Example:

How many milliliters of 6.00M hydrochloric acid is needed to completely react with 25.0 g of sodium carbonate?

$$2HCl(aq) + Na_2CO_3(s) \longrightarrow H_2O(l) + CO_2(g) + 2NaCl(aq)$$

1 - Convert 25.0 g of sodium carbonate to moles. Use formula weight.
2 - Convert moles sodium carbonate to moles HCl. Use ratio from the chemical equation.
3 - Convert moles hydrochloric acid to volume using molar concentration (6.00 moles/L)

(1) Na_2CO_3 - Na: 2×22.99

C: 1×12.01

O: 3×16.00

$105.99\,g\ Na_2CO_3 = mol\ Na_2CO_3$

$$25.0\,g\ Na_2CO_3 \times \frac{mol\ Na_2CO_3}{105.99\,g\ Na_2CO_3} = 0.2358713086\ mol\ Na_2CO_3$$

(2) $2\ mol\ HCl = mol\ Na_2CO_3$

$$0.2358713086\ mol\ Na_2CO_3 \times \frac{2\ mol\ HCl}{mol\ Na_2CO_3} = 0.4717426172\ mol\ HCl$$

145 Example:

How many milliliters of 6.00M hydrochloric acid is needed to completely react with 25.0 g of sodium carbonate?

$$2HCl(aq) + Na_2CO_3(s) \longrightarrow H_2O(l) + CO_2(g) + 2NaCl(aq)$$

1 - Convert 25.0 g of sodium carbonate to moles. Use formula weight.

2 - Convert moles sodium carbonate to moles hydrochloric acid. Use chemical equation

3 - Convert moles hydrochloric acid to volume. Use molarity (6.00 mol HCl = L)

(3) 6.00 mol HCl = L $\quad$ mL = 10^{-3} L

$$0.4717426172 \text{ mol HCl} \times \frac{\text{L}}{6.00 \text{ mol HCl}} \times \frac{\text{mL}}{10^{-3}\text{ L}} = 78.6 \text{ mL of } 6.00\text{ M HCl}$$

EXAMPLE PROBLEM:

$$2\,Na(s) + Cl_2(g) \rightarrow 2\,NaCl(s)$$

How many grams of sodium metal is required to completely react with 2545 grams of chlorine gas?

1 - Convert 2545 grams of chlorine gas to moles. Use formula weight.
2 - Convert moles chlorine gas to moles sodium. Use chemical equation.
3 - Convert moles sodium to grams using formula weight.

① Cl_2: $2 \times 35.45 = 70.90\ g\ Cl_2 = mol\ Cl_2$

② $2\ mol\ Na = mol\ Cl_2$

③ $22.99\ g\ Na = mol\ Na$

$$2545\ g\ Cl_2 \times \frac{mol\ Cl_2}{70.90\ g\ Cl_2} \times \frac{2\ mol\ Na}{mol\ Cl_2} \times \frac{22.99\ g\ Na}{mol\ Na} = \boxed{1650.\ g\ Na}$$

$1.650 \times 10^3\ g\ Na$

 EXAMPLE PROBLEM:

$$H_2SO_4(aq) + 2\,NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

How many mL of 0.250 M sodium hydroxide is required to completely react with 15.0 mL of 2.00 M sulfuric acid?

1 - Convert 15.0 mL of sulfuric acid solution to moles (use concentration: 2.00 mol/L)
2 - Convert moles sulfuric acid to moles sodium hydroxide using chemical equation
3 - Convert moles sodium hydroxide to volume (use concentration: 0.250 mol/L)

(1) 2.00 mol H_2SO_4 = L mL = 10^{-3} L | (2) mol H_2SO_4 = 2 mol NaOH

(3) 0.250 mol NaOH = L mL = 10^{-3} L

$$15.0\text{ mL} \times \frac{10^{-3}\text{ L}}{\text{mL}} \times \frac{2.00\text{ mol } H_2SO_4}{\text{L}} \times \frac{2\text{ mol NaOH}}{\text{mol } H_2SO_4} \times \frac{\text{L}}{0.250\text{ mol NaOH}} \times \frac{\text{mL}}{10^{-3}\text{ L}} =$$

(1) (2) (3)

= 240. mL of 0.250 M NaOH

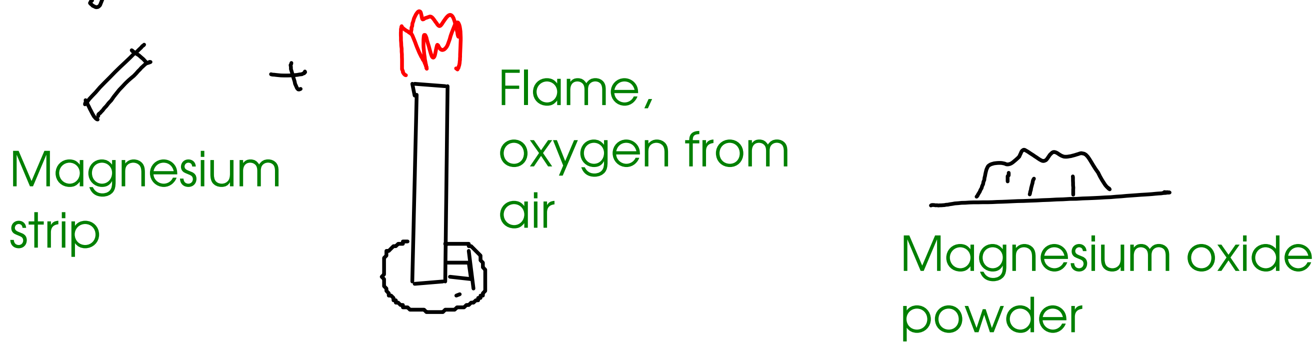
Shortcut: Use MILLIMOLES instead of MOLES for the calculation.

$$15.0\text{ mL} \times \frac{2.00\text{ mol } H_2SO_4}{\text{L}} \times \frac{2\text{ mol NaOH}}{\text{mol } H_2SO_4} \times \frac{\text{L}}{0.250\text{ mol NaOH}} = 240\text{ mL}$$

CONCEPT OF LIMITING REACTANT

- When does a chemical reaction STOP?

$$2Mg(s) + O_2(g) \xrightarrow{\Delta} 2MgO(s)$$



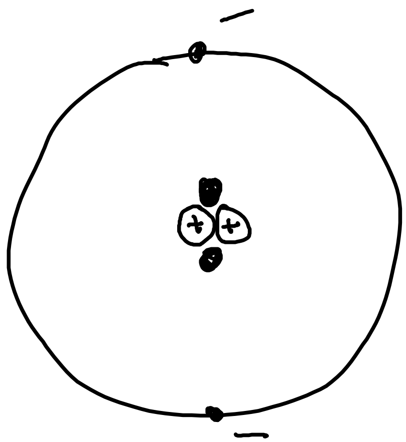
- When does this reaction stop? When burned in open air, this reaction stops when all the MAGNESIUM STRIP is gone. We say that the magnesium is LIMITING.

- This reaction is controlled by the amount of available magnesium

- At the end of a chemical reaction, the LIMITING REACTANT will be completely consumed, but there may be some amount of OTHER reactants remaining. We do chemical calculations in part to minimize these "leftovers".

- Reactants that are left at the end of a chemical reaction (in other words, they are NOT the limiting reactant) are often called "excess". So reacting magnesium with "excess oxygen" means that magnesium is limiting.

STRUCTURE OF THE ELECTRON CLOUD



The nuclear model describes atoms as consisting of a NUCLEUS containing protons and neutrons and an ELECTRON CLOUD containing electrons.

The ELECTRON CLOUD is described as being a diffuse (lots of empty space) region of the atom. Nothing else about it is part of the nuclear model.

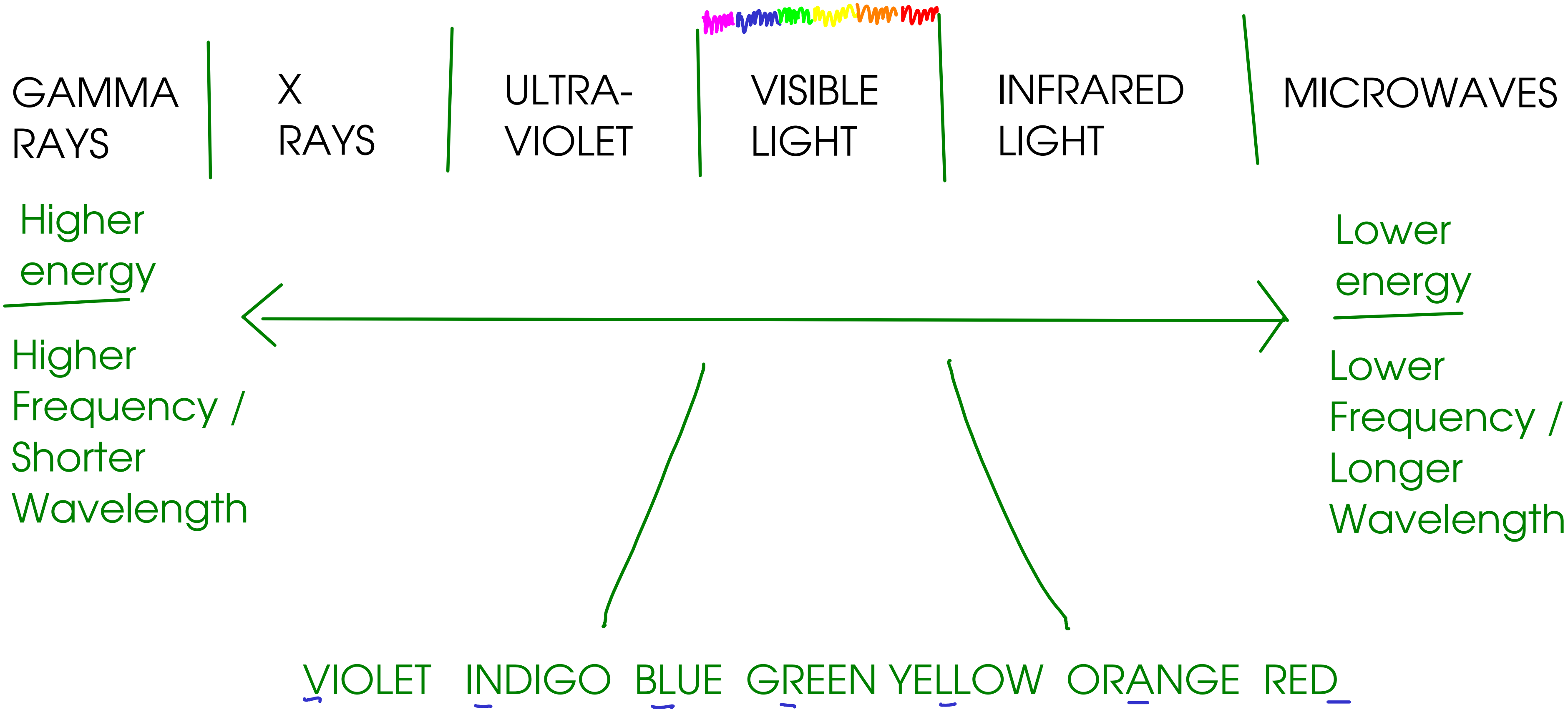
... but the nuclear model is not useful to explain several things:

- Does not explain why atoms react differently from one another
- Does not explain how atoms emit and absorb light (atomic line spectra)

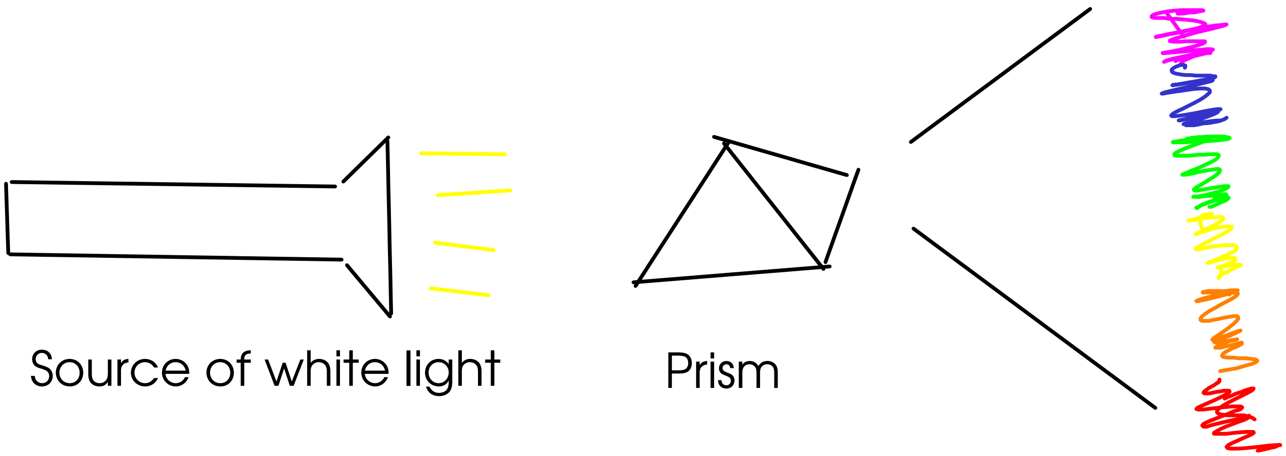
ELECTROMAGNETIC SPECTRUM

(see p324-326)

- Different kinds of "light" have different energy contents



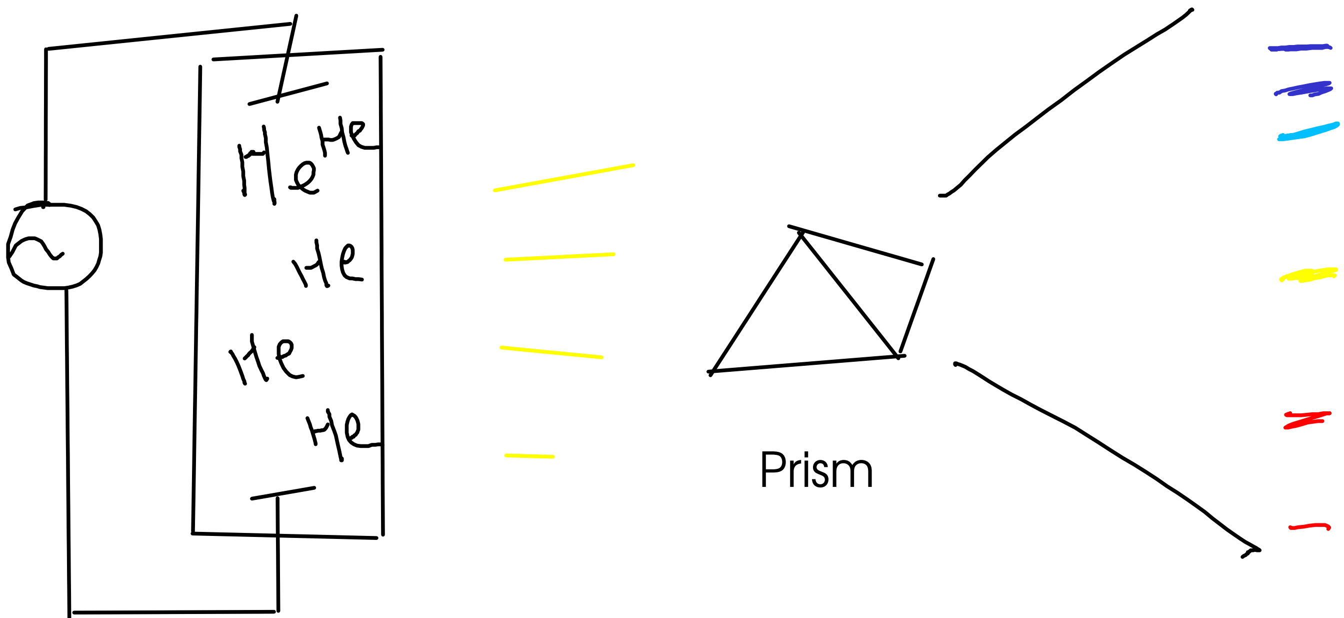
- Different colors of visible light correspond to different amounts of energy



Source of white light

Prism

Rainbow (all colors represented)



Prism

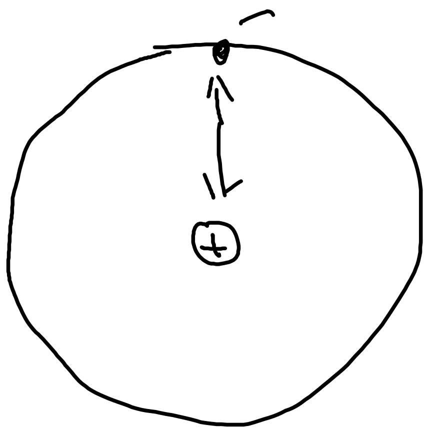
Gaseous Helium excited
by electricity

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

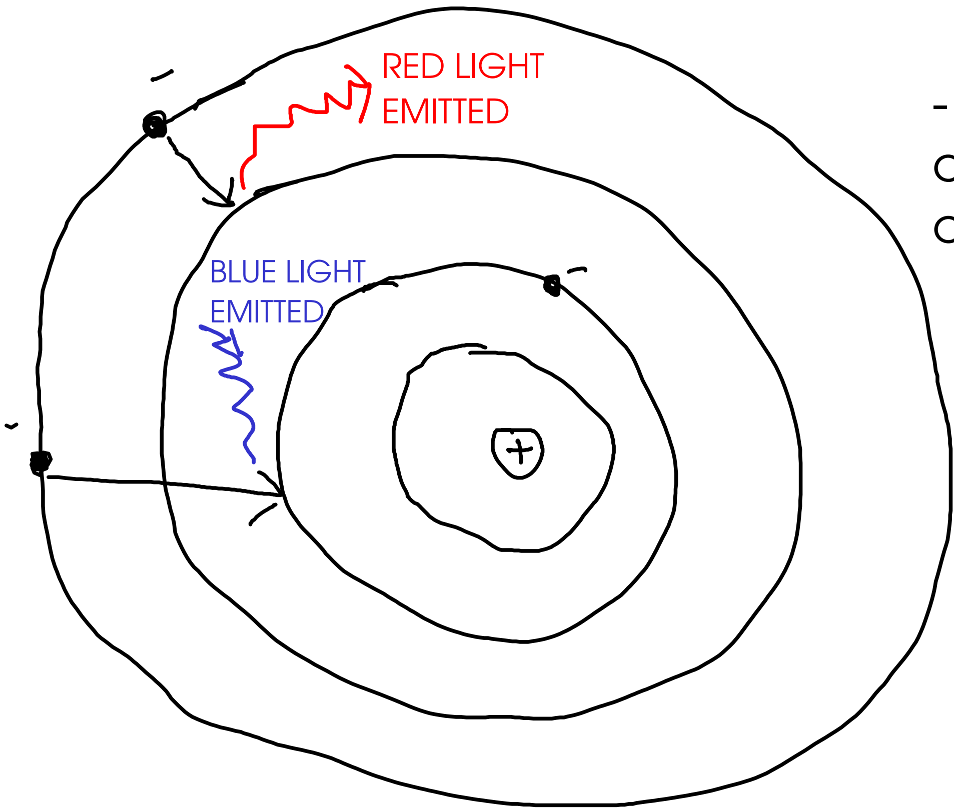
An orbit that is CLOSER to the nucleus means that the electron has LESS energy

- 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.

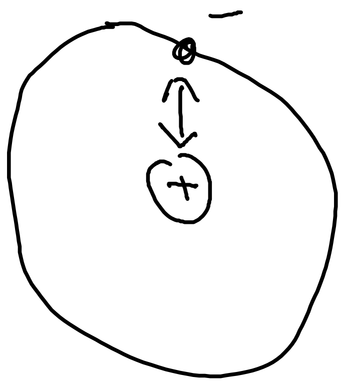


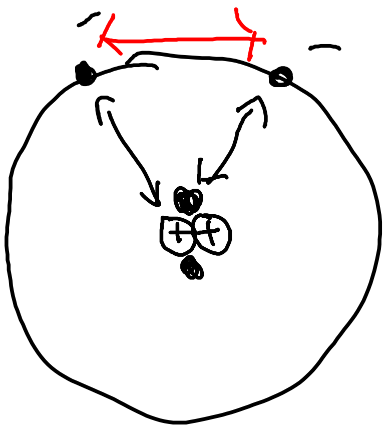
- theorized that electrons could only be certain distances from the nucleus. In other words, they could only have certain values for ENERGY.
  - Electrons could move only from one "energy level" to another DIRECTLY by giving up or abosrbing a photon (light) that was equal in energy to the distance between 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

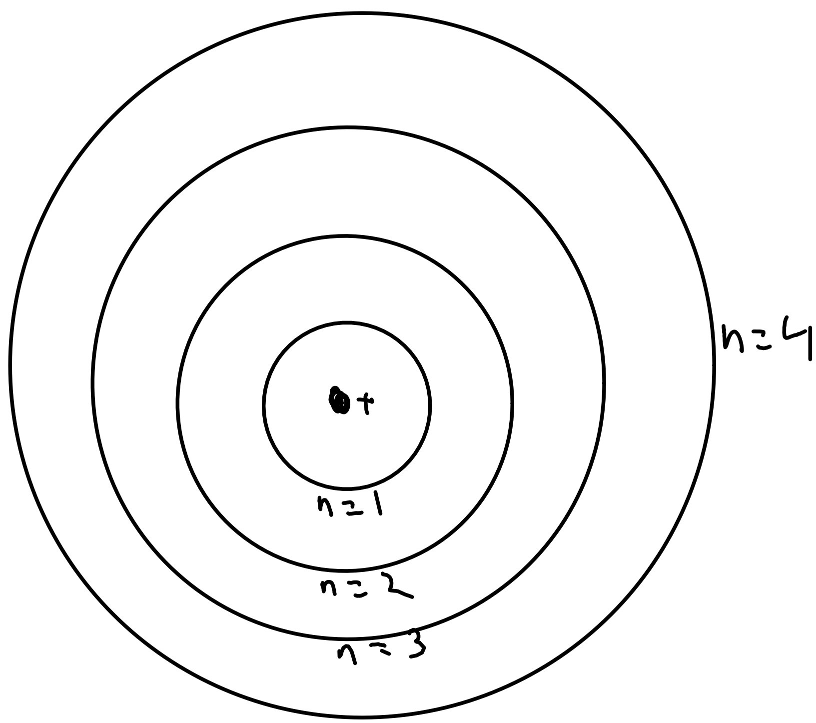
- Bohr's model predicted that energy levels (called SHELLS) were enough to describe completely how electrons were arranged around an atom. But there's more to it!

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)

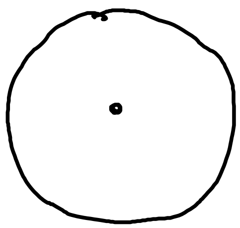
- Higher numbers correspond to greater distance from the nucleus and greater energy, and larger size!

- Higher shells can hold more electrons than lower shells!

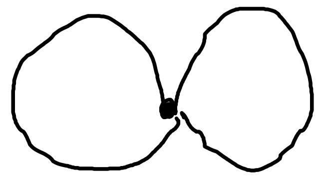


SUBSHELLS: Within a SHELL, electrons may move in different ways around the nucleus! These different "paths" are called SUBSHELLS

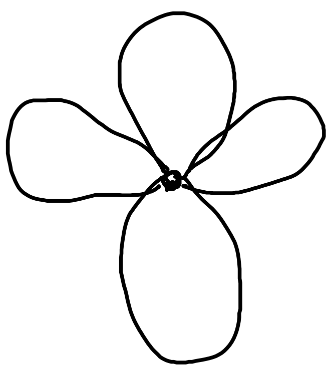
- SHAPES of regions of space that electrons are able to exist in.



"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

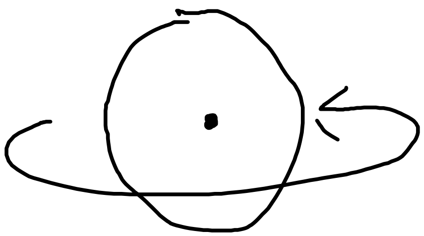
- The SHAPE of an orbital is defined by the SUBSHELL it is in
- The ENERGY of an orbital is defined by both the SHELL the orbital is in AND the kind of SUBSHELL it is in
- Each orbital may, at most, contain TWO ELECTRONS

ARRANGEMENT OF SHELLS, SUBSHELLS, AND ORBITALS

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

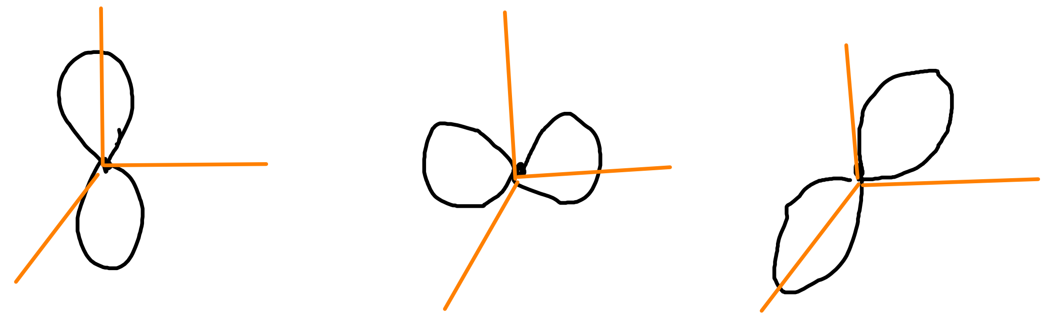
1st shell: ONE possible subshell (s)
2nd shell: TWO possible subshells (s, p)
3rd shell: THREE possible subshells (s, p, d)
4th shell: FOUR possible subshells (s, p, d, f)
... and so on

- 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!

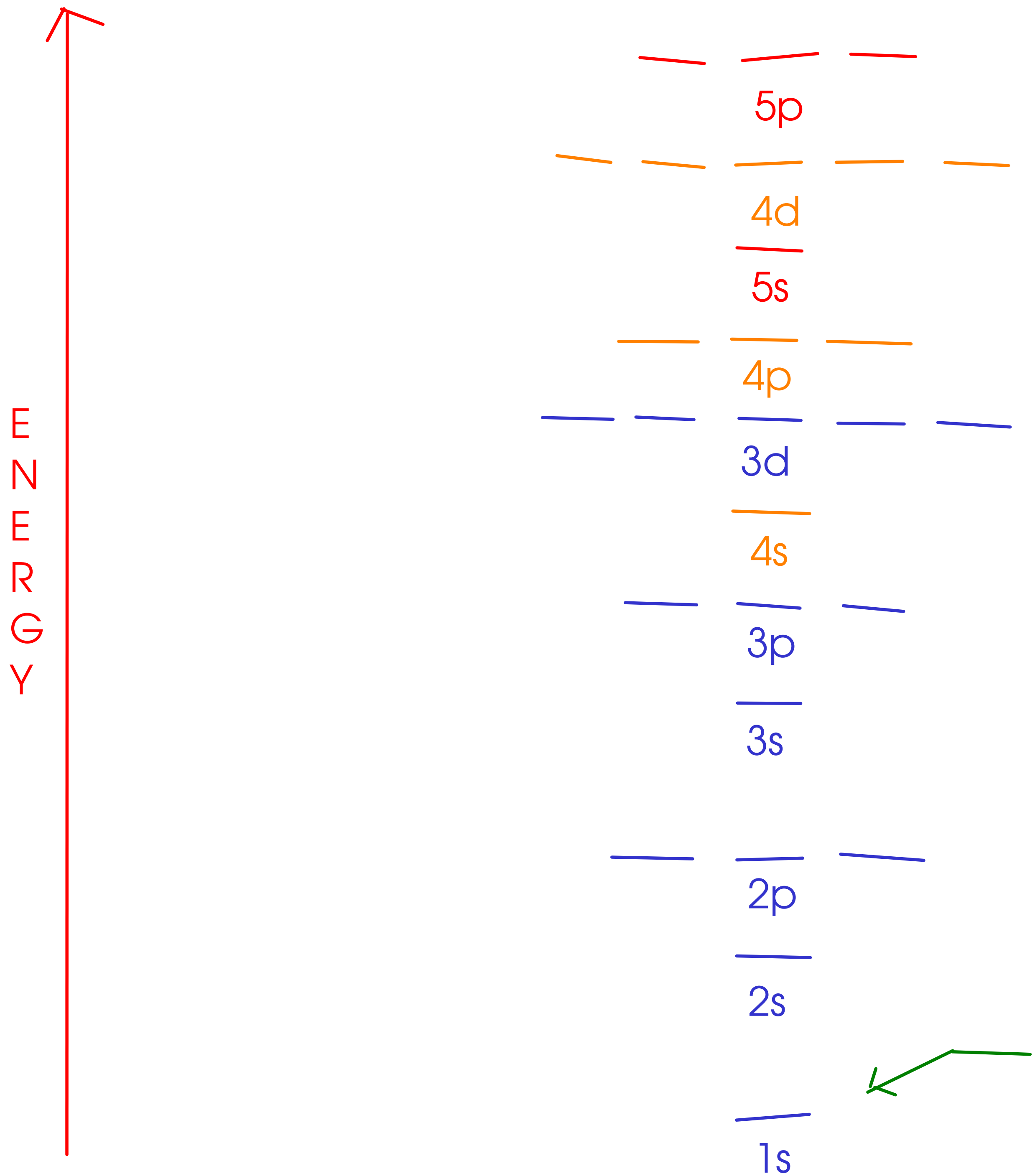
Maximum 10 electrons in 5 orbitals

(see p 335)

Maximum 14 electrons in 7 orbitals

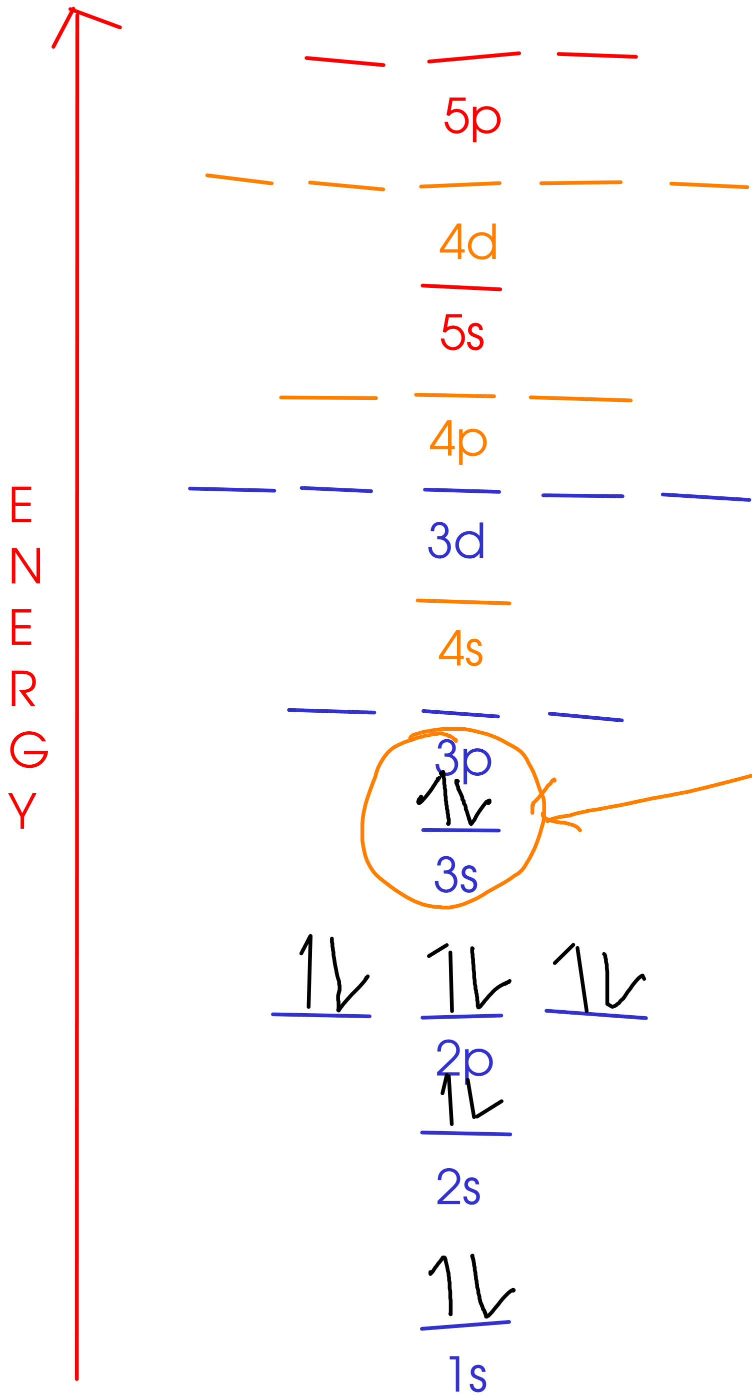
ENERGY DIAGRAM

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



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

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



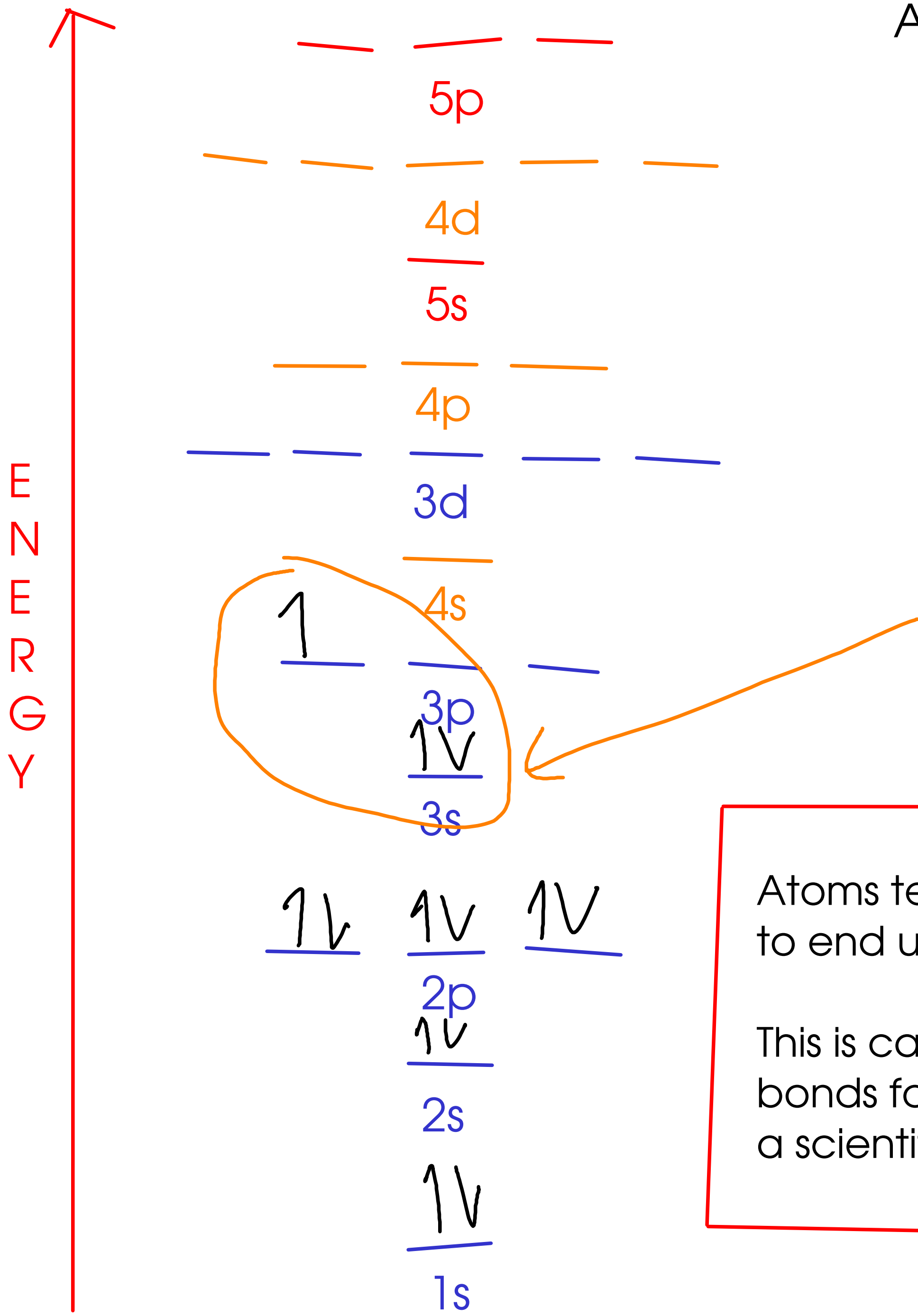
Let's look at some example atoms:

Magnesium: Z=12 , 12 electrons

↑ Z: atomic #

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

Aluminum: Z = 13

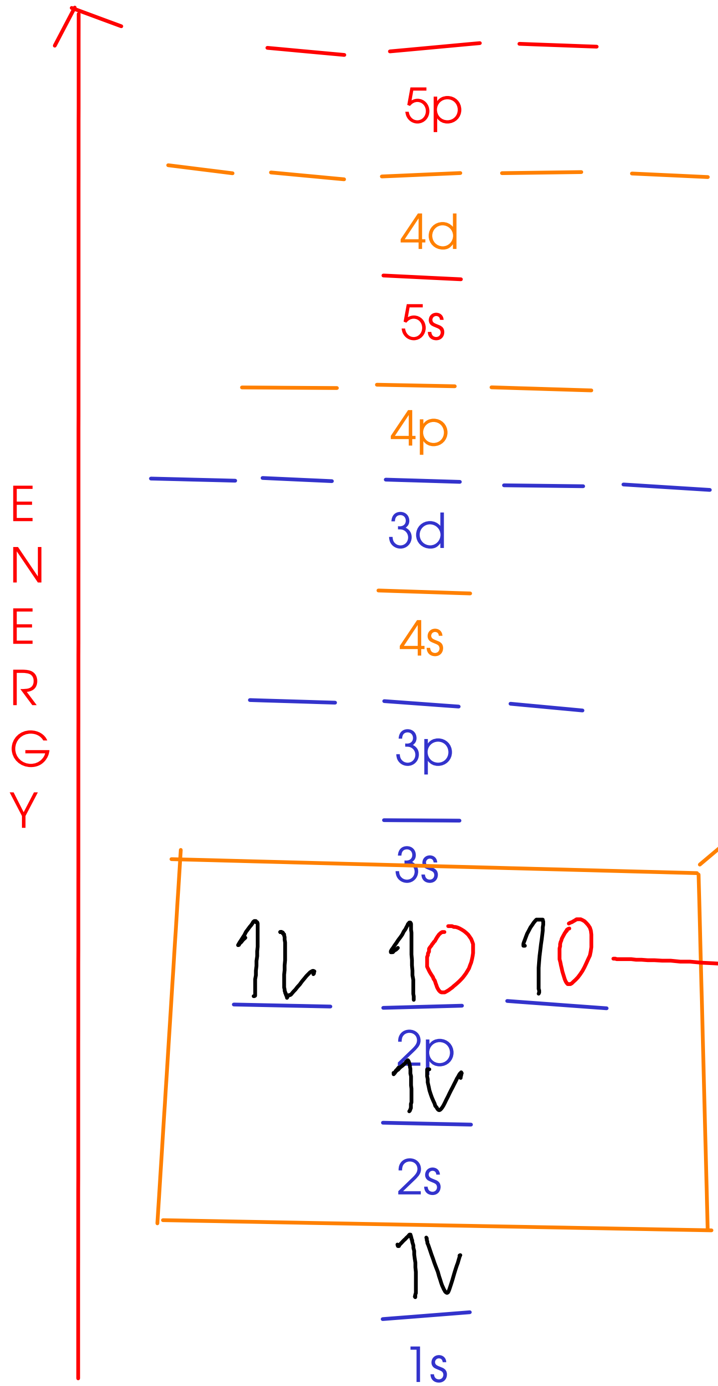


Aluminum has THREE valence electrons!
(All electrons in the outer shell are valence electrons!)

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!

ELECTRON CONFIGURATION

- A shorthand way to write about electron arrangement around an atom.

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

$1s$ — Shell and subshell

^{4} — Number of electrons in the subshell!

Mg: $1s^2 2s^2 2p^6 \underline{3s^2}$

Al: $1s^2 2s^2 2p^6 \underline{3s^2 3p^1}$

Valence electrons are the ones in the outermost SHELL, not just the last subshell. Aluminum has THREE valence electrons.

ELECTRON CONFIGURATION AND THE PERIODIC TABLE

two elements wide

Helium is part of the "s" block!

six elements wide

* "d" block is offset one period down

ten elements wide

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIIIB	VIIIB	VIIIB	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	H																	He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg											Al	Si	P	S	Cl	Ar
4	K	Ca	3d Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	4d Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	5d* La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra	6d* Ac	Rf	Db	Sg	Bh	Hs	Mt									

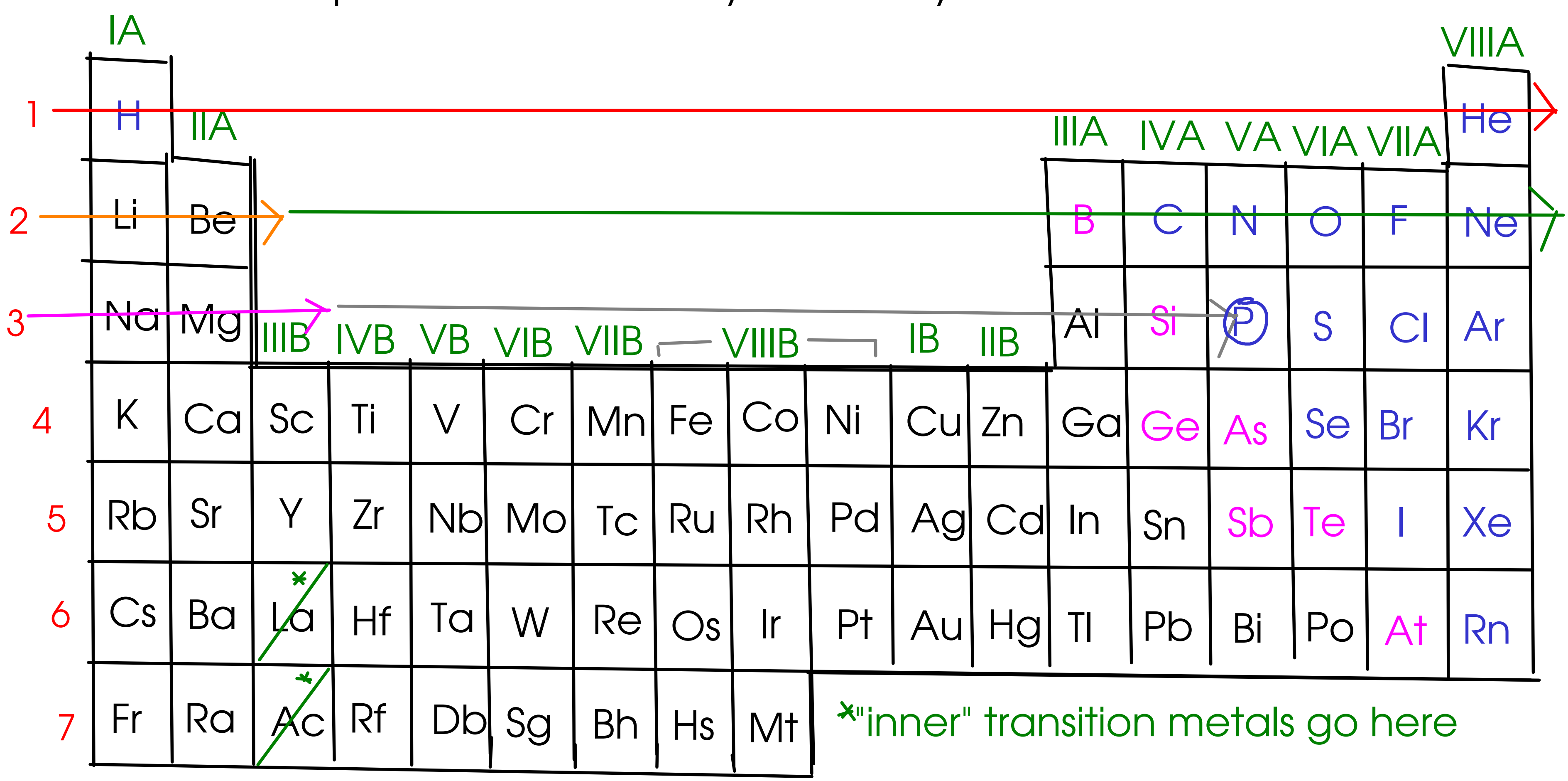
*"inner" transition metals go here

"s" block: last electron in these atoms is in an "s" orbital!

"p" block: last electron in these atoms is in a "p" orbital!

"d" block: last electron is these atoms is in a "d" orbital

- To write an electron configuration using the periodic table, start at hydrogen, and count up the electrons until you reach your element!



Example: Phosphorus (P): $1s^2 2s^2 2p^6 3s^2 3p^3$

VALENCE ELECTRONS (outermost electrons) of phosphorus. 5 electrons

EXAMPLES:

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

Remember - valence electrons are ALL of the electrons in the outermost SHELL! (may have more than one SUBSHELL)!

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

TITANIUM is a transition metal that commonly forms either +2 or +4 cations. The 4s electrons are lost when the +2 ion forms, while the 4s AND 3d electrons are lost to form the +4!

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

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

You can order the subshells in numeric order OR in filling order

Ti $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$ or $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

or [Ar] $3d^2 4s^2$

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

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

Noble gas core notation. Use the previous noble gas on the table, then add the electrons that it doesn't have to the end.

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

You are responsible for writing electron configurations up to Z=18, Argon. These are here to illustrate other points!