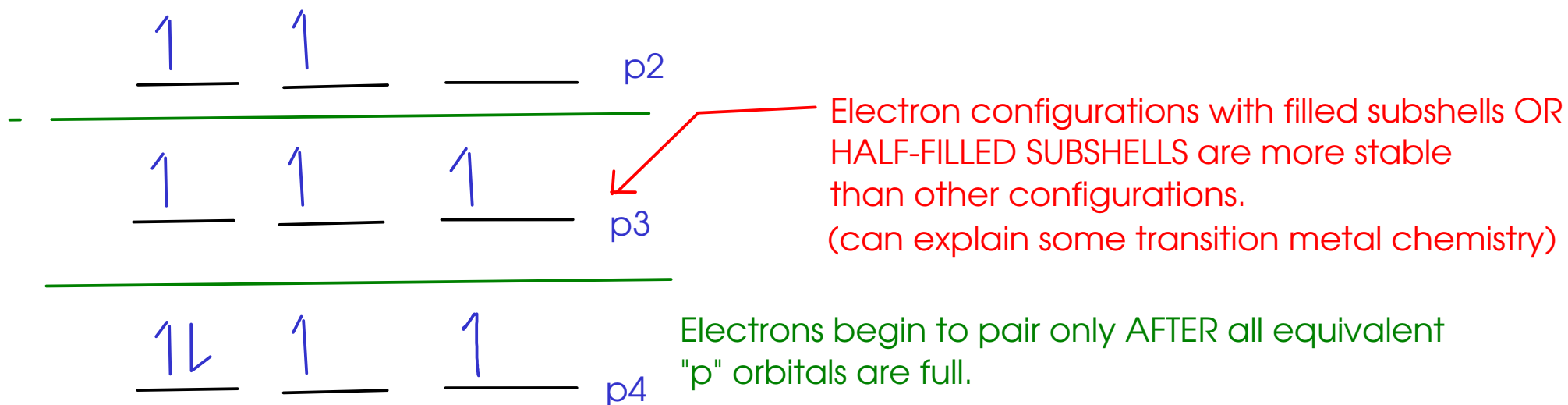


Hund's Rule

- 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.

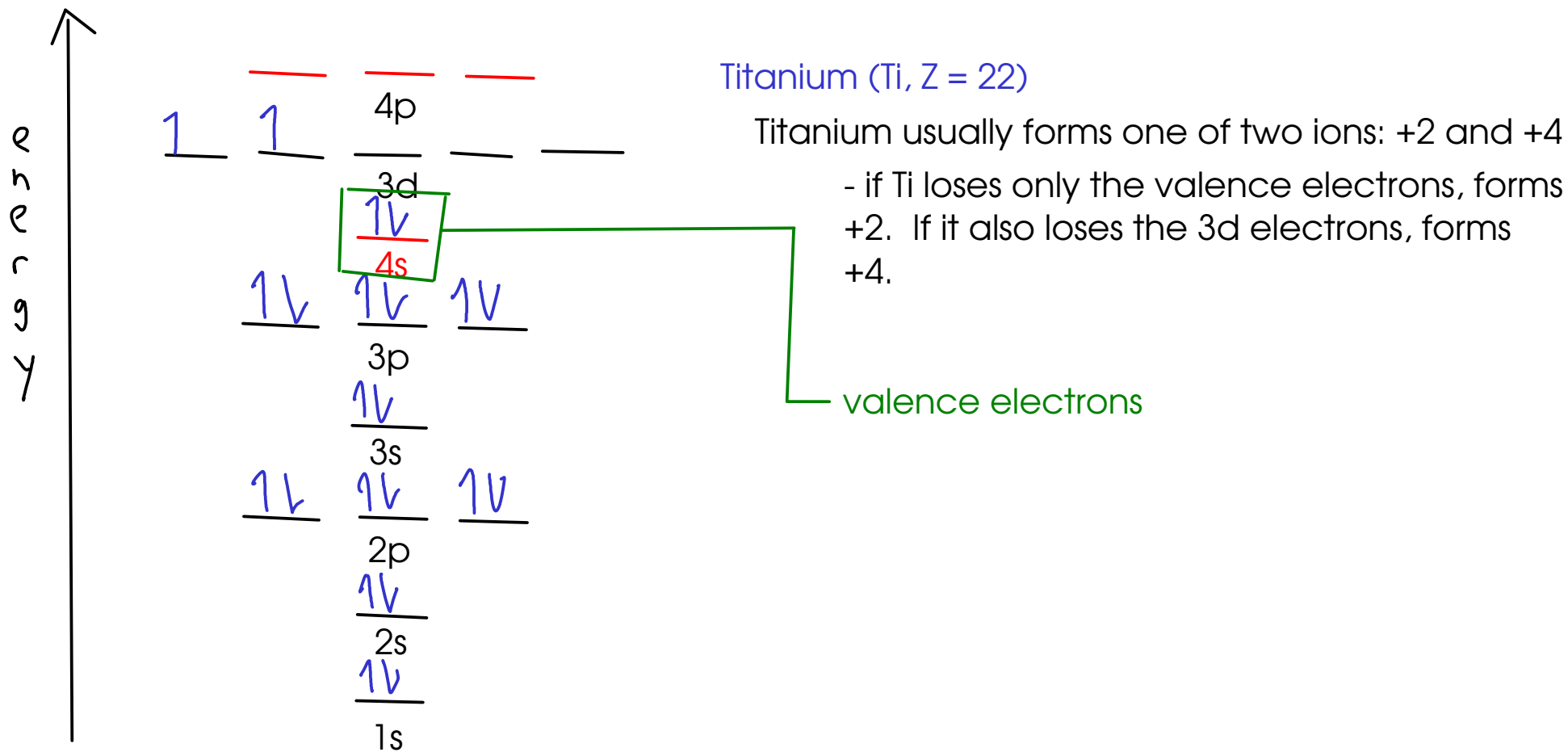


Experimental evidence for Hund's rule:

"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.

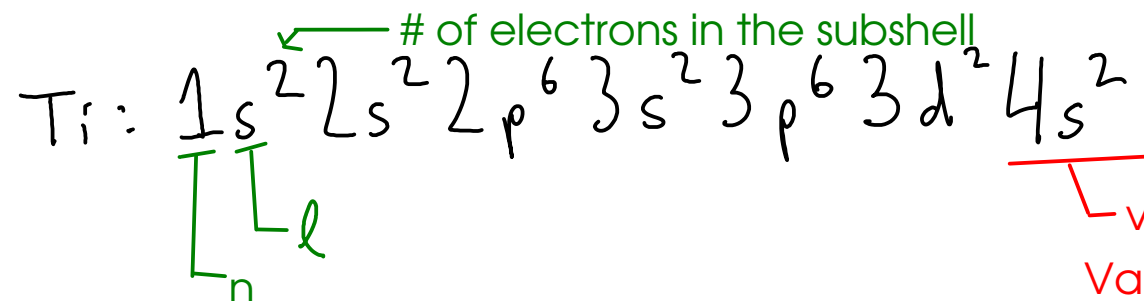
A little bit about transition metals...



- 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													
IA	IIA											III A	IV A	V A	VI A	VII A	VIII A								
H	He																								
Li	Be											B	C	N	O	F	Ne								
Na	Mg	III B	IV B	V B	V I B	V II B	V III B	I B	I I B											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 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!

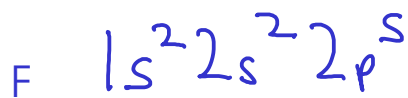
1	IA H											VIIIA He						
2	Li	IIA Be											IIIA B	IVA C	V N	VIA O	VIIA F	VIIIA Ne
3	Na	Mg	IIIB	IVB	VB	VIB	VIIB	VIIIB	IB	IIB	Al	Si	P	S	Cl	Ar		
4	K _{4s}	Ca	Sc _{3d}	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga _{4p}	Ge	As	Se	Br	Kr
5	Rb	Sr	Y _{4d}	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	La _{5d}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

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

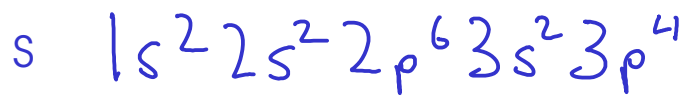
Example: Phosphorus (P): $1s^2 2s^2 2p^6 3s^2 3p^3$

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

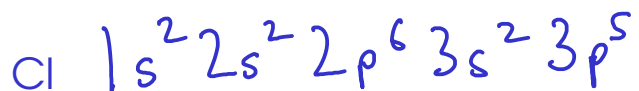
EXAMPLES:



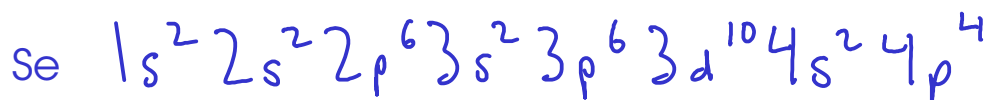
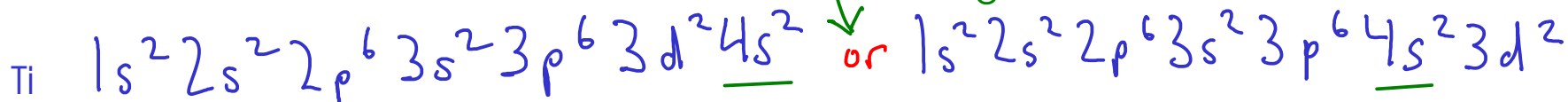
Remember - valence electrons are ALL of the electrons in the outermost SHELL (n)! More that one subshell (l) may be included in the valence electrons



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!



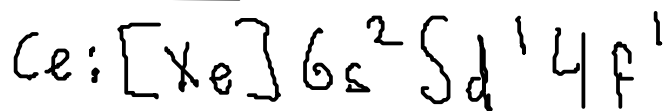
You can order the subshells in numeric order OR in filling order



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

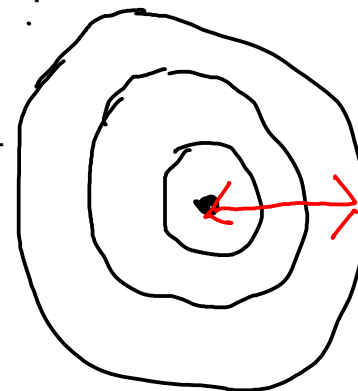


PERIODIC TRENDS

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

ATOMIC RADIUS

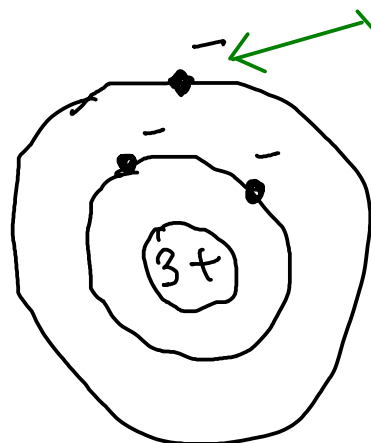
- The distance between the nucleus of the atoms and the outermost shell of the electron cloud.
- Relates to the size of the atom.
- As you go DOWN A GROUP (↓), the atomic radius INCREASES.



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

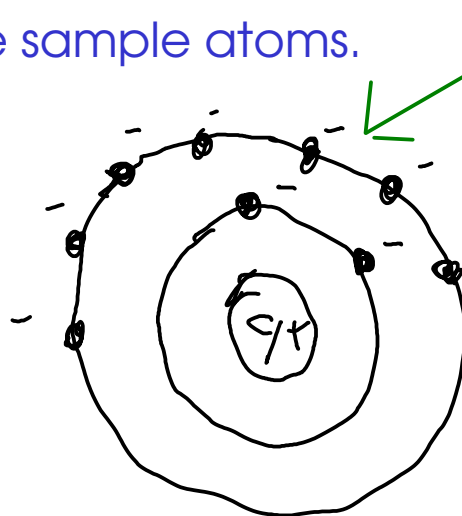
- As you go ACROSS A PERIOD (→), the atomic radius DECREASES

Why? Let's look at some sample atoms.



lithium $1s^2 2s^1$

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



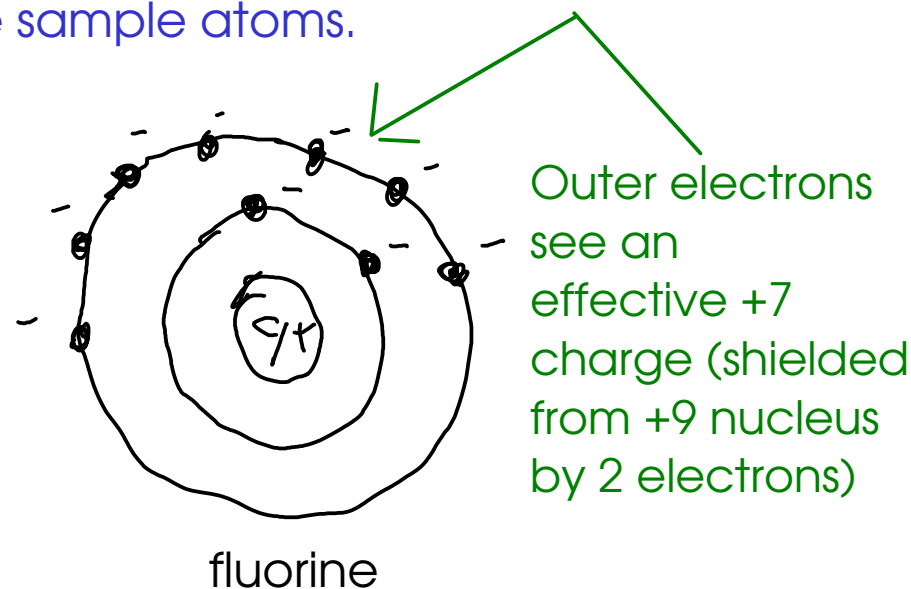
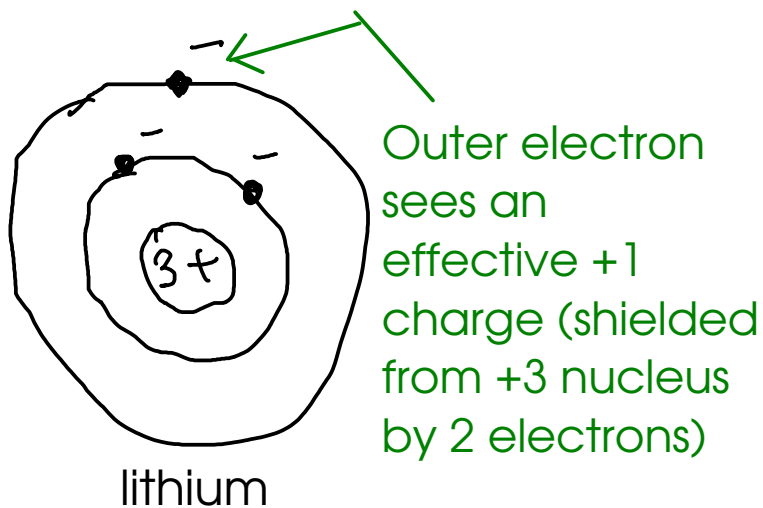
fluorine $1s^2 2s^2 2p^5$

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

... 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.



... 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

LARGER
IONIZATION
ENERGYSMALLER
RADIUS

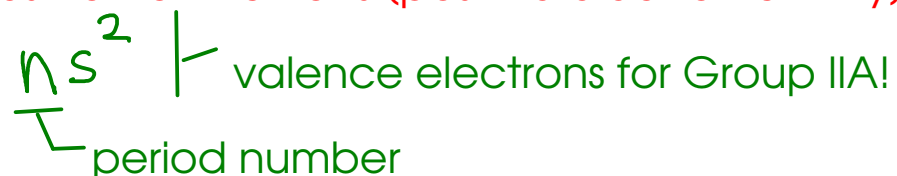
IA																			VIIIA
H																			He
	IIA											IIIA	IVA	VA	VIA	VIIA			
Li	Be											B	C	N	O	F			Ne
Na	Mg	IIIB	IVB	VB	VIB	VII B	VIII B	IB	IIB			Al	Si	P	S	Cl			Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br			Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I			Xe
Cs	Ba	La*	Hf	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										

LARGER
RADIUS
SMALLER
IONIZATION
ENERGY

- 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.

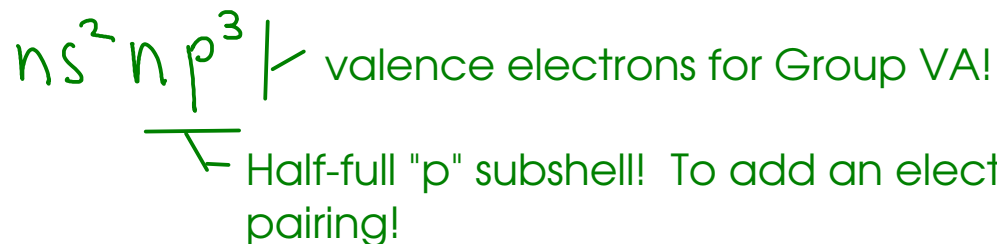
EXCEPTIONS

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



- 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



- Group VIIIA (noble gases) does not form anions



"MAIN" or "REPRESENTATIVE" GROUPS OF THE PERIODIC TABLE

	IA		IIA
1	H		
2	Li		Be
3	Na		Mg
4	K		Ca
5	Rb		Sr
6	Cs		Ba
7	Fr		Ra

Alkali metals

Alkaline earth metals

Read about these in Section 8.7 of the Ebbing textbook!

						VIIIA
III A	IV A	V A	VI A	VII A		He
B	C	N	O	F		Ne
Al	Si	P	S	Cl		Ar
Ga	Ge	As	Se	Br		Kr
In	Sn	Sb	Te	I		Xe
Tl	Pb	Bi	Po	At		Rn

Chalcogens

Halogens

Noble/Inert gases

The representative (main) groups

GROUP IA - the alkali metals

valence electrons:



- React with water to form HYDROXIDES



alkali metals form BASES when
put into water!

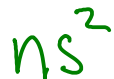
- Alkali metal OXIDES also form bases when put into water. (This is related to METALLIC character. The more metallic something is, the more basic its oxide. Nonmetals have ACIDIC oxides!)



- Physical properties: All of these elements are soft metals with relatively low melting points.

GROUP IIA - the alkaline earth metals

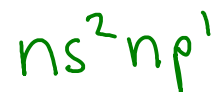
valence electrons:



- May react with water in a reaction similar to the alkali metals, producing hydroxides and hydrogen gas. For some of the alkaline earth metals, this reaction takes place at a significant rate only at high temperatures..
- Form basic oxides, formula: M_2O
- These elements are soft and low-melting ... but harder and higher melting than alkali metals.
- The name "alkaline earth" comes from the observation that the "earths" (oxides) of these metals are basic.

GROUP IIIA

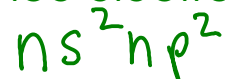
valence electrons:



- most of the elements in this group are metals, but there is also a semiconductor (boron).
- The oxides of these elements are of the form M_2O_3
 - oxides of boron are acidic (metalloids tend to behave more like nonmetals in the acidity of their oxides).
 - Aluminum and gallium have AMPHOTERIC oxides (react as acids or bases), and the larger Group IIIA oxides are basic
- These elements do not react directly with water to make hydroxides, unlike Groups IA and IIA.

GROUP IVA

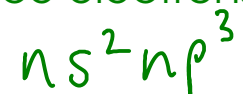
valence electrons



- contains some elements of each type: nonmetal, metalloid, and metal.
- oxides range from acidic to amphoteric, with formulas MO_2 or MO (C, Pb form both!)
- don't react with water to make hydroxides

GROUP VA

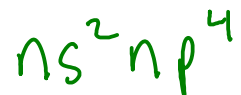
valence electrons



- range from nonmetal to metallic, but with only one metal (bismuth).
- Oxides of group VA nonmetals are acidic, while the group VA metalloids have amphoteric oxides. Bismuth's oxide is basic
- Formulas of these oxides vary considerably, but the most common variants are: RO_2, RO_3

GROUP VIA - the chalcogens

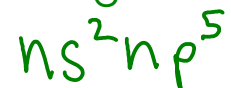
valence electrons



- Like Group VA, formulas of oxides of these elements vary. Common ones are: RO_2, RO_3
- mostly nonmetals/metalloids, plus one metal (polonium). Oxides range from acidic to amphoteric.
- This group's name means - "ore producers" Many metal ores contain oxygen and/or sulfur!

GROUP VIIA - the halogens

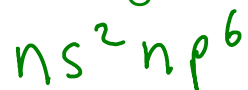
electron configuration:



- react with water, but form ACIDS when they do so! (ex: chlorine and water make HCl and HOCl).
- Oxides of the halogens are not very stable, but they are acidic.
- nonmetals, exist primarily as DIATOMIC MOLECULES.
- halogens are very similar in their chemical reactions, even though their physical appearance varies considerably!
- This group's name means "salt formers" (think sodium chloride)

GROUP VIIIA - the noble or inert gases

electron configuration:



- characterized by their lack of chemical reactivity. The lighter noble gases have no known compounds, while the heavier ones sometimes form molecules with reactive elements like oxygen and fluorine.
- exist primarily as single (uncombined) atoms - NOT diatomic molecules like the halogens.