ELECTRON CONFIGURATIONS

VISUAL CHEM CARDS

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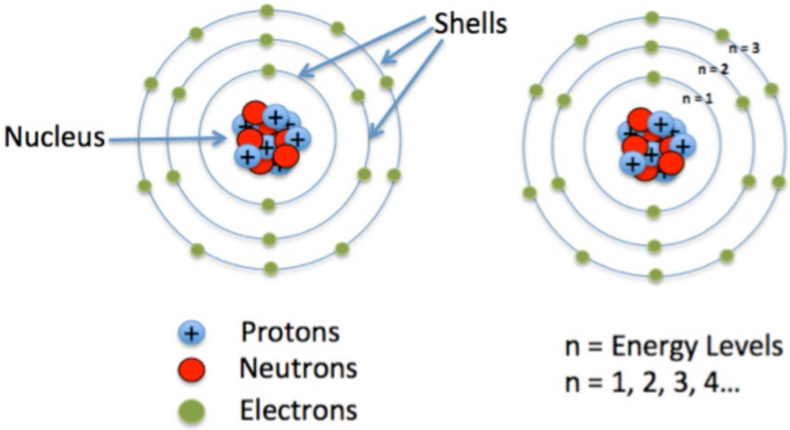
Electrons

Mass:	9.11×10^{-31} kg ($\frac{1}{1836}$ the mass of a proton)
Charge:	-1 (-1.602×10^{-19} C)
Symbol:	e-, β-



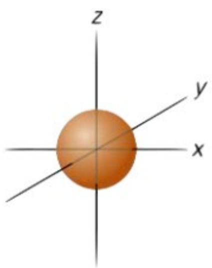
Electrons occupy shells around the nucleus

Different shells can hold different maximum numbers of electrons.

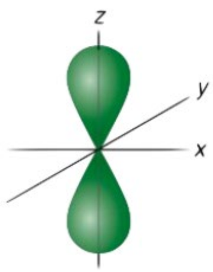
Electron shells have one or more electron **subshells (s, p, d, f).**

Subshells made from **atomic orbitals** – each orbital can hold a maximum of 2 electrons.

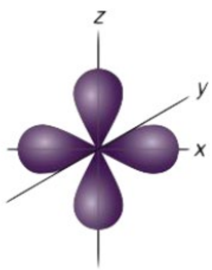
Sub Shell	Number of Atomic Orbitals	Maximum No. of Electrons
s	1	2
p	3	6
d	5	10
f	7	14



S

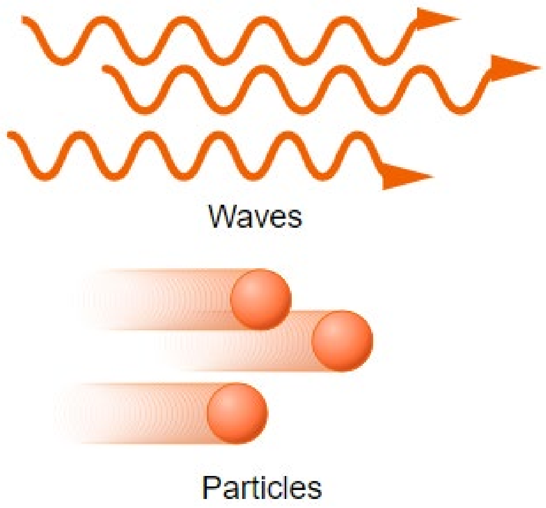


p



d

Electrons



Electrons exhibit properties of both particles and wave.

Electron configuration – arrangement of electrons in an atom.

Three (3) rules govern electron configurations:

Aufbau Principle

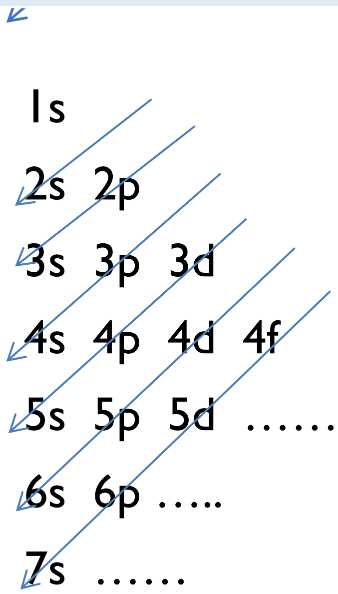
Pauli Exclusion Principle

Hund's Rule

Aufbau Principle

“Electrons in an atom occupy first the lowest possible energy levels and/or orbitals.”

Order of filling: 1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p



Note: 4s atomic orbital is of lower energy than 3p atomic orbital and hence is occupied first.

Notation

Longhand configuration

S $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

core electrons **valence electrons**

Shorthand configuration

S [Ne] $3s^2\ 3p^4$

Valence electrons: electrons located in the outermost shell (valence shell) of the atom.

Lewis symbol

● valence electron

Electron Configurations

Electron arrangements using s, p, d notation for the first 18 elements.

Element	Symbol	Proton No.	Electron Configuration	s, p, d Electron Configuration	Number of electrons in the valence shell
Hydrogen	H	1	1	$1s^1$	1
Helium	He	2	2	$1s^2$	2
Lithium	Li	3	2:1	$1s^2 2s^1$	1
Beryllium	Be	4	2:2	$1s^2 2^2 s^2$	2
Boron	B	5	2:3	$1s^2 2s^2 2p^1$	3
Carbon	C	6	2:4	$1s^2 2s^2 2p^2$	4
Nitrogen	N	7	2:5	$1s^2 2s^2 2p^3$	5
Oxygen	O	8	2:6	$1s^2 2s^2 2p^4$	6
Fluorine	F	9	2:7	$1s^2 2s^2 2p^5$	7
Neon	Ne	10	2:8	$1s^2 2s^2 2p^6$	8
Sodium	Na	11	2:8:1	$1s^2 2s^2 2p^6 3s^1$	1
Magnesium	Mg	12	2:8:2	$1s^2 2s^2 2p^6 3s^2$	2
Aluminum	Al	13	2:8:3	$1s^2 2s^2 2p^6 3s^2 3p^1$	3
Silicon	Si	14	2:8:4	$1s^2 2s^2 2p^6 3s^2 3p^1$	4
Phosphorus	P	15	2:8:5	$1s^2 2s^2 2p^6 3s^2 3p^3$	5
Sulfur	S	16	2:8:6	$1s^2 2s^2 2p^6 3s^2 3p^4$	6
Chlorine	Cl	17	2:8:7	$1s^2 2s^2 2p^6 3s^2 3p^5$	7

Quantum Numbers

Pauli Exclusion Principle: *no two electrons in an atom can have the same four electronic quantum numbers.*

There are **four quantum numbers**:

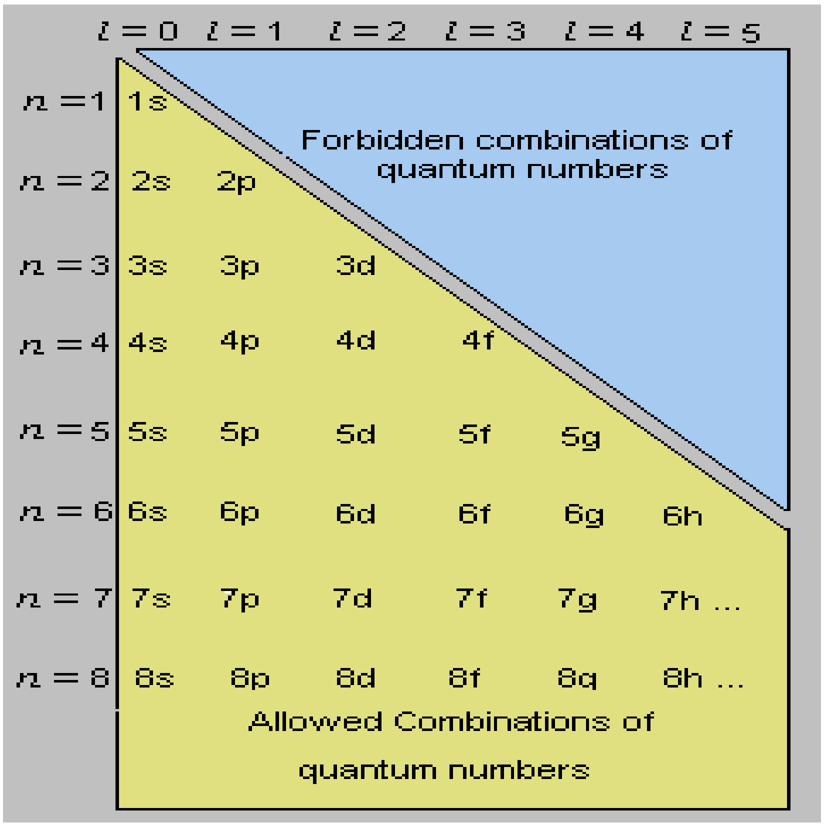
n - principal quantum number - describes the energy level.

ℓ - azimuthal or angular momentum quantum number - describes the subshell.

m_ℓ or m - magnetic quantum number - describes the orbital of the subshell.

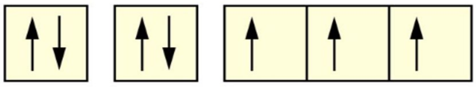
m_s or **s - spin quantum number** - describes the spin.

Na: $1s^2\ 2s^2\ 2p_x^2\ 2p_y^2\ 2p_z^2\ 3s^1$



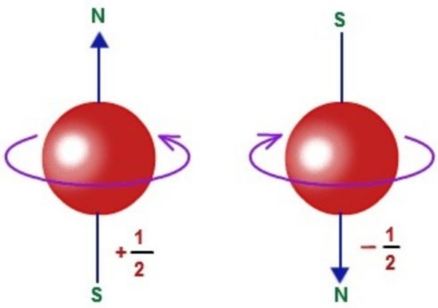
Hundt's Rule

Electrons occupy all empty orbitals in a sub shell with parallel spins before pairing up.

N 

Orbital Box diagram. Direction of arrows indicate electron spin. Opposing spins are paired.

N $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$



Electrons can spin in two directions:

- Spin up
- Spin down

Cr and Cu Electron Configurations

There are two main exceptions to Aufbau Principle: chromium and copper.

Using the Aufbau principle, you would write the following electron configurations:

Cr = [Ar] $4s^2$ $3d^4$
Cu = [Ar] $4s^2$ $3d^9$

The actual electron configurations are:

Cr = [Ar] $4s^1$ $3d^5$
Cu = [Ar] $4s^1$ $3d^{10}$

This results from the fact that:

- completely filled sublevels are more stable than partially filled sublevels.

- half filled is more stable than a partially filled sublevel which is not half full.