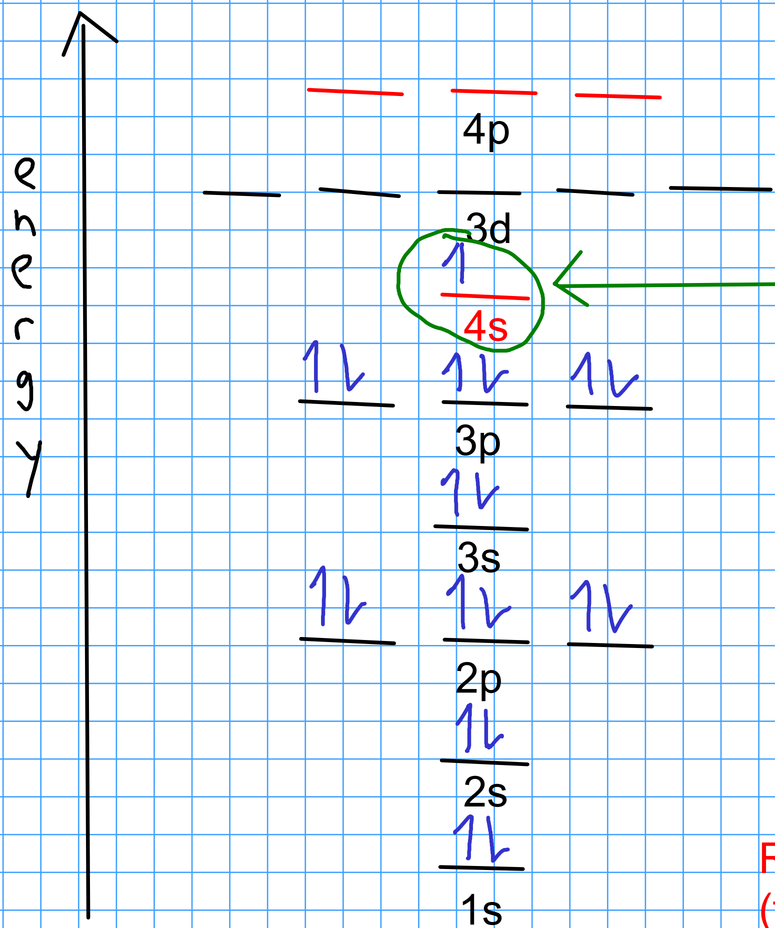
How would an orbital diagram for the element POTASSIUM look?



K , $\underline{Z} = 19$

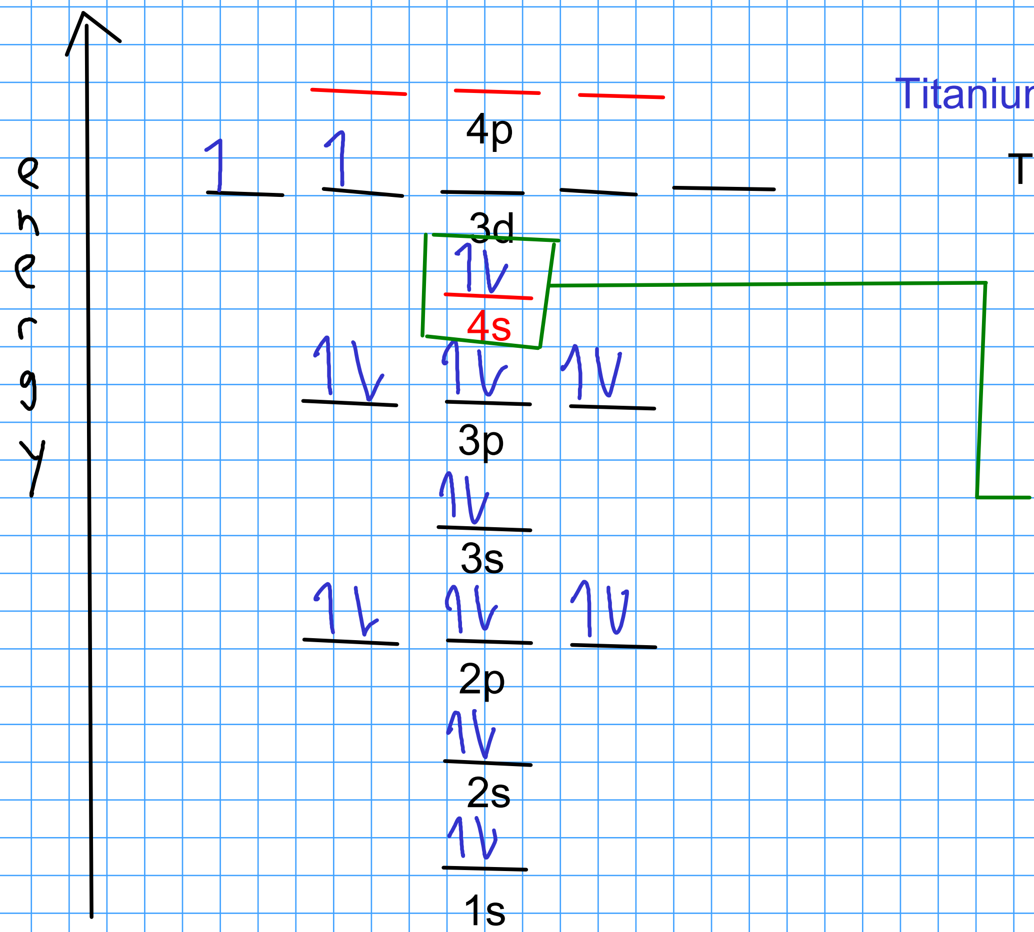
atomic number

Electrons in the outermost shell of an atom are called VALENCE electrons. THESE electrons are normally involved in chemical bonding.

Remember: Potassium tends to lose a single electron (forming a cation) in chemical reactions.

A little bit about transition metals...

Titanium (Ti, Z = 22)



Titanium usually forms one of two ions: +2 and +4

- if Ti loses only the valence electrons, forms +2. If it also loses the 3d electrons, forms +4.

- Most transition metals have TWO valence electrons (in an "s" subshell), and the other ions they form come from electron loss in "d" subshells.

BONDING AND ELECTRON CONFIGURATION

- Filled and half-filled subshells seem to be preferred by atoms.

ELECTRON CONFIGURATION (SHORT FORM)

- We can represent the electron configuration without drawing a diagram or writing down pages of quantum numbers every time. We write the "electron configuration".

# of electrons in the subshell

$$\text{Ti}: \underbrace{1}_{n}\underbrace{s}_{\ell}{}^{2}2s^{2}2p^{6}3s^{2}3p^{6}3d^{2}\underbrace{4s^{2}}_{\text{valence electrons}}$$

$$\text{Ti}: \underbrace{[\text{Ar}]}\,3d^{2}\underbrace{4s^{2}}_{\text{valence electrons}}$$

"noble gas core". We're saying that titanium has the same electron configuration as argon does, with the addition of the electrons that follow. This is a useful shorthand, since the "core" electrons generally don't get involved in bonding.

ELECTRON CONFIGURATION AND THE PERIODIC TABLE

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

"d" block: The d block is shifted DOWN.!

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	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 4s	Ca	Sc 3d	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga 4p	Ge	As	Se	Br	Kr
	Rb	Sr	Y 4d	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
	Cs	Ba	* La 5d	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
	Fr	Ra	* Ac 6d	Rf	Db	Sg	Bh	Hs	Mt									

*"inner" transition metals go here

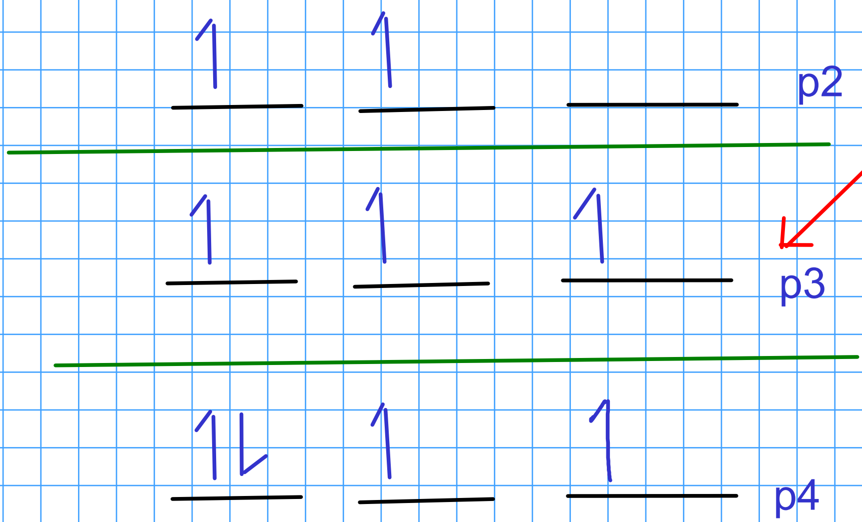
Example: Phosphorus (P): $1s^2 2s^2 2p^6 3s^2 3p^3$

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

Phosphorus has five valence electrons!

Hund's Rule

- When you have two or more orbitals with equivalent energy, electrons will go into each equivalent orbital BEFORE pairing. Pairing costs a bit of energy - less than going to a higher-energy orbital, but more than going to another equivalent orbital.



Electron configurations with filled subshells OR HALF-FILLED SUBSHELLS are more stable than other configurations.

(can explain some transition metal chemistry)

Electrons begin to pair only AFTER all equivalent "p" orbitals are full.

Experimental evidence for Hund's rule:

"Paramagnetism" - attraction of an atom to a magnetic field

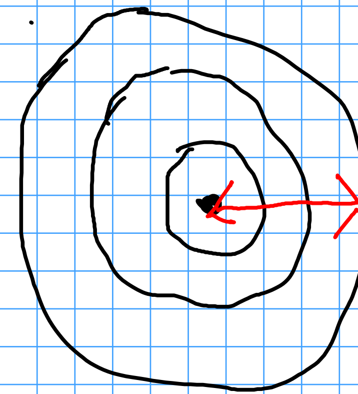
* Spinning electrons are magnetic, but OPPOSITE spins cancel each other out.

* Atoms with unpaired electrons are paramagnetic, while atoms containing only paired electrons are not.

PERIODIC TRENDS

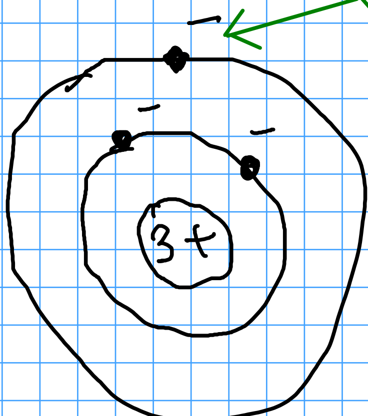
- Some properties of elements can be related to their positions on the periodic table.

ATOMIC RADIUS



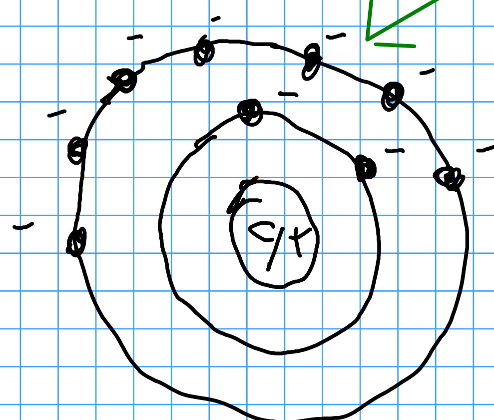
- The distance between the nucleus of the atoms and the outermost shell of the electron cloud.
- Relates to the size of the atom.
- As you go DOWN A GROUP ($\downarrow$), the atomic radius INCREASES.
    - Why? As you go down a period, you are ADDING SHELLS!
- As you go ACROSS A PERIOD ($\longrightarrow$), the atomic radius DECREASES

Why? Let's look at some sample atoms.



Outer electron sees an effective +1 charge (shielded from +3 nucleus by 2 electrons)

lithium



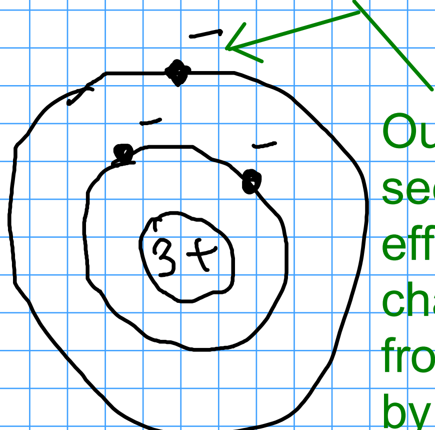
Outer electrons see an effective +7 charge (shielded from +9 nucleus by 2 electrons)

fluorine

... so fluorine's outer shell is pulled closer to the nucleus than lithium's!

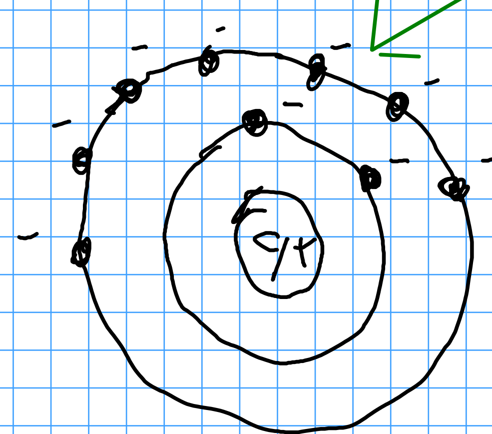
(FIRST) IONIZATION ENERGY

- The amount of energy required to remove a single electron from the outer shell of an atom.
- Relates to reactivity for metals. The easier it is to remove an electron, the more reactive the metal.
- As you go DOWN A GROUP ($\downarrow$), the ionization ebergy DECREASES.
  - Why? As you go down a period, you are ADDING SHELLS. Since the outer electrons are farther friom the nucleus and charge attraction lessens with distance, this makes electrons easier to remove as the atoms get bigger!
- As you go ACROSS A PERIOD ($\longrightarrow$), the ionization energy INCREASES.
  - Why? Let's look at some sample atoms.



lithium

Outer electron sees an effective +1 charge (shielded from +3 nucleus by 2 electrons)

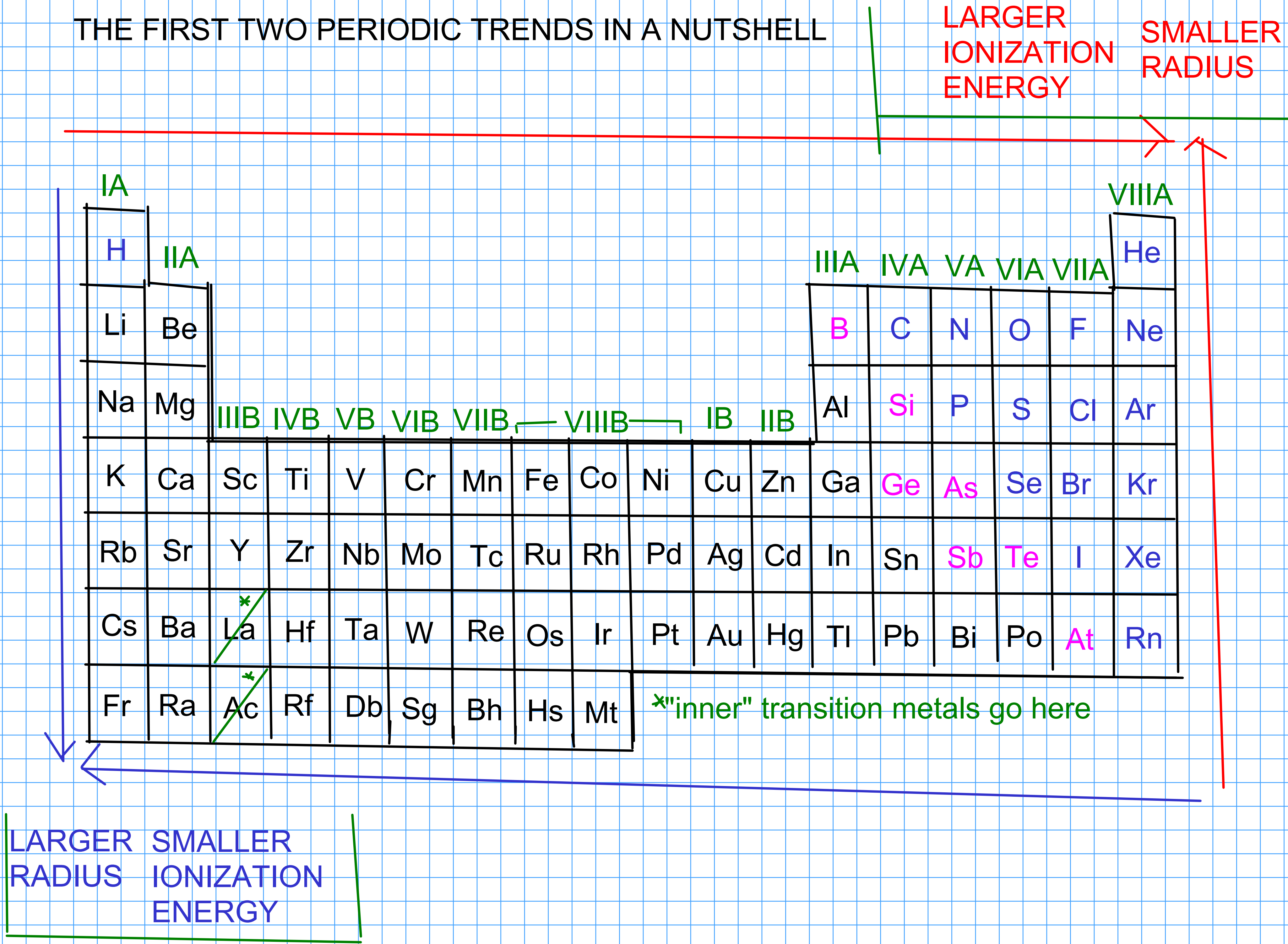


fluorine

Outer electrons see an effective +7 charge (shielded from +9 nucleus by 2 electrons)

... since fluorine's outer electrons are held on by a larger effective charge, they are more difficult to remove than lithium's.

THE FIRST TWO PERIODIC TRENDS IN A NUTSHELL



ELECTRON AFFINITY

- the electron affinity is the ENERGY CHANGE on adding a single electron to an atom.
    - Atoms with a positive electron affinity cannot form anions.
    - The more negative the electron affinity, the more stable the anion formed!

- General trend: As you move to the right on the periodic table, the electron affinity becomes more negative.

EXCEPTIONS

- Group IIA does not form anions (positive electron affinity)!

ns^2 valence electrons for Group IIA!

n: period number

- To add an electron, the atom must put it into a higher-energy (p) subshell.

- Group VA: can form anions, but has a more POSITIVE electron affinity than IVA

ns^2np^3 valence electrons for Group VA!

np^3: Half-full "p" subshell! To add an electron, must start pairing!

- Group VIIIA (noble gases) does not form anions

ns^2np^6: full "s" and "p" subshells!