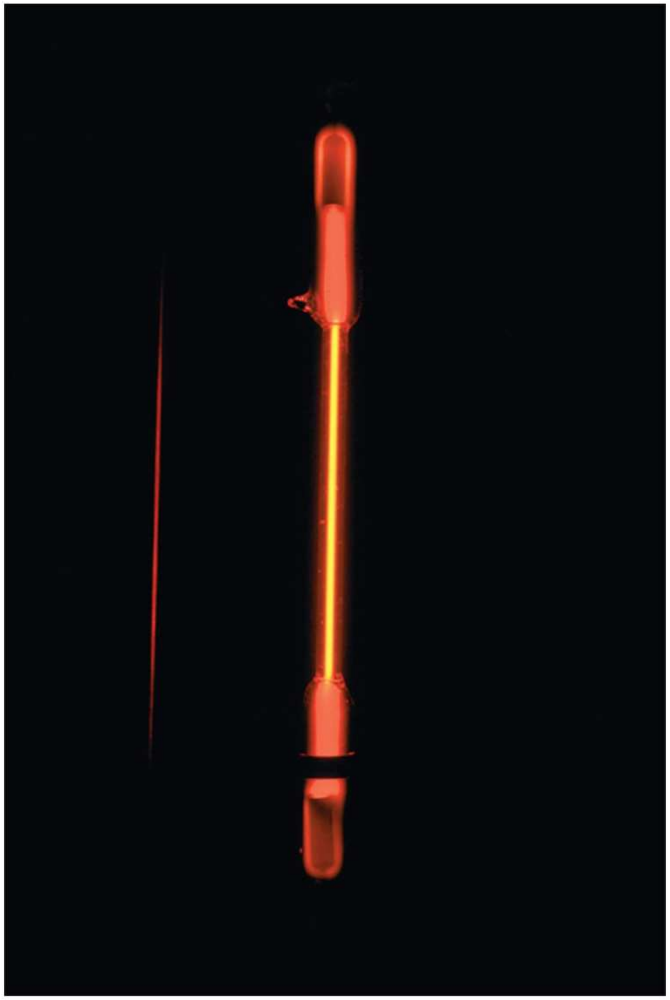
- Einstein used Planck's theory to explain photoelectric effect.
- He said that the radiant energy striking the metal surface is behaving like a stream of tiny energy packets.

- These packets are photons and behave like tiny particles
- Einstein also deduced that each photon must have the energy equivalent to the Planck's constant times the frequency of that light.

$$E = h\nu$$

- Under the right conditions, a photon can strike the metal surface and be absorbed. Then it transfers it’s energy to an electron in the metal. Once enough energy is absorbed, the electron is emitted from the metal
- The minimum energy to release the electron is different for different metals

The Nature of Energy



Another mystery involved the emission spectra observed from energy emitted by atoms and molecules.

Electronic Structure of Atoms

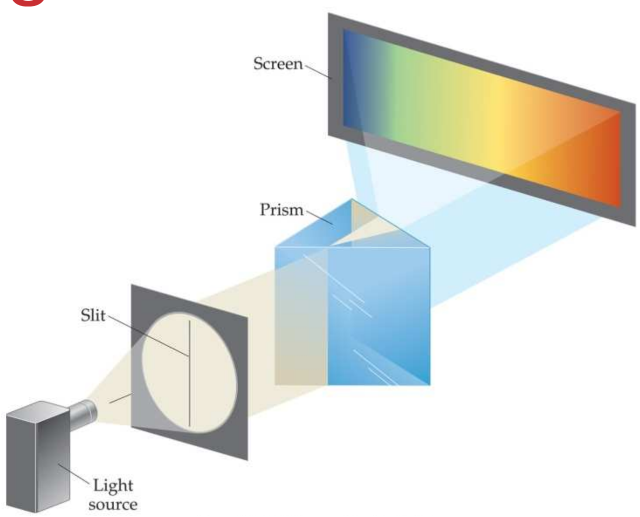
- Before we go any further …..

Monochromatic light:

The radiation has one single wavelength.

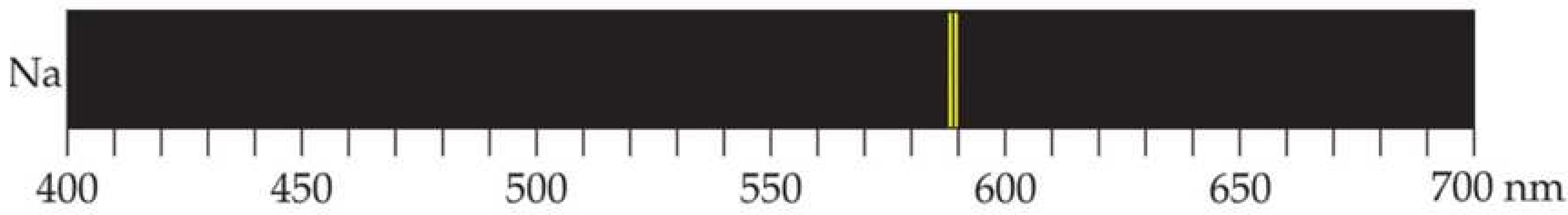
Spectrum:

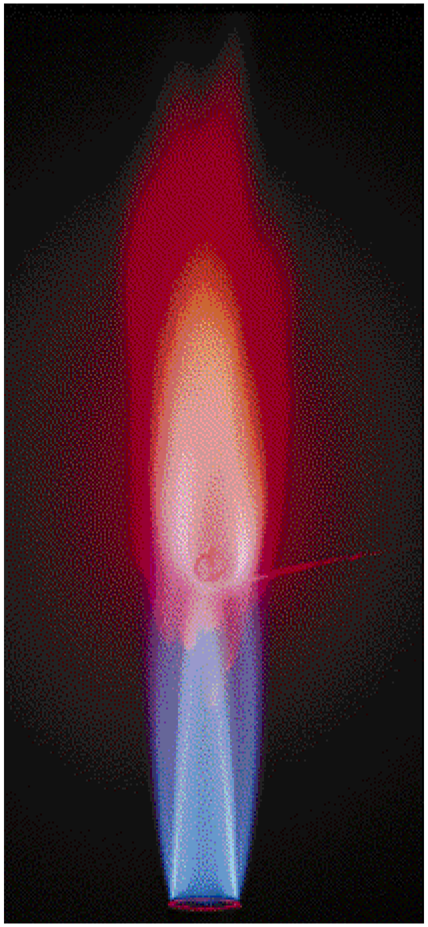
when the radiation has many wavelengths.

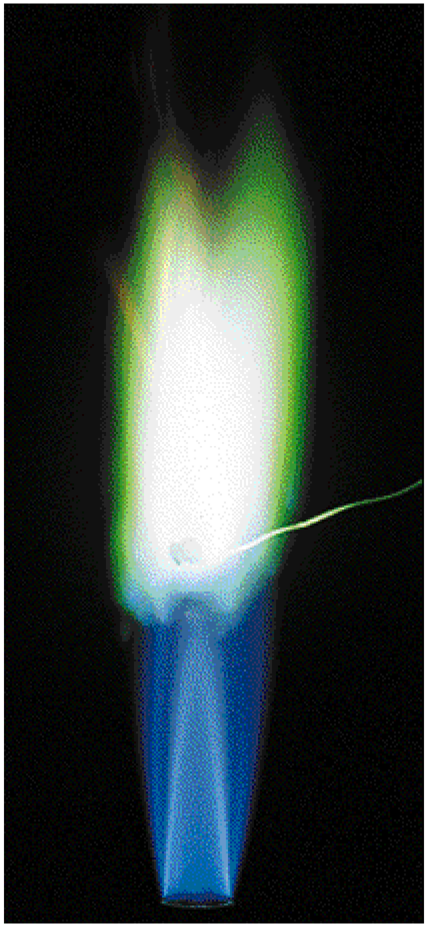


- When high voltage is applied to a tube containing a gas at reduced pressure, the gas emits different colors of light

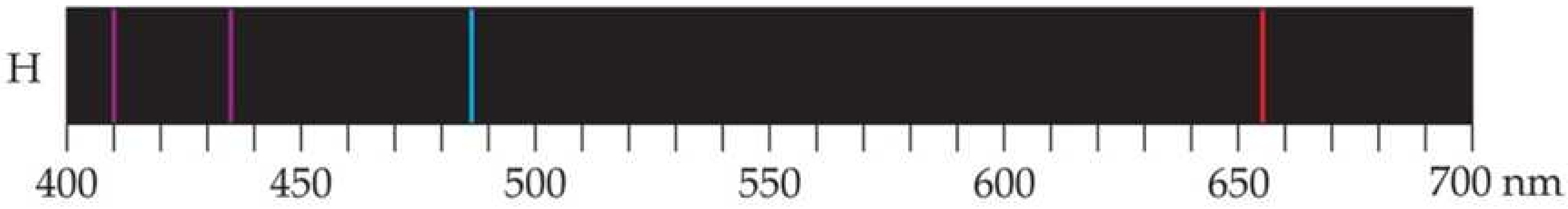
Na vapor emits a light which when passed through a prism gives a single line at 590 nm wavelength in the yellow range.







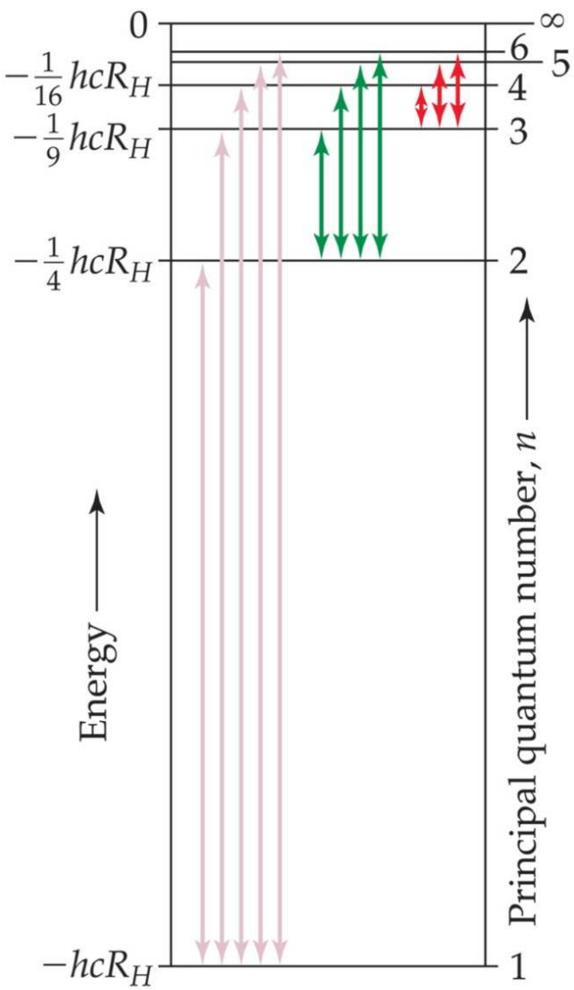
- For Hydrogen the spectrum has four distinct lines



Bohr’s Model:

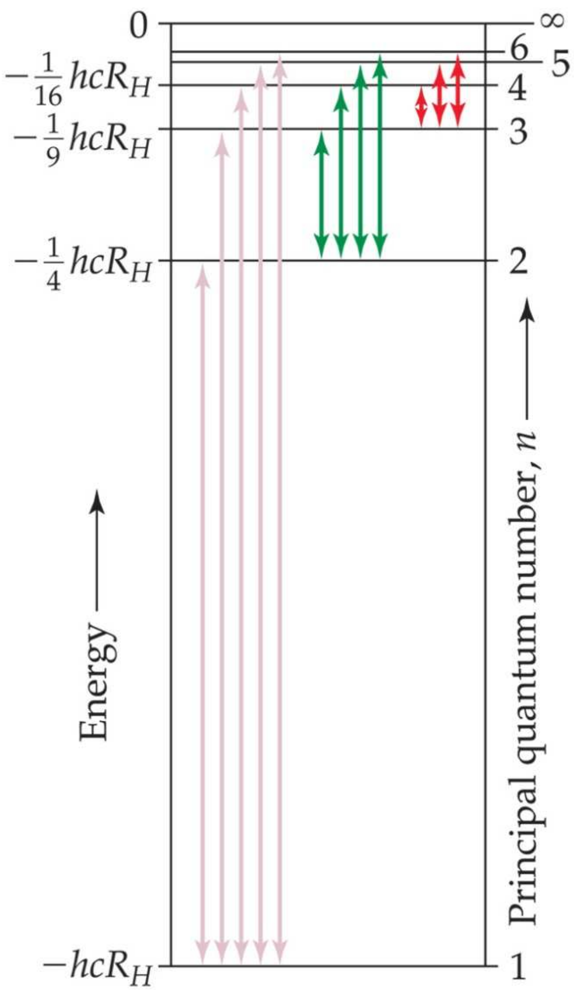
According to classical physics, electrically charged particles that are moving in a circular orbit should continuously be losing energy and collapse into the nucleus but this does not seem to happen. Bohr adapted Planck’s idea that energies are quantized.

The Nature of Energy



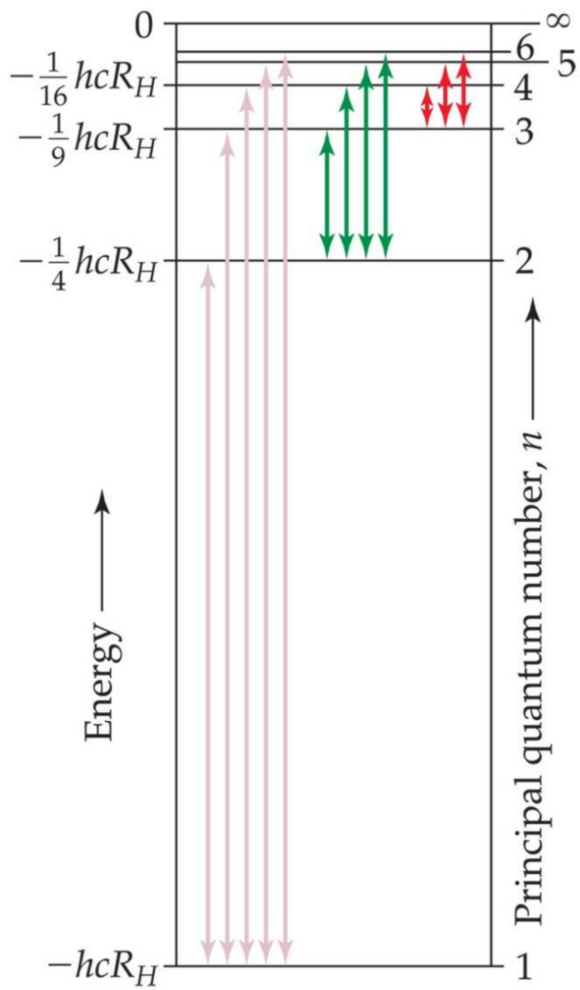
- Niels Bohr adopted Planck's assumption and explained these phenomena in this way:
  1. Electrons in an atom can only occupy certain orbits (corresponding to certain energies).

The Nature of Energy



- Niels Bohr adopted Planck's assumption and explained these phenomena in this way:
  2. Electrons in permitted orbits have specific, "allowed" energies; these energies will not be radiated from the atom.

The Nature of Energy



- Niels Bohr adopted Planck's assumption and explained these phenomena in this way:
  3. Energy is only absorbed or emitted in such a way as to move an electron from one "allowed" energy state to another; the energy is defined by

$$E = h\nu$$

Bohr calculated the energy corresponding to each each orbit

$$E = (-hcR_H)\left(\frac{1}{n^2}\right)$$

Where h = planck's Constant
c = speed of Light
R_H = Rydberg's Constant

And the product of all the three is 2.18×10^{-18} J

Electronic Structure of Atoms

n = principal quantum number and can have the values $1 - \infty$

Bohr assumed that electons could jump from one energy state to another

a radient energy is emitted when electron jumps from a higher to a lower energy state

$$\Delta E = E_f - E_i = E_{photon} = h\nu$$

Electronic Structure of Atoms

$$\Delta E = h\nu = \frac{hc}{\lambda} = (-2.18 \times 10^{-18} J)\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

So if $n_i = 3$ and $n_f = 1$

$$\Delta E = (-2.18 \times 10^{-18} J)\left(\frac{1}{1^2} - \frac{1}{3^2}\right) = -2.18 \times 10^{-18} J \left(1 - \frac{1}{9}\right)$$

$$= -2.18 \times 10^{-18} J \left(\frac{8}{9}\right)$$

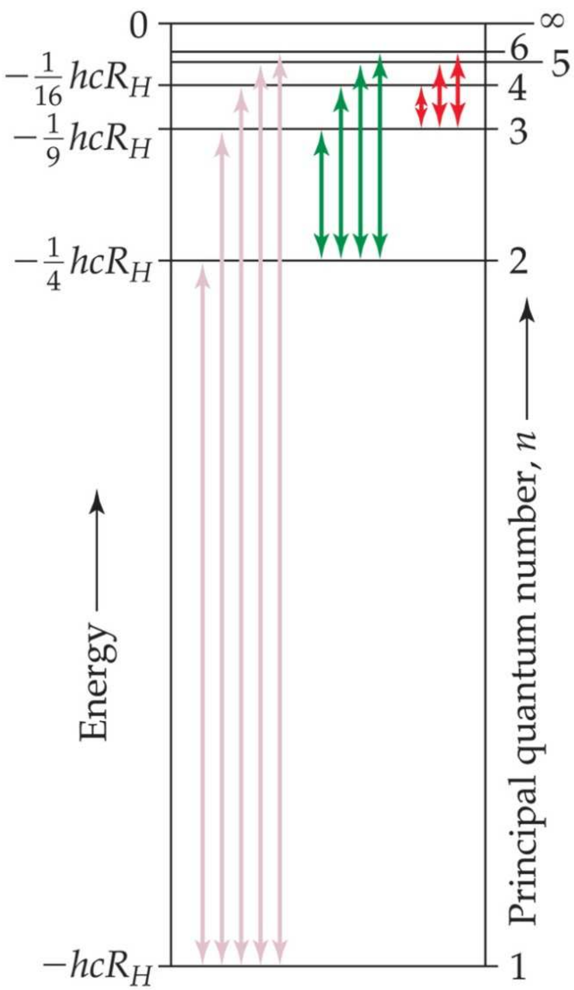
$$= -1.94 \times 10^{-18} J$$

Electronic Structure of Atoms

$$\lambda = \frac{c}{\nu} = \frac{hc}{\Delta E} = \frac{(6.63 \times 10^{-34}\ \text{J-s})(3 \times 10^{8}\ \text{m/s})}{1.94 \times 10^{-18}\ \text{J}}$$

$$= 1.03 \times 10^{-7}\ \text{m}.$$

The Nature of Energy



The energy absorbed or emitted from the process of electron promotion or demotion can be calculated by the equation:

$$E = -hcR_H\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

where R_H is the Rydberg constant, 2.18×10^{-18} J, and n_i and n_f are the initial and final energy levels of the electron.

imitations of Bohr’s model

- It applied only to hydrogen atom

The Wave Nature of Matter

- Louis de Broglie posited that if light can have material properties, matter should exhibit wave properties.
- He demonstrated that the relationship between mass and wavelength was

$$\lambda = \frac{h}{mv}$$

The Uncertainty Principle

- Heisenberg showed that the more precisely the momentum of a particle is known, the less precisely is its position known:

$$(\Delta x)\,(\Delta mv) \geq \frac{h}{4\pi}$$

- In many cases, our uncertainty of the whereabouts of an electron is greater than the size of the atom itself!

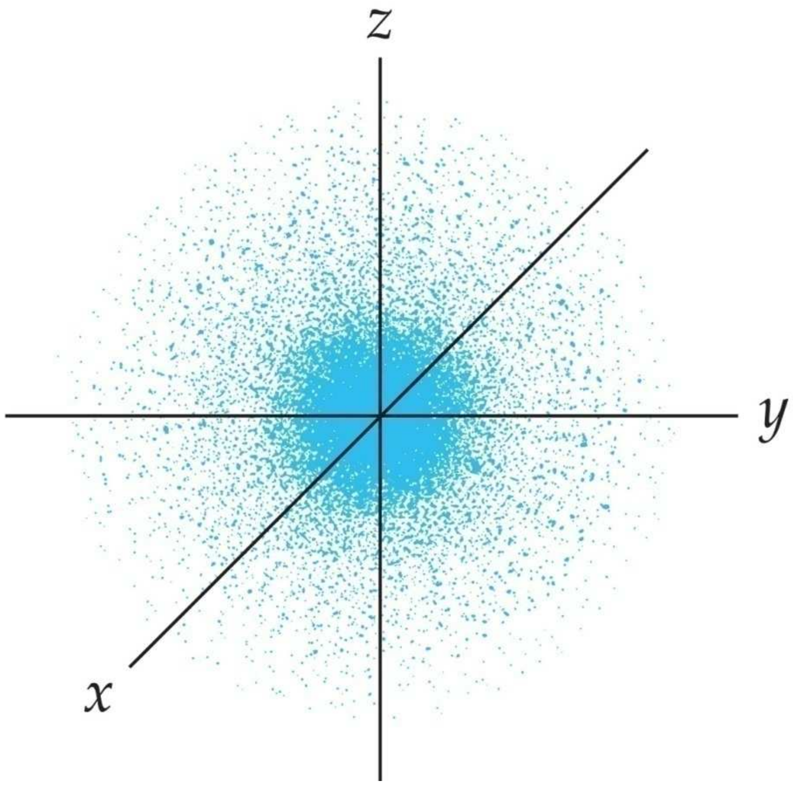
- Nora Tobin has completed her orientation and Student Employment paperwork and is now ready to tutor Chem 115 and 116. Students can register for tutoring in the Academic Support Programs Office (CC-1-1300).

- With DeBrogie’s correlation of electron and wave and Heisenberg's uncertainty principal the wave nature of electron was accepted

- Now the location of electrons is described in probability and the energy of an electron can be calculated.

Quantum Mechanics

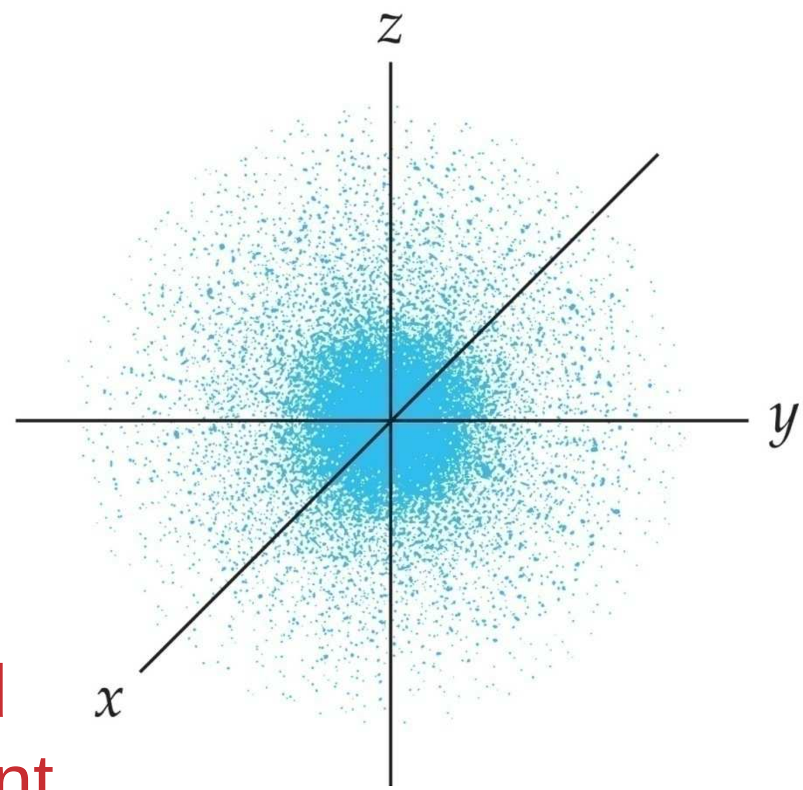
- Erwin Schrödinger developed a mathematical treatment into which both the wave and particle nature of matter could be incorporated.
- It is known as quantum mechanics.



Electronic Structure of Atoms

Quantum Mechanics

- The wave function is designated with a lower case Greek *psi* (ψ).
- The square of the wave equation, ψ^2, gives a probability density map of where an electron has a certain statistical likelihood of being at any given instant in time.



Quantum Numbers

- Solving the wave equation gives a set of wave functions, or **orbitals**, and their corresponding energies.
- Each orbital describes a spatial distribution of electron density.
- An orbital is described by a set of three **quantum numbers**.

- Remember now the term being used will be orbital, not orbit the term used by Bohr
- Bore had only the quantum number n for orbit.
- The quantum mechanics uses four quantum numbers.

Principal Quantum Number, *n*

- The principal quantum number, *n*, describes the energy level on which the orbital resides.
- The values of *n* are integers 1, 2,3 and so forth

Azimuthal Quantum Number, l

- This quantum number defines the shape of the orbital.
- Allowed values of l are integers ranging from 0 to $n - 1$.
- We use letter designations to communicate the different values of l and, therefore, the shapes and types of orbitals.

Azimuthal Quantum Number, *l*

Value of *l*	0	1	2	3
Type of orbital	*s*	*p*	*d*	*f*

Electronic
Structure
of Atoms

Magnetic Quantum Number, m_l

- Describes the three-dimensional orientation of the orbital.
- Values are integers ranging from -*l* to *l*:

$$-l \leq m_l \leq l.$$

- Therefore, on any given energy level, there can be up to 1 *s* orbital, 3 *p* orbitals, 5 *d* orbitals, 7 *f* orbitals, etc.

- n=1 , l = 0, m = 0

- n=2, l = 0 m = 0

- n=2 l = 1 m = -1
- n=2 l = 1 m = 0
- n=2 l = 1 m = +1
-

Electronic Structure of Atoms

$$n=3 \qquad l=0 \qquad m=0 \qquad s$$

$$\left.\begin{cases} n=3 & l=1 & m=-1 \\ n=3 & l=1 & m=0 \\ n=3 & l=1 & m=+1 \end{cases}\right\} p$$

$$\left.\begin{cases} n=3 & l=2 & m=-2 \\ n=3 & l=2 & m=-1 \\ n=3 & l=2 & m=0 \\ n=3 & l=2 & m=+1 \\ n=3 & l=2 & m=+2 \end{cases}\right\} d$$

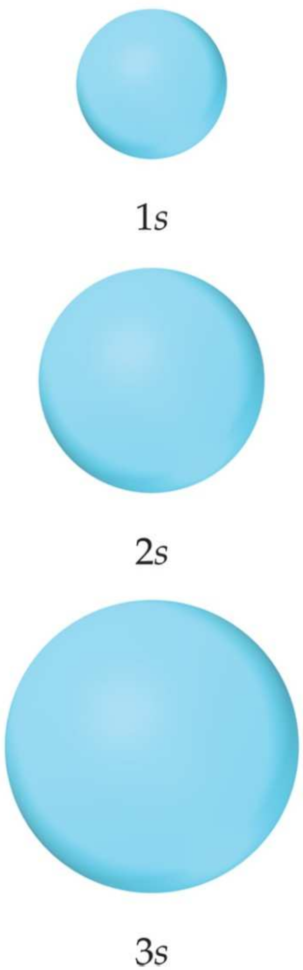
Electronic Structure of Atoms

Magnetic Quantum Number, m_l

- Orbitals with the same value of n form a **shell**.
- Different orbital types within a shell are **subshells**.
- Magnetic quantum numbers represent orbitals.

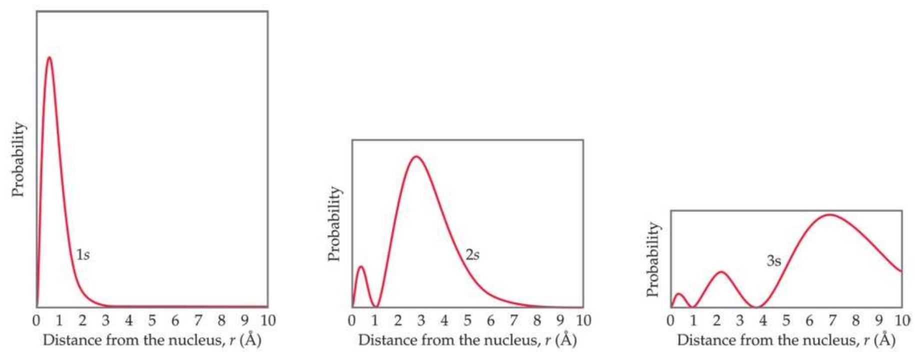
n	Possible Values of l	Subshell Designation	Possible Values of m_l	Number of Orbitals in Subshell	Total Number of Orbitals in Shell
1	0	1*s*	0	1	1
2	0	2*s*	0	1	
	1	2*p*	1, 0, −1	3	4
3	0	3*s*	0	1	
	1	3*p*	1, 0, −1	3	
	2	3*d*	2, 1, 0, −1, −2	5	9
4	0	4*s*	0	1	
	1	4*p*	1, 0, −1	3	
	2	4*d*	2, 1, 0, −1, −2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

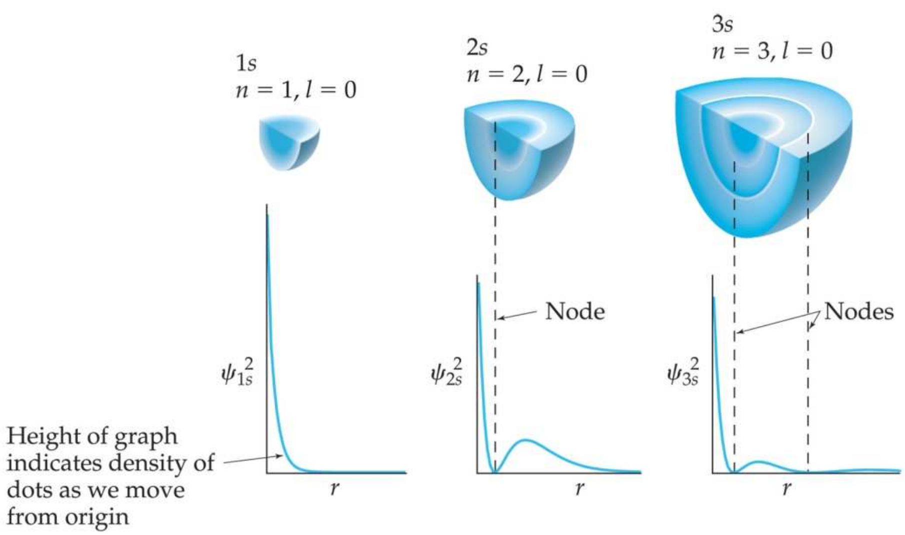
s Orbitals



- Value of *l* = 0.
- Spherical in shape.
- Radius of sphere increases with increasing value of *n*.

s Orbitals

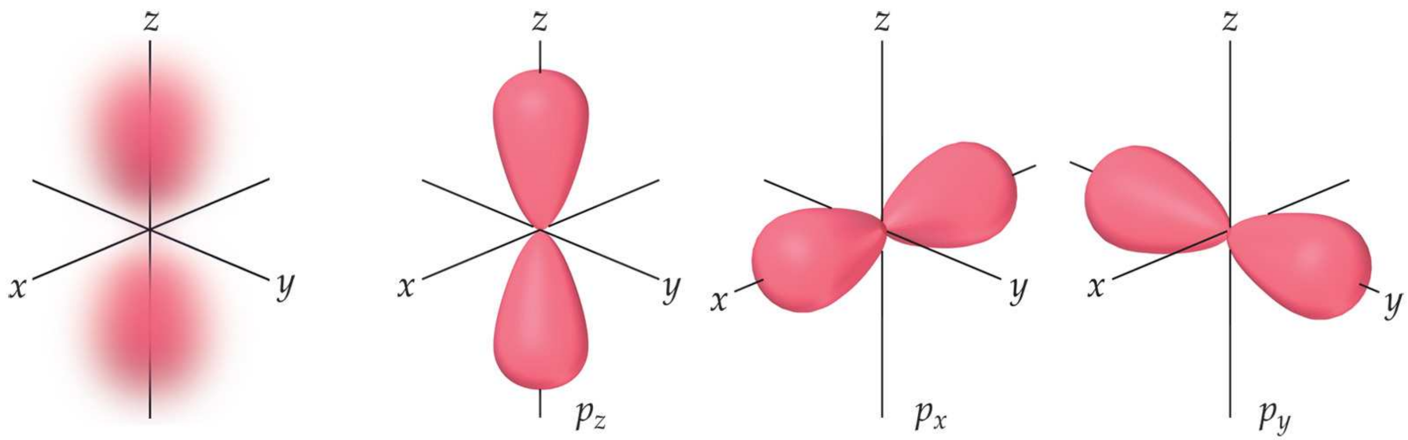




Observing a graph of probabilities of finding an electron versus distance from the nucleus, we see that *s* orbitals possess *n*−1 nodes, or regions where there is 0 probability of finding an electron.

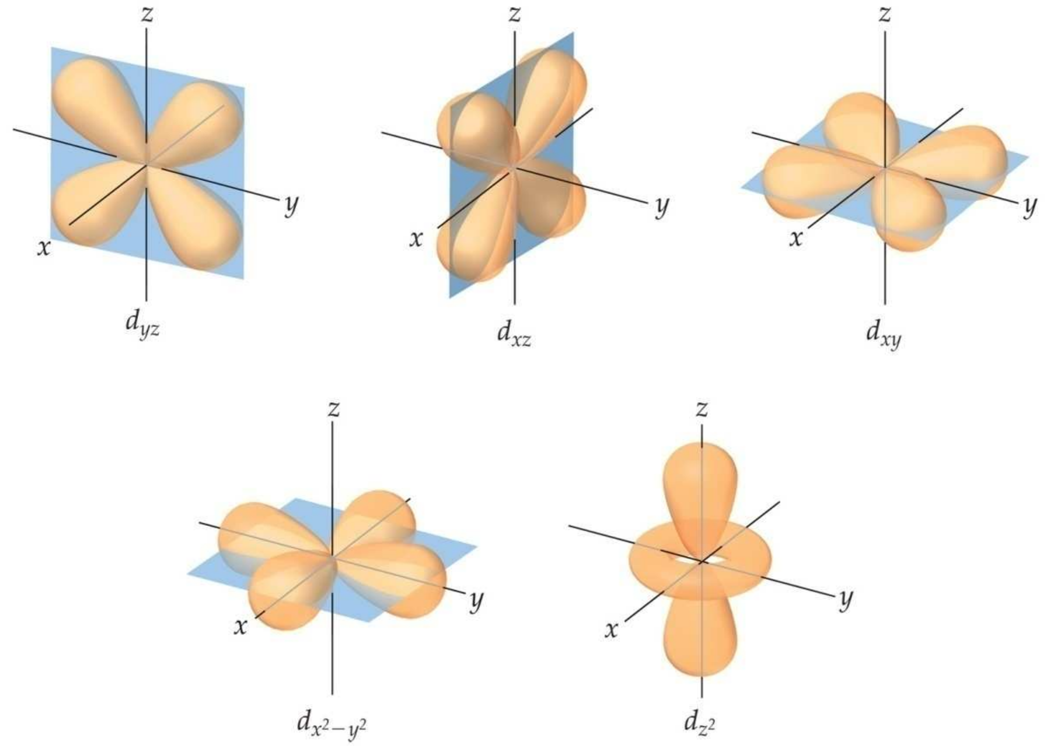
p Orbitals

- Value of *l* = 1.
- Have two lobes with a node between them.



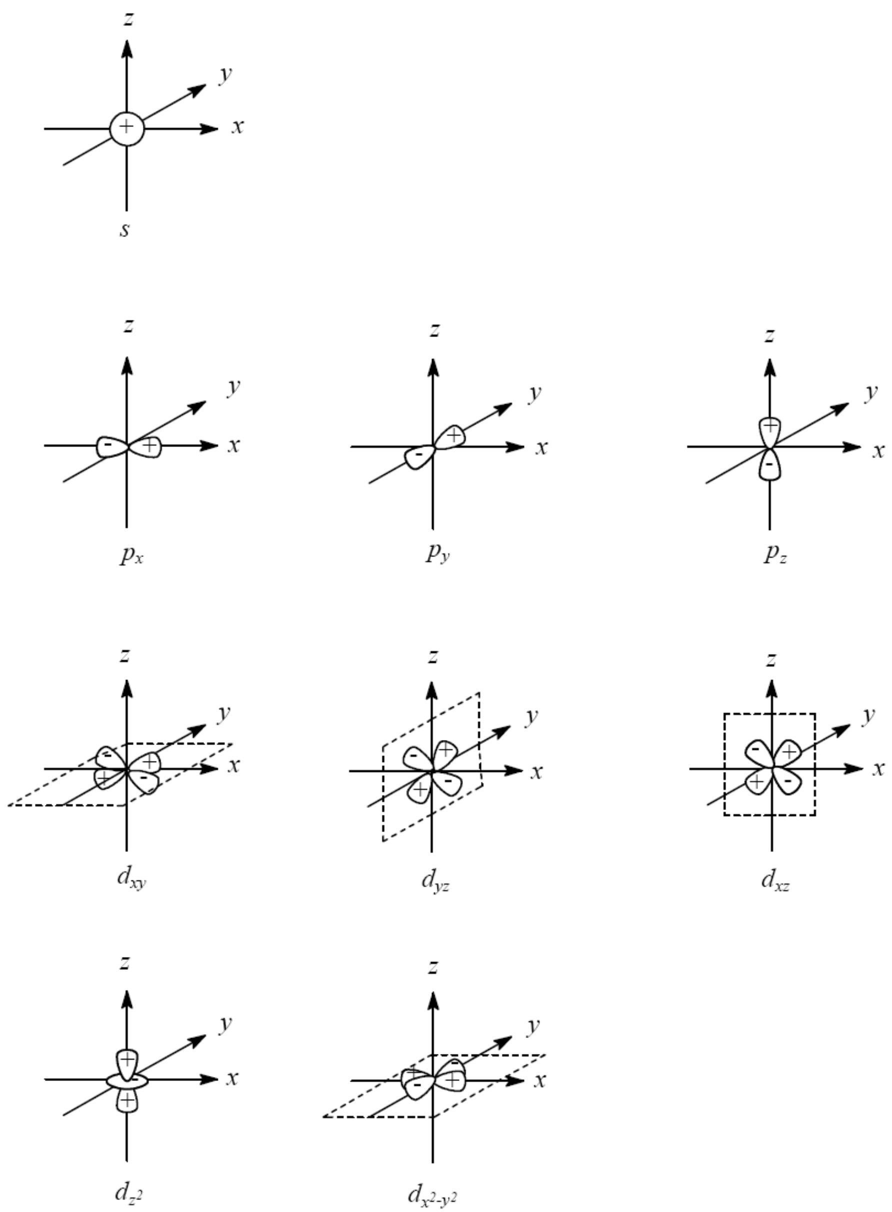
http://www.d.umn.edu/~pkiprof/ChemWebV2/AOs/ao2.html

d Orbitals



- Value of *l* is 2.
- Four of the five orbitals have 4 lobes; the other resembles a *p* orbital with a doughnut around the center.

"Balloon" Models of Atomic Orbitals



From Dr Carter's website.

Electronic Structure of Atoms

Shells

- n is a shell

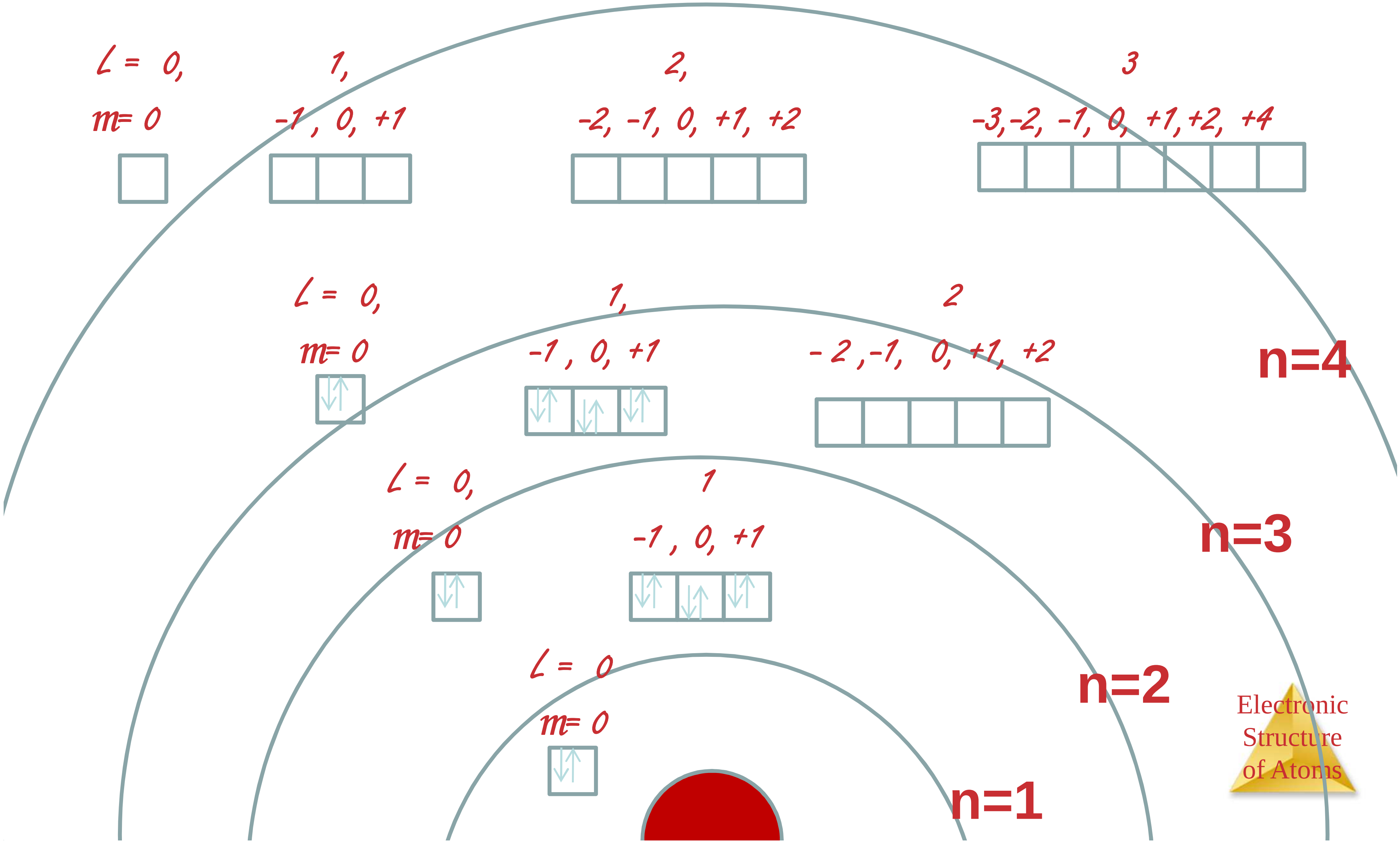
Subshells

- s, p d and f are sub shells

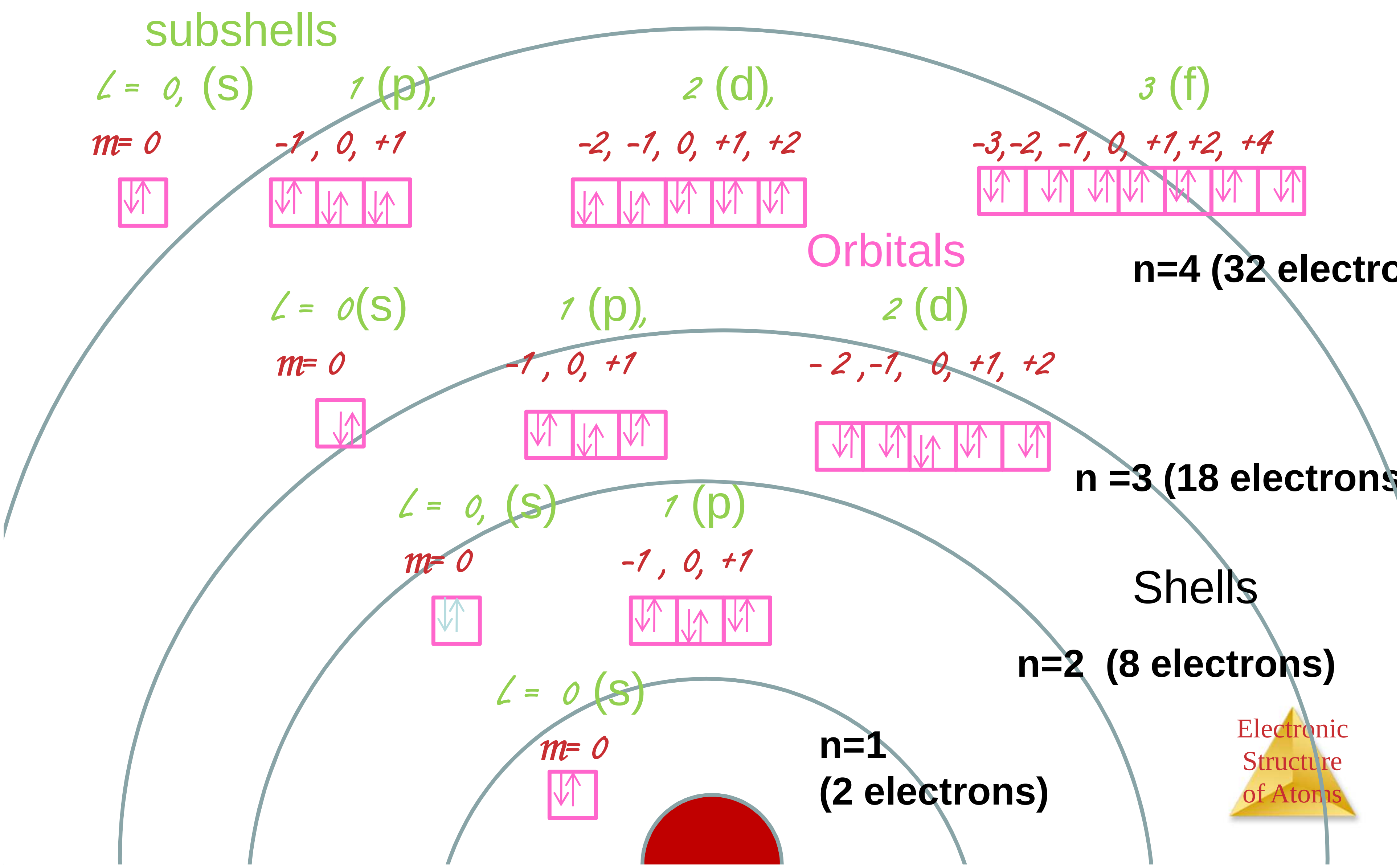
Orbitals

- s has 1,
- p has 3,
- d has 5 and
- f has 7 orbitals.

- The number of orbitals in a each shell is n^2
- n=1 orbital $1^2 = 1$
- n=2 orbital $2^2 = 4$ 1s and 3p
- n=3 orbital $3^2 = 9$ 1s, 3p and 5d
- n=4 orbital $4^2 = 16$ 1s, 3p, 5d and 7f

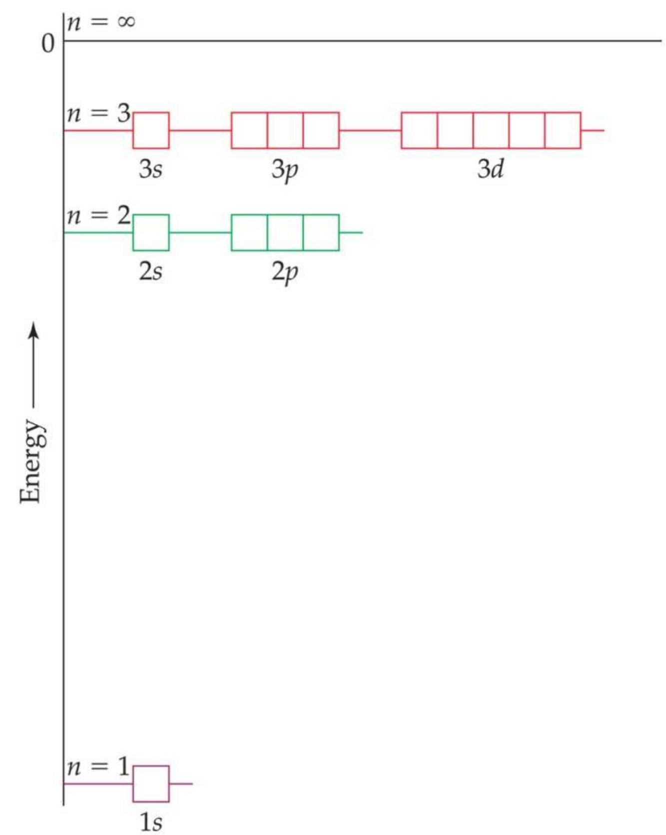


l = 0,
1,
2,
3
m= 0
-1 , 0, +1
-2, -1, 0, +1, +2
-3,-2, -1, 0, +1,+2, +4
l = 0,
1,
2
m= 0
-1 , 0, +1
- 2 ,-1, 0, +1, +2
n=4
l = 0,
1
m= 0
-1 , 0, +1
n=3
l = 0
m= 0
n=2
n=1
Electronic Structure of Atoms


subshells
l = 0, (s)
1 (p),
2 (d),
3 (f)
m= 0
-1 , 0, +1
-2, -1, 0, +1, +2
-3,-2, -1, 0, +1,+2, +4
Orbitals
n=4 (32 electro
l = 0(s)
1 (p),
2 (d)
m= 0
-1 , 0, +1
- 2 ,-1, 0, +1, +2
n =3 (18 electrons
l = 0, (s)
1 (p)
m= 0
-1 , 0, +1
Shells
n=2 (8 electrons)
l = 0 (s)
m= 0
n=1
(2 electrons)
Electronic
Structure
of Atoms

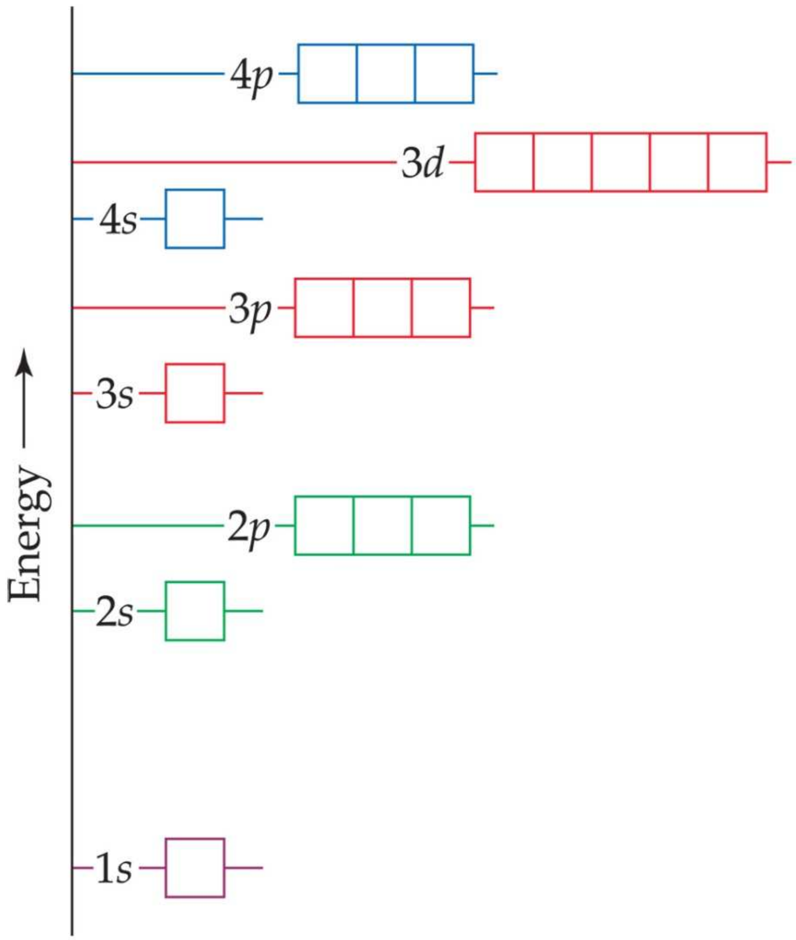
Energies of Orbitals

- For a one-electron hydrogen atom, orbitals on the same energy level have the same energy.
- That is, they are degenerate.

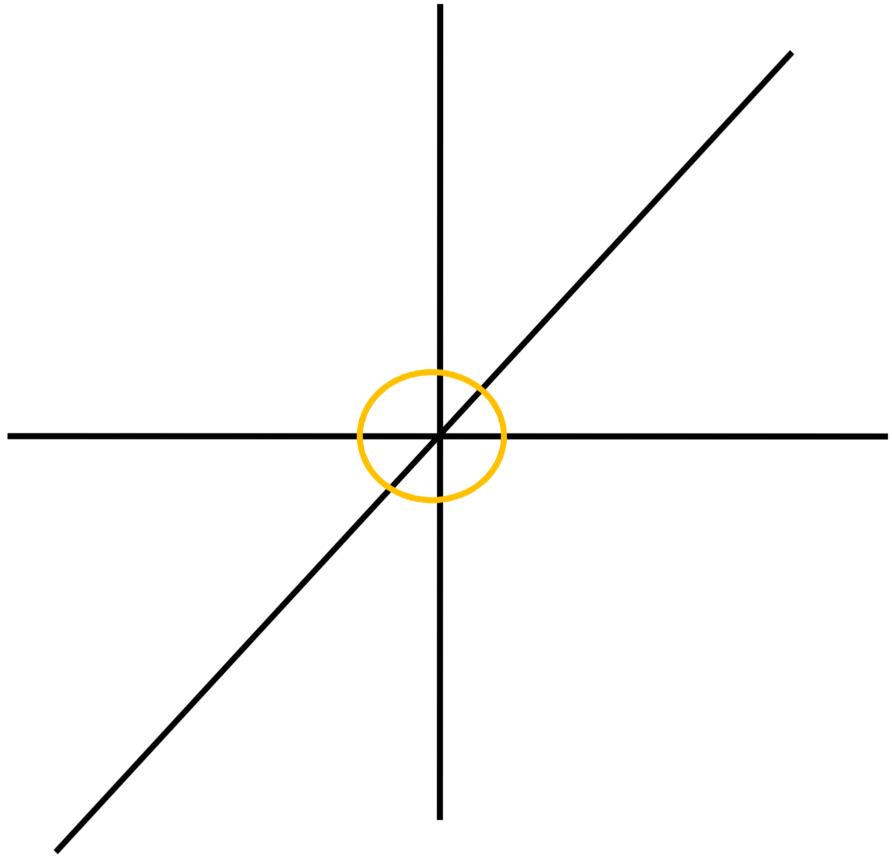


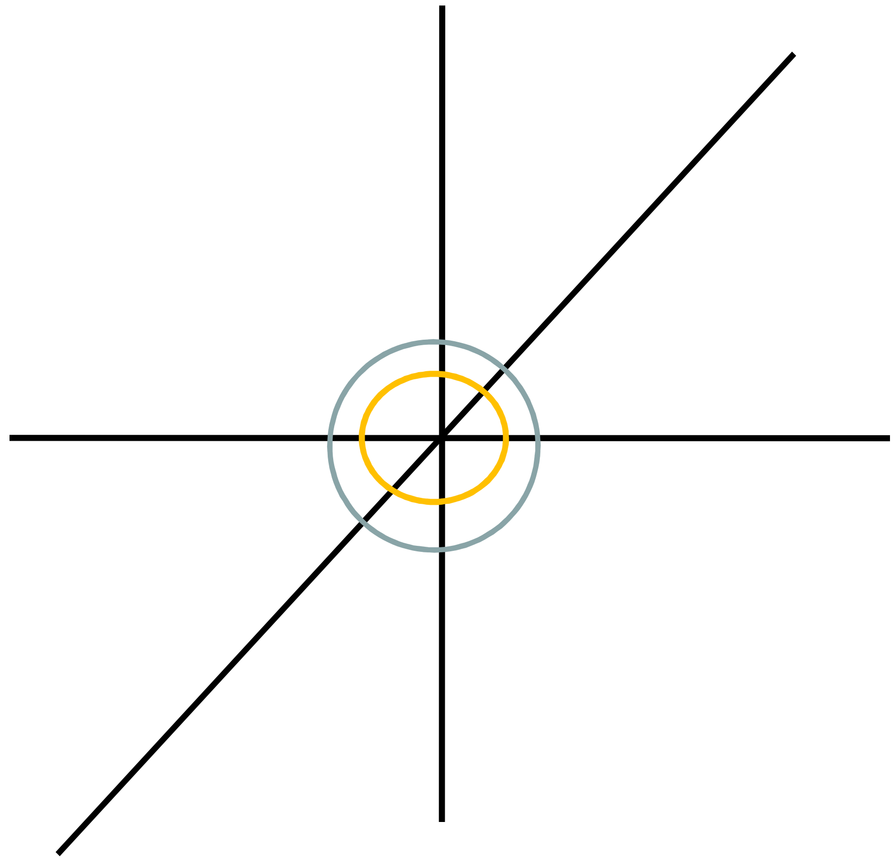
Electronic Structure of Atoms

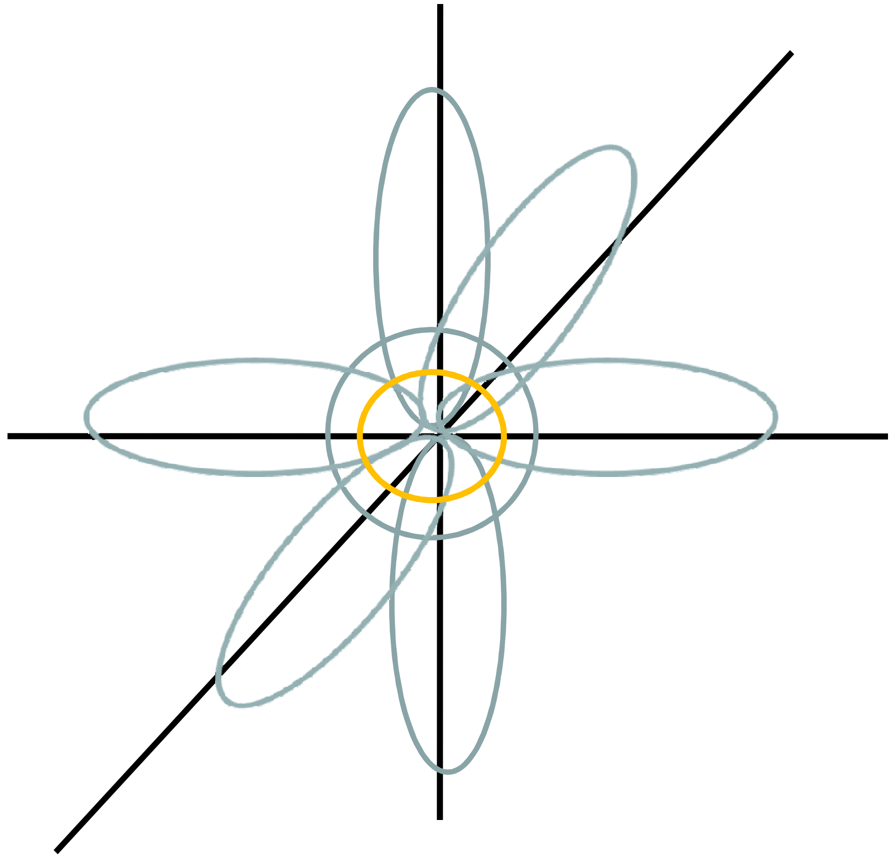
Energies of Orbitals

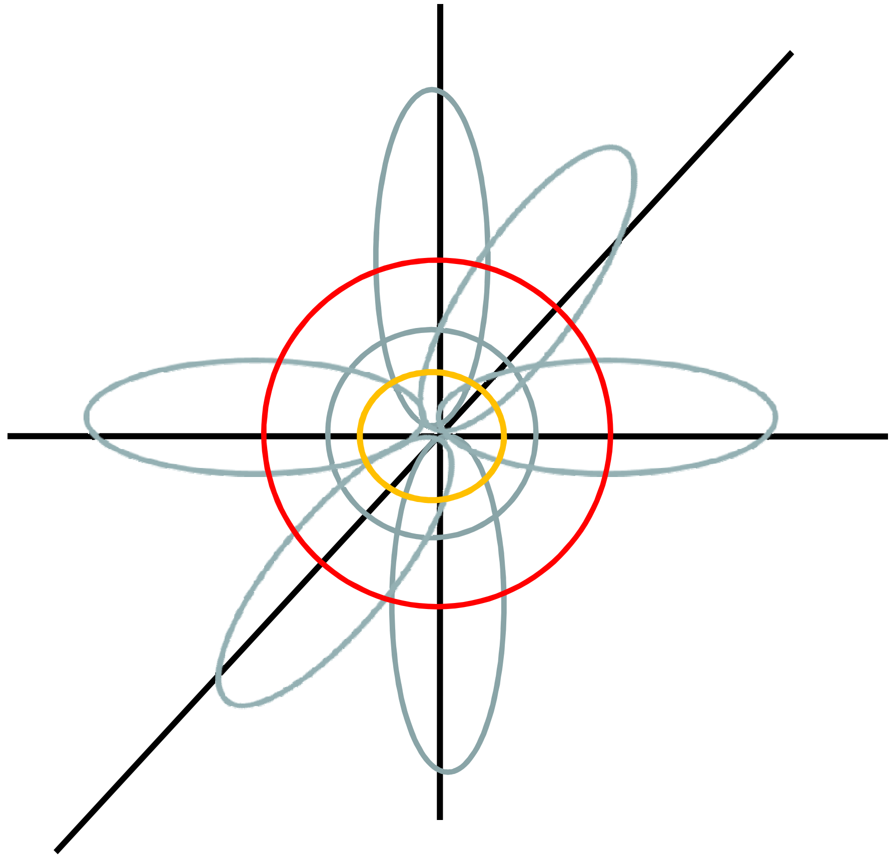


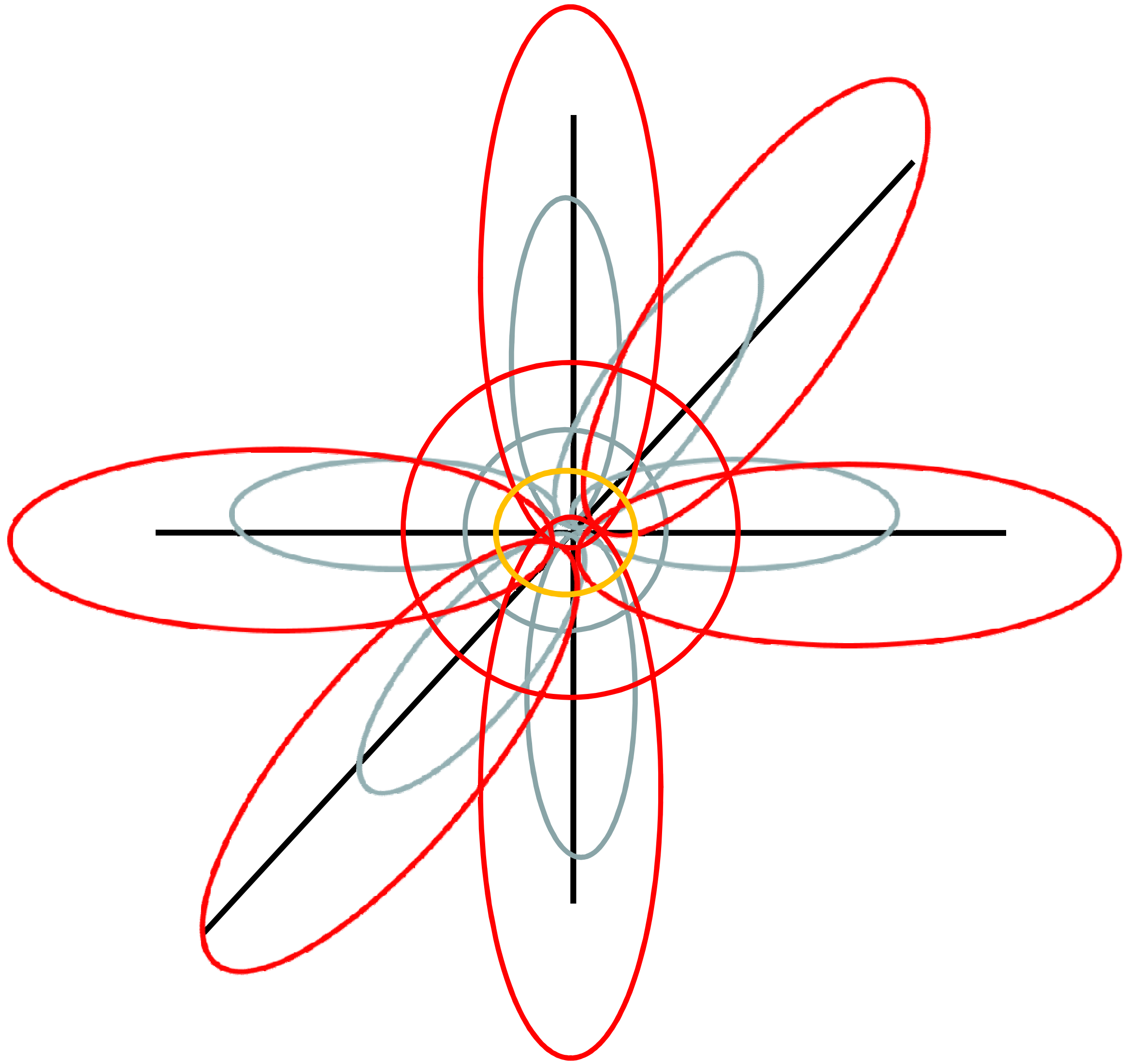
- As the number of electrons increases, though, so does the repulsion between them.
- Therefore, in many-electron atoms, orbitals on the same energy level are no longer degenerate.



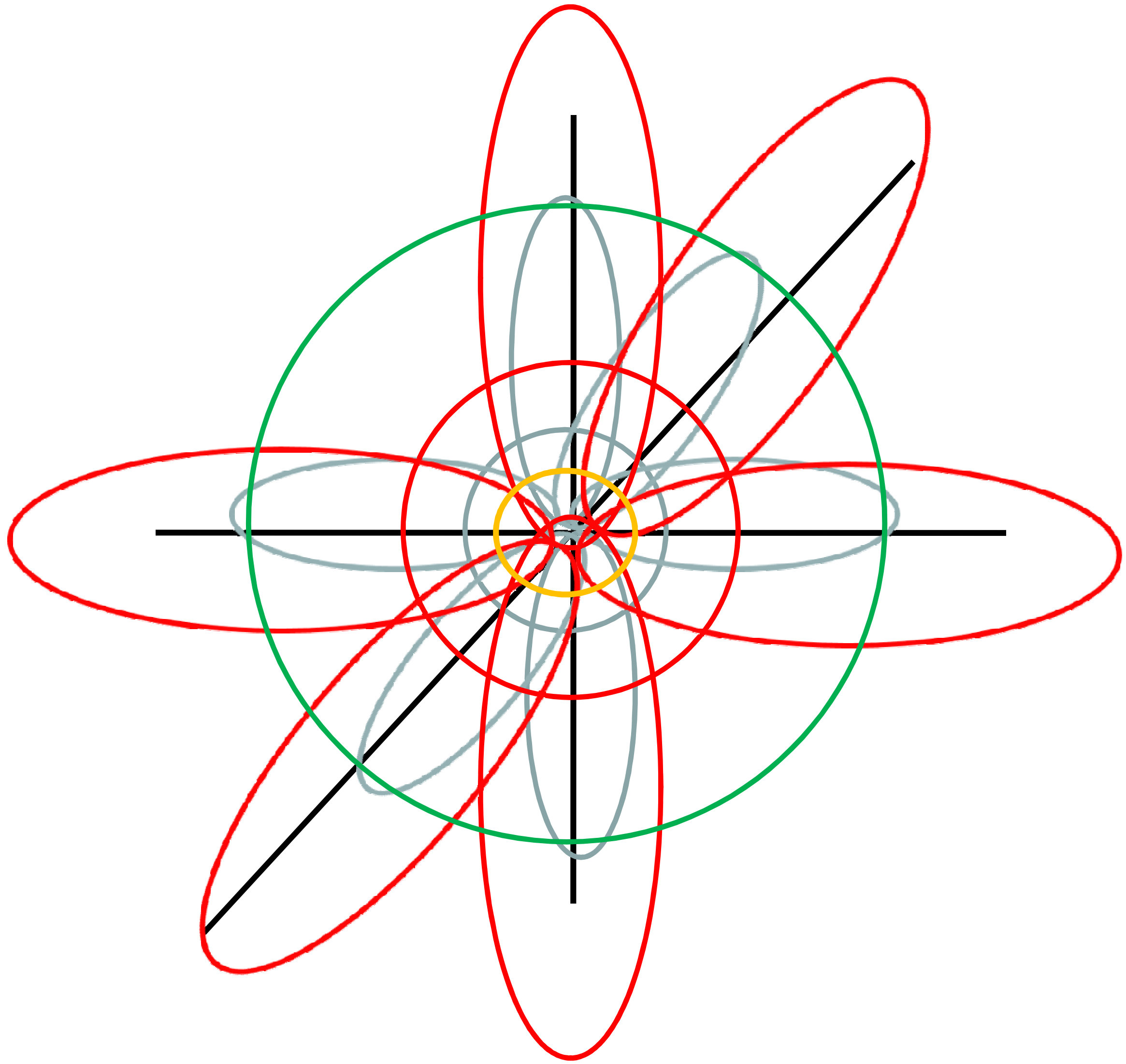


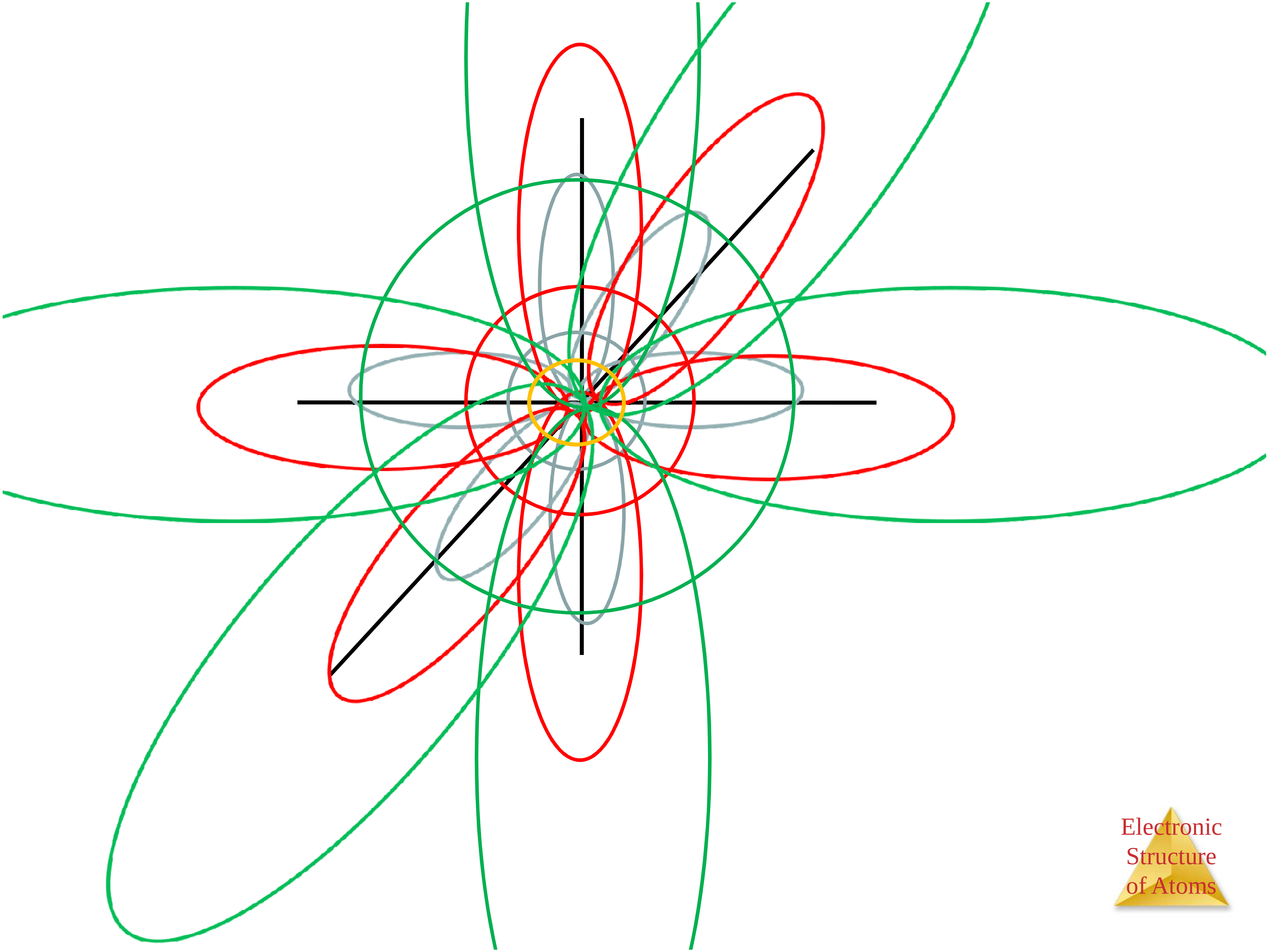






Electronic
Structure
of Atoms

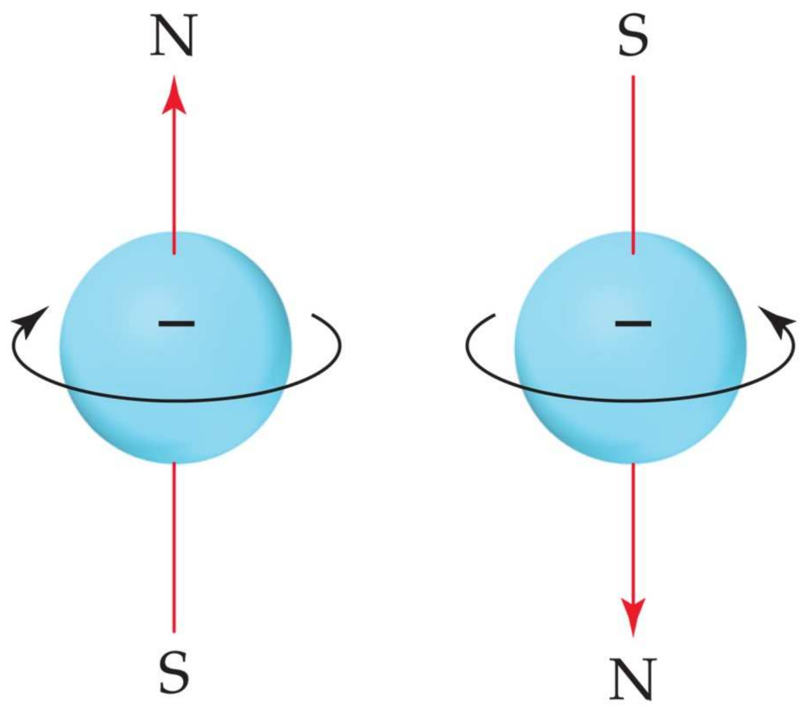




Electronic
Structure
of Atoms

Spin Quantum Number, m_s

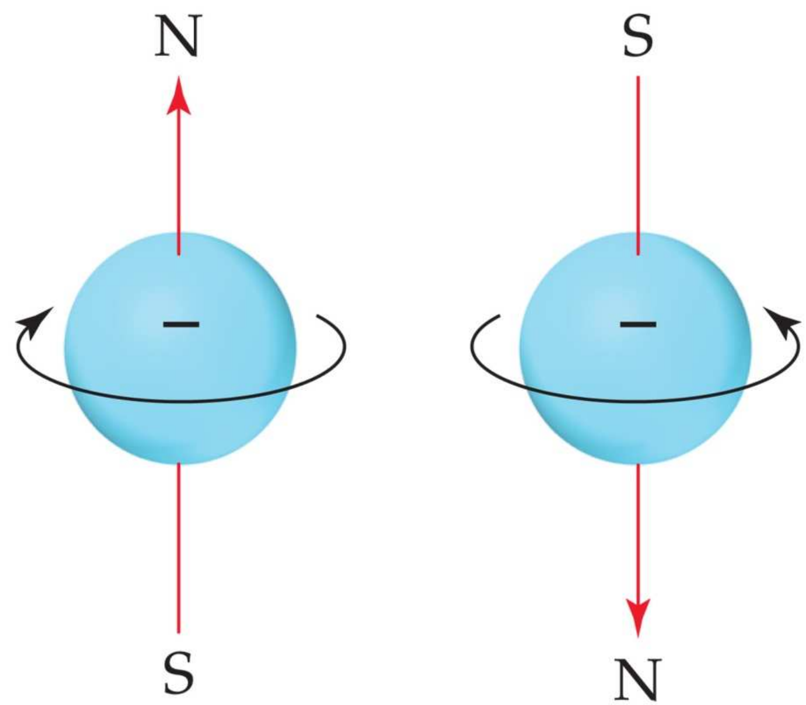
- In the 1920s, it was discovered that two electrons in the same orbital do not have exactly the same energy.
- The “spin” of an electron describes its magnetic field, which affects its energy.



Electronic Structure of Atoms

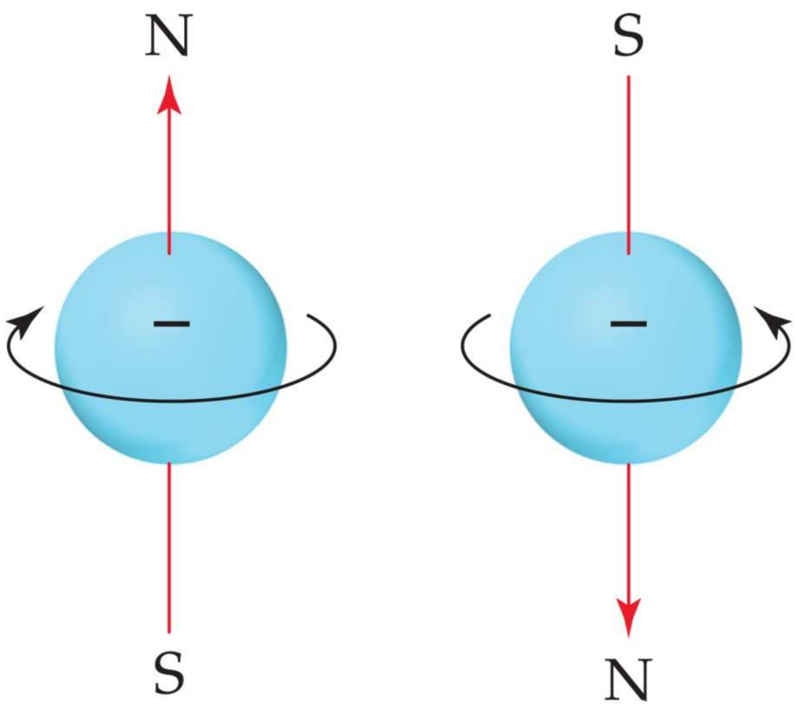
Spin Quantum Number, m_s

- This led to a fourth quantum number, the spin quantum number, m_s.
- The spin quantum number has only 2 allowed values: +1/2 and −1/2.



Pauli Exclusion Principle

- No two electrons in the same atom can have exactly the same energy.
- For example, no two electrons in the same atom can have identical sets of quantum numbers.



Electron Configurations

$4p^5$

- Distribution of all electrons in an atom
- Consist of
  - Number denoting the energy level

Electron Configurations

$4p^5$

- Distribution of all electrons in an atom
- Consist of
  - Number denoting the energy level
  - Letter denoting the type of orbital

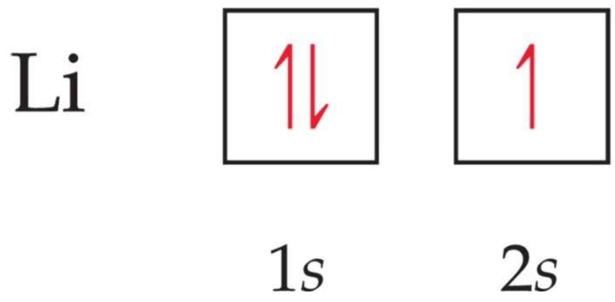
Electron Configurations

$4p^5$

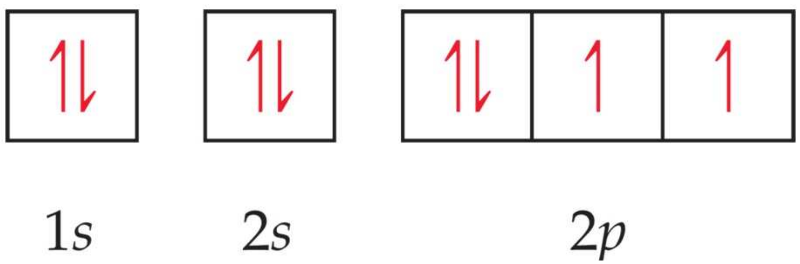
- Distribution of all electrons in an atom.
- Consist of
  - Number denoting the energy level.
  - Letter denoting the type of orbital.
  - Superscript denoting the number of electrons in those orbitals.

Orbital Diagrams

- Each box represents one orbital.
- Half-arrows represent the electrons.
- The direction of the arrow represents the spin of the electron.

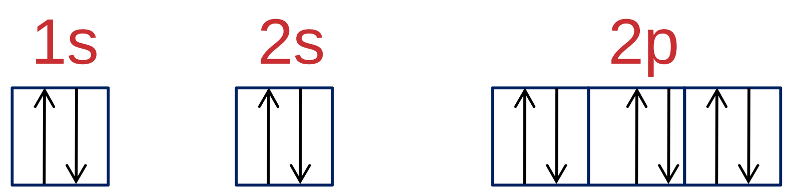


Hund's Rule



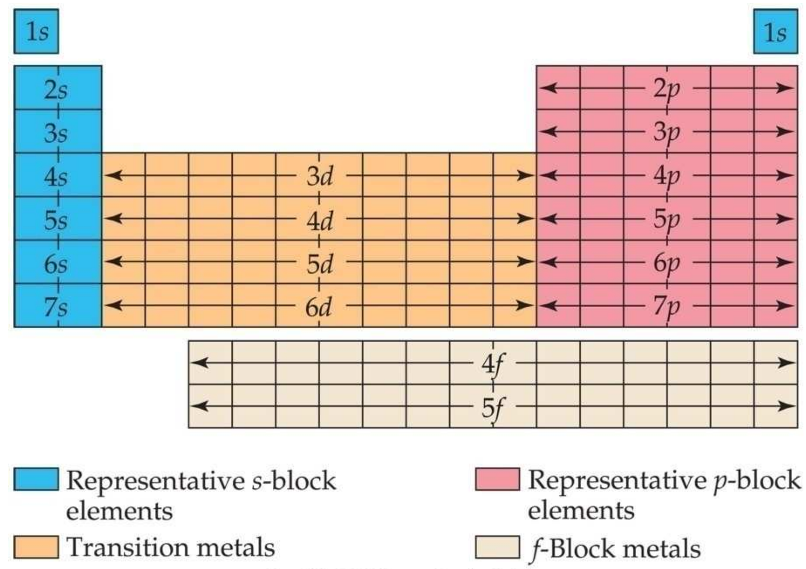
"For degenerate orbitals, the lowest energy is attained when the number of electrons with the same spin is maximized."

The sequence of filling of electrons in atoms will be:



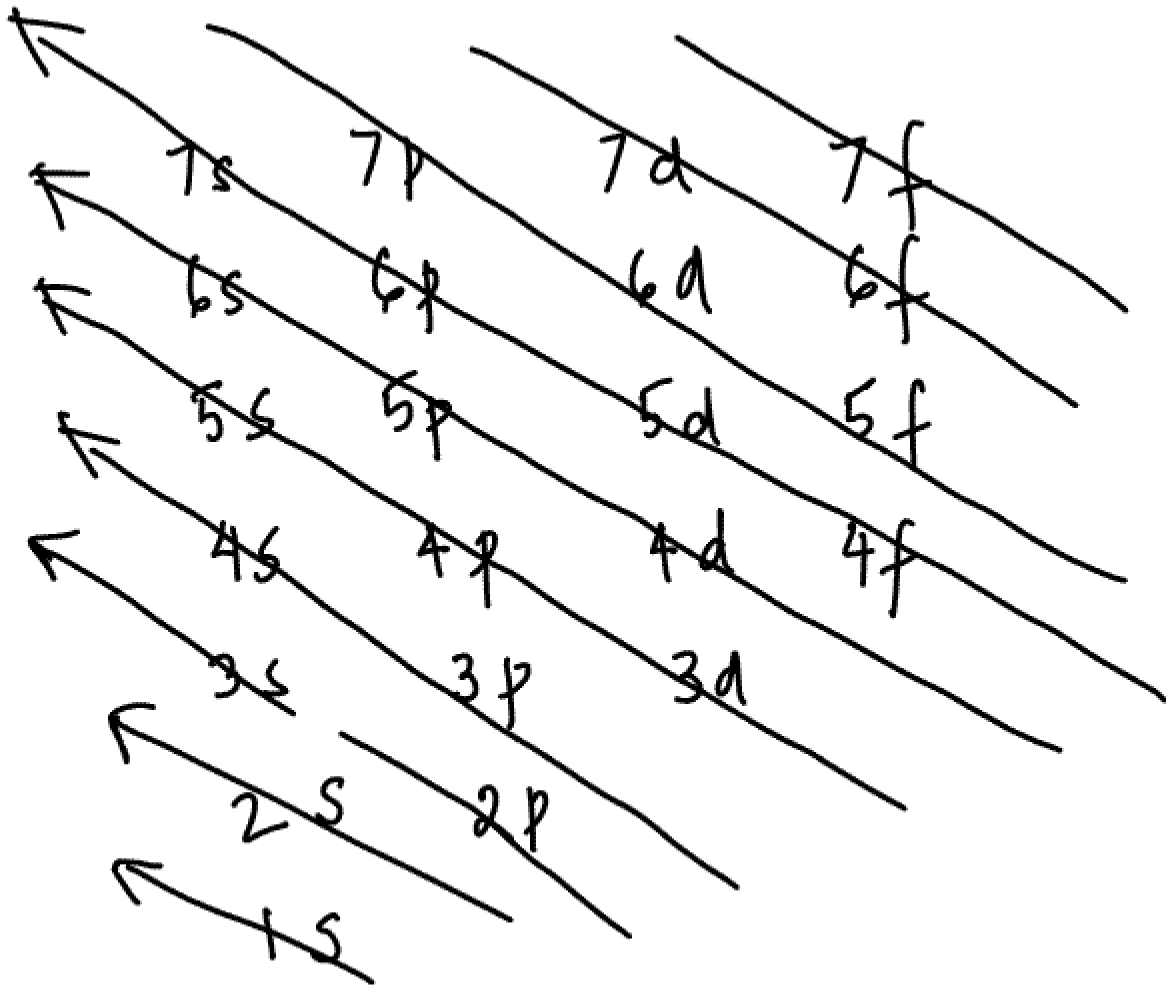
Periodic Table

- We fill orbitals in increasing order of energy.
- Different blocks on the periodic table, then correspond to different types of orbitals.

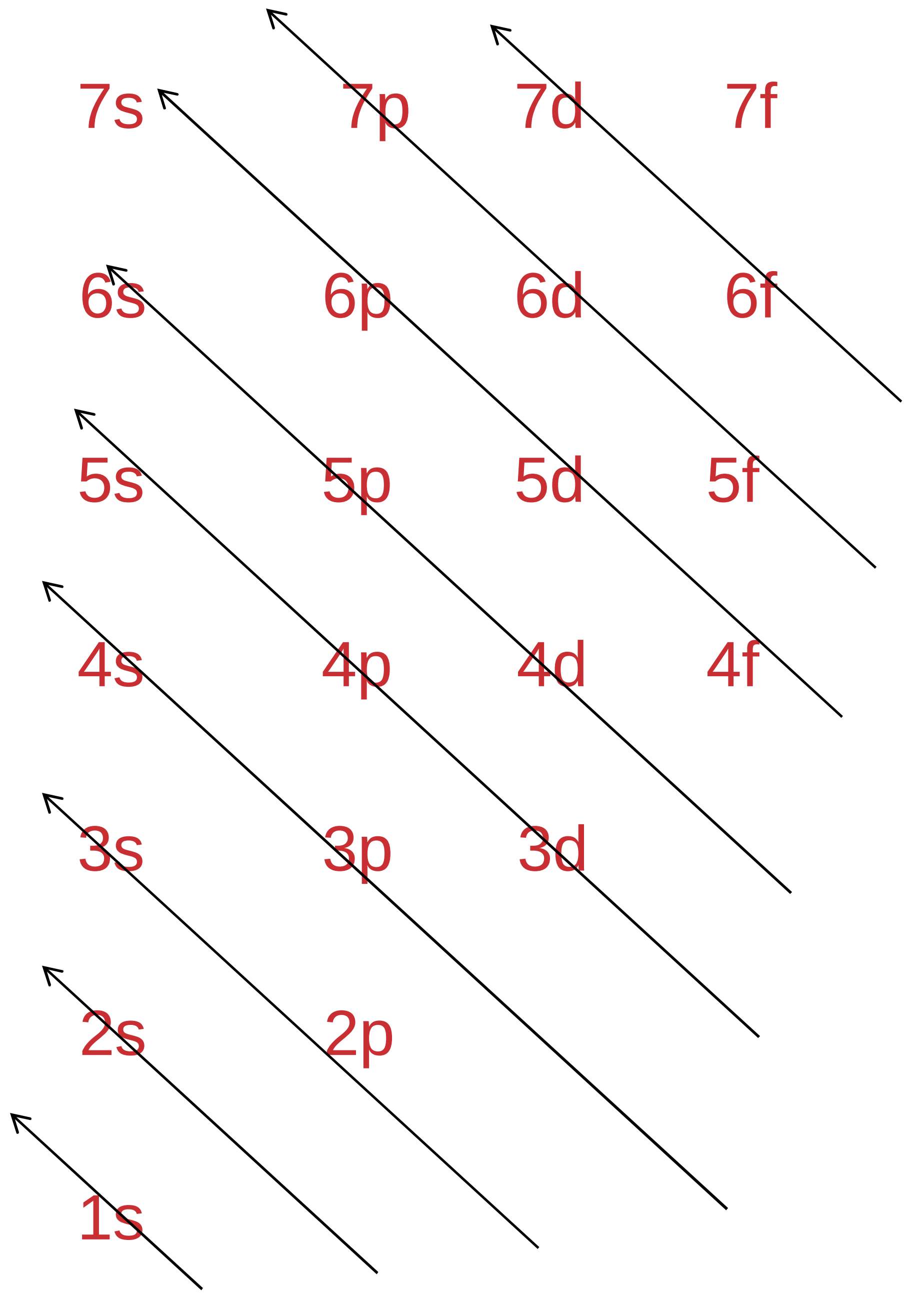


Aufbau Principle

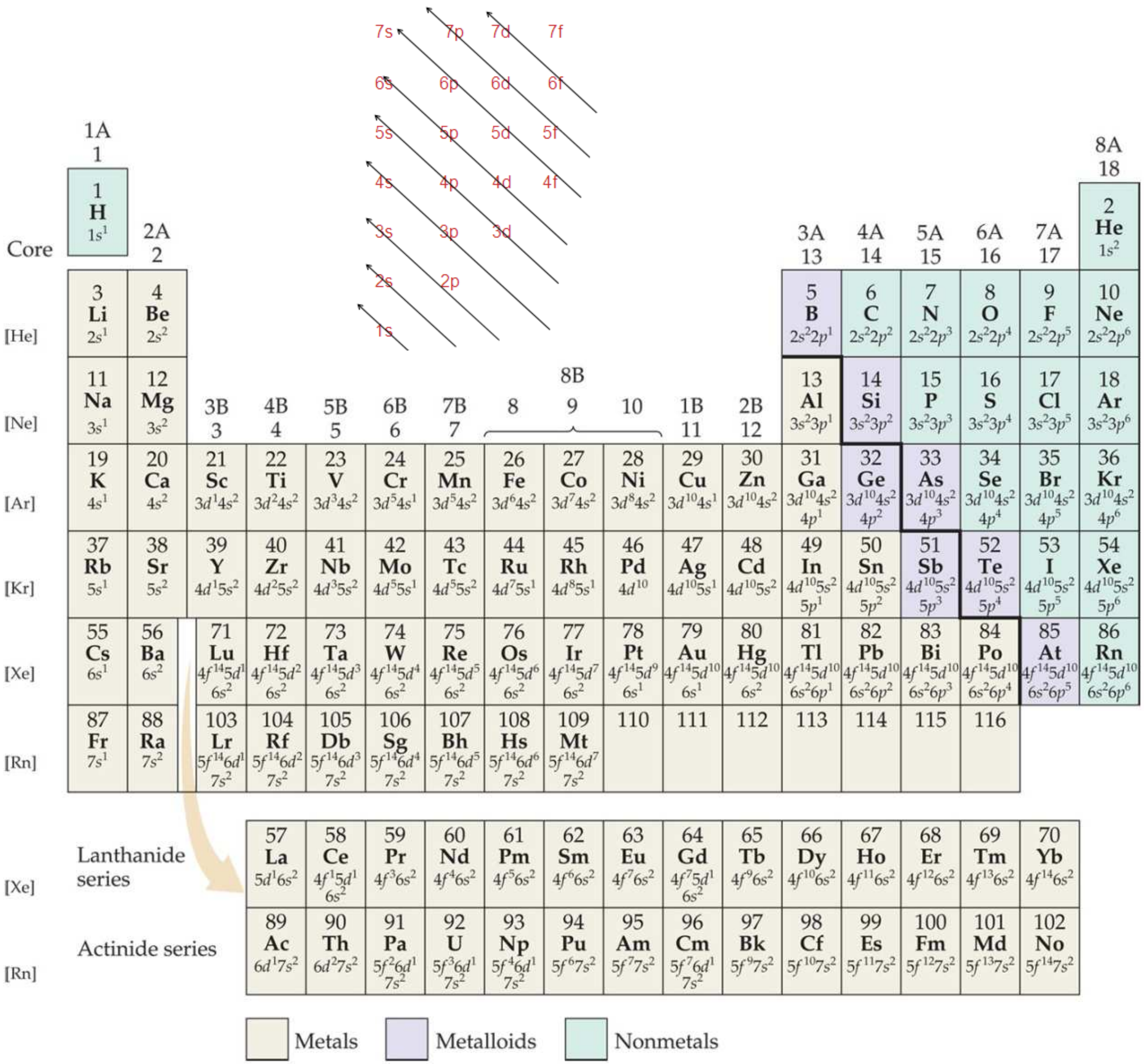
- This is the building up principle of the electrons.
- Aufbau in German means building up.



Electronic
Structure
of Atoms



Electronic Structure of Atoms



1s	1s
2s2p	2sp
3s3p	3spd
4s3d4p	4spdf
5a4d4p	5spdf
6s4f5d6p	6spdf
7s5f6d	7spdf

Electronic Structure of Atoms

- Electronic configuration
- Condensed electronic configuration
- Orbital Diagram

B $1s^2\,2s^2\,2p^1$ or $\underset{1s^2}{\underline{\uparrow\downarrow}}\quad\underset{2s^2}{\underline{\uparrow\downarrow}}\quad\underbrace{\underline{\uparrow}\ \underline{}\ \underline{}}_{2p^1}$ or [He] $2s^2\,2p^1$

C $1s^2\,2s^2\,2p^2$ or $\underset{1s^2}{\underline{\uparrow\downarrow}}\quad\underset{2s^2}{\underline{\uparrow\downarrow}}\quad\underbrace{\underline{\uparrow}\ \underline{\uparrow}\ \underline{}}_{2p^2}$ or [He] $2s^2\,2p^2$

N $1s^2\,2s^2\,2p^3$ or $\underset{1s^2}{\underline{\uparrow\downarrow}}\quad\underset{2s^2}{\underline{\uparrow\downarrow}}\quad\underbrace{\underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}}_{2p^3}$ or [He] $2s^2\,2p^3$

O $1s^2\,2s^2\,2p^4$ or $\underset{1s^2}{\underline{\uparrow\downarrow}}\quad\underset{2s^2}{\underline{\uparrow\downarrow}}\quad\underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow}\ \underline{\uparrow}}_{2p^4}$ or [He] $2s^2\,2p^4$

F $1s^2\,2s^2\,2p^5$ or $\underset{1s^2}{\underline{\uparrow\downarrow}}\quad\underset{2s^2}{\underline{\uparrow\downarrow}}\quad\underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow}}_{2p^5}$ or [He] $2s^2\,2p^5$

Ne $1s^2\,2s^2\,2p^6$ or $\underset{1s^2}{\underline{\uparrow\downarrow}}\quad\underset{2s^2}{\underline{\uparrow\downarrow}}\quad\underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}}_{2p^6}$

and 8A Elements

GROUP	ELEMENT	ATOMIC NUMBER	VALENCE-SHELL ELECTRON CONFIGURATION
1A	Li (lithium)	3	$2s^1$
2A			
7A			
8A			

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and 8A Elements

GROUP	ELEMENT	ATOMIC NUMBER	VALENCE-SHELL ELECTRON CONFIGURATION
1A	Li (lithium)	3	$2s^1$
	Na (sodium)	11	$3s^1$
2A			
7A			
8A			

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and 8A Elements

GROUP	ELEMENT	ATOMIC NUMBER	VALENCE-SHELL ELECTRON CONFIGURATION
1A	Li (lithium)	3	$2s^1$
	Na (sodium)	11	$3s^1$
	K (potassium)	19	$4s^1$
2A			
7A			
8A			

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and 8A Elements

GROUP	ELEMENT	ATOMIC NUMBER	VALENCE-SHELL ELECTRON CONFIGURATION
1A	Li (lithium)	3	$2s^1$
	Na (sodium)	11	$3s^1$
	K (potassium)	19	$4s^1$
	Rb (rubidium)	37	$5s^1$
2A			
7A			
8A			

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GROUP	ELEMENT	ATOMIC NUMBER	VALENCE-SHELL ELECTRON CONFIGURATION
1A	Li (lithium)	3	$2s^1$
	Na (sodium)	11	$3s^1$
	K (potassium)	19	$4s^1$
	Rb (rubidium)	37	$5s^1$
	Cs (cesium)	55	$6s^1$
2A			
7A			
8A			

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and 8A Elements

GROUP	ELEMENT	ATOMIC NUMBER	VALENCE-SHELL ELECTRON CONFIGURATION
1A	Li (lithium)	3	$2s^1$
	Na (sodium)	11	$3s^1$
	K (potassium)	19	$4s^1$
	Rb (rubidium)	37	$5s^1$
	Cs (cesium)	55	$6s^1$
2A	Be (beryllium)	4	$2s^2$
7A			
8A			

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and 8A Elements

GROUP	ELEMENT	ATOMIC NUMBER	VALENCE-SHELL ELECTRON CONFIGURATION
1A	Li (lithium)	3	$2s^1$
	Na (sodium)	11	$3s^1$
	K (potassium)	19	$4s^1$
	Rb (rubidium)	37	$5s^1$
	Cs (cesium)	55	$6s^1$
2A	Be (beryllium)	4	$2s^2$
	Mg (magnesium)	12	$3s^2$
7A			
8A			

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and 8A Elements

GROUP	ELEMENT	ATOMIC NUMBER	VALENCE-SHELL ELECTRON CONFIGURATION
1A	Li (lithium)	3	$2s^1$
	Na (sodium)	11	$3s^1$
	K (potassium)	19	$4s^1$
	Rb (rubidium)	37	$5s^1$
	Cs (cesium)	55	$6s^1$
2A	Be (beryllium)	4	$2s^2$
	Mg (magnesium)	12	$3s^2$
	Ca (calcium)	20	$4s^2$
7A			
8A			

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and 2A Elements

GROUP	ELEMENT	ATOMIC NUMBER	VALENCE-SHELL ELECTRON CONFIGURATION
1A	Li (lithium)	3	$2s^1$
	Na (sodium)	11	$3s^1$
	K (potassium)	19	$4s^1$
	Rb (rubidium)	37	$5s^1$
	Cs (cesium)	55	$6s^1$
2A	Be (beryllium)	4	$2s^2$
	Mg (magnesium)	12	$3s^2$
	Ca (calcium)	20	$4s^2$
	Sr (strontium)	38	$5s^2$
	Ba (barium)	56	$6s^2$

4 A have 4 electrons
5A 5 electrons
6A 6 electrons
7 A 7 electrons
8A 8 electrons
in their outer most shell

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and 8A Elements

GROUP	ELEMENT	ATOMIC NUMBER	VALENCE-SHELL ELECTRON CONFIGURATION
1A	Li (lithium)	3	$2s^1$
	Na (sodium)	11	$3s^1$
	K (potassium)	19	$4s^1$
	Rb (rubidium)	37	$5s^1$
	Cs (cesium)	55	$6s^1$
2A	Be (beryllium)	4	$2s^2$
	Mg (magnesium)	12	$3s^2$
	Ca (calcium)	20	$4s^2$
	Sr (strontium)	38	$5s^2$
	Ba (barium)	56	$6s^2$
7A	F (fluorine)	9	$2s^2\ 2p^5$
	Cl (chlorine)	17	$3s^2\ 3p^5$
	Br (bromine)	35	$4s^2\ 4p^5$
	I (iodine)	53	$5s^2\ 5p^5$
8A	He (helium)	2	$1s^2$
	Ne (neon)	10	$2s^2\ 2p^6$
	Ar (argon)	18	$3s^2\ 3p^6$
	Kr (krypton)	36	$4s^2\ 4p^6$
	Xe (xenon)	54	$5s^2\ 5p^6$

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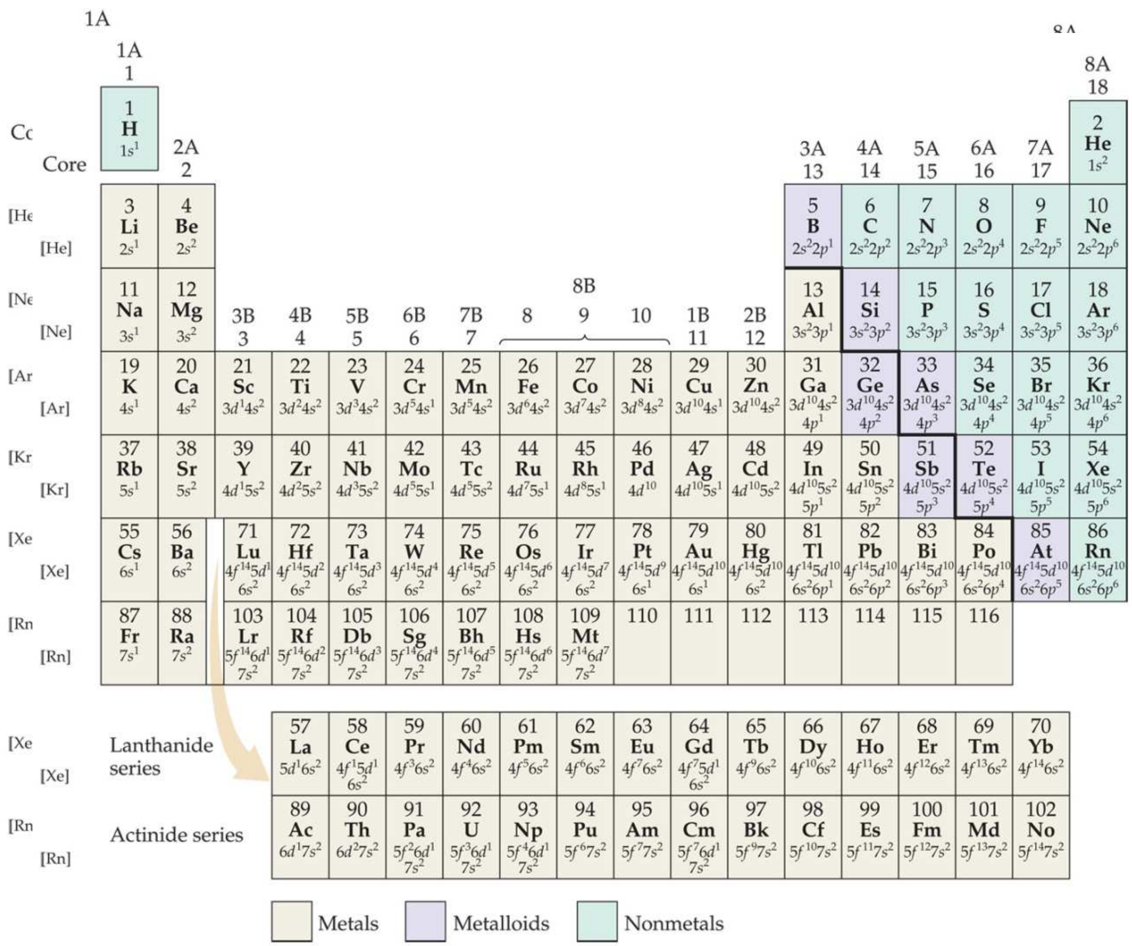
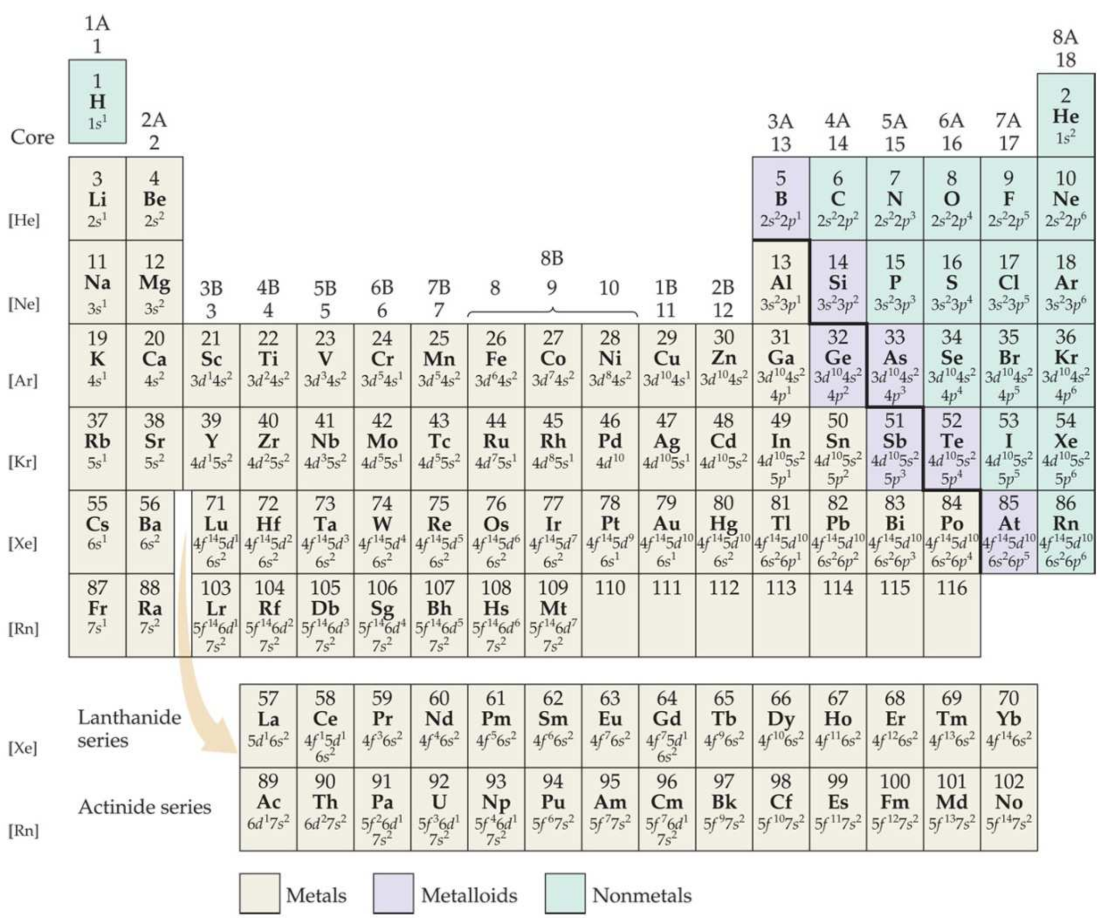
Some Anomalies



Some irregularities occur when there are enough electrons to half-fill *s* and *d* orbitals on a given row.

Some Anomalies



For instance, the electron configuration for copper is

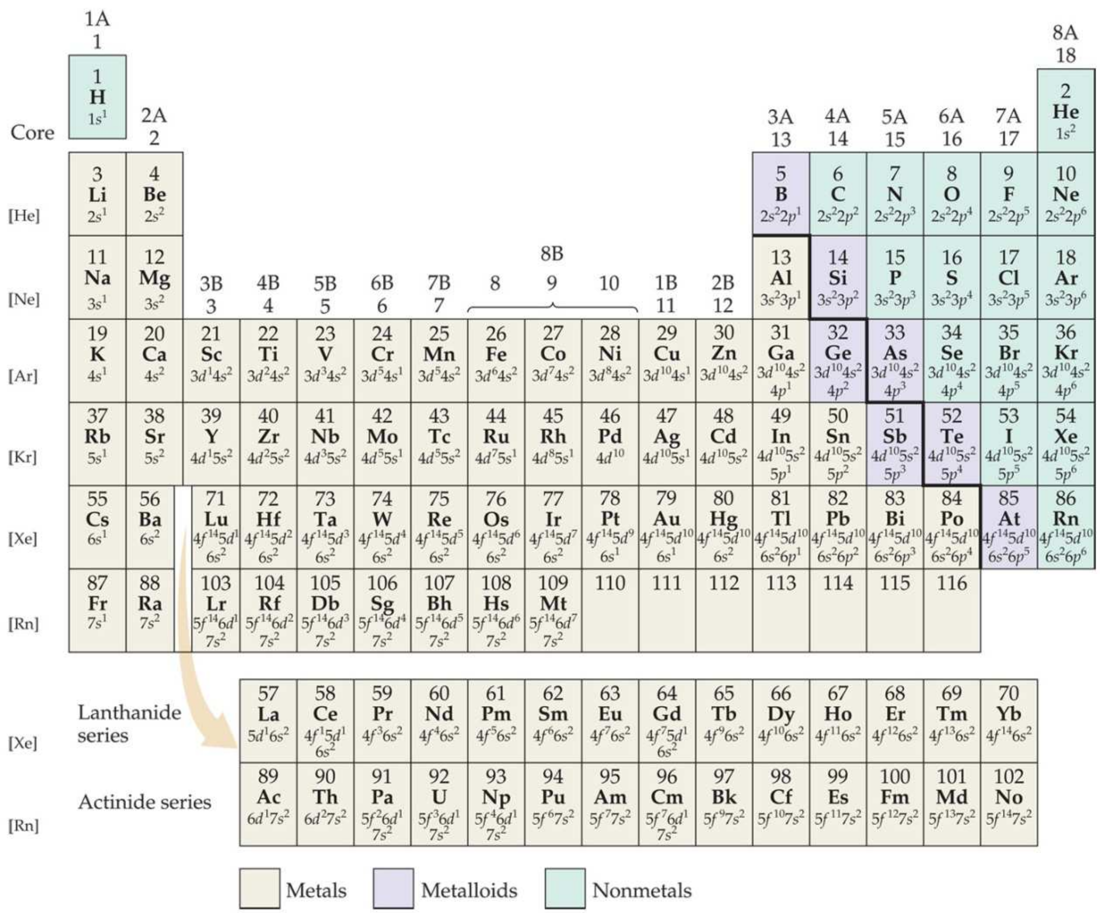
[Ar] $4s^1\ 3d^{10}$

rather than the expected

[Ar] $4s^2\ 3d^9$.

Electronic Structure of Atoms

Some Anomalies



- This occurs because the 4*s* and 3*d* orbitals are very close in energy.
- Mo, Pd and Ag also show these anomalies.
- .

Electronic Structure of Atoms

	Actual	Not
• Mo	[Kr] $4d^5\ 5s^1$	[Kr] $4d^4\ 5s^2$
• Pd	[Kr] $4d^{10}$	[Kr] $4d^8\ 5s^2$
• Ag	[Kr] $4d^{10}\ 5s^1$	[Kr] $4d^9\ 5s^2$

- There are some anomalies in the f block elements too.
- These anomalies though interesting do not have great chemical significance.