

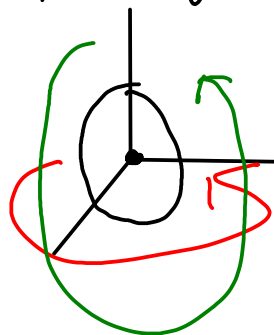
③ MAGNETIC QUANTUM NUMBER m_l

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

$$m_l = -l \text{ to } +l, \text{ integers}$$

$$l = 0, m_l = 0$$

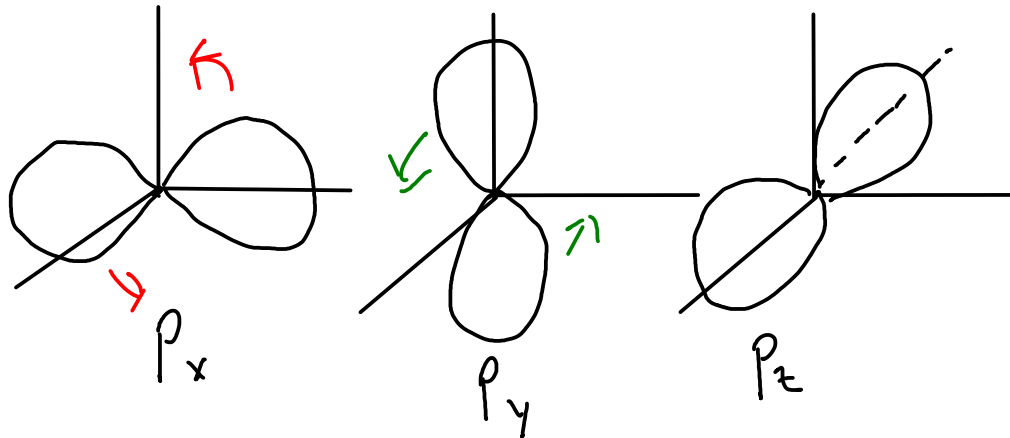
"s"



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

$$l = 1, m_l = -1, 0, 1$$

"p"



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

$$l = 2, m_l = -2, -1, 0, 1, 2 \text{ (five orientations)}$$

"d"

picture
p305

$$l = 3, m_l = -3, -2, -1, 0, 1, 2, 3 \text{ (seven orientations)}$$

"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.

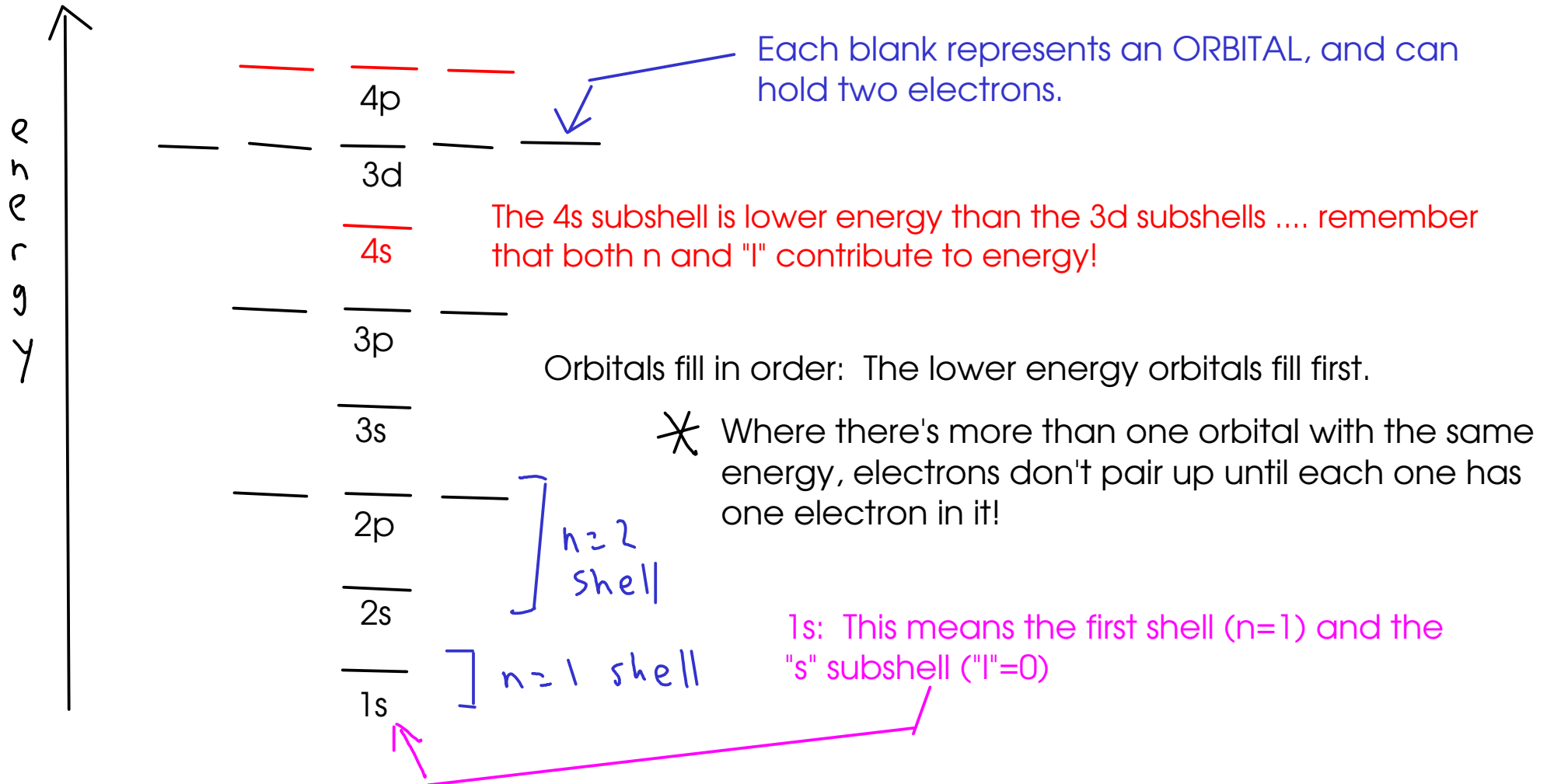
(4) (MAGNETIC) SPIN QUANTUM NUMBER: m_s

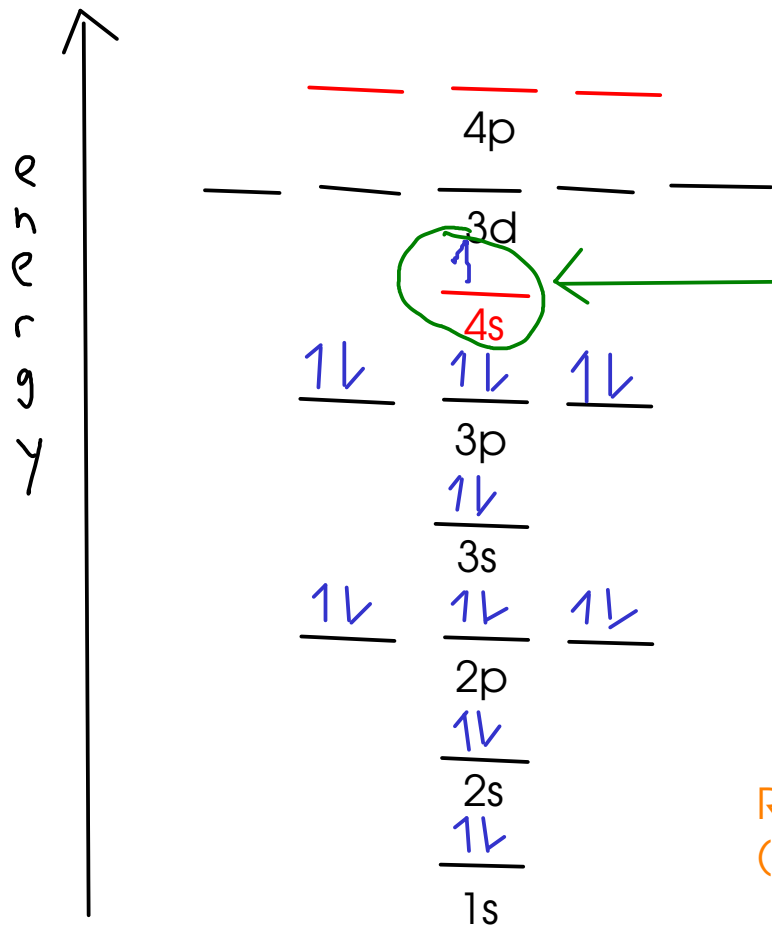
$$m_s = -\frac{1}{2} \text{ OR } +\frac{1}{2} \quad \text{"spin down" or "spin up"}$$

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

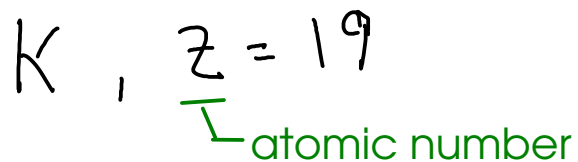
ORBITAL DIAGRAM

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





How would an orbital diagram for the element POTASSIUM look?



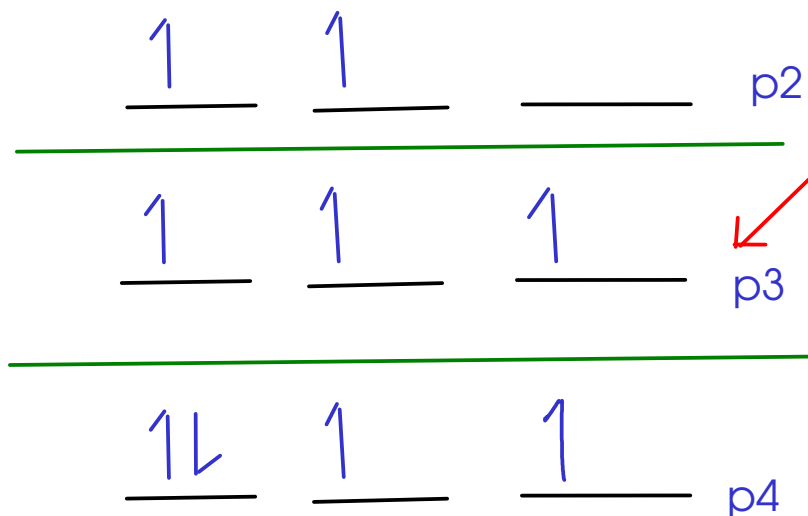
Electrons in the outermost shell of an atom are called VALENCE electrons. THESE electrons are normally involved in chemical bonding.

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)

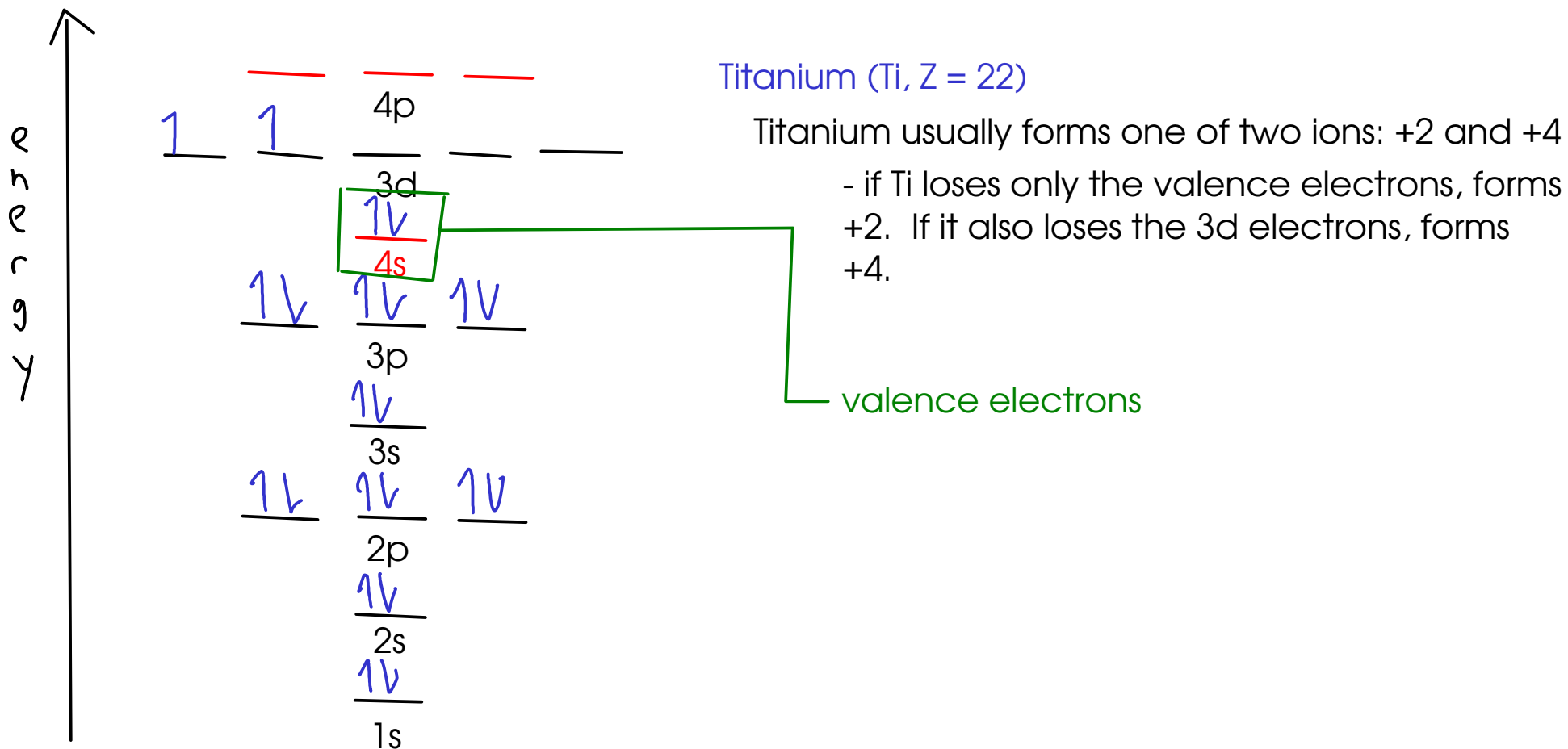
Electrons begin to pair only AFTER all equivalent "p" orbitals are full.

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.

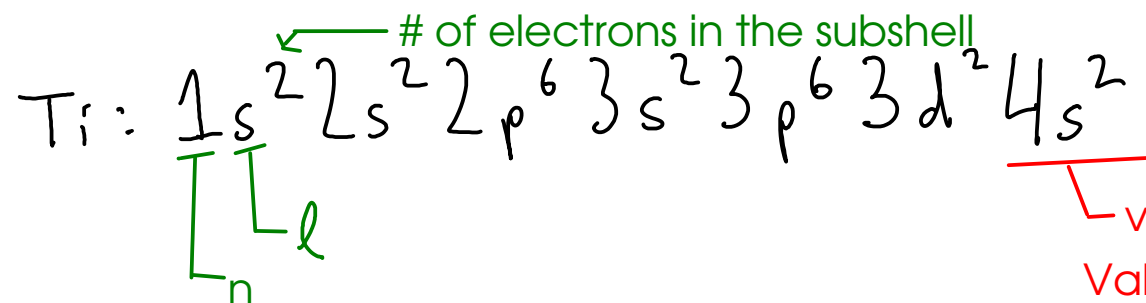
A little bit about transition metals...



- 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

Valence electrons have the largest value for "n"!



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												VIII A					
I A	II A											III A	IV A	V A	VII A	VII A	He
H	He											B	C	N	O	F	Ne
Li	Be											Al	Si	P	S	Cl	Ar
Na	Mg	III B	IV B	V B	VII B	VIII B	IX B	X B	IB	IIB	Ga	Ge	As	Se	Br	Kr	
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	In	Sn	Sb	Te	I	Xe
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	Tl	Pb	Bi	Po	At	Rn
Cs	Ba	La*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg						
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 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!

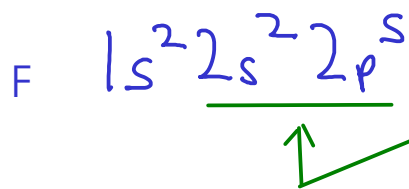
1	H																	He
2	Li	Be										B	C	N	O	F		Ne
3	Na	Mg										Al	Si	P	S	Cl		Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	*"inner" transition metals go here								

Labels above the table: IA, IIA, IIIA, IVA, VA, VIA, VIIA, VIIIA. Labels below the table: IIIB, IVB, VB, VIB, VIIB, VIIIB, IB, IIB. Handwritten notes: "d" block: The d block is shifted DOWN! (in red); "inner" transition metals go here (in green). Arrows indicate the order of filling: red (1s), orange (2s), purple (3s), red (4s), red (3d), red (4p), red (5s), red (4d), red (5p), red (6s), red (4f), red (5d), red (6p), red (7s), red (5f), red (6d), red (7p).

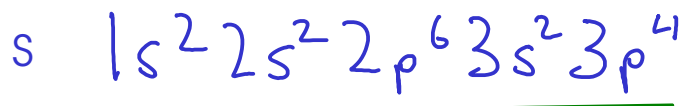
Example: Phosphorus (P): $1s^2 2s^2 2p^6 3s^2 3p^3$

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

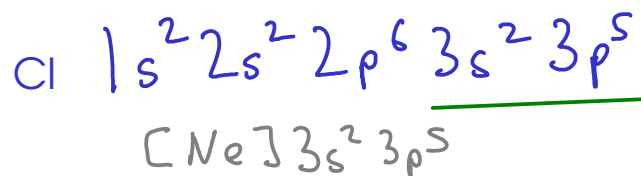
EXAMPLES:



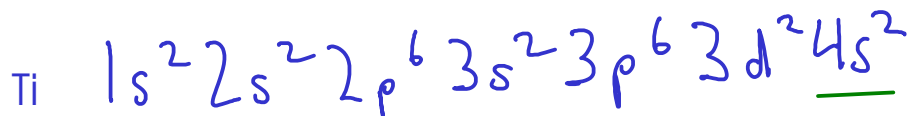
Remember - valence electrons are ALL of the electrons in the outermost SHELL (n)! More than one subshell (l) may be included in the valence electrons



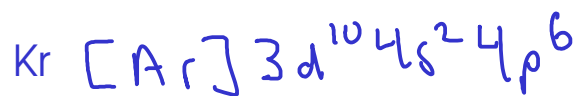
TITANIUM is a transition metal that commonly forms either +2 or +4 cations. The 4s electrons are lost when the +2 ion forms, while the 4s AND 3d electrons are lost to form the +4!



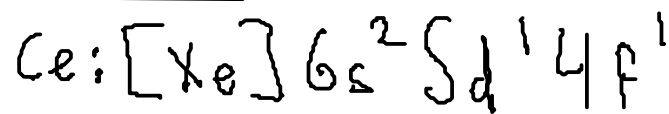
You can order the subshells in numeric order OR in filling order



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



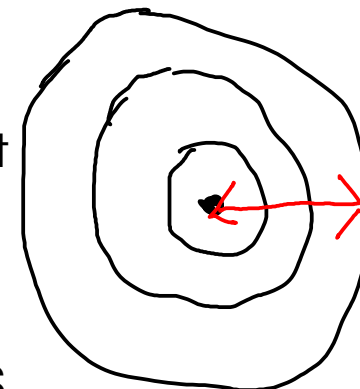
Sample f-block element



PERIODIC TRENDS

ATOMIC RADIUS

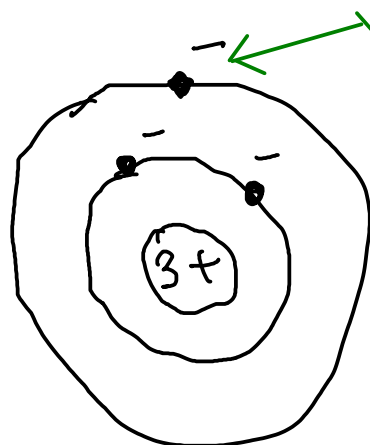
- The distance between the nucleus of the atoms and the outermost shell of the electron cloud.
- Relates to the size of the atom.
- As you go DOWN A GROUP (↓), the atomic radius INCREASES.



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

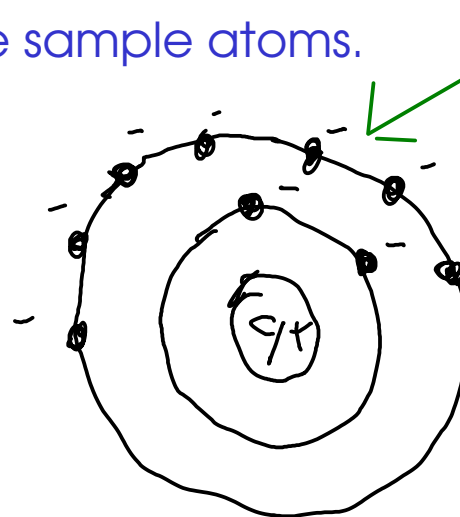
- As you go ACROSS A PERIOD (→), the atomic radius DECREASES

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)

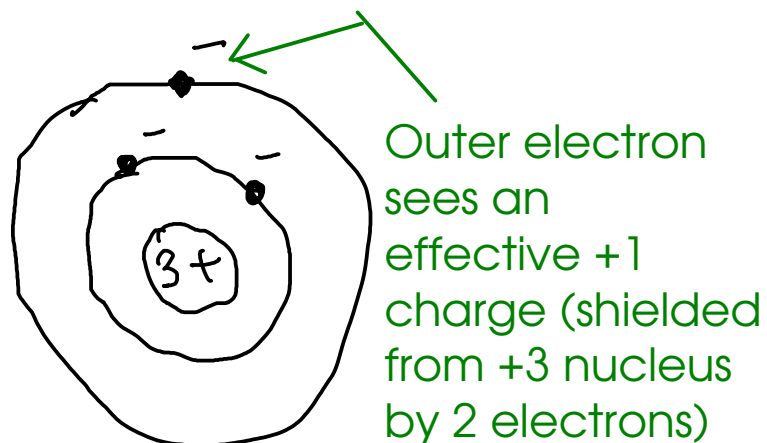
fluorine $1s^2 2s^2 2p^5$

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

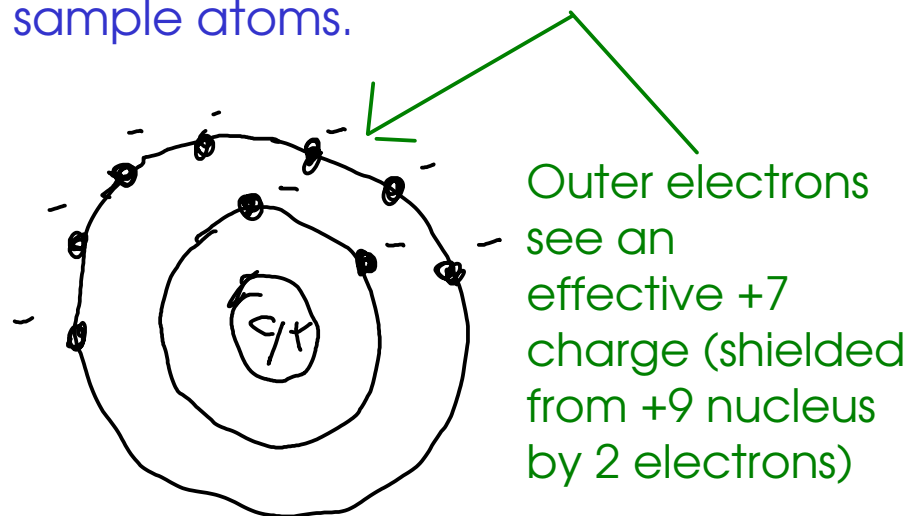
(FIRST) IONIZATION ENERGY

- 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 (↓), the ionization energy DECREASES.
 - Why? As you go down a period, you are ADDING SHELLS. Since the outer electrons are farther from 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 (→), the ionization energy INCREASES.

- Why? Let's look at some sample atoms.



lithium



fluorine

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