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Structure 1.3

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The Electromagnetic spectrum

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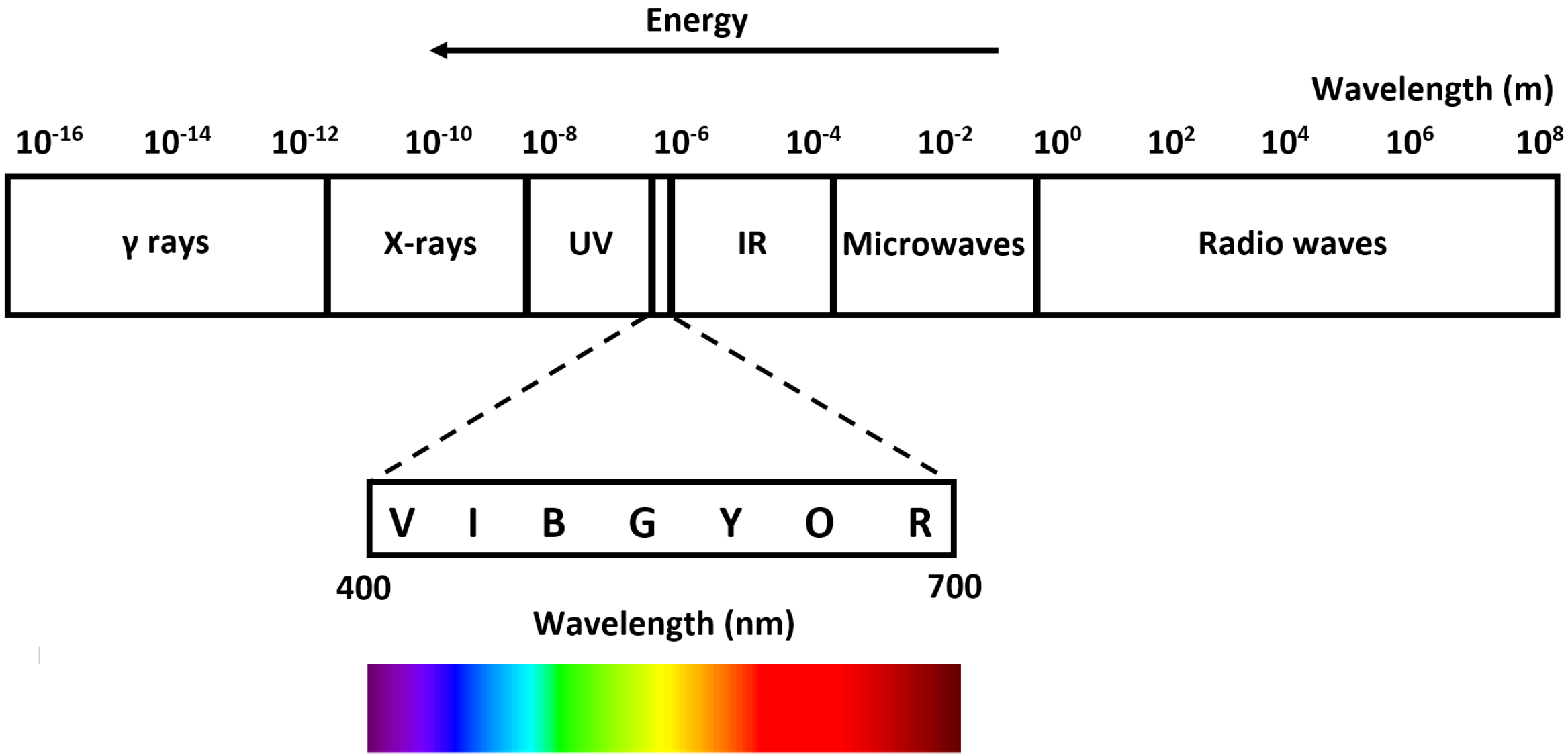
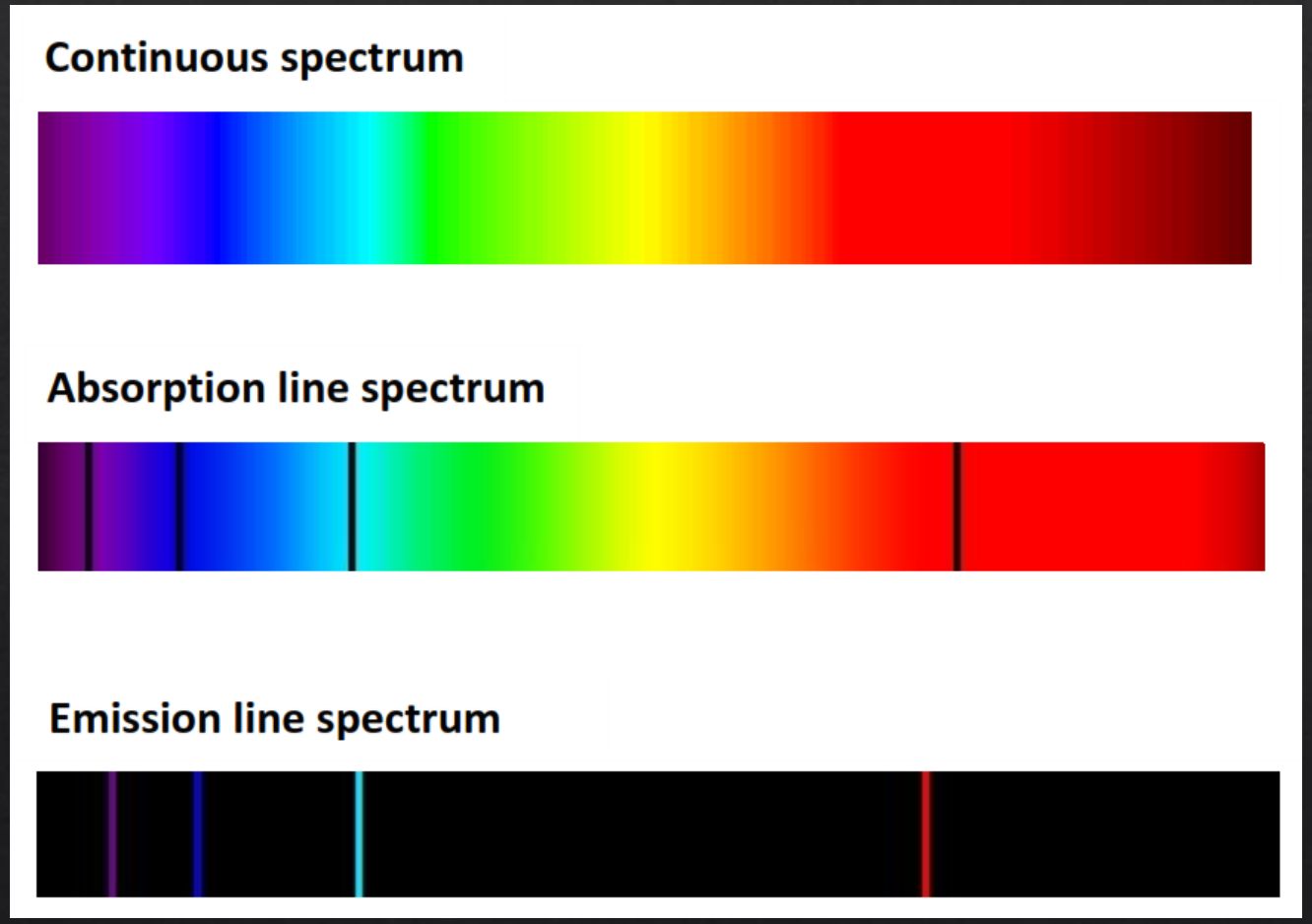
Electromagnetic spectrum



Electromagnetic spectrum

Region	Use/application
Radio waves	^{1}H NMR
Microwaves	Used in a microwave oven
Infrared	IR spectroscopy
Visible light	Enables us to see
Ultraviolet	Used to kill bacteria
X-rays	X-ray crystallography
Gamma rays	Used to treat cancer

Line spectra



- Continuous spectrum: shows all the wavelengths of visible light.
- Absorption line spectrum: black lines on coloured background (certain wavelengths missing).
- Emission line spectrum: coloured lines on a black background (only certain wavelengths visible).

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Hydrogen emission spectrum

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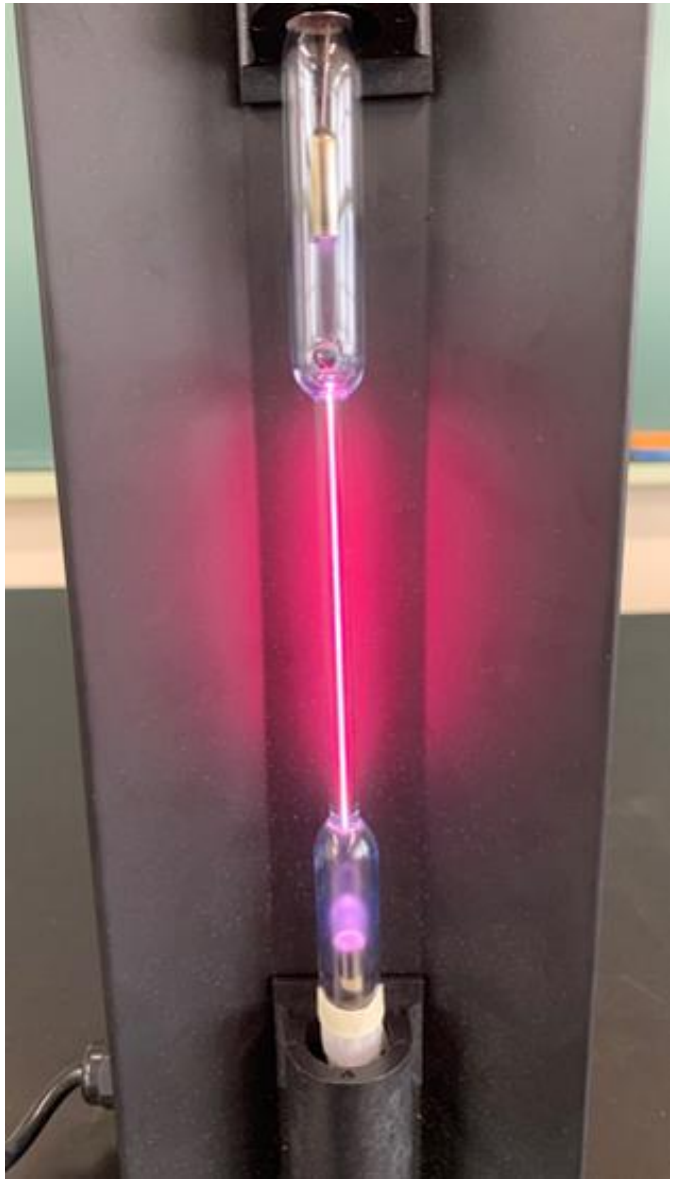
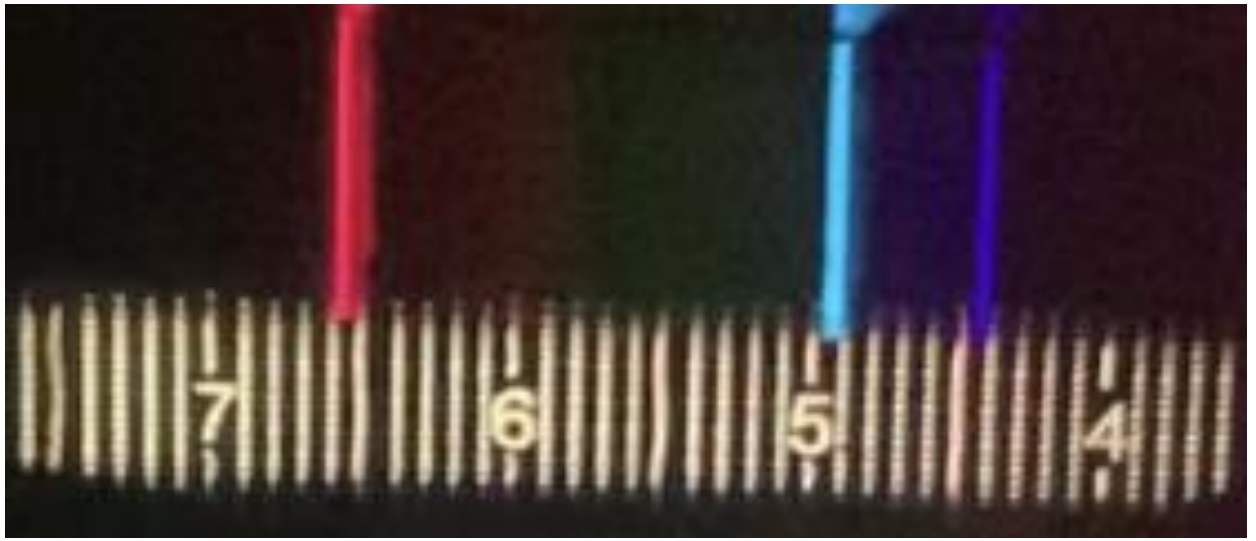
Hydrogen emission spectrum





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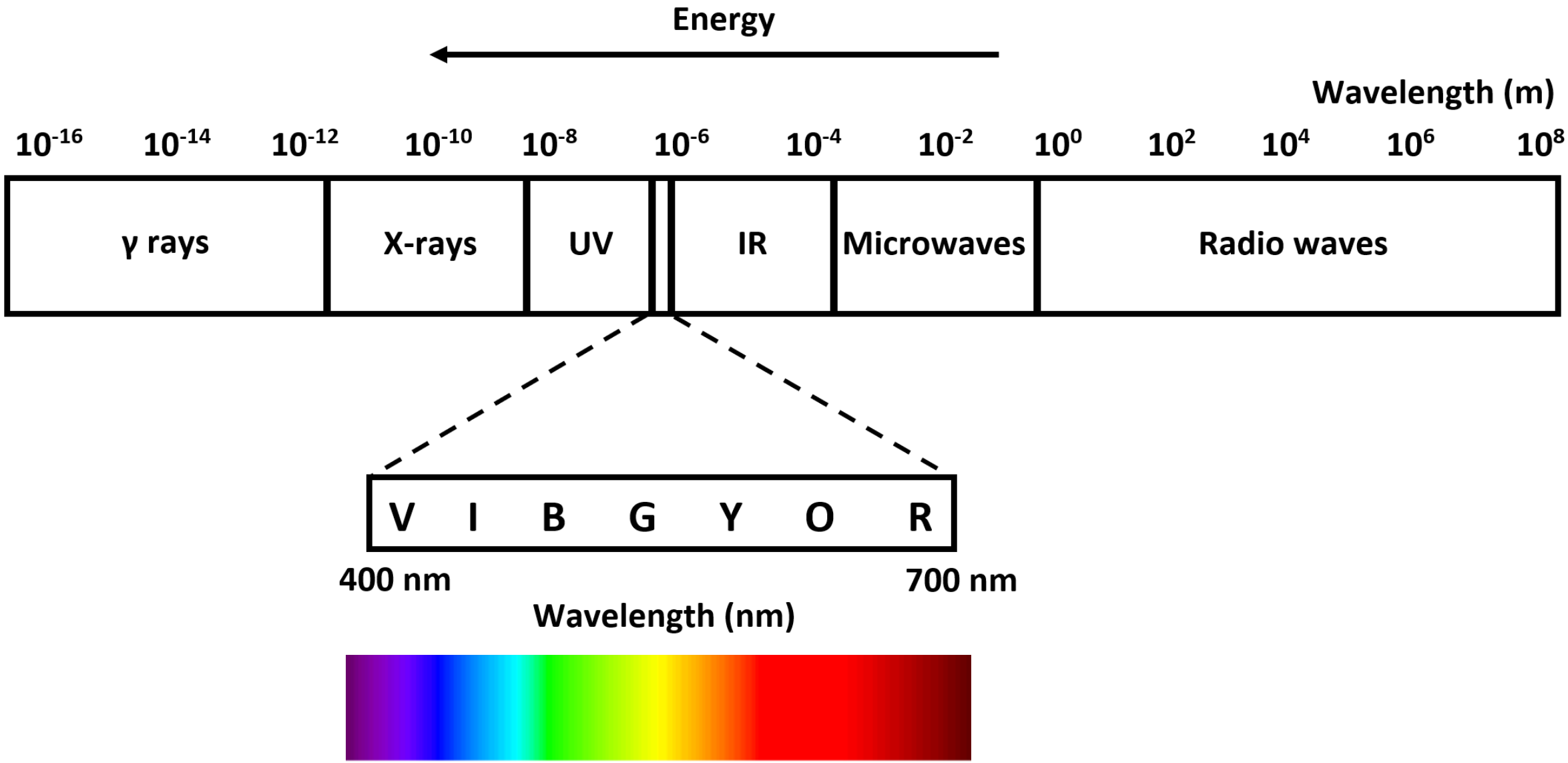
Hydrogen emission spectrum



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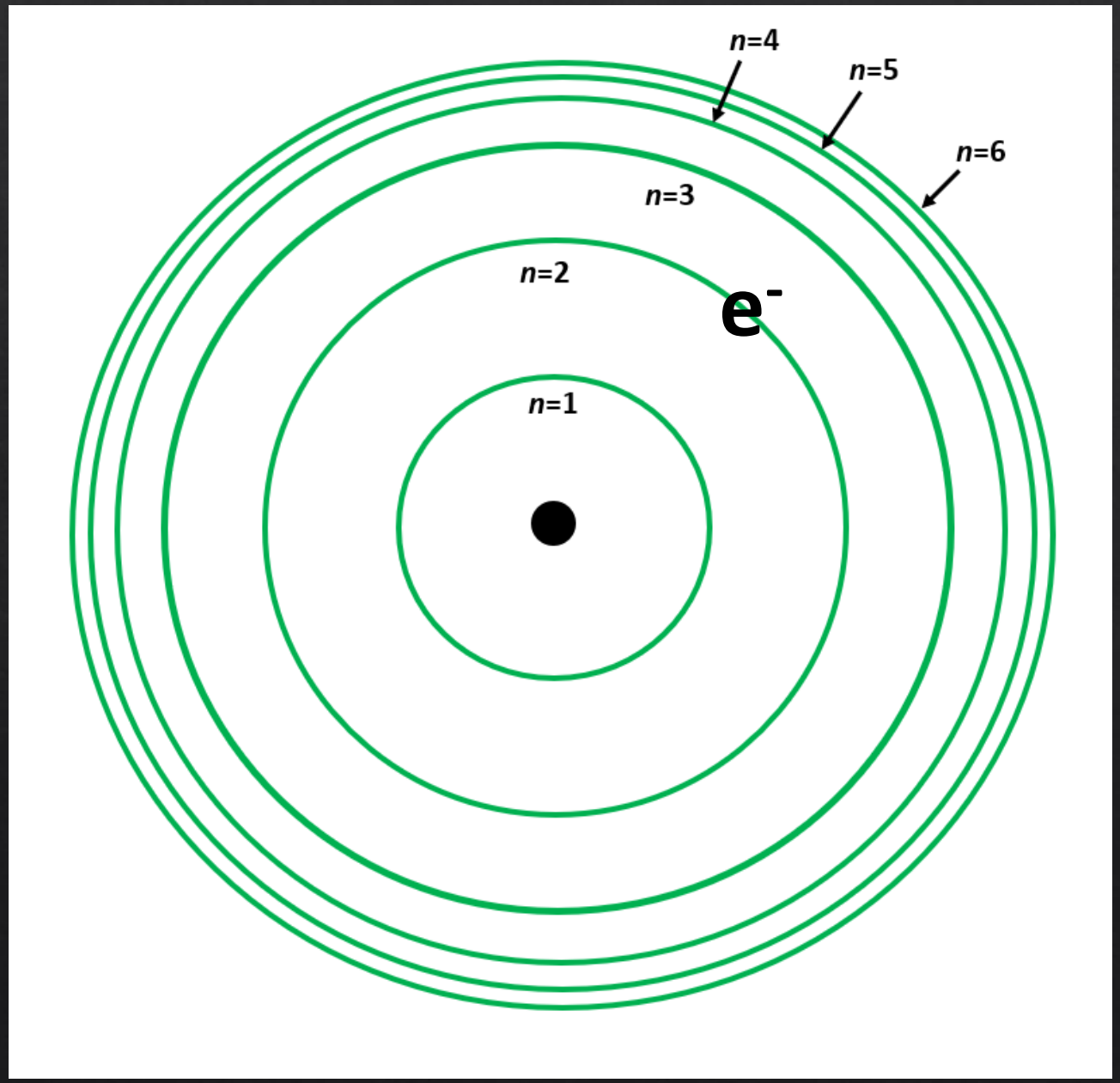
Hydrogen emission spectrum



Electrons can transition between energy levels by either absorbing or emitting photons of energy.

When electrons absorb photon of energy, they transition to higher energy levels.

When electrons transition to lower energy levels, they emits photons of energy.

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Hydrogen emission spectrum

Electron transition	Wavelength (nm)	Energy (J)
n=6 to n=2	410	4.85×10^{-19}
n=5 to n=2	434	4.58×10^{-19}
n=4 to n=2	486	4.09×10^{-19}
n=3 to n=2	656	3.03×10^{-19}

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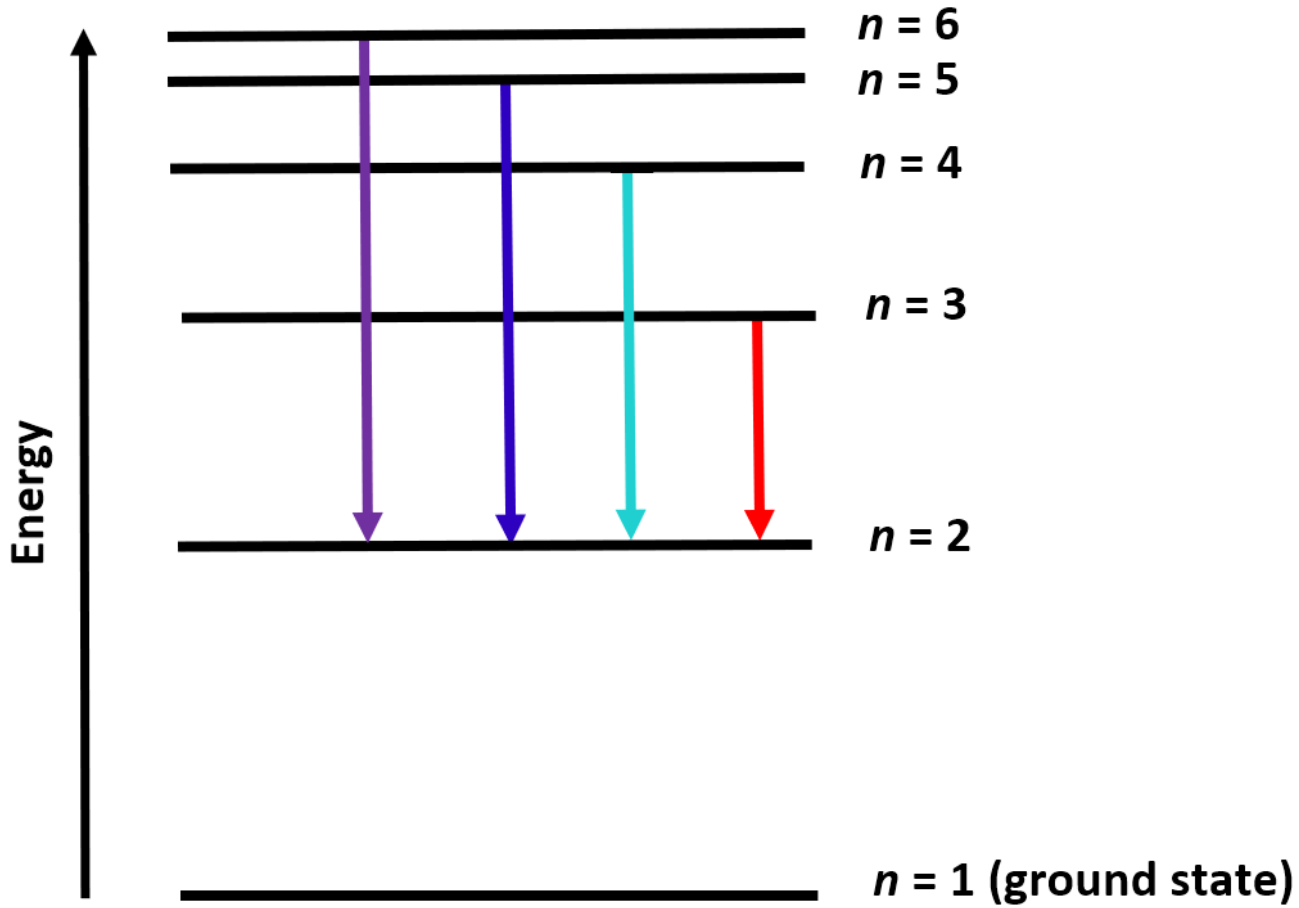
Hydrogen emission spectrum



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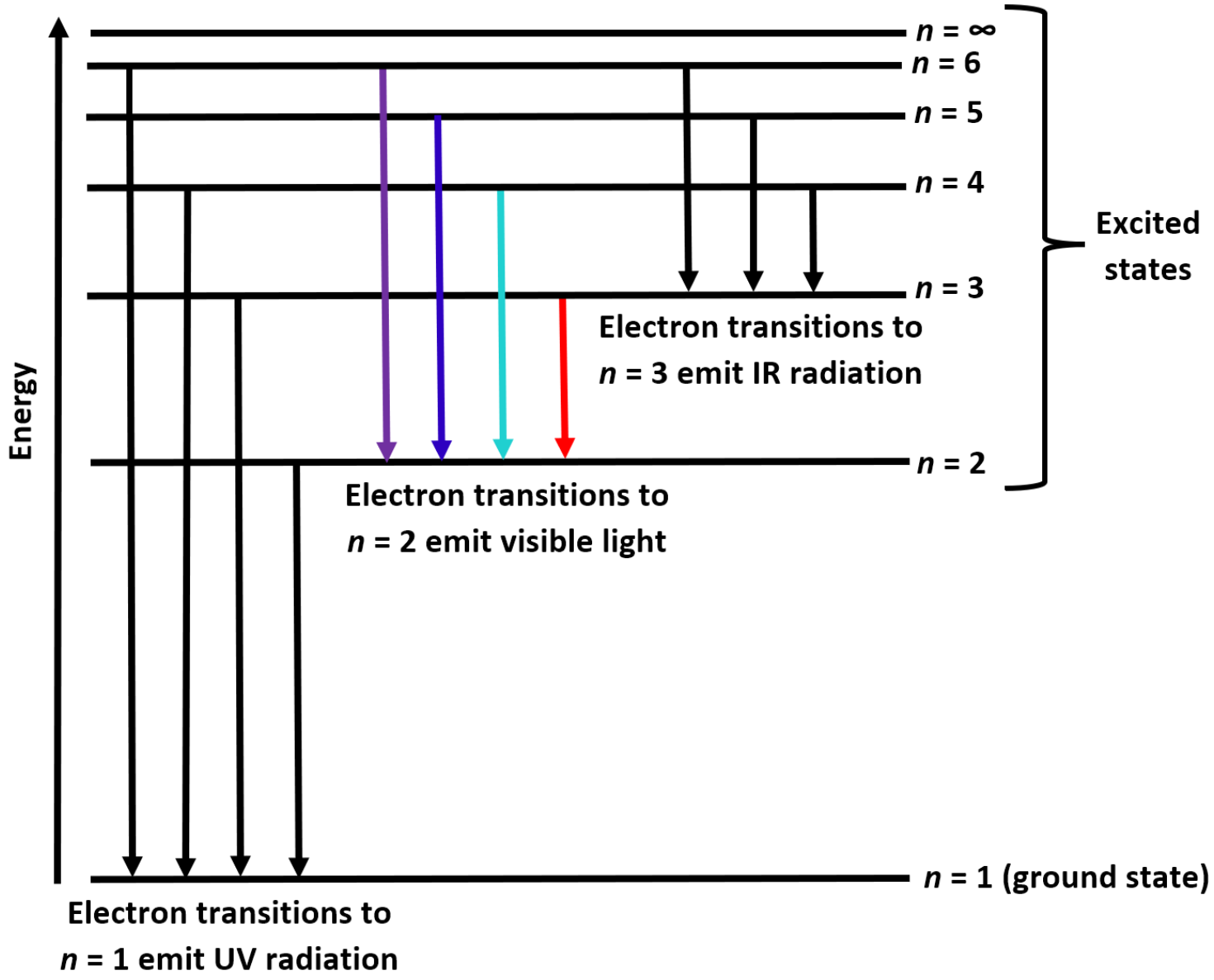
Hydrogen emission spectrum



Electron transitions to *n*=1 emit UV radiation.

Electron transitions to *n*=2 emit visible light.

Electron transitions to *n*=3 emit infrared radiation.

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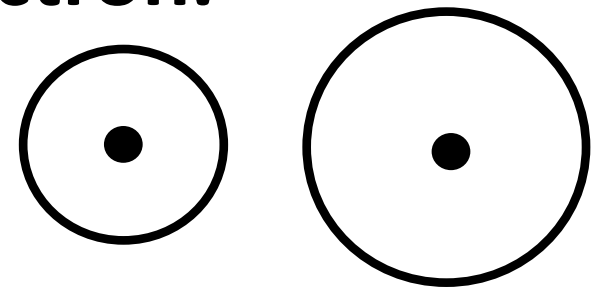
Atomic orbitals and sub-levels

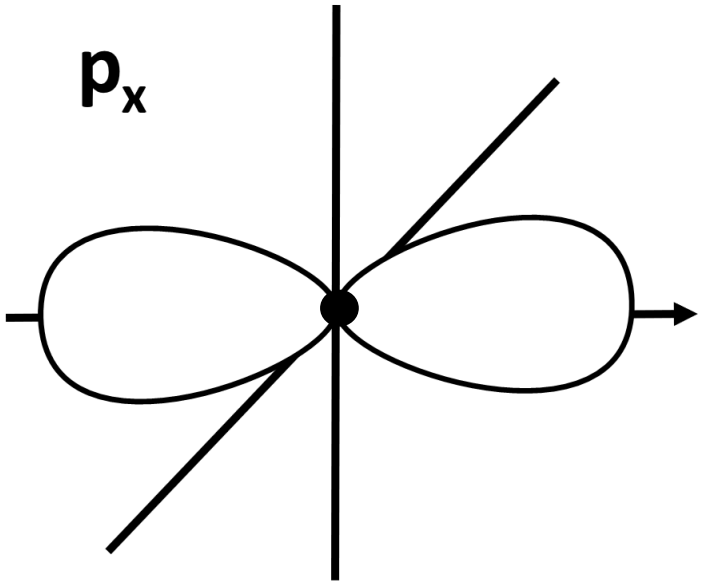
Atomic orbitals

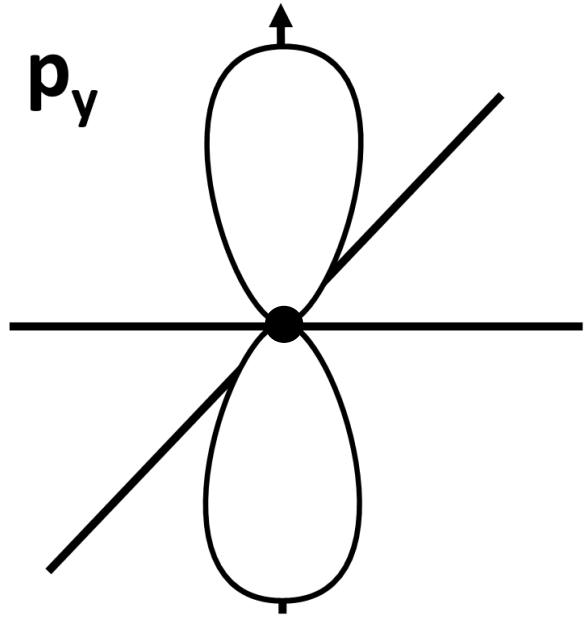
An atomic orbital is a region of space where there is a high probability of finding an electron.

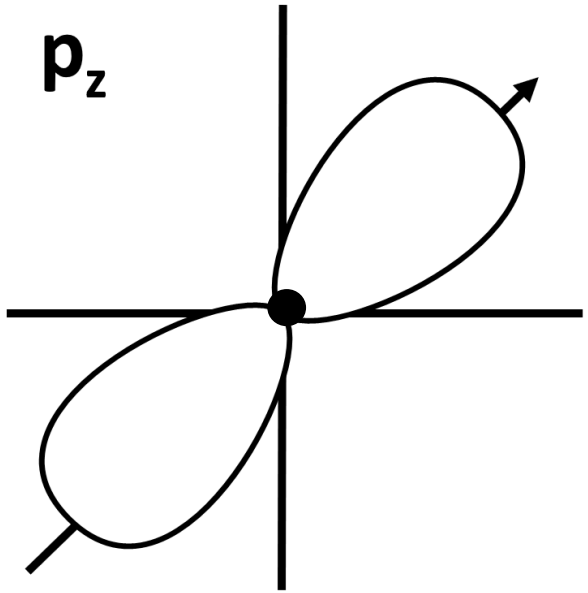
s orbitals are spherical.



p orbitals are dumbbell shaped.







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Sub-levels

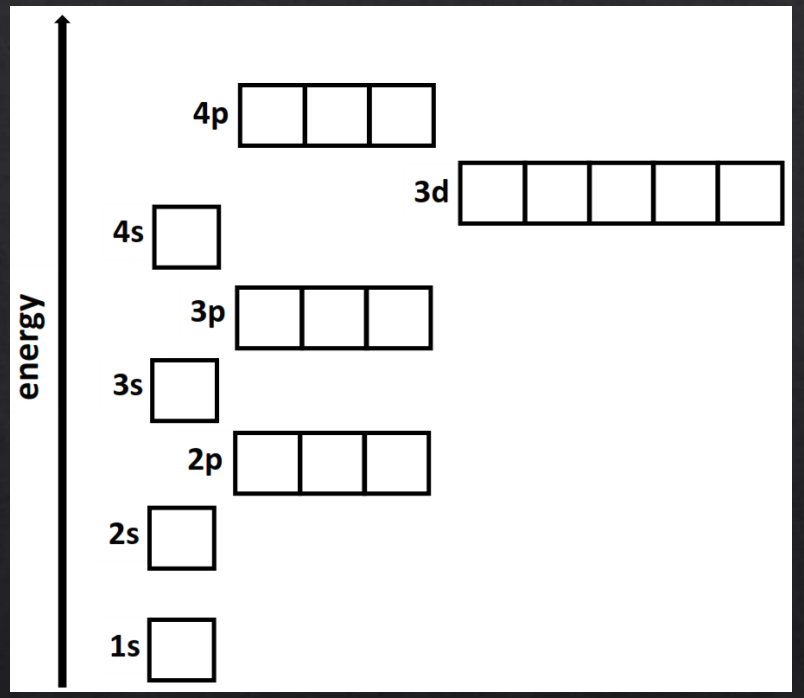
Each principal energy level is split into sub-levels.

n=1 has 1 sub-level (1s)

n=2 has 2 sub-levels (2s, 2p)

n=3 has 3 sub-levels (3s, 3p, 3d)

n=4 has 4 sub-levels (4s, 4p, 4d, 4f)



Within a main energy level, the order of energy is:

s < p < d < f

Sub-levels

Principal energy level (*n*)	Sub-levels	Number of electrons in sub-level	Number of orbitals	Total number of electrons in energy level
1	1s	2	1	2
2	2s	2	4	8
	2p	6		
3	3s	2	9	18
	3p	6		
	3d	10		
4	4s	2	16	32
	4p	6		
	4d	10		
	4f	14		

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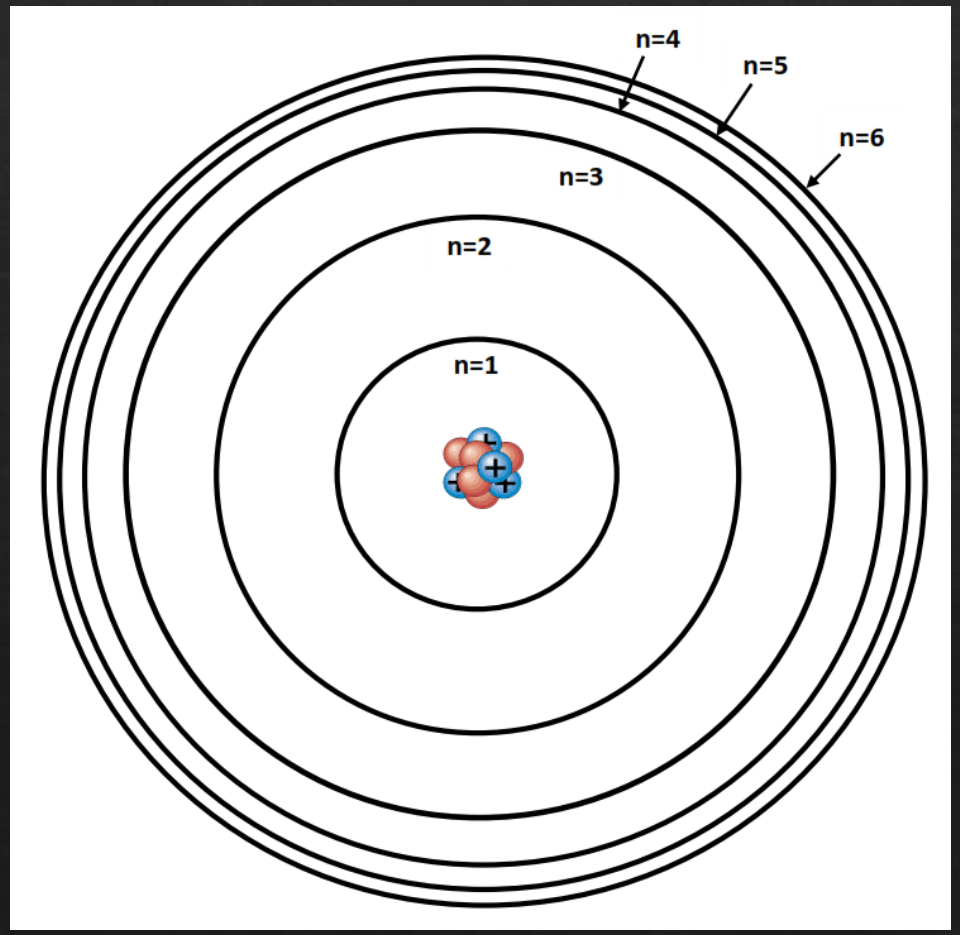
Principal energy levels

Principal energy levels (n) in an atom



Electrons are located in principal energy levels (main energy levels). The first main energy level (n=1) has the lowest energy and energy increases as the value of n increases. Each main energy level can hold a maximum of $2n^2$ electrons.

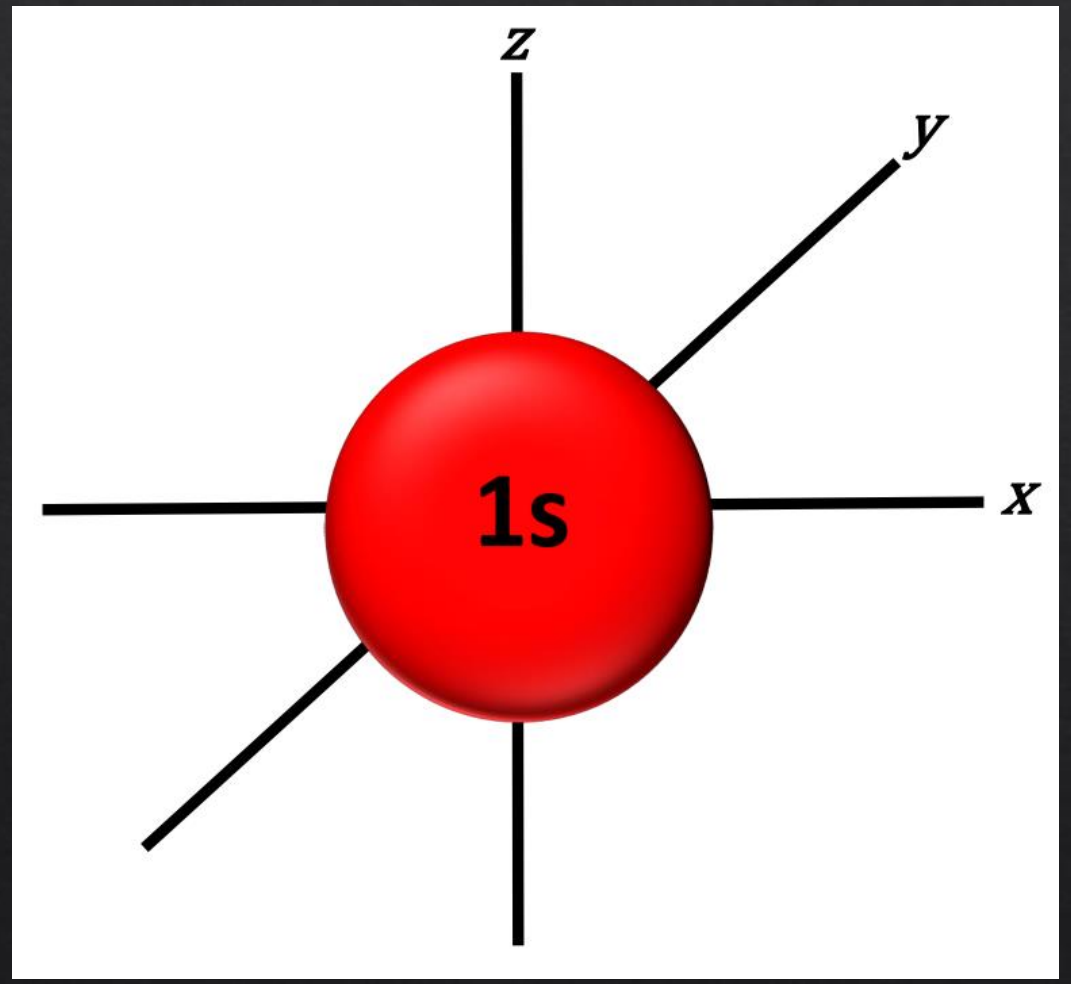
n=2 $(2 \times 2^2) = 8$ electrons

n=3 $(2 \times 3^2) = 18$ electrons

Sub-levels in the atom

Principal energy level, *n*	Sub-levels	Number of orbitals	Number of electrons
1	1s	1	2
2	2s, 2p	4	8
3	3s, 3p, 3d	9	18
4	4s, 4p, 4d, 4f	16	32

Atomic orbitals



s orbitals are spherical. Each s orbital can hold a maximum of 2 electrons with opposite spins, according to the Pauli exclusion principle.

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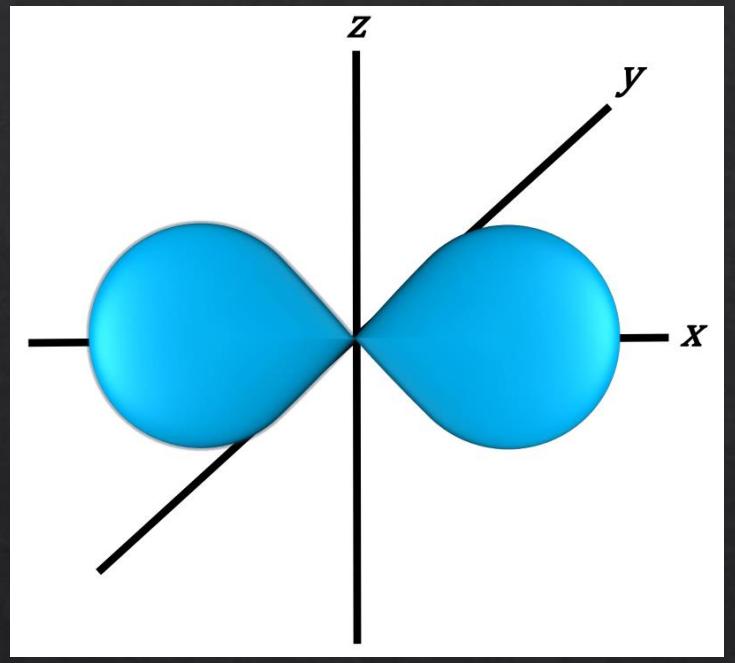
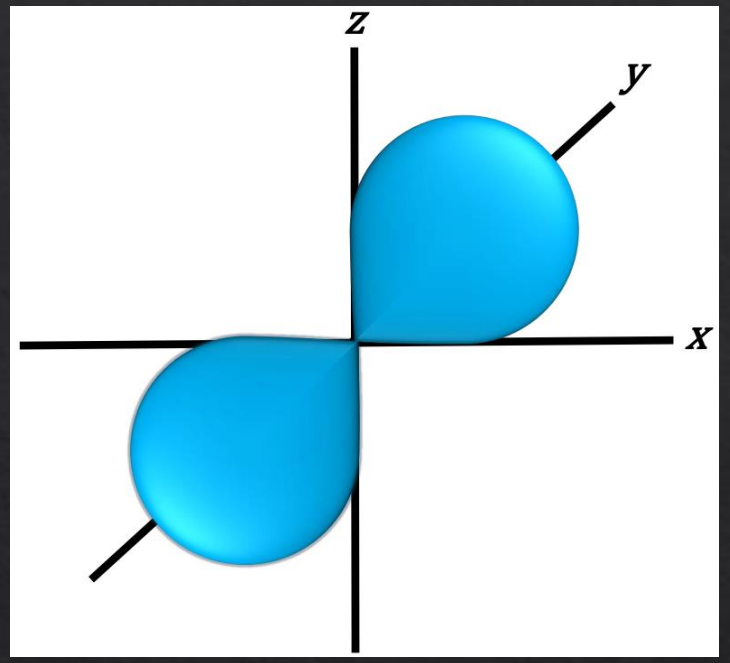
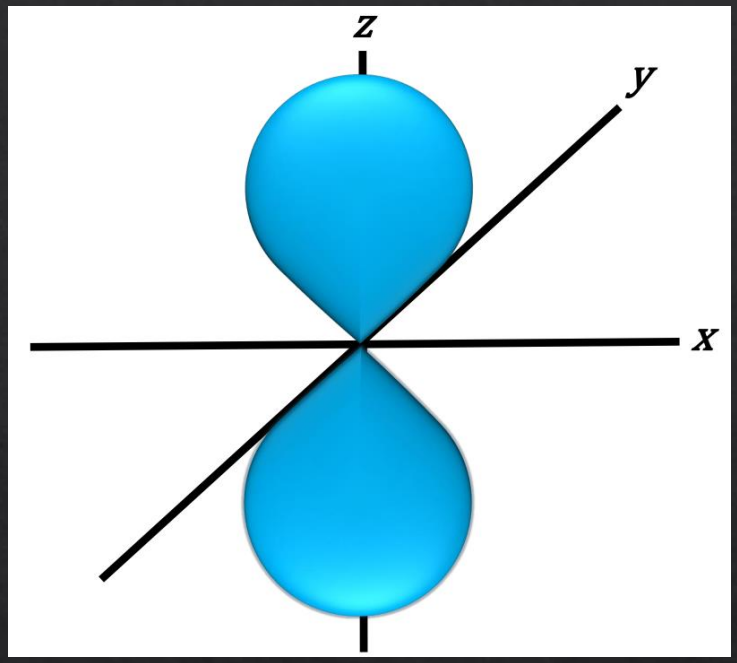
Atomic orbitals



p_x



p_y



p_z

The p orbitals are 'dumbbell' shaped.
Each p orbital can hold a maximum of 2 electrons with opposite spins (total of 6 electrons in the p sub-level).

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Electron configurations and the Aufbau principle

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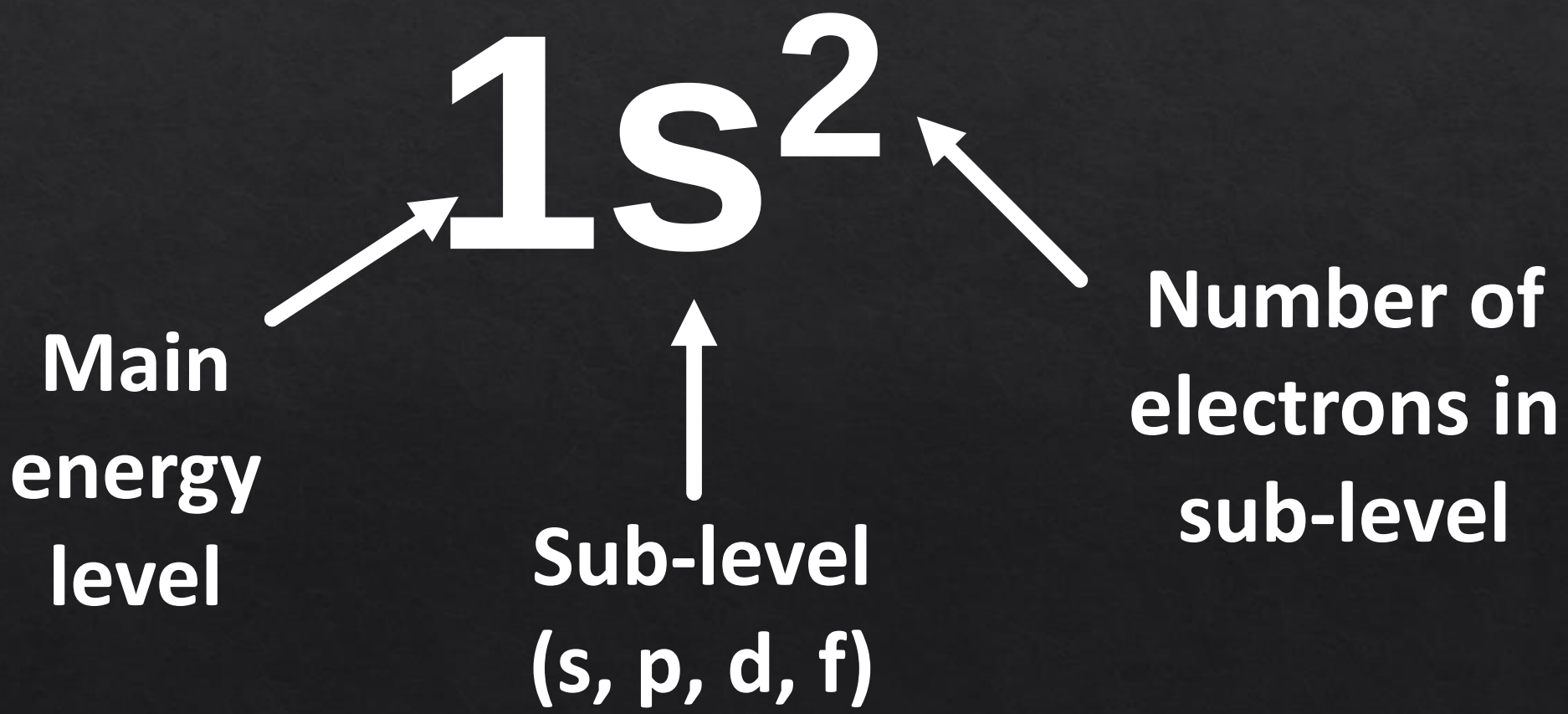
Electron configurations

An electron configuration is a representation of the arrangement of electrons within an atom or ion.



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Electron configurations

Electron configuration of carbon, C

$1s^2\ 2s^2\ 2p^2$ $[He]\ 2s^2\ 2p^2$

Electron configuration of sodium, Na

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

Electron configuration of bromine, Br

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$

$[Ar]\ 4s^2\ 3d^{10}\ 4p^5$

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Electron configurations

Notation	Electron configuration (core electrons)
[He]	$1s^2$
[Ne]	$1s^2\ 2s^2\ 2p^6$
[Ar]	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
[Kr]	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

Electron configuration for sodium, Na

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

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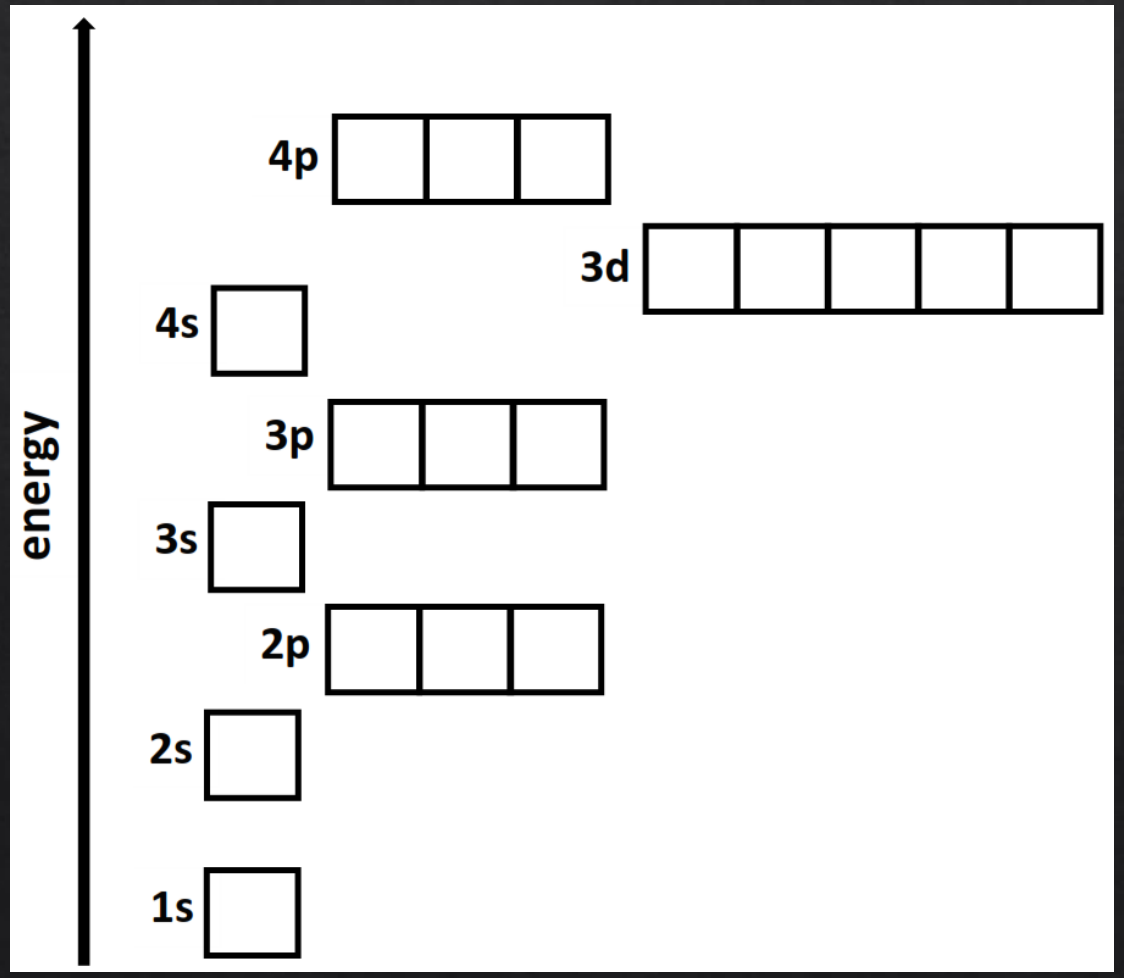
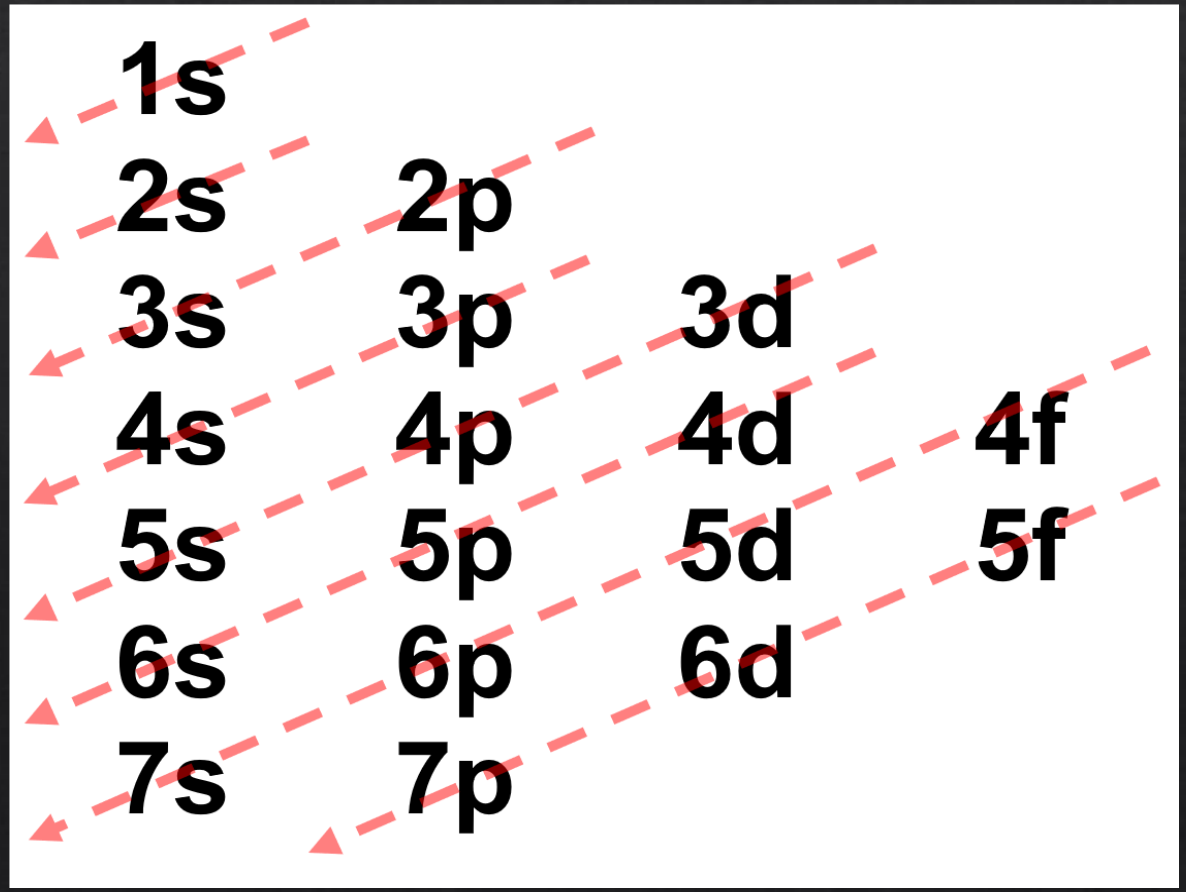
The Aufbau principle

The Aufbau principle is used to determine the electron configuration of an atom or ion.





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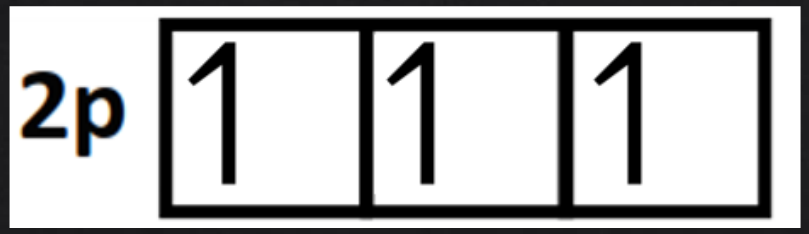
The Aufbau principle

Pauli exclusion principle

An atomic orbital can hold a maximum of two electrons and they must have opposite spins.

Hund's rule

Degenerate orbitals are filled singly with the same spin before being doubly occupied.



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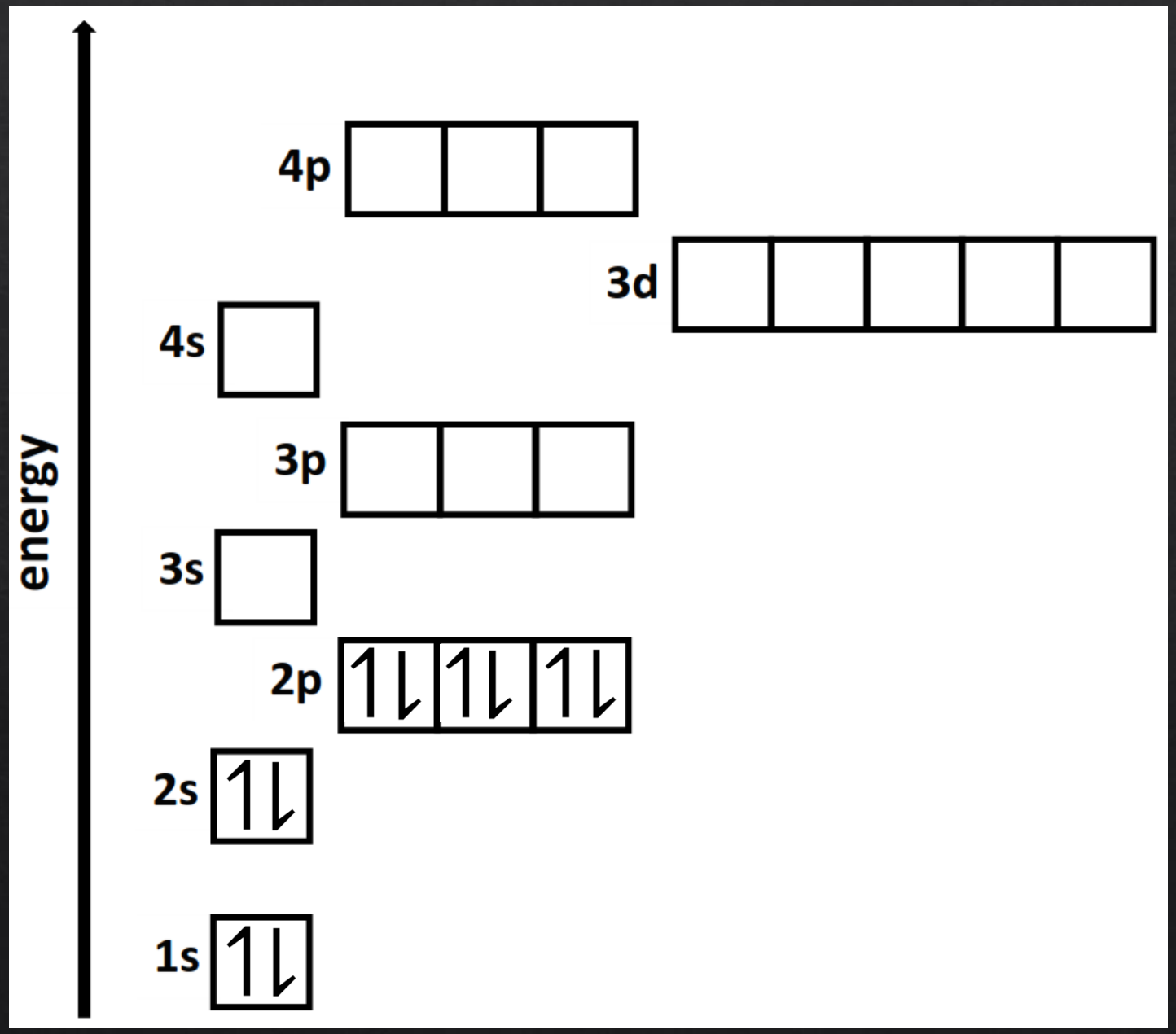
Electron configurations



H	$1s^1$
He	$1s^2$
Li	[He] $2s^1$
Be	[He] $2s^2$
B	[He] $2s^2\ 2p^1$
C	[He] $2s^2\ 2p^2$
N	[He] $2s^2\ 2p^3$
O	[He] $2s^2\ 2p^4$
F	[He] $2s^2\ 2p^5$
Ne	[He] $2s^2\ 2p^6$

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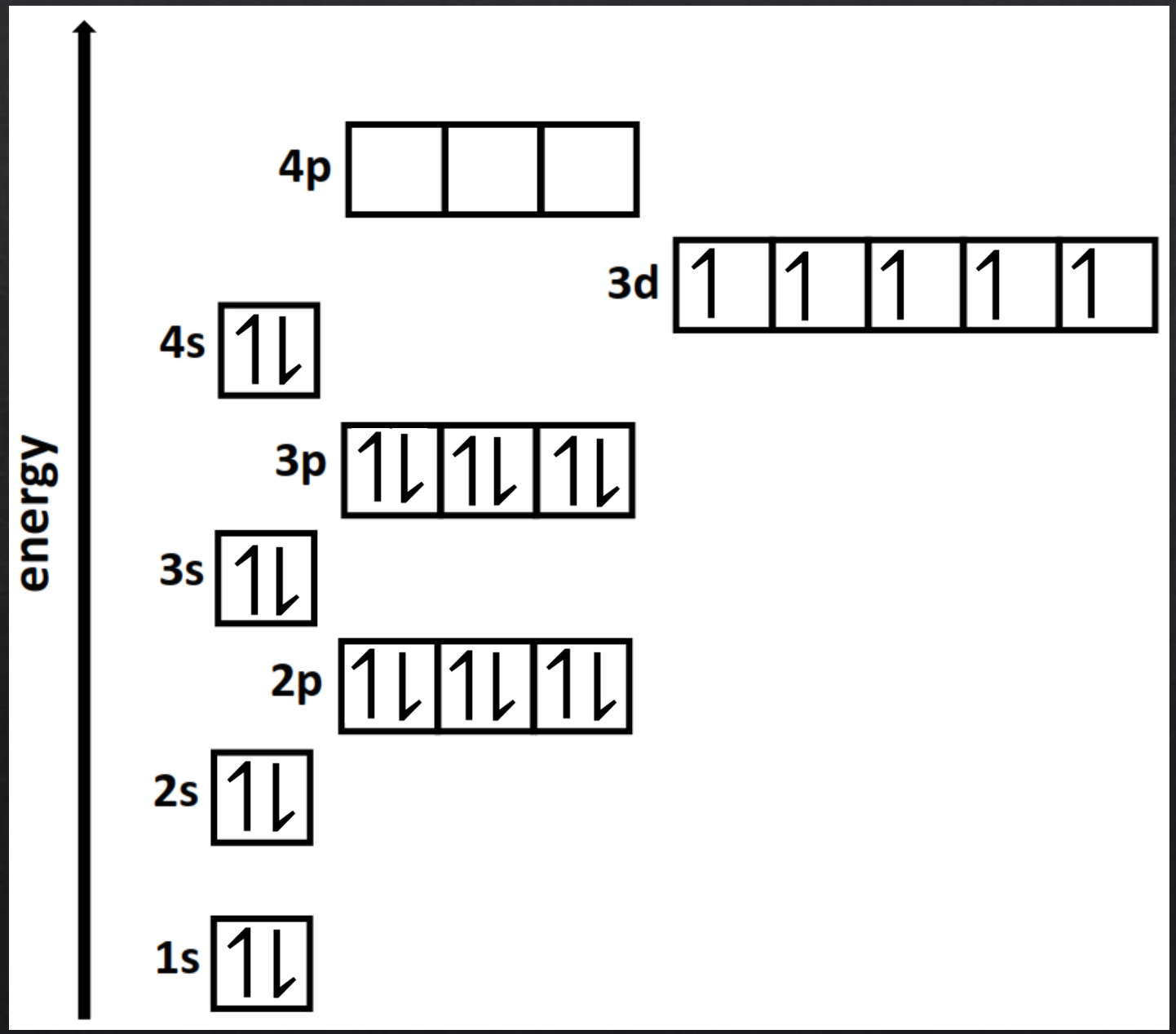
Electron configurations



Na [Ne] $3s^1$

Mg [Ne] $3s^2$

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

Si [Ne] $3s^2\ 3p^2$

P [Ne] $3s^2\ 3p^3$

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

Cl [Ne] $3s^2\ 3p^5$

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

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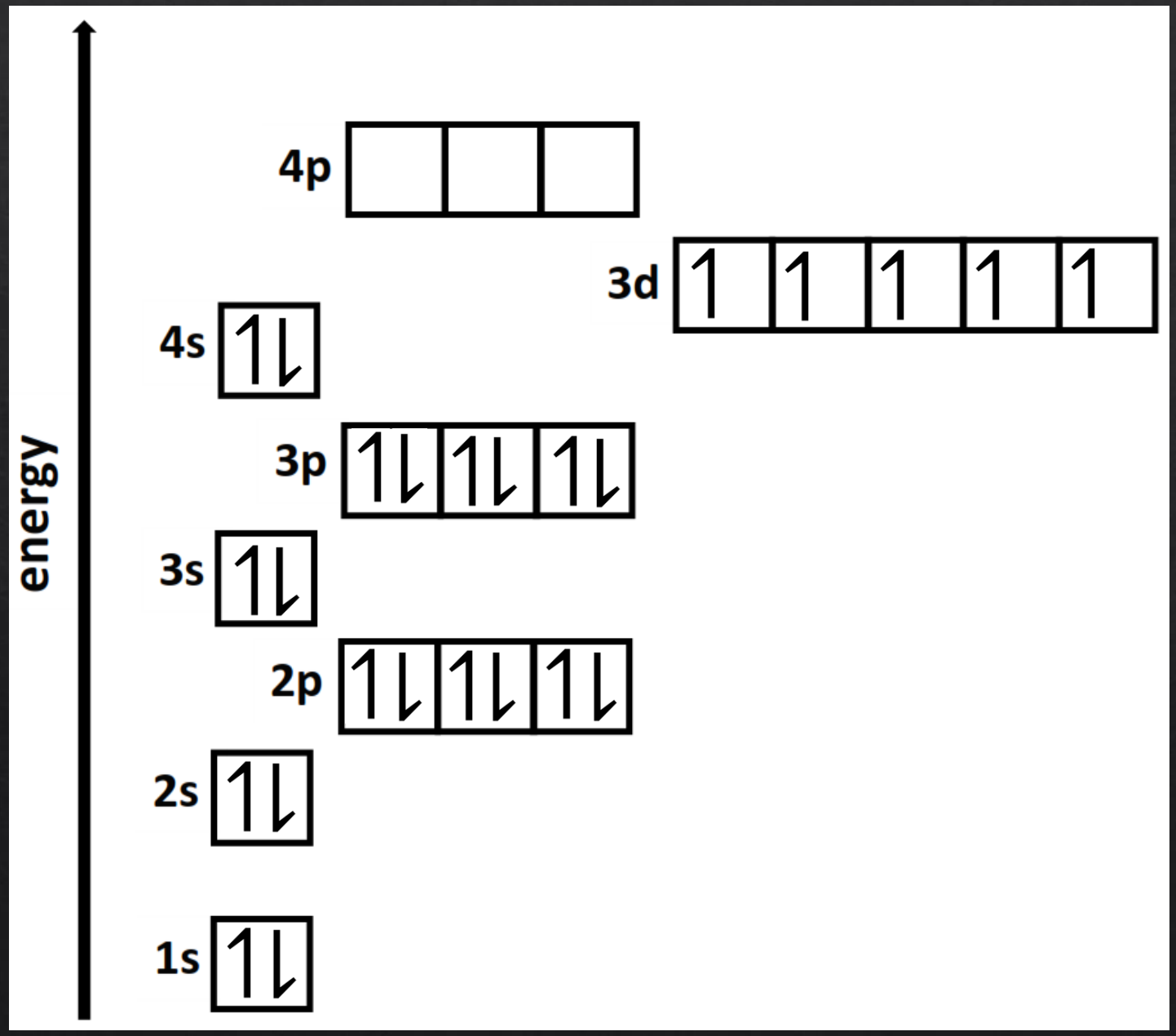
Electron configurations



K	[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] $4s^1$ $3d^5$
Mn	[Ar] $4s^2$ $3d^5$

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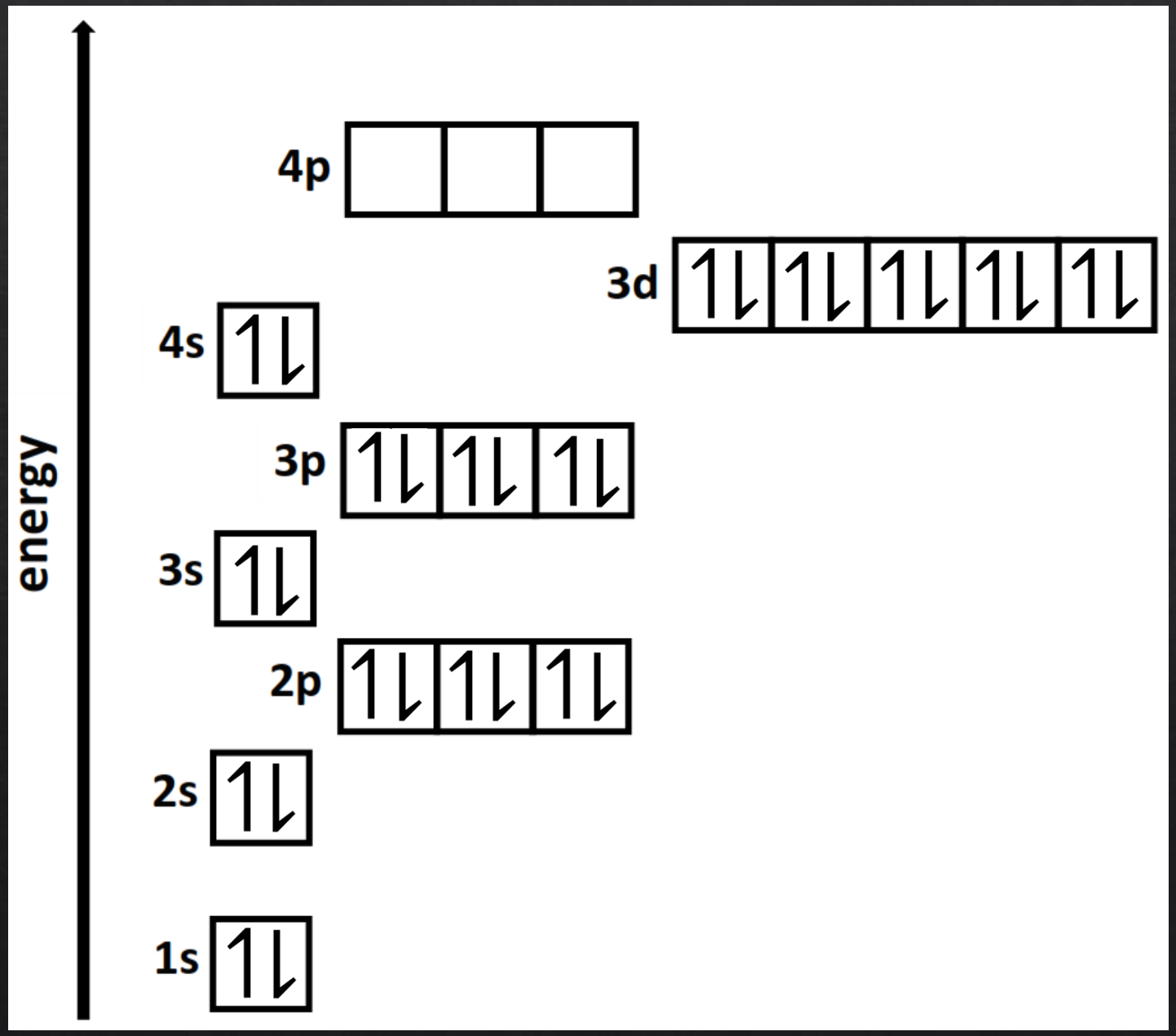
Electron configurations



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

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

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

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

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

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Electron configurations

Ga [Ar] $4s^2\ 3d^{10}\ 4p^1$

Ge [Ar] $4s^2\ 3d^{10}\ 4p^2$

As [Ar] $4s^2\ 3d^{10}\ 4p^3$

Se [Ar] $4s^2\ 3d^{10}\ 4p^4$

Br [Ar] $4s^2\ 3d^{10}\ 4p^5$

Kr [Ar] $4s^2\ 3d^{10}\ 4p^6$

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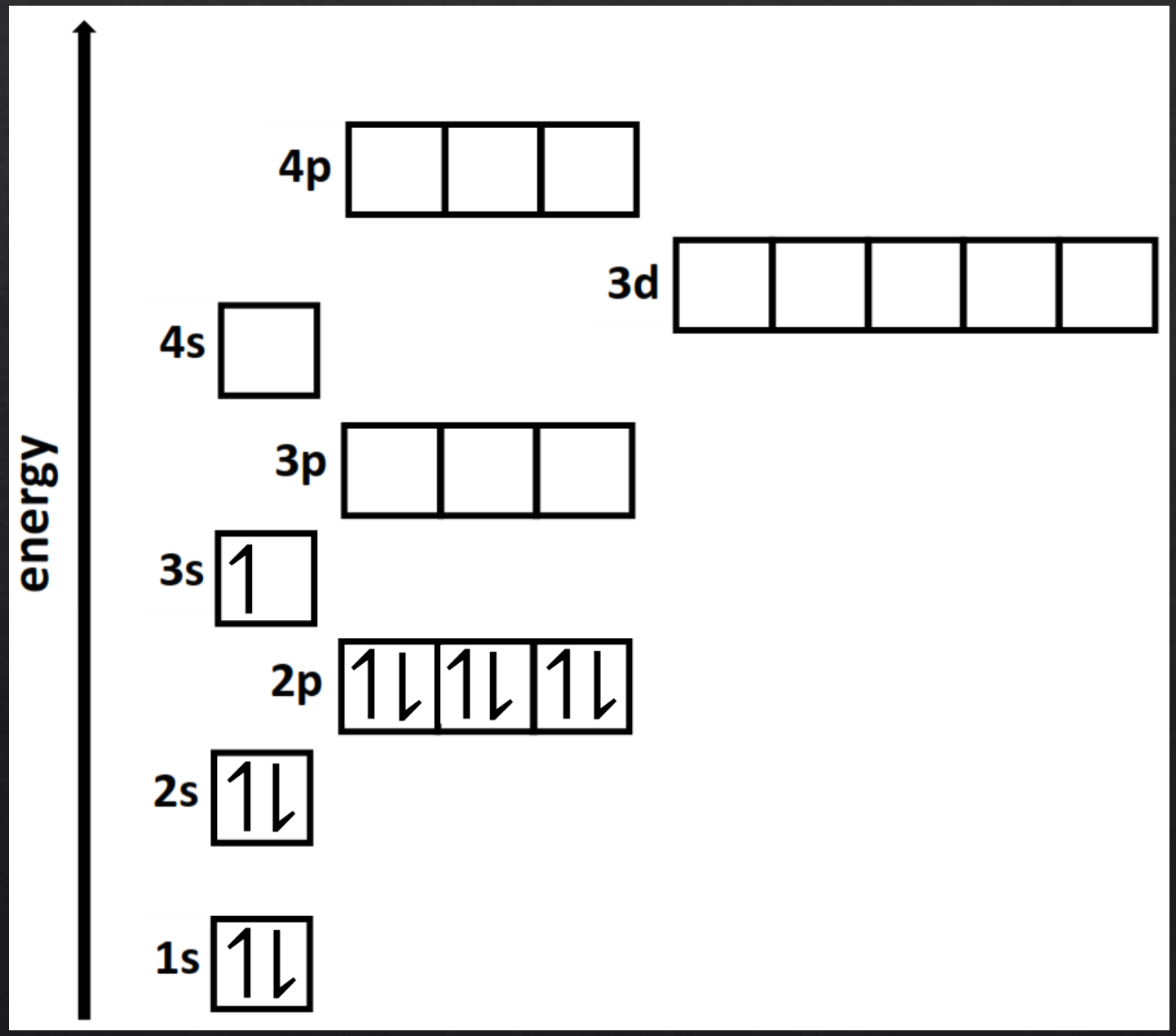
The Aufbau principle



Sodium (Na)

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

Pauli exclusion principle – an atomic orbital can hold two electrons with opposite spins.

Hund's rule – degenerate orbitals in a sub-level are singly occupied before being doubly occupied, and the electrons have the same spin.

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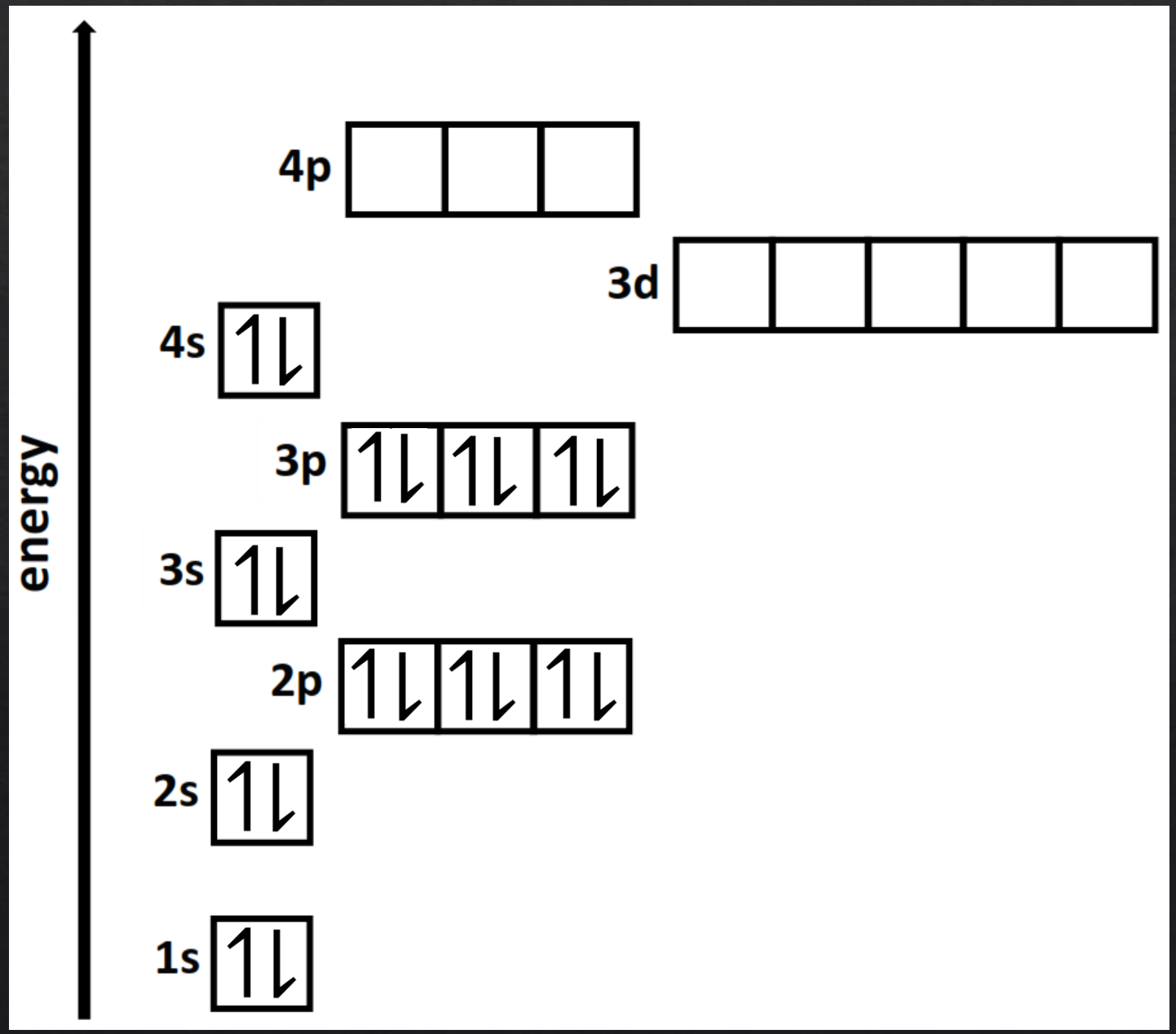
The Aufbau principle



Calcium (Ca)

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

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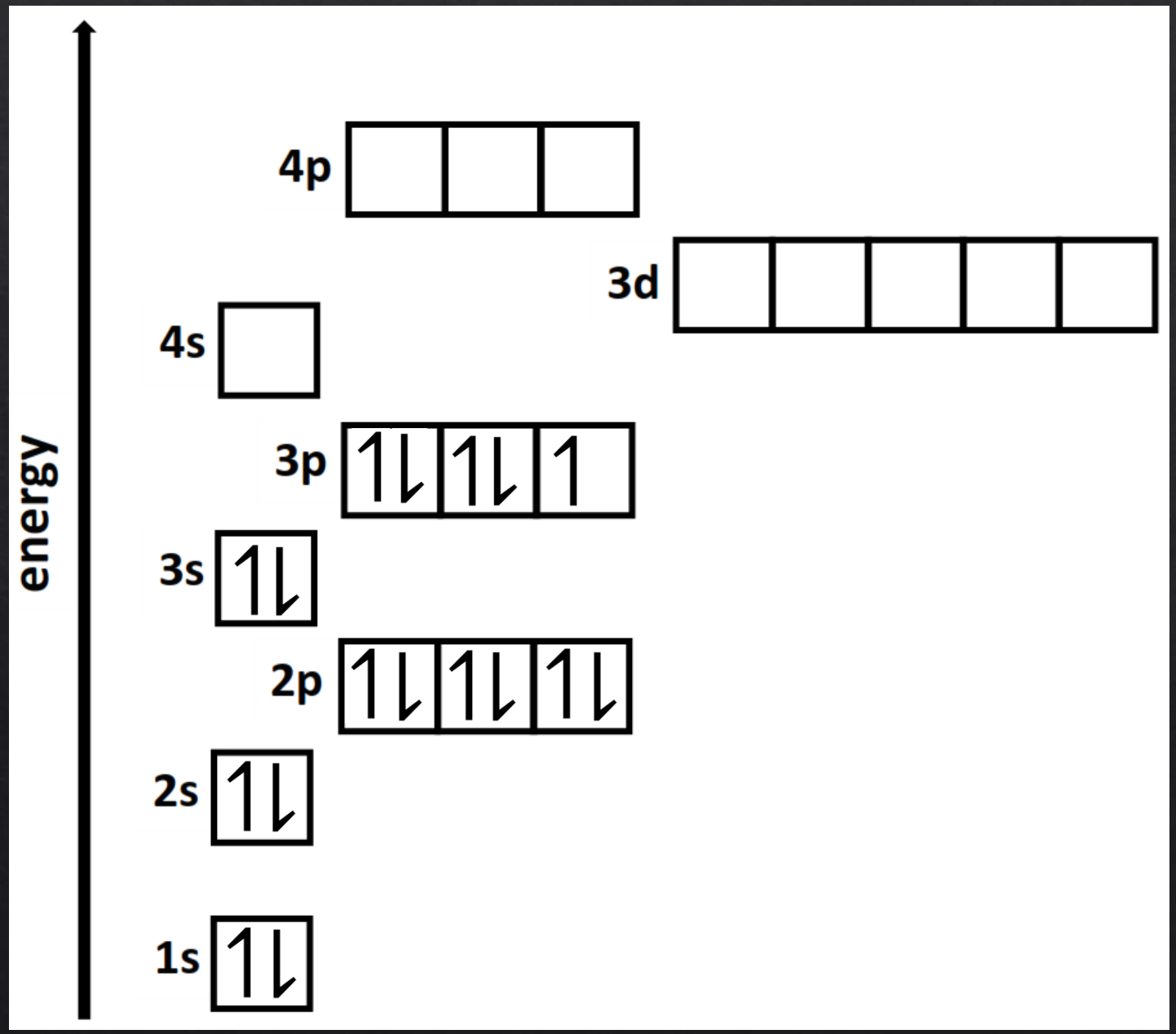
The Aufbau principle



Chlorine (Cl)

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

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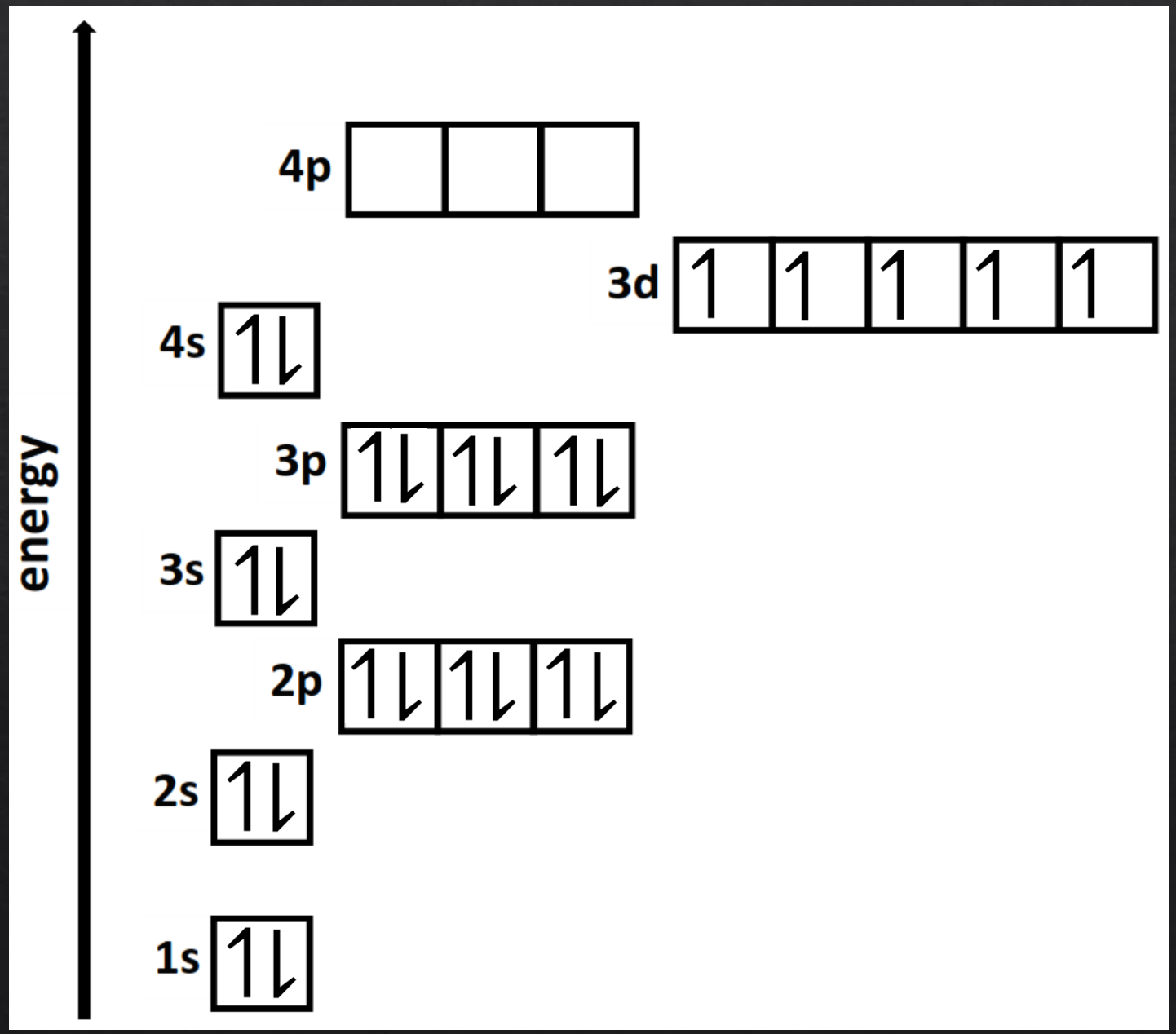
The Aufbau principle



Manganese (Mn)

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$

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Electron configurations for elements 1-36

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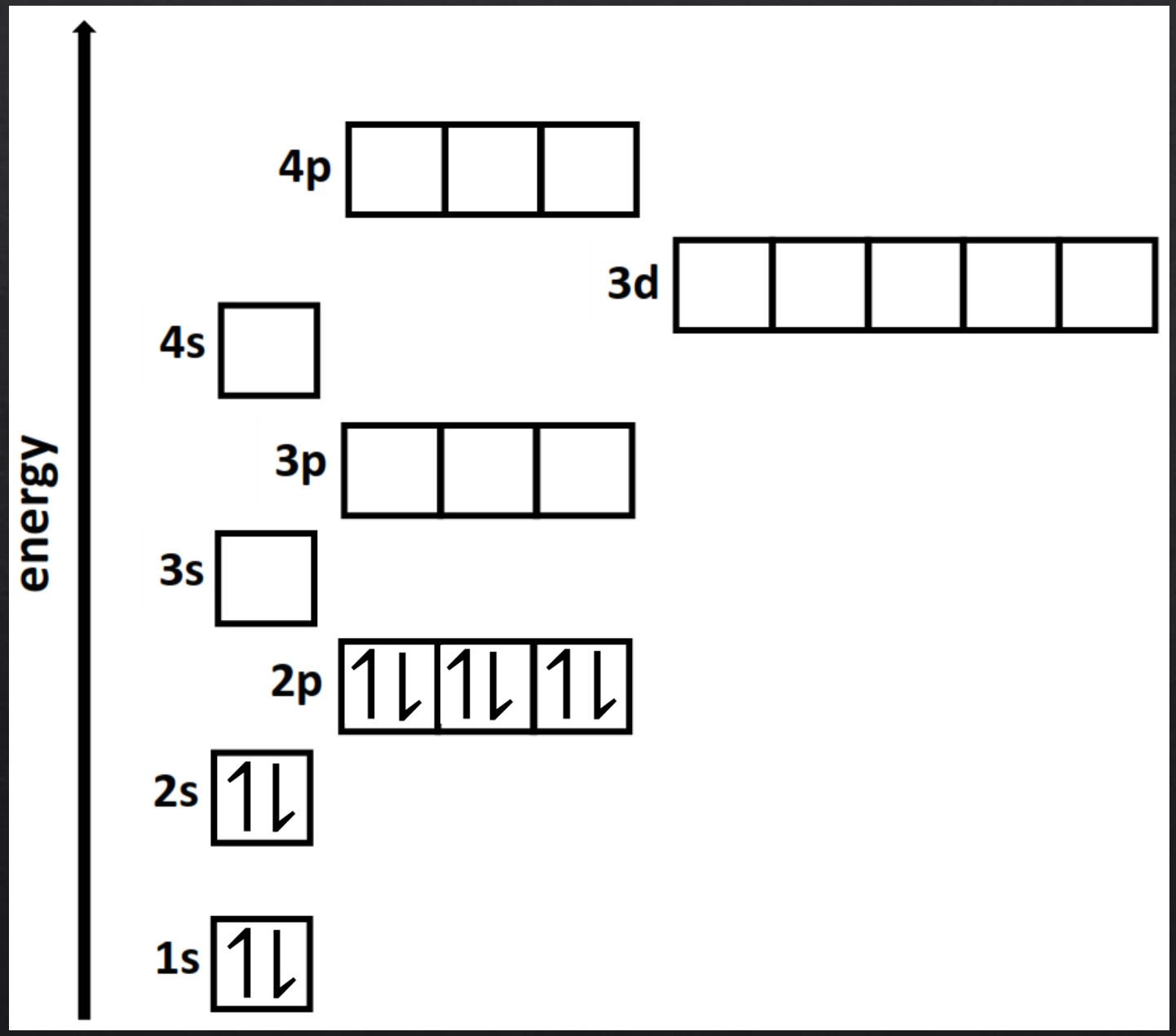
Electron configurations



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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	$1s^2\ 2s^2\ 2p^6$

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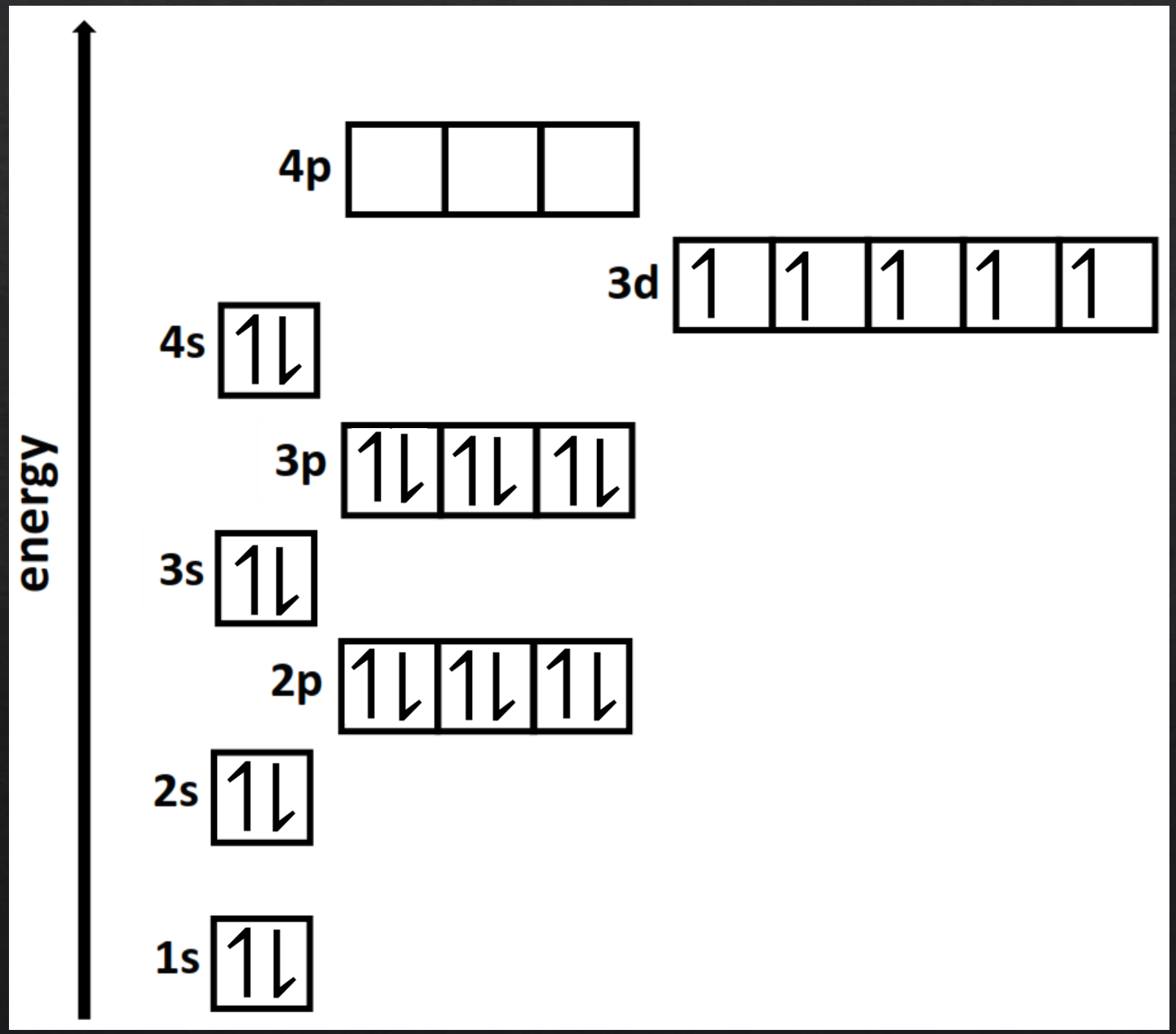
Electron configurations



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

Si [Ne] $3s^2\ 3p^2$

P [Ne] $3s^2\ 3p^3$

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

Cl [Ne] $3s^2\ 3p^5$

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

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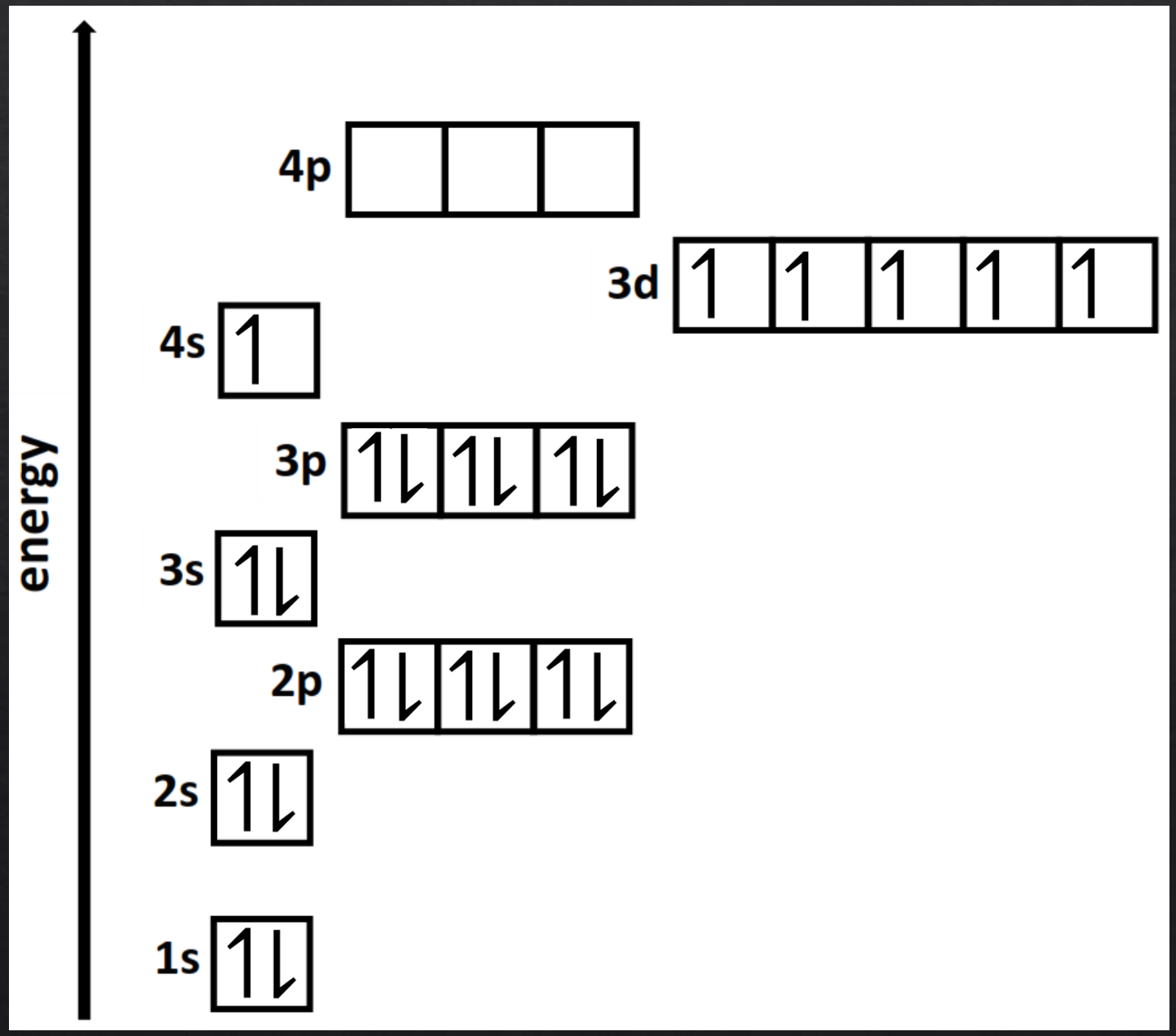
Electron configurations



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Sc	[Ar] $4s^2$ $3d^1$
Ti	[Ar] $4s^2$ $3d^2$
V	[Ar] $4s^2$ $3d^3$
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Mn	[Ar] $4s^2$ $3d^5$

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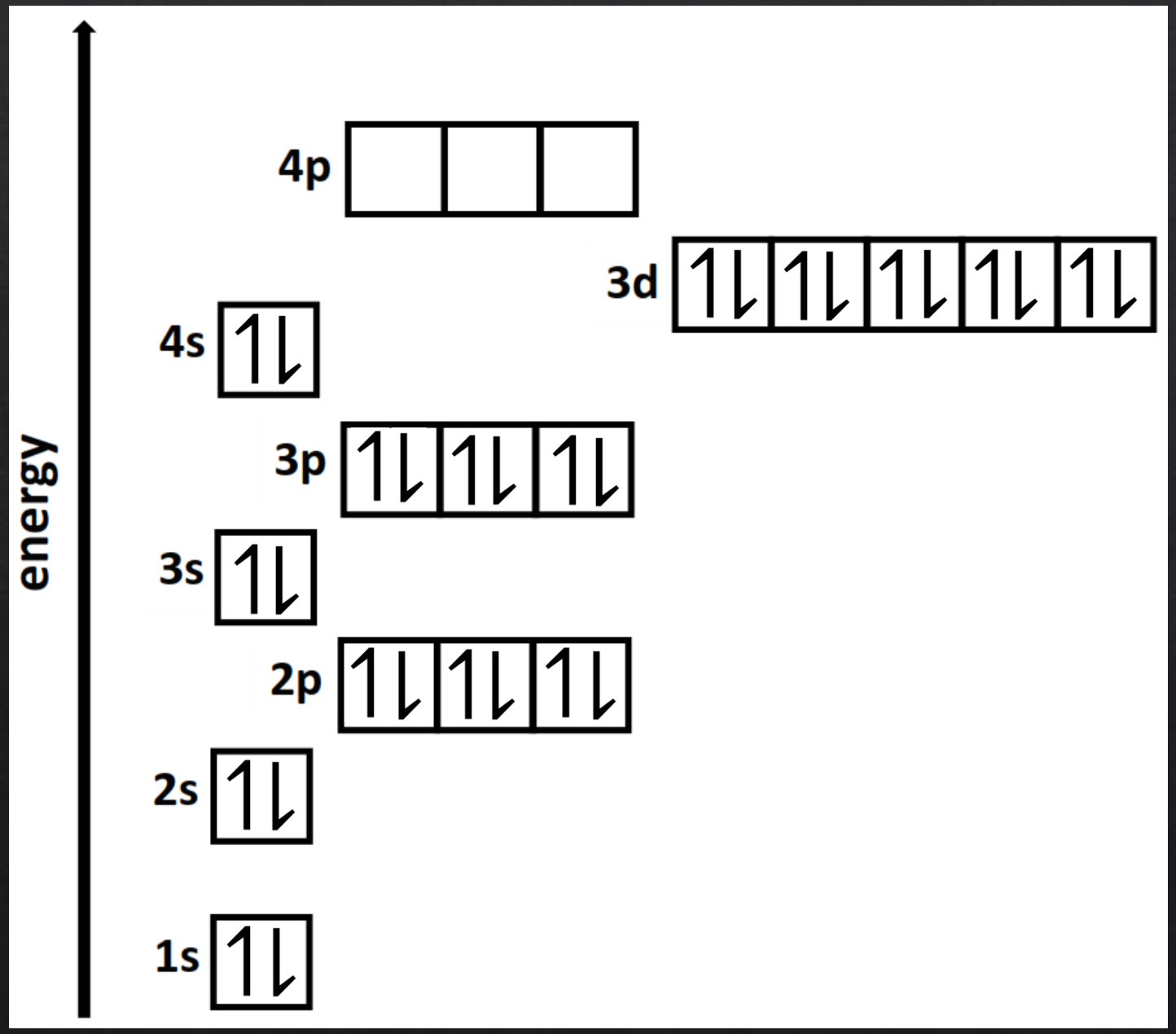
Electron configurations



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

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

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

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

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

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Electron configurations

Ga [Ar] $4s^2\ 3d^{10}\ 4p^1$

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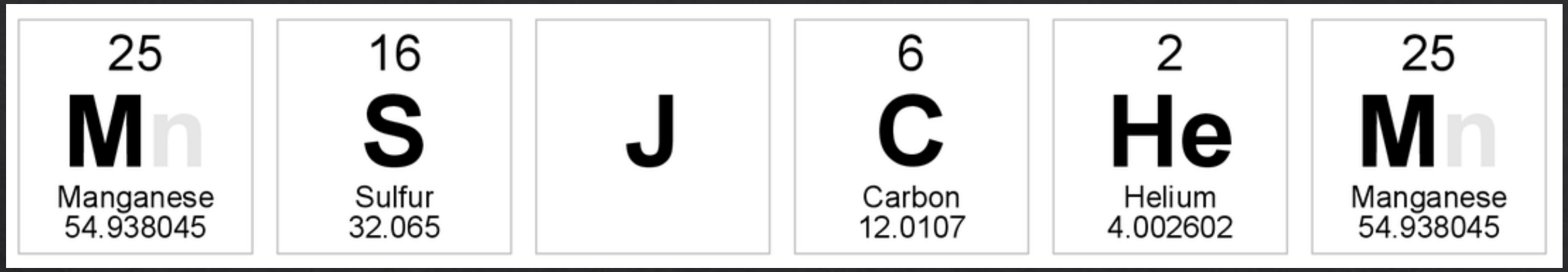
Se [Ar] $4s^2\ 3d^{10}\ 4p^4$

Br [Ar] $4s^2\ 3d^{10}\ 4p^5$

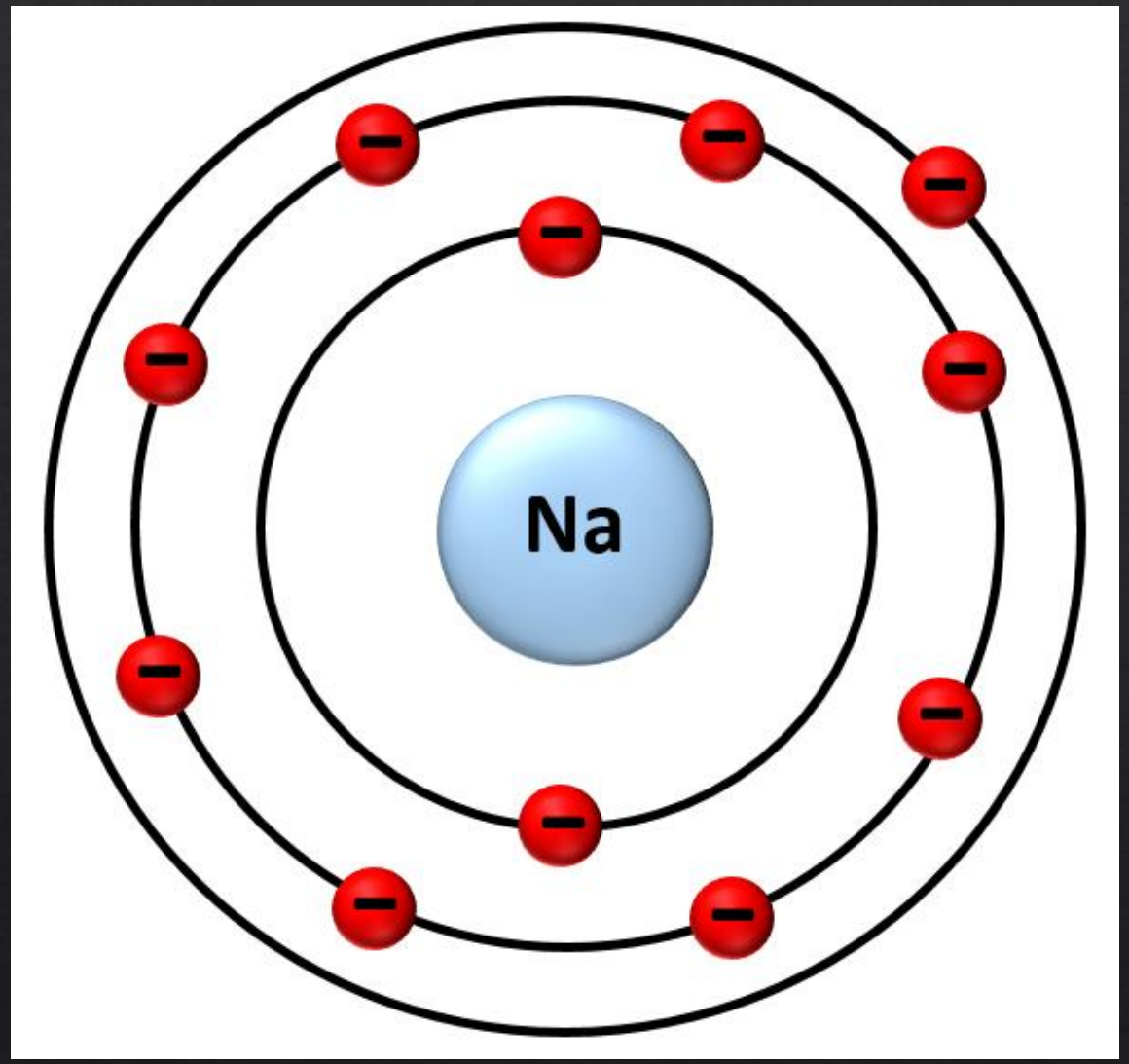
Kr [Ar] $4s^2\ 3d^{10}\ 4p^6$



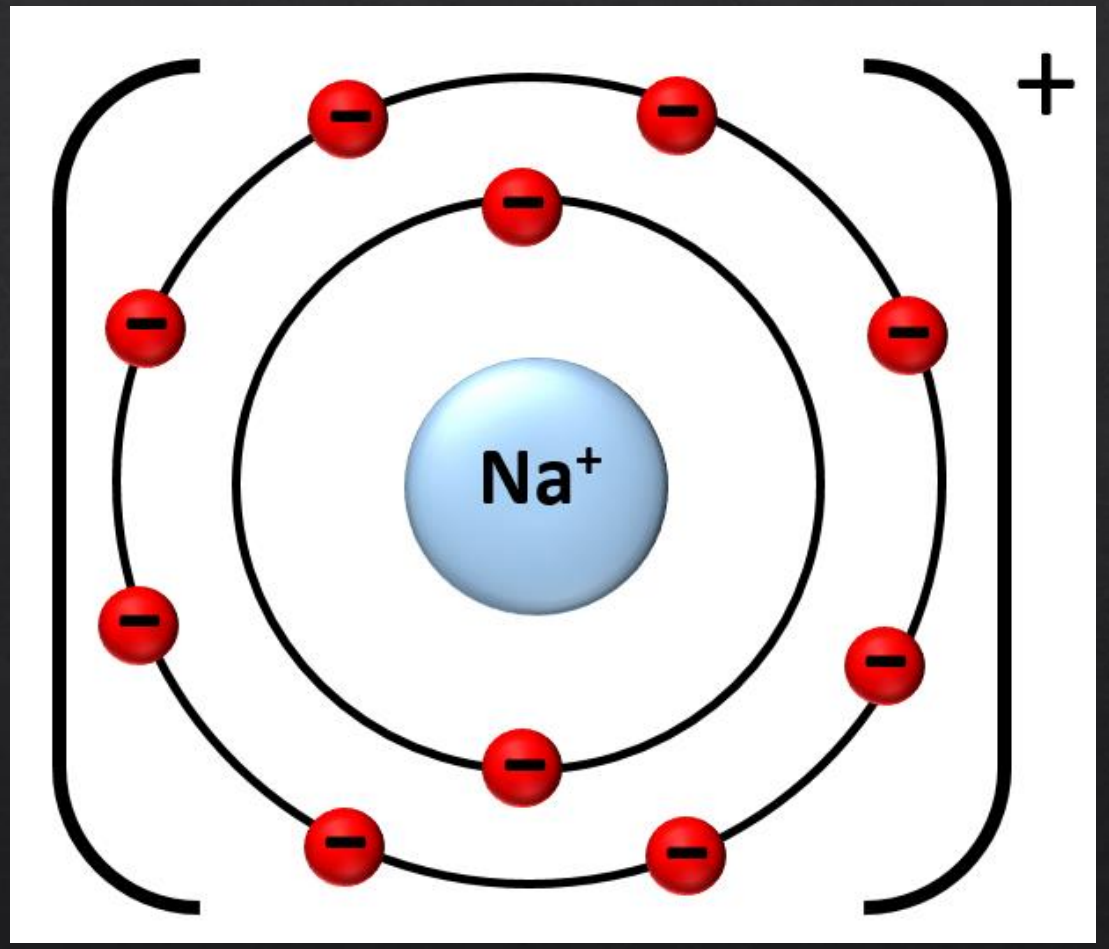
Electron configurations of ions

25	16		6	2	25
Mn	S	J	C	He	Mn
Manganese 54.938045	Sulfur 32.065		Carbon 12.0107	Helium 4.002602	Manganese 54.938045

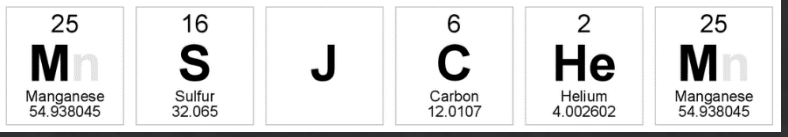
Positive ions (cations)



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



$1s^2\ 2s^2\ 2p^6$

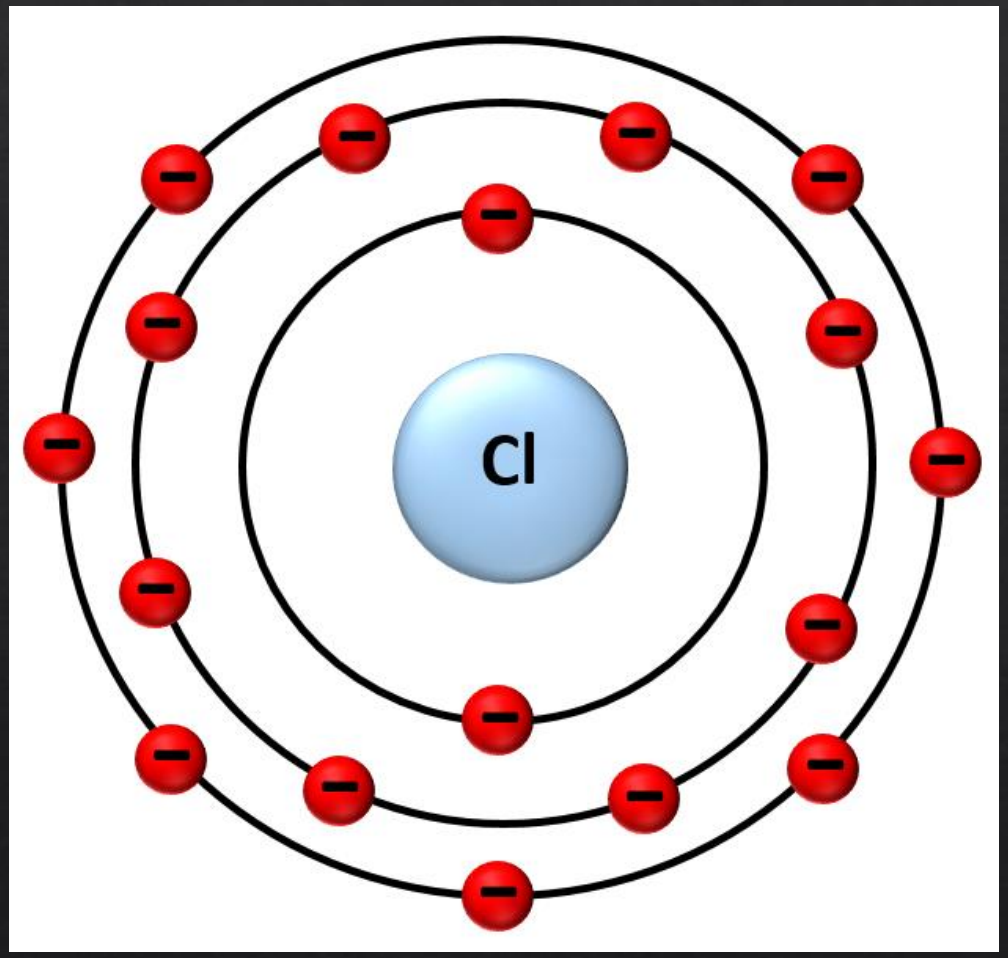


Electron configurations

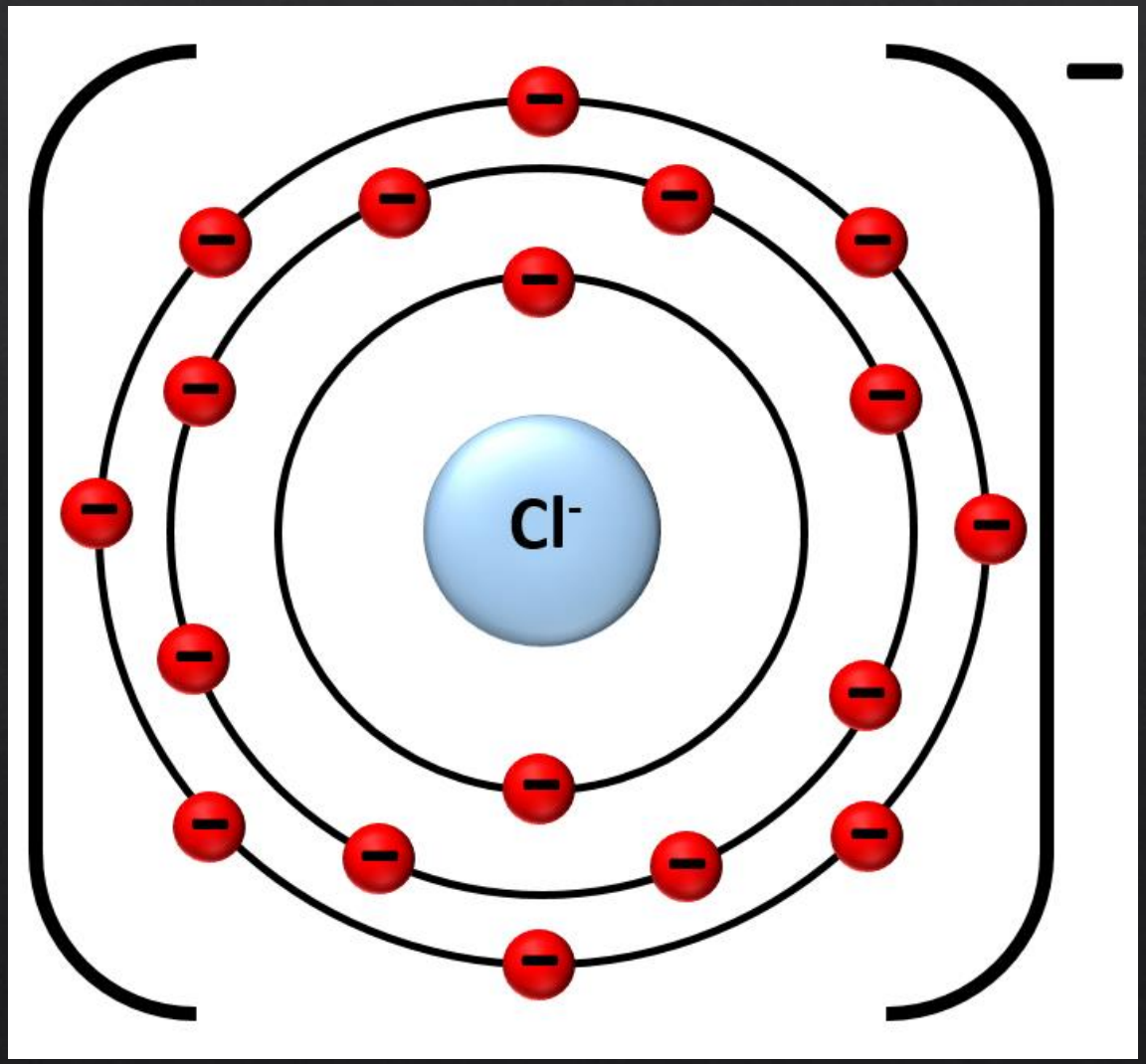
Parent atom	Electron configuration	Ion	Electron configuration
Sodium (Na)	[Ne] $3s^1$	Sodium (Na^+)	[He] $2s^2$ $2p^6$
Magnesium (Mg)	[Ne] $3s^2$	Magnesium (Mg^{2+})	[He] $2s^2$ $2p^6$
Aluminium (Al)	[Ne] $3s^2$ $3p^1$	Aluminium (Al^{3+})	[He] $2s^2$ $2p^6$
Titanium (Ti)	[Ar] $4s^2$ $3d^2$	Titanium (Ti^{2+})	[Ar] $3d^2$
Chromium (Cr)	[Ar] $4s^1$ $3d^5$	Chromium (Cr^{3+})	[Ar] $3d^3$
Nickel (Ni)	[Ar] $4s^2$ $3d^8$	Nickel (Ni^{2+})	[Ar] $3d^8$

25	16		6	2	25
Mn	S	J	C	He	Mn
Manganese 54.938045	Sulfur 32.065		Carbon 12.0107	Helium 4.002602	Manganese 54.938045

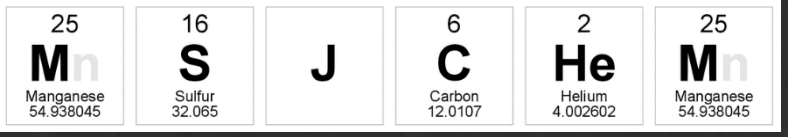
Negative ions (anions)



$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$



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



Electron configurations

Parent atom	Electron configuration	Ion	Electron configuration
Nitrogen (N)	[He] $2s^2\ 2p^3$	Nitride (N^{3-})	[He] $2s^2\ 2p^6$
Oxygen (O)	[He] $2s^2\ 2p^4$	Oxide (O^{2-})	[He] $2s^2\ 2p^6$
Fluorine (F)	[He] $2s^2\ 2p^5$	Fluoride (F^-)	[He] $2s^2\ 2p^6$
Phosphorus (P)	[Ne] $3s^2\ 3p^3$	Phosphide (P^{3-})	[Ne] $3s^2\ 3p^6$
Sulfur (S)	[Ne] $3s^2\ 3p^4$	Sulfide (S^{2-})	[Ne] $3s^2\ 3p^6$
Chlorine (Cl)	[Ne] $3s^2\ 3p^5$	Chloride (Cl^-)	[Ne] $3s^2\ 3p^6$