ELECTROMAGNETIC SPECTRUM

(see p324-326)

- Different kinds of "light" have different energy contents

GAMMA RAYS | X RAYS | ULTRA-VIOLET | VISIBLE LIGHT | INFRARED LIGHT | MICROWAVES

Higher energy

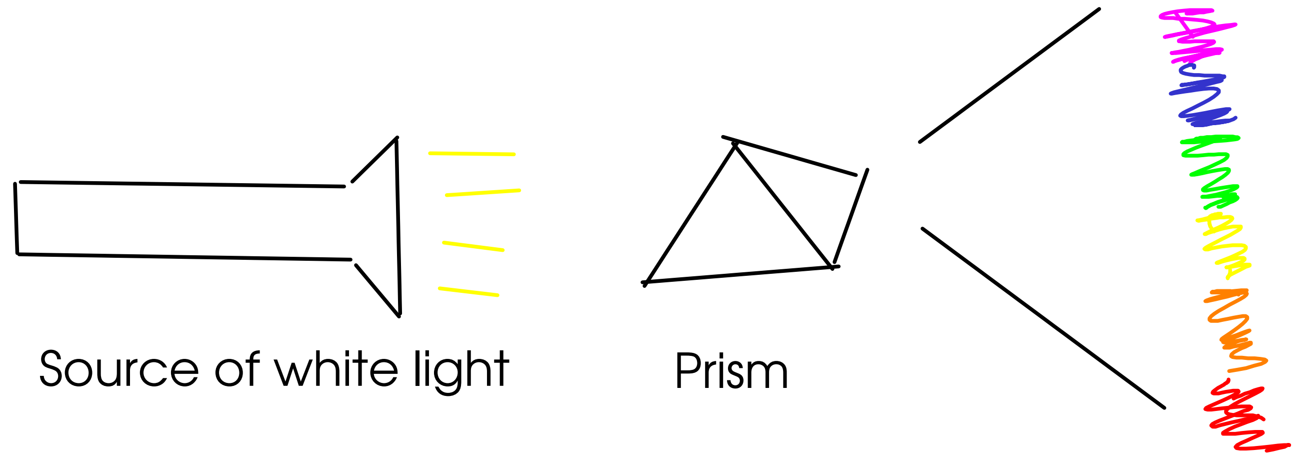
Higher Frequency / Shorter Wavelength

Lower energy

Lower Frequency / Longer Wavelength

VIOLET INDIGO BLUE GREEN YELLOW ORANGE RED

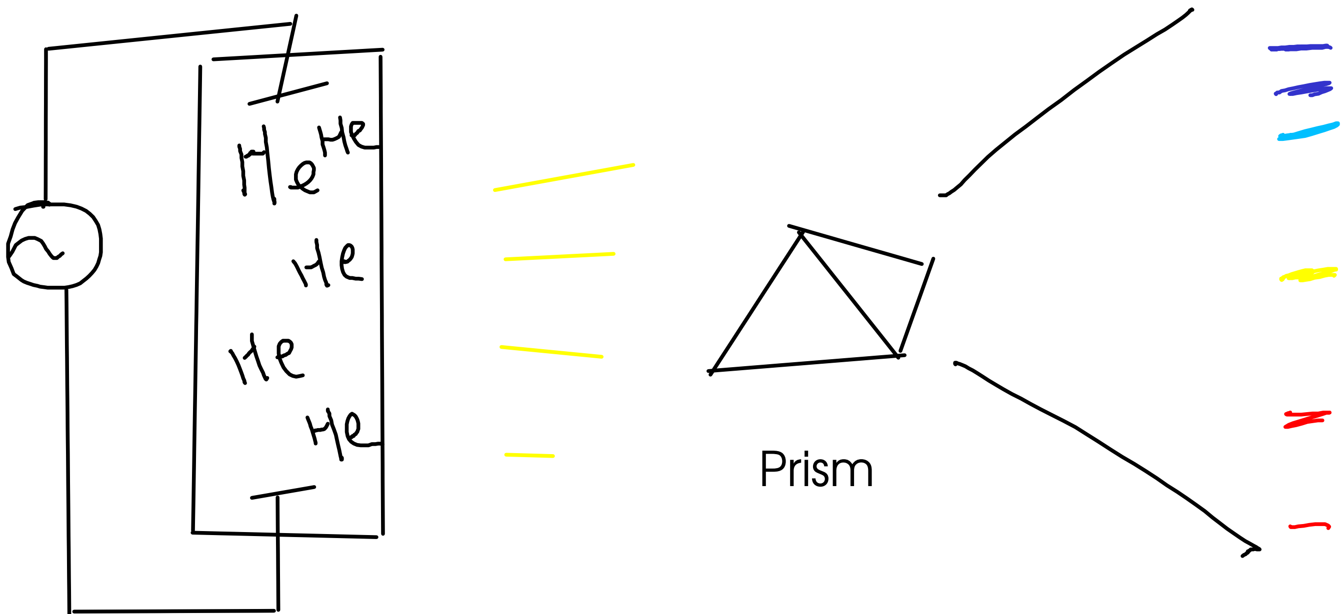
- Different colors of visible light correspond to different amounts of energy



Source of white light

Prism

Rainbow (all colors represented)



Prism

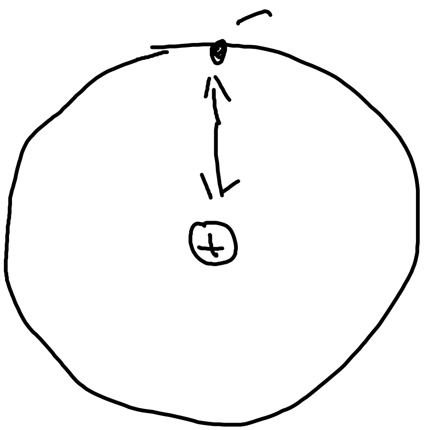
Gaseous Helium excited
by electricity

LINE SPECTRUM - only
a few specific colors appear!
(see p329 for example)

- Atomic line spectra are UNIQUE to each element. They're like atomic "fingerprints".

p329:

- Problem was that the current model of the atom completely failed to explain why atoms emitted these lines.



An orbit that is FARTHER from the nucleus means that the electron has MORE energy

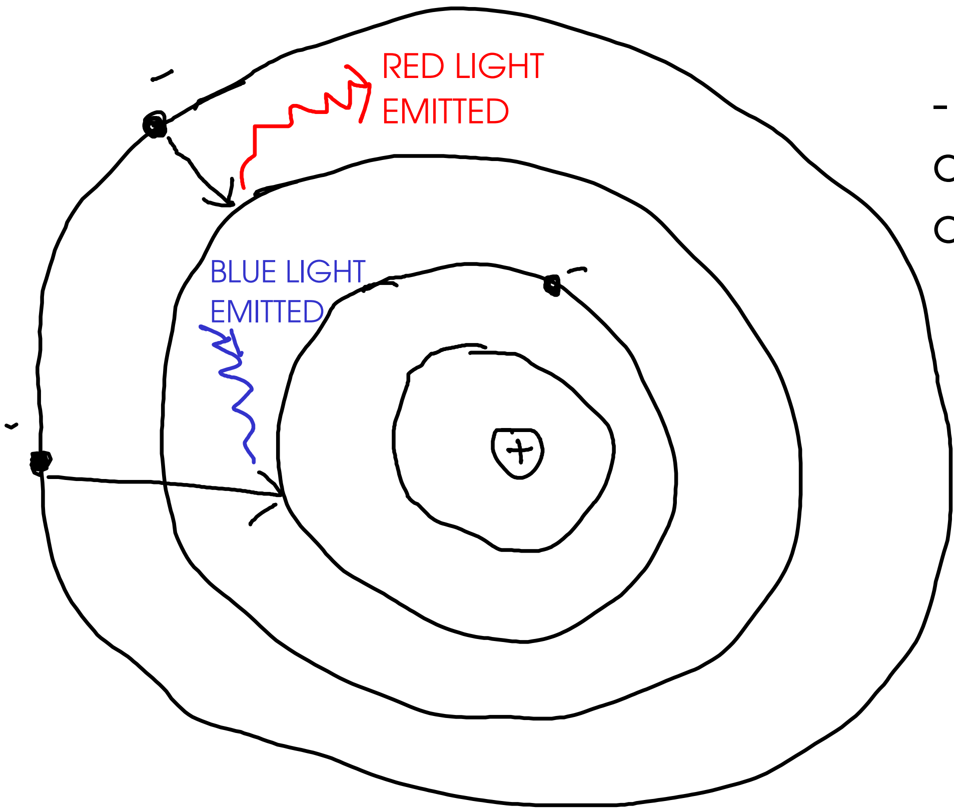
An orbit that is CLOSER to the nucleus means that the electron has LESS energy

- Electrons may gain or lose energy by either ABSORBING (to gain) or EMITTING (to lose) a PHOTON of light. (Photon = particle or "packet" of energy.)

- If the electrons can gain or lose ANY amount of energy, then each atom would emit a RAINBOW rather than an LINE SPECTRUM.

BOHR MODEL

- Theorized that electrons couldn't be just ANYWHERE around the nucleus. There must be restrictions on the motion of electrons that traditional physics did not explain.



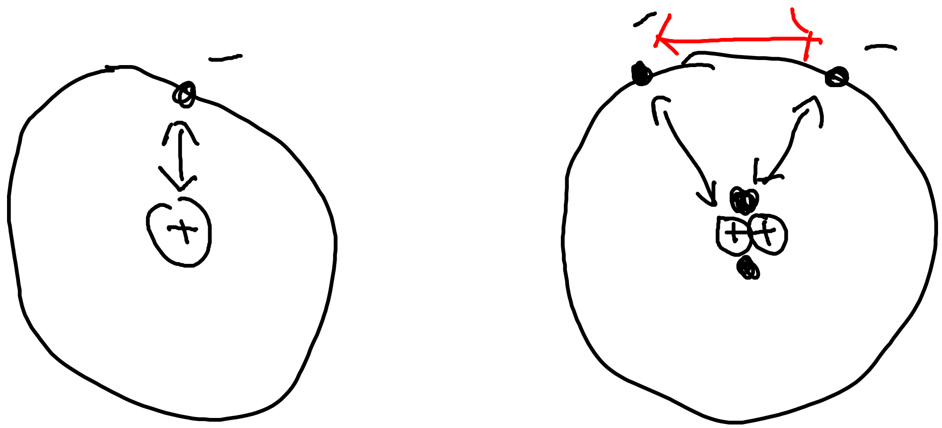
- theorized that electrons could only be certain distances from the nucleus. In other words, they could only have certain values for ENERGY.

    - Electrons could move only from one "energy level" to another DIRECTLY by giving up or abosrbing a photon (light) that was equal in energy to the distance between the energy levels.

- The restrictions on where electrons could be in Bohr's model predicted that atoms would give LINE SPECTRA.

- Bohr's model accurately described the line spectrum of hydrogen (first time this had been done!)

- For other atoms, Bohr's model predicted a line spectrum, but the lines weren't the right colors!



Bohr's model didn't account for electron-electron interactions (which didn't exist in HYDROGEN)

- To account for this added complexity, a more sophisticated model had to be devised: QUANTUM THEORY. Quantum theory is the modern picture of the atom and its electron cloud.

SHELLS, SUBSHELLS, AND ORBITALS

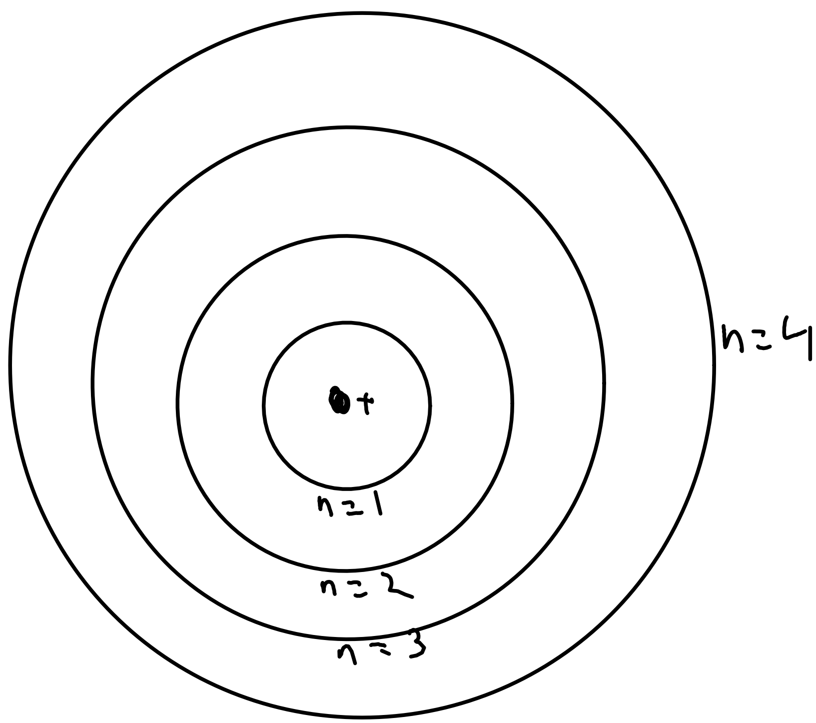
- Bohr's model predicted that energy levels (called SHELLS) were enough to describe completely how electrons were arranged around an atom. But there's more to it!

SHELL: Equivalent to Bohr's energy levels. Electrons in the same SHELL are all the same distance from the nucleus. They all have SIMILAR (but not necessarily the SAME) energy.

- Shells are numbered (1-... - Elements on the periodic table have shells numbered from 1 to 7)

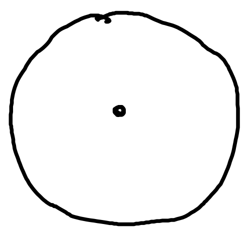
- Higher numbers correspond to greater distance from the nucleus and greater energy, and larger size!

- Higher shells can hold more electrons than lower shells!

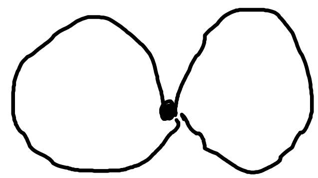


SUBSHELLS: Within a SHELL, electrons may move in different ways around the nucleus! These different "paths" are called SUBSHELLS

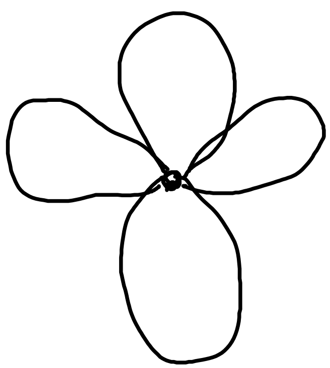
- SHAPES of regions of space that electrons are able to exist in.



"s" subshell
(a spherical region)



"p" subshell
(a dumbbell shaped region)



"d" subshell

- Some atoms also have "f" subshells (not pictured)

See p 334-335 for nicer drawings of the subshells.

ORBITALS - are specific regions of space where electrons may exist

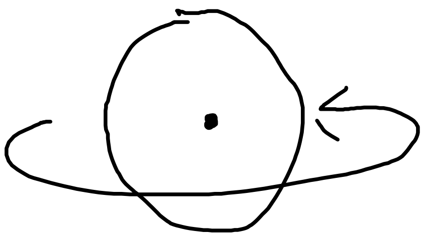
- The SHAPE of an orbital is defined by the SUBSHELL it is in
- The ENERGY of an orbital is defined by both the SHELL the orbital is in AND the kind of SUBSHELL it is in
- Each orbital may, at most, contain TWO ELECTRONS

ARRANGEMENT OF SHELLS, SUBSHELLS, AND ORBITALS

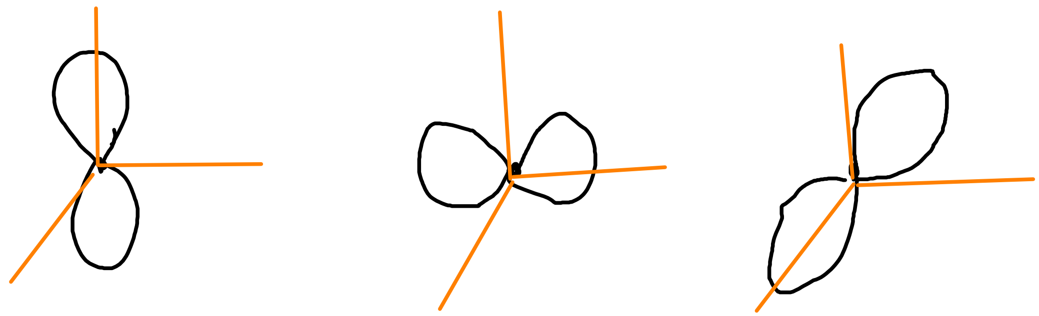
- Shells are numbered. Each shell can contain the same number of SUBSHELLS as its number:

1st shell: ONE possible subshell (s)
2nd shell: TWO possible subshells (s, p)
3rd shell: THREE possible subshells (s, p, d)
4th shell: FOUR possible subshells (s, p, d, f)
... and so on

- Each subshell can contain one or more ORBITALS, depending on how many different ways there are to arrange an orbital of that shape around the nucleus.



"s" subshell
One possible orientation
Maximum 2 electrons in 1 orbital



"p" subshell: Three possible orientations
Maximum 6 electrons in 3 orbitals

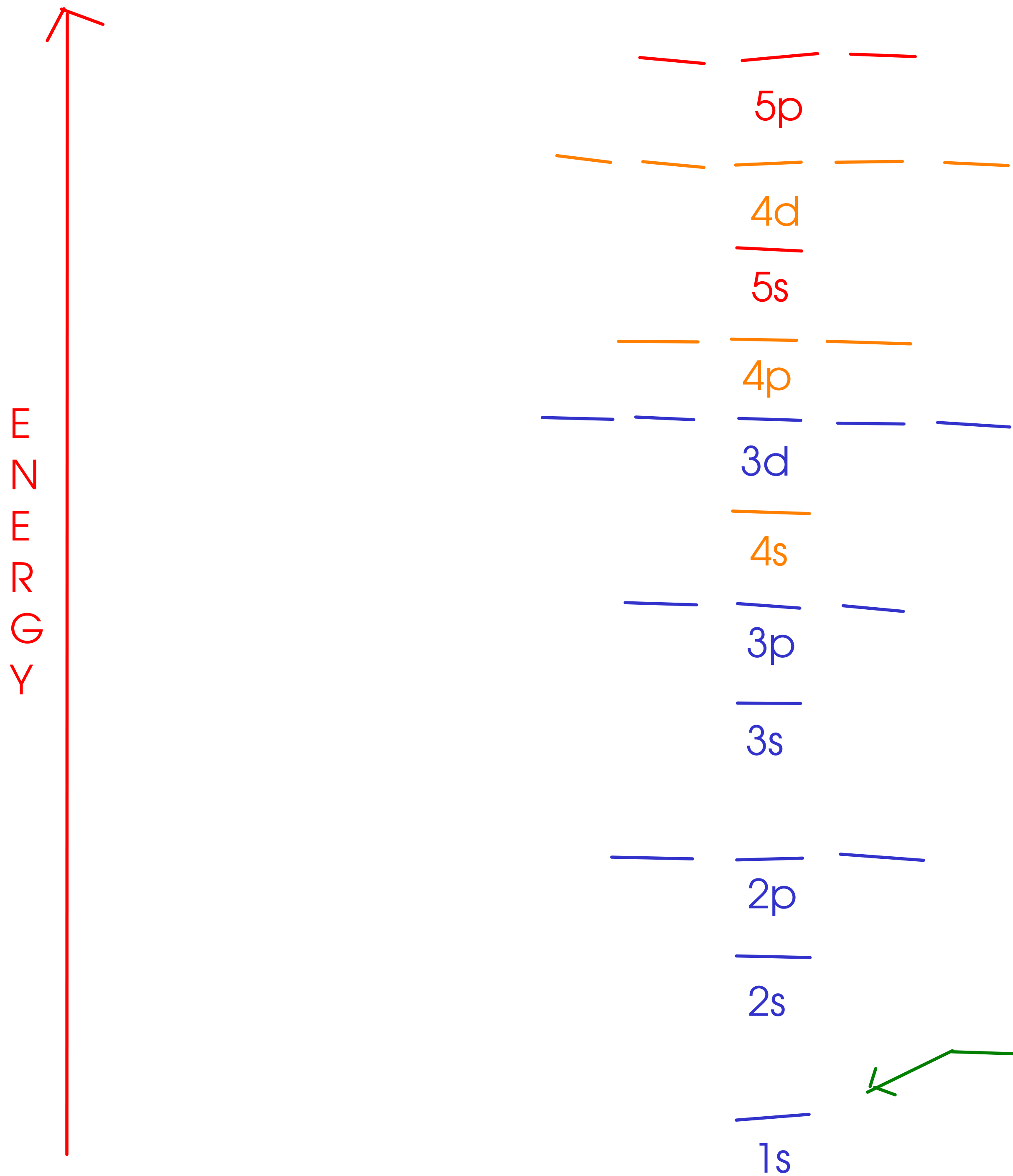
- There are five possible orbitals in a "d" subshell, and 7 possible orbitals in an "f" subshell!

Maximum 10 electrons in 5 orbitals
(see p 335)

Maximum 14 electrons in 7 orbitals

ENERGY DIAGRAM

- We can map out electrons around an atom using an energy diagram:



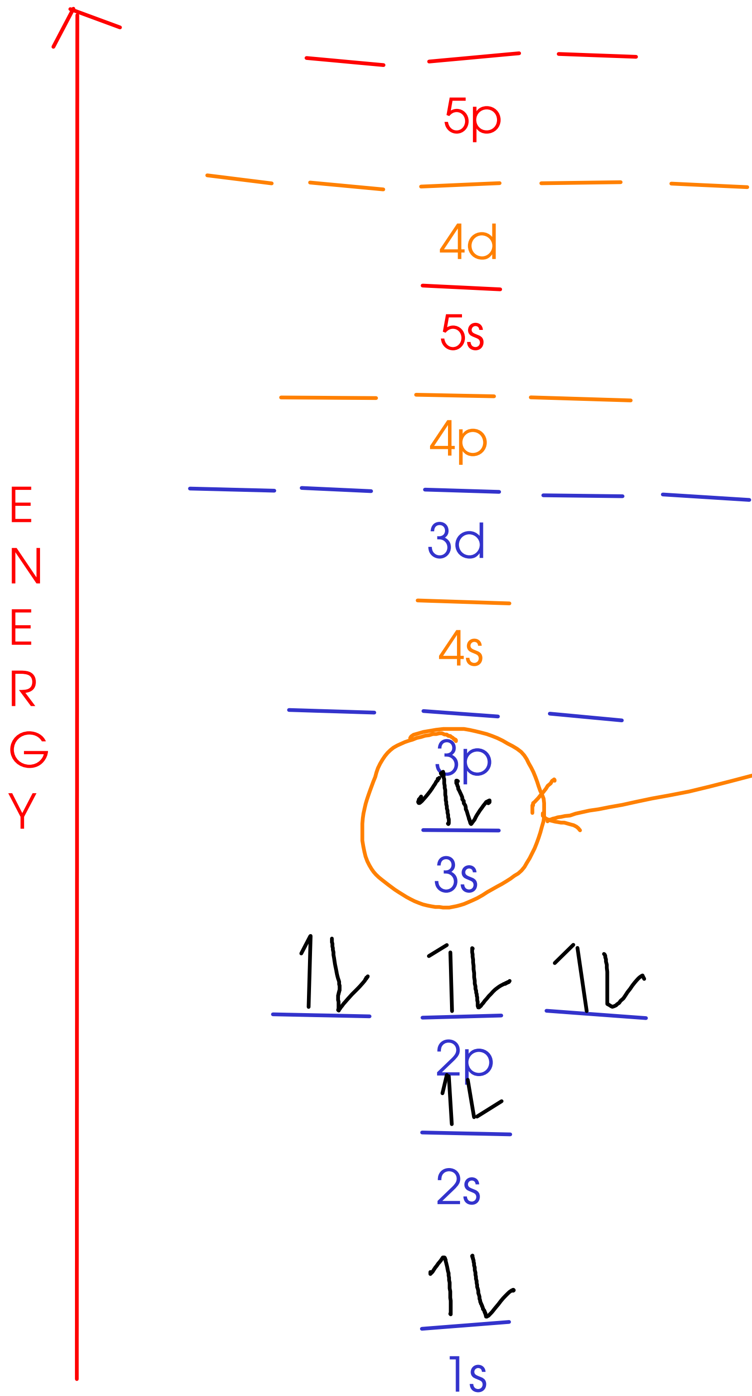
"1s" means first shell,
"s" subshell

Each blank represents an ORIBITAL which can hold up to TWO electrons

Let's look at some example atoms:

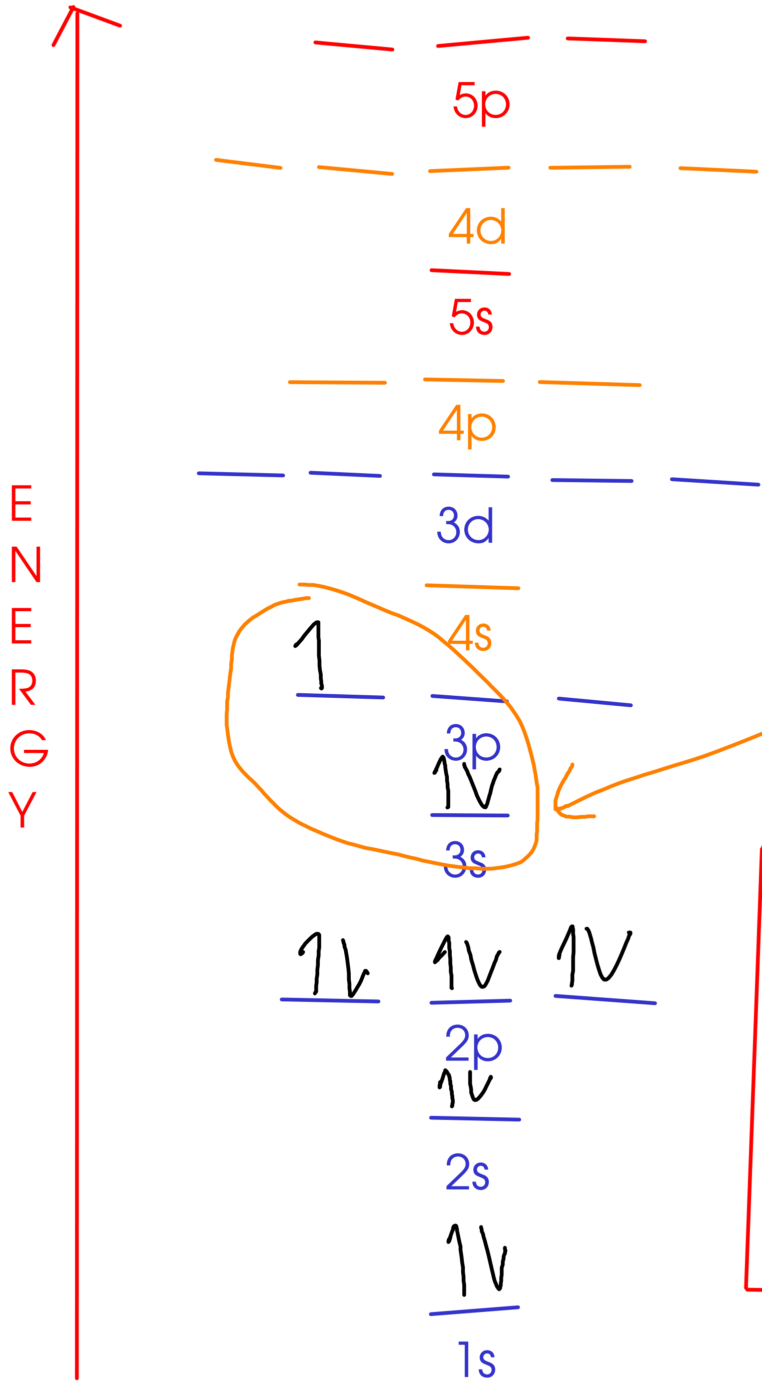
Magnesium: Z=12 , 12 electrons

Z: atomic #



Outermost electrons of magnesium "valence electrons". These electrons are involved in chemical bonding!

Aluminum: Z = 13

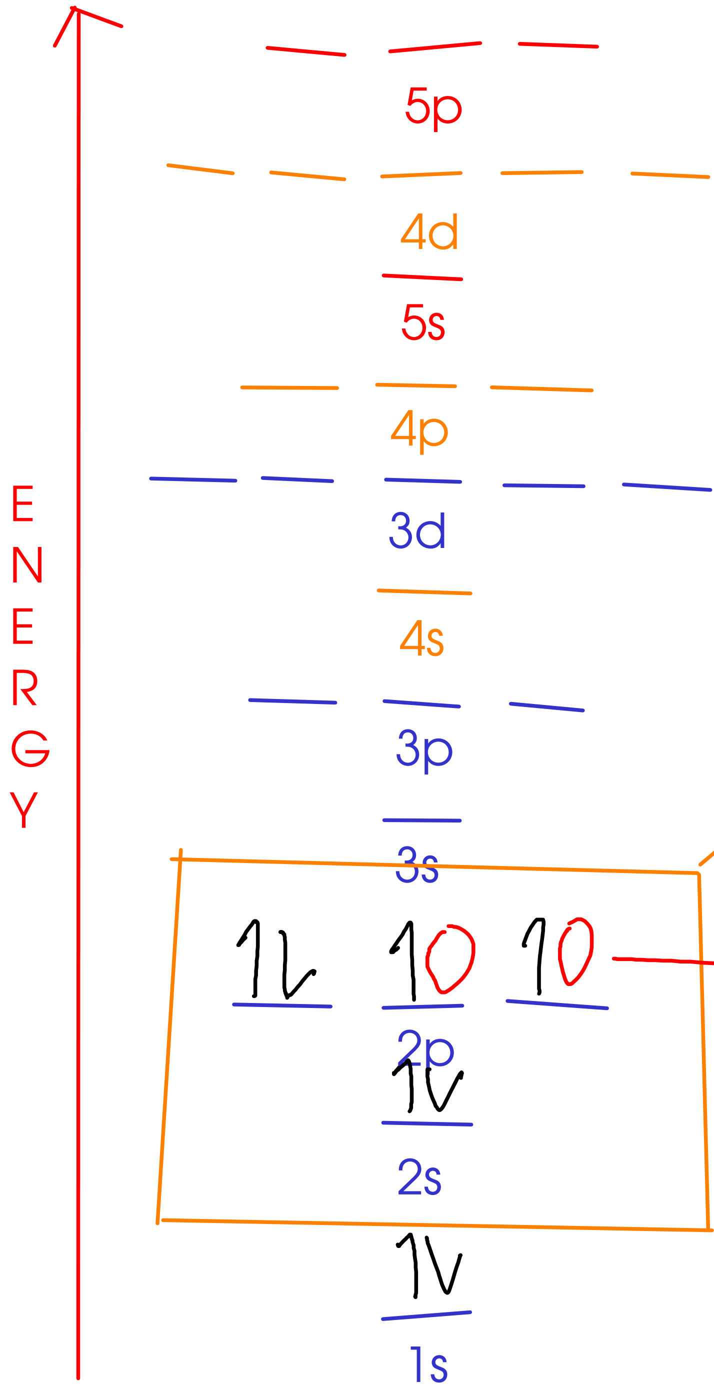


Aluminum has THREE valence electrons!
(All electrons in the outer shell are valence electrons!)

Atoms tend to form ions or chemical bonds in order to end up with filled outer "s" and "p" subshells.

This is called the "octet" rule. (Not all chemical bonds follow this - it's a RULE OF THUMB, not a scientific law!)

Example: Oxygen, Z = 8



Valence electrons for oxygen. (6 electrons)

Oxygen needs two more electrons to complete its outer "p" subshell!

In ionic compounds, oxygen has gained two electrons to become the oxide ion (2- charge). In molecular compounds, oxygen shares electrons with other atoms so that it has a share in eight electrons in its outer shell!