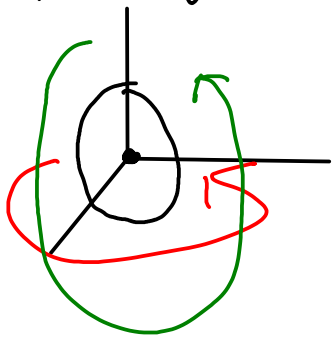
③ MAGNETIC QUANTUM NUMBER m_ℓ

- Represents the ORIENTATION of a subshell in 3D space.

$m_\ell = -\ell$ to $+\ell$, integers

$\ell = 0$, $m_\ell = 0$

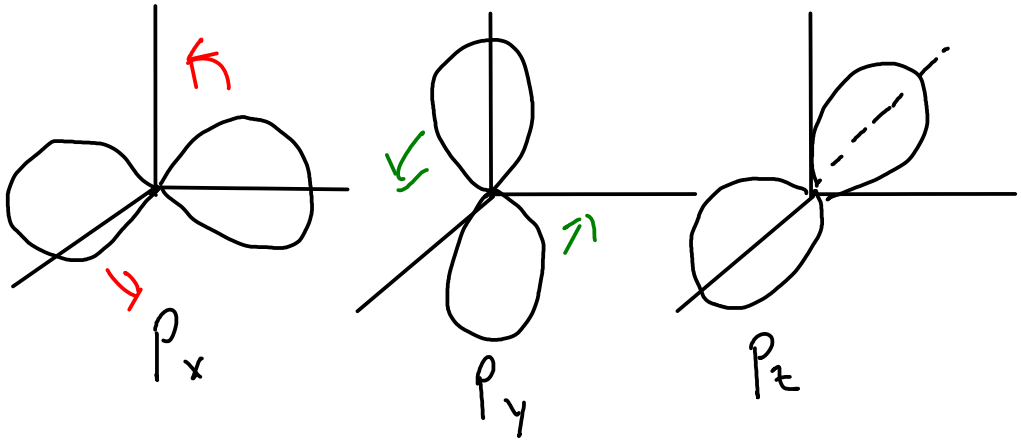
"s"



There is only one possible orientation for an "s" subshell!

$\ell = 1$; $m_\ell = -1, 0, 1$

"p"



There are THREE possible orientations for a "p" subshell!

$\ell = 2$, $m_\ell = -2, -1, 0, 1, 2$ (five orientations)

"d"

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$\ell = 3$, $m_\ell = -3, -2, -1, 0, 1, 2, 3$ (seven orientations)

"f"

... all the arrangements of a single subshell have the same energy. The magnetic quantum number DOESN'T contribute to the energy of an electron.

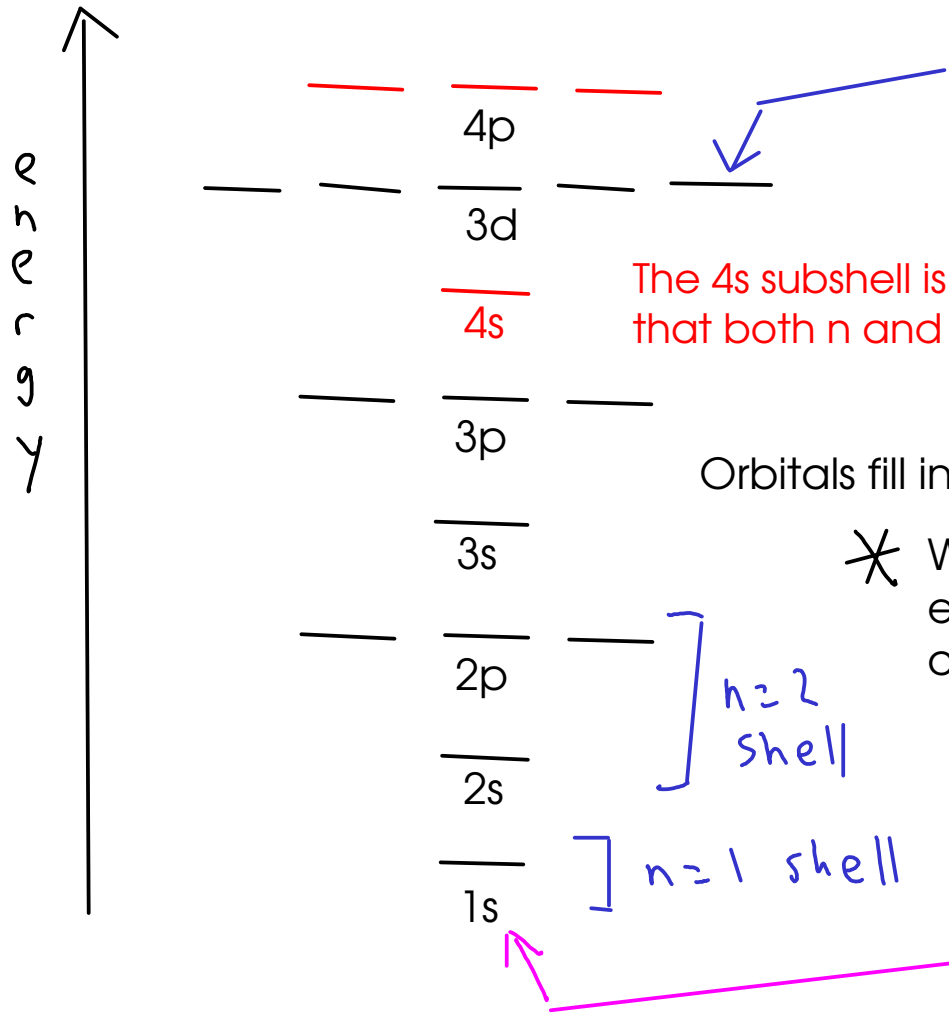
(4) (MAGNETIC) SPIN QUANTUM NUMBER: m_s

$m_s = -1/2$ OR $+1/2$ "spin down" or "spin up"

- An ORBITAL (region with fixed "n", "l" and "ml" values) can hold TWO electrons.

ORBITAL DIAGRAM

- A graphical representation of the quantum number "map" of electrons around an atom.



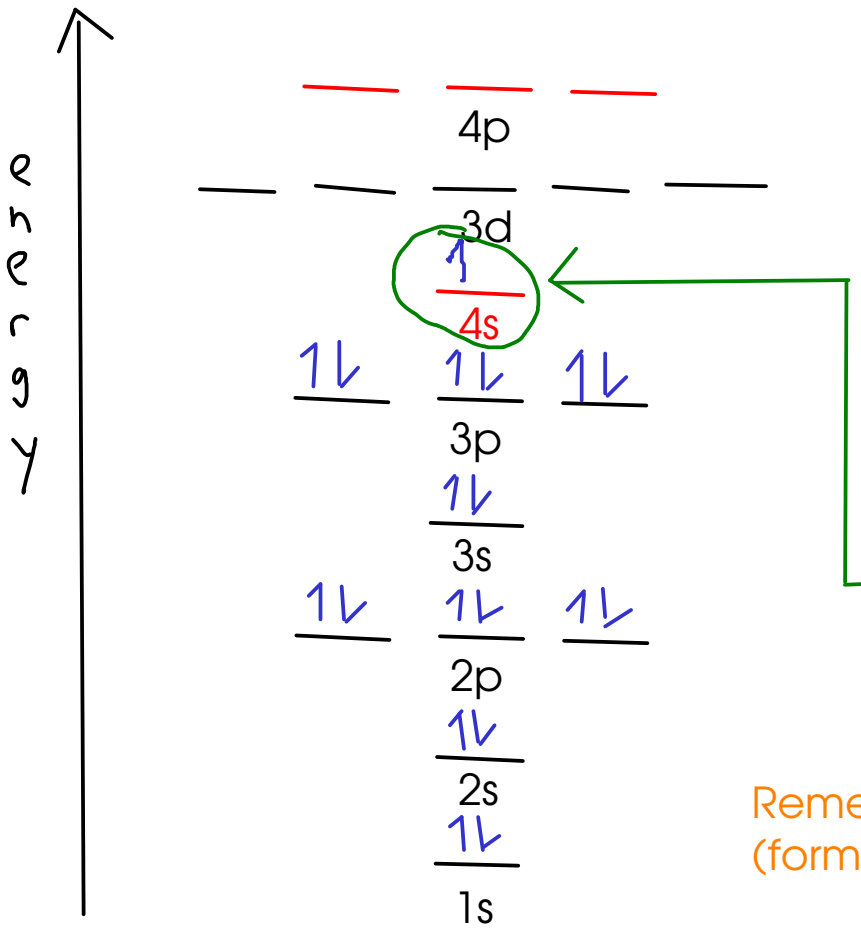
Each blank represents an ORBITAL, and can hold two electrons.

The 4s subshell is lower energy than the 3d subshells .... remember that both n and "l" contribute to energy!

Orbitals fill in order: The lower energy orbitals fill first.

* Where there's more than one orbital with the same energy, electrons don't pair up until each one has one electron in it!

1s: This means the first shell (n=1) and the "s" subshell ("l"=0)



How would an orbital diagram for the element POTASSIUM look?

K , Z = 19

atomic number

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

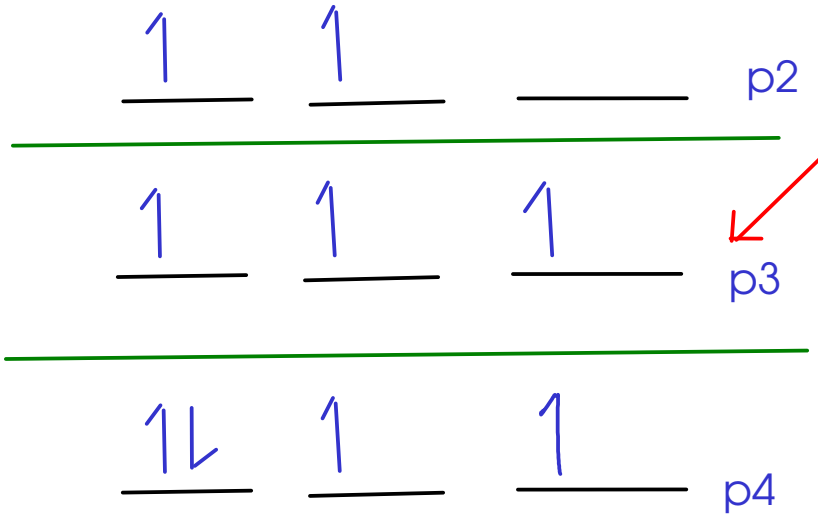
Remember: Potassium tends to lose a single electron (forming a cation) in chemical reactions. K^+

A note on chemical bonding and electron arrangement:

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

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)

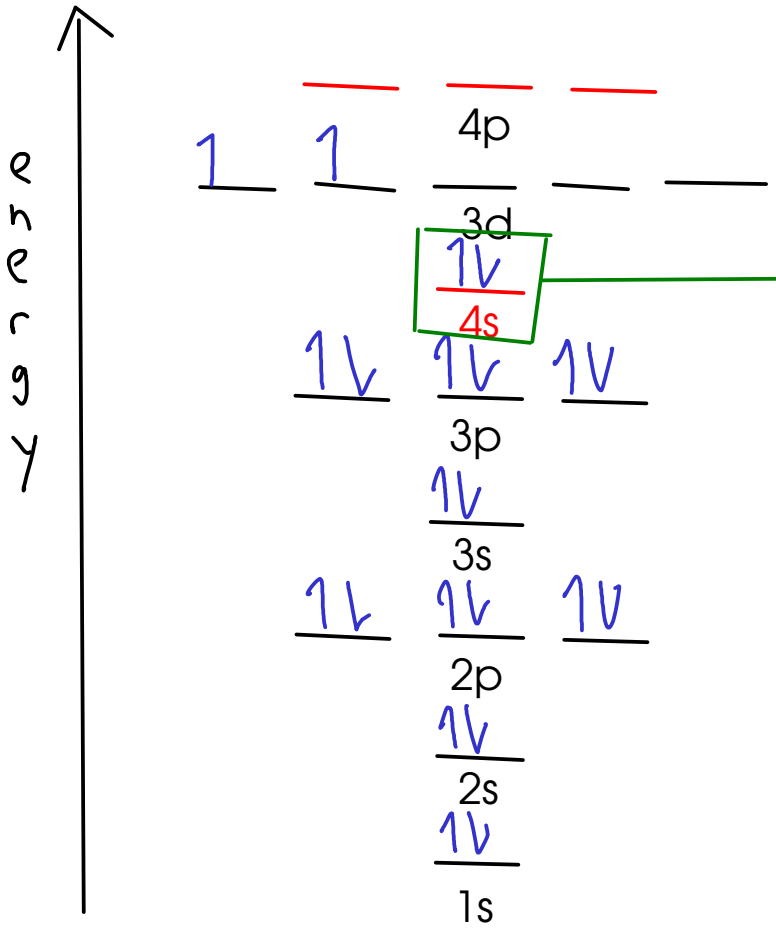
Electrons begin to pair only AFTER all equivalent "p" orbitals are full.

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.

A little bit about transition metals...



Titanium (Ti, Z = 22)

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

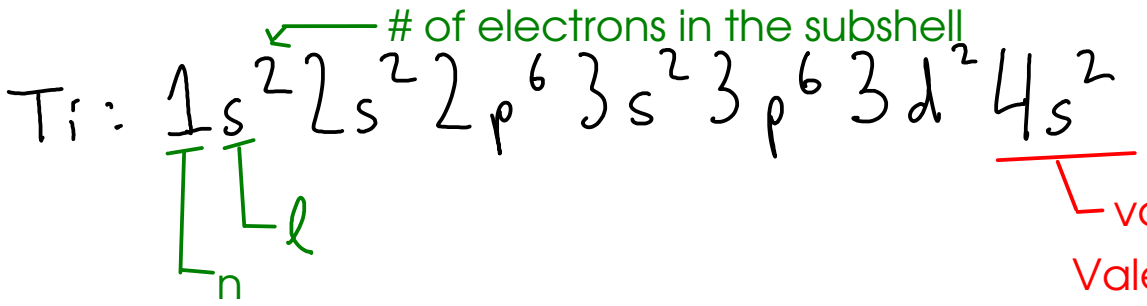
- if Ti loses only the valence electrons, forms +2. If it also loses the 3d electrons, forms +4.

valence electrons

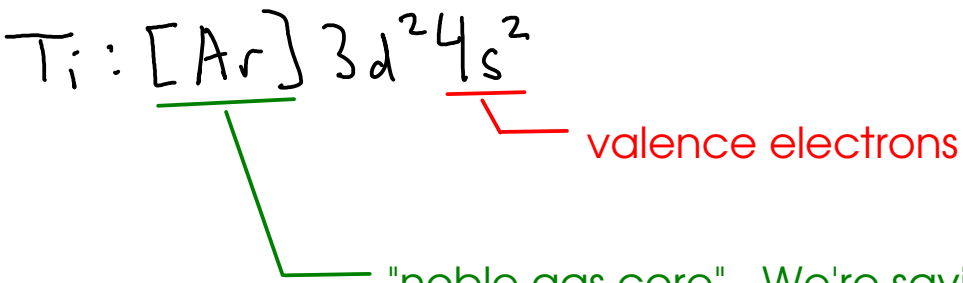
- Most transition metals have TWO valence electrons (in an "s" subshell), and the other ions they form come from electron loss in "d" subshells.

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".

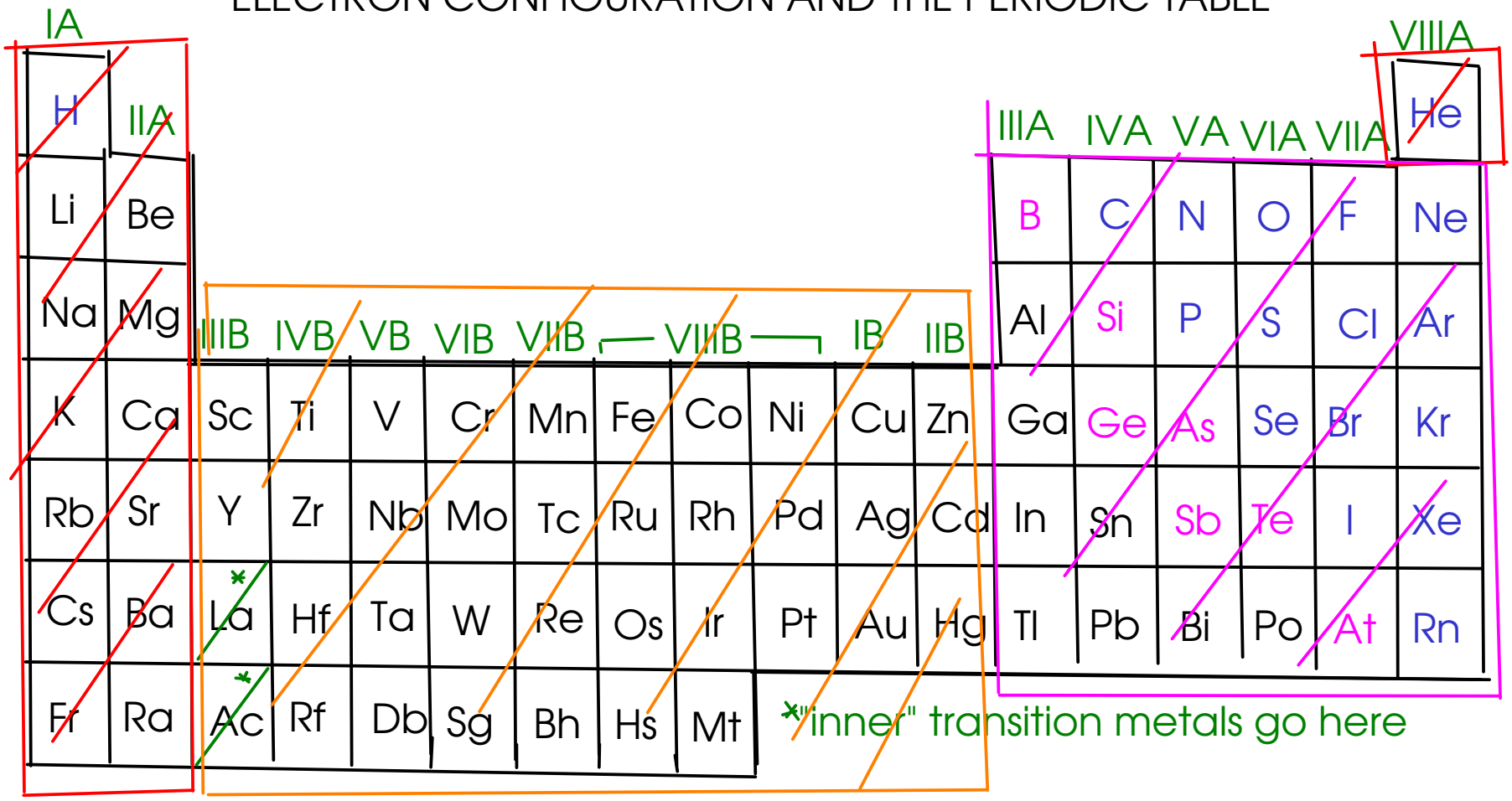


Valence electrons have the largest value for "n"!



"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

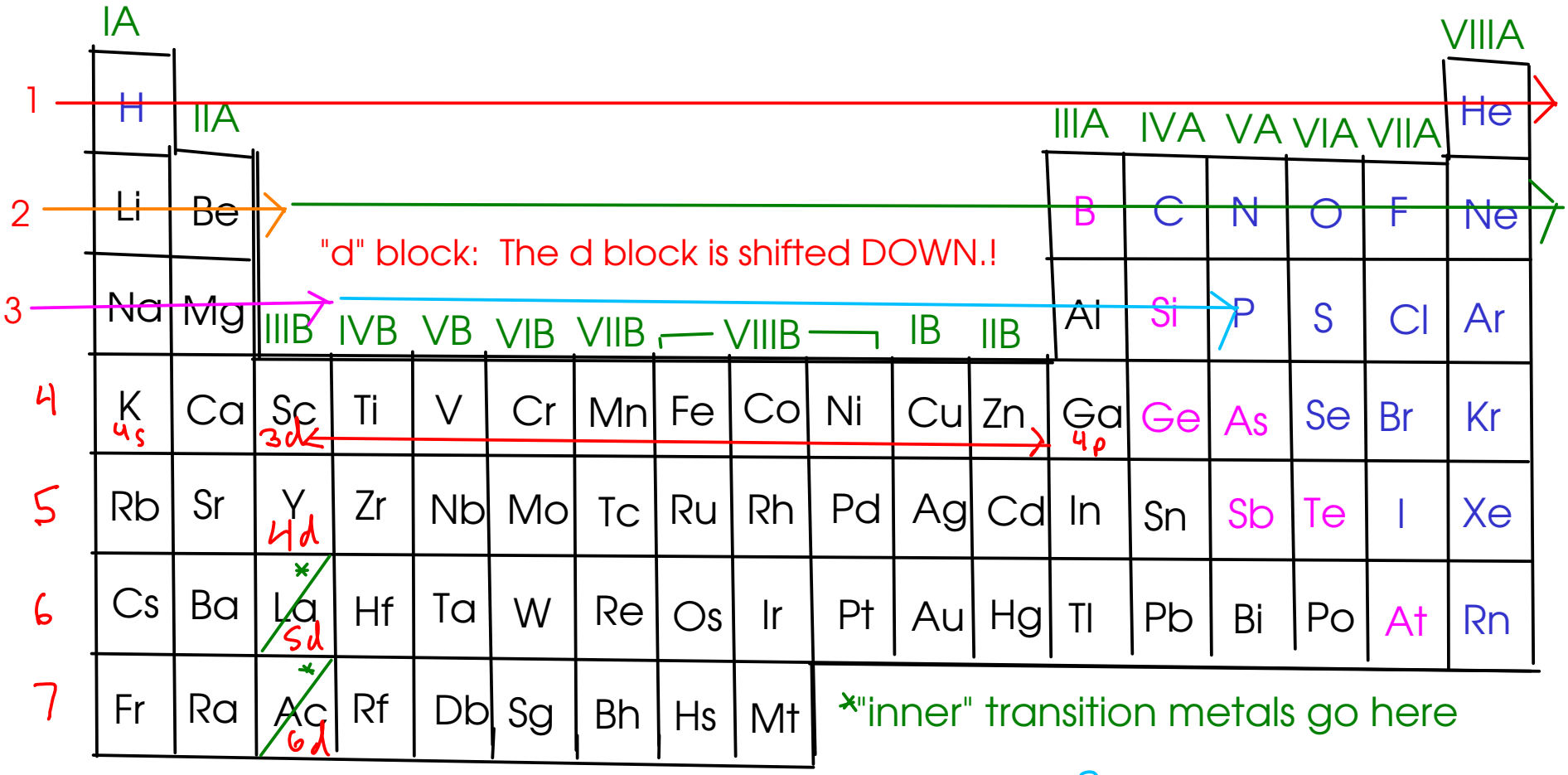


"s" block: last electron in these atoms is in an "s" orbital!

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

"d" block: last electron in 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): $1s^2 2s^2 2p^6 3s^2 3p^3$

Noble gas core notation for P: $[Ne]\,3s^2 3p^3$

EXAMPLES:

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

Remember - valence electrons are ALL of the electrons in the outermost SHELL (n)! More that one subshell (l) may be included in the valence electrons

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

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

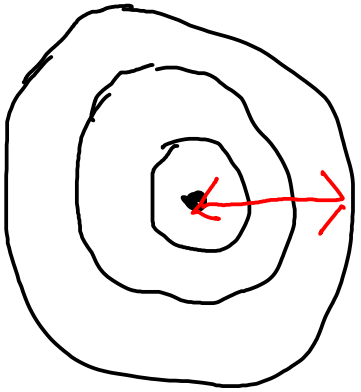
~~[Kr]~~

Sample f-block element

Ce: $[Xe]6s^2 5d^1 4f^1$

PERIODIC TRENDS

ATOMIC RADIUS



- The distance between the nucleus of the atoms and the outermost shell of the electron cloud.

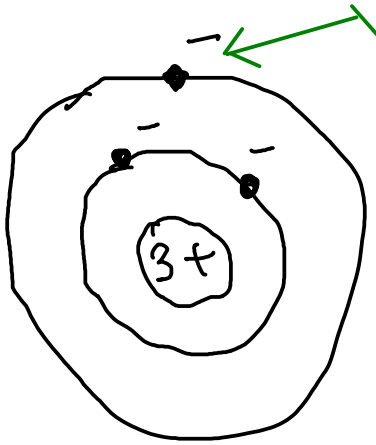
- Relates to the size of the atom.

- As you go DOWN A GROUP (↓), the atomic radius INCREASES.

- Why? As you go down a group, you are ADDING SHELLS!

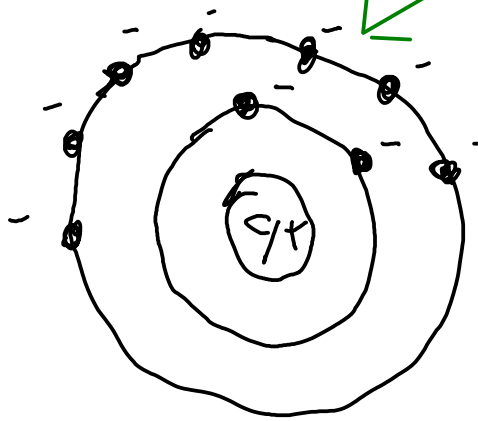
- As you go ACROSS A PERIOD (→), the atomic radius DECREASES

Why? Let's look at some sample atoms.



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

lithium $1s^2 2s^1$



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

fluorine $1s^2 2s^2 2p^5$

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