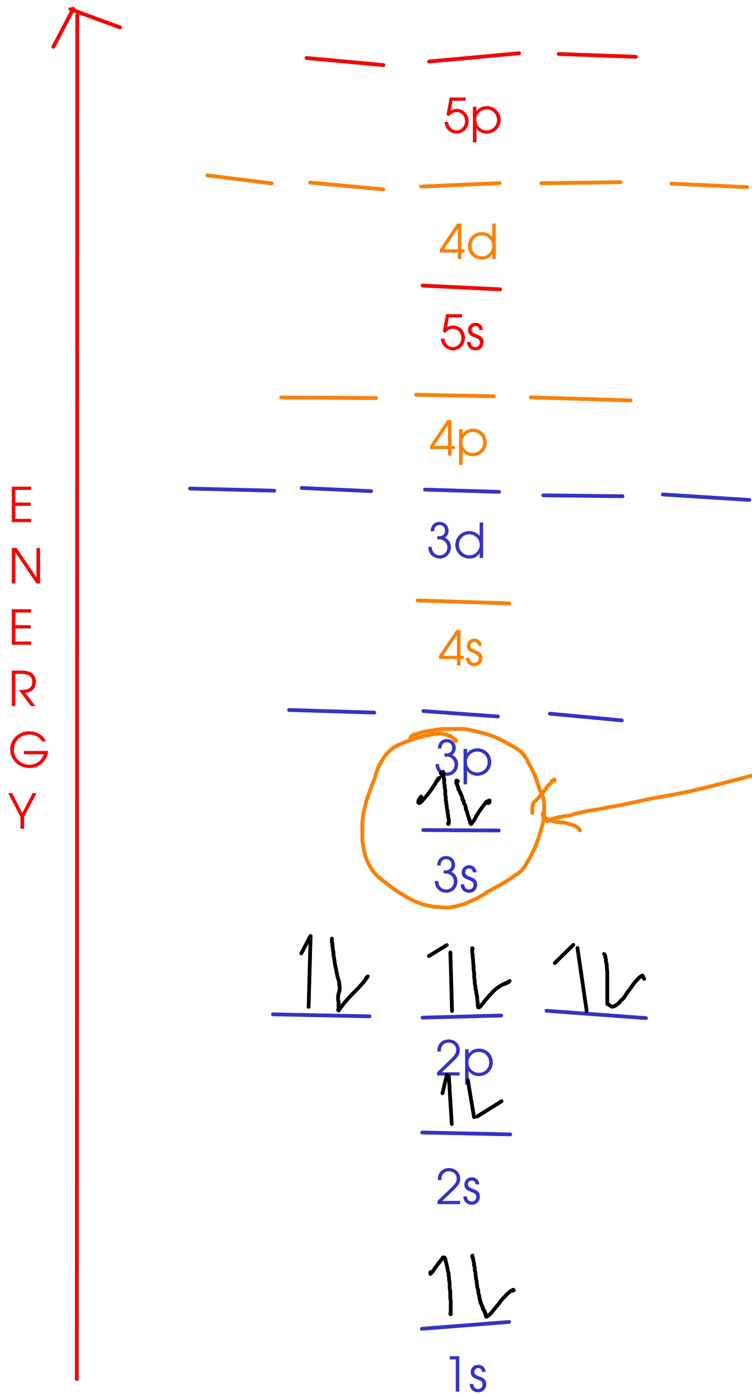
Let's look at some example atoms:

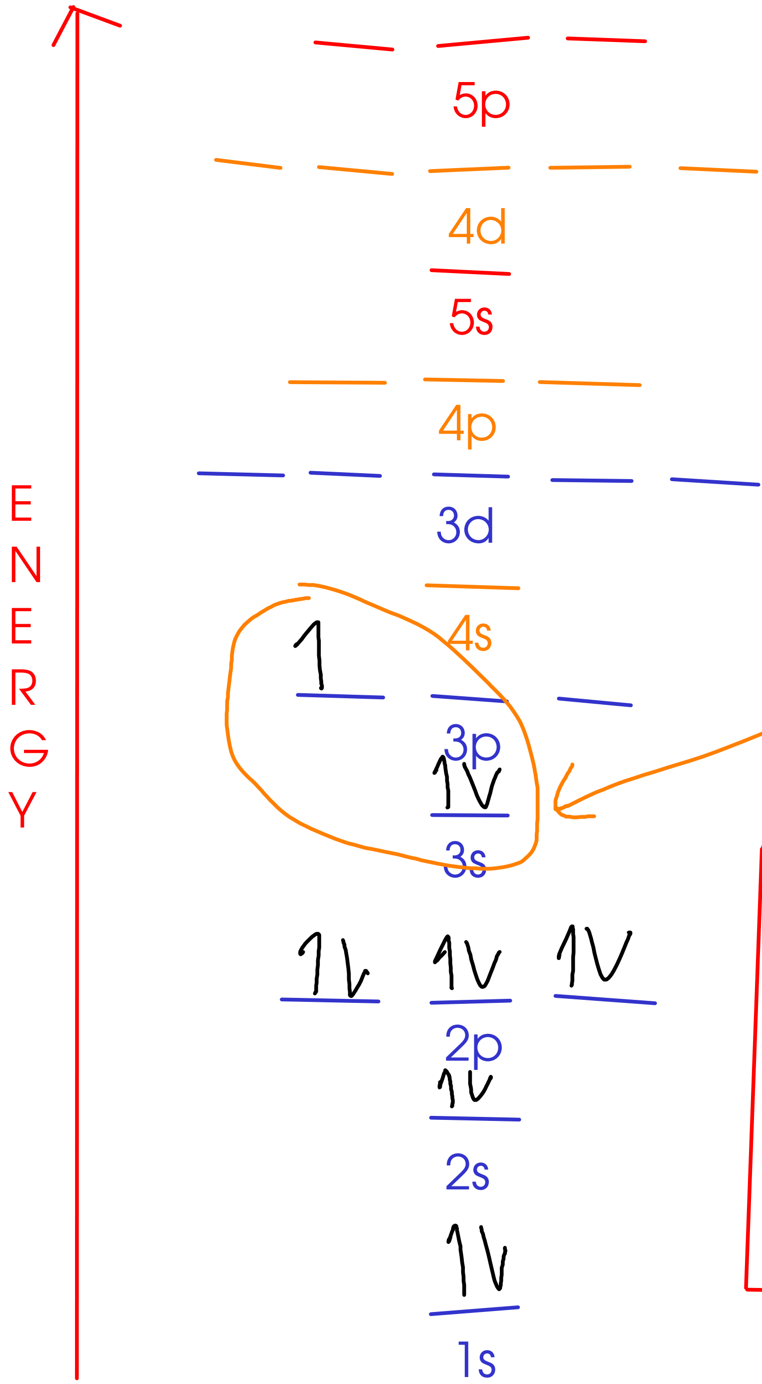
Magnesium: Z=12 , 12 electrons

Z: atomic #



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

Aluminum: Z = 13

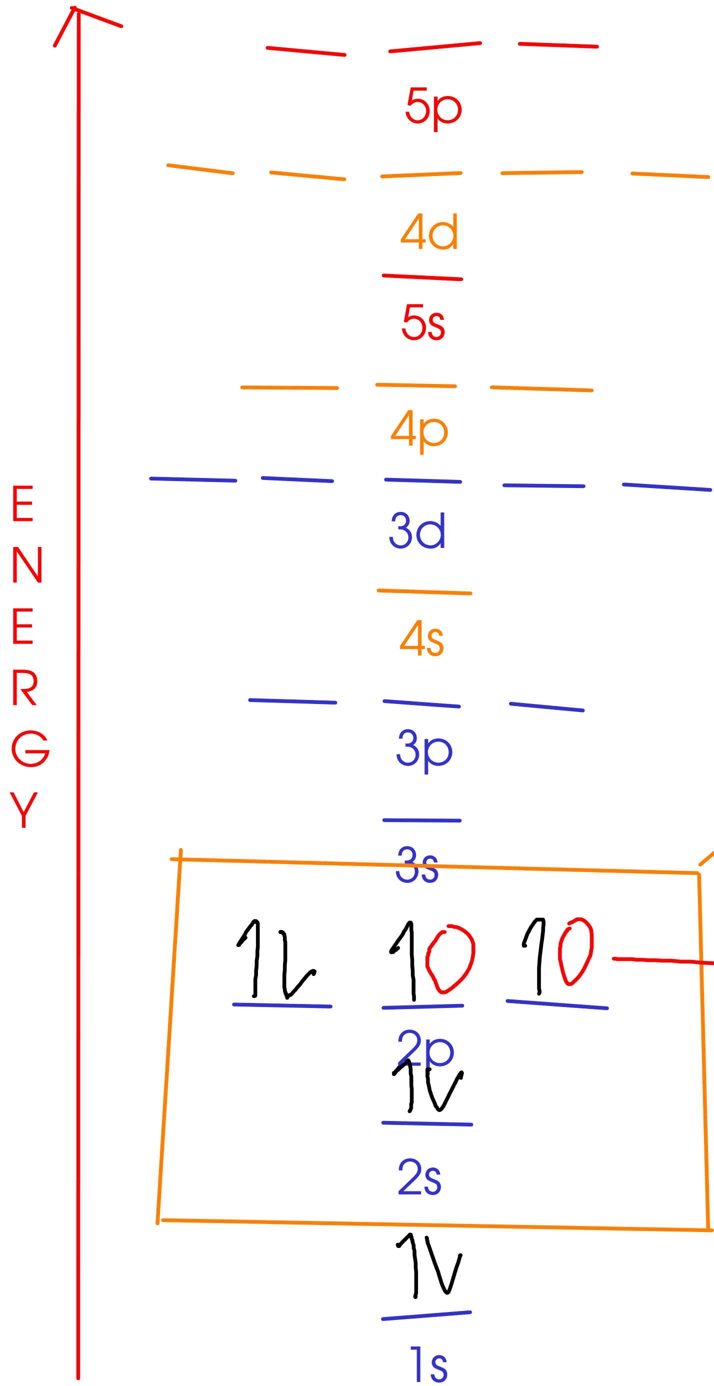


Aluminum has THREE valence electrons!
(All electrons in the outer shell are valence electrons!)

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.

$$O: \underbrace{1s}_{\text{Shell and subshell}}{}^{2}\,2s^{2}\,2p^{\boxed{4}}$$

Number of electrons in the subshell!

Shell
and
subshell

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

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

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

two elements wide

Helium is part of the "s" block!

ELECTRON CONFIGURATION AND THE PERIODIC TABLE

six elements wide

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	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

"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!

	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	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

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

Shortcut: You may use "noblre gas core" notation - which starts from the previous noble gas rather than hydrogen. This is useful for big atoms.

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

EXAMPLES:

F $1s^2 \underline{2s^2 2p^5}$

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

S $1s^2 2s^2 2p^6 \underline{3s^2 3p^4}$

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 $1s^2 2s^2 2p^6 \underline{3s^2 3p^5}$

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

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

Ti $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 \underline{4s^2}$ or $1s^2 2s^2 2p^6 3s^2 3p^6 \underline{4s^2} 3d^2$

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.

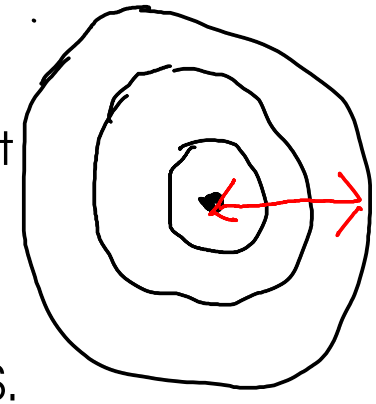
Kr $[Ar]3d^{10} \underline{4s^2 4p^6}$

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

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.

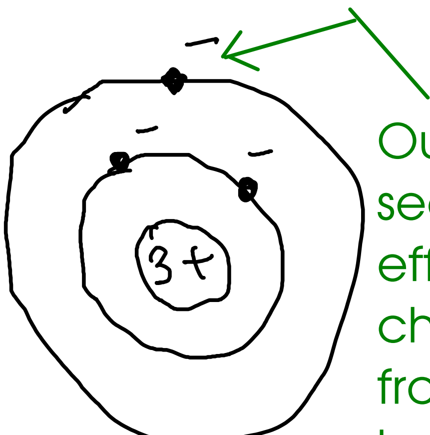
- Relates to the size of the atom.

- As you go DOWN A GROUP ($\downarrow$), the atomic radius INCREASES.

  - Why? As you go down a group (from one period to the next) , you are ADDING SHELLS!

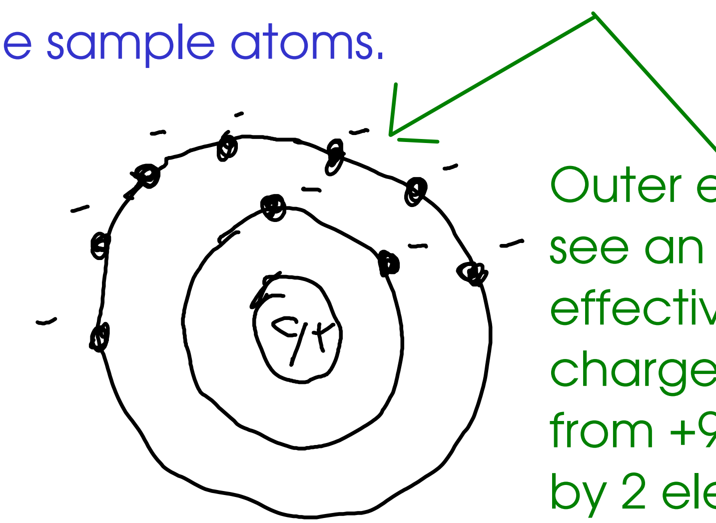
- As you go ACROSS A PERIOD ($\longrightarrow$), the atomic radius DECREASES

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!

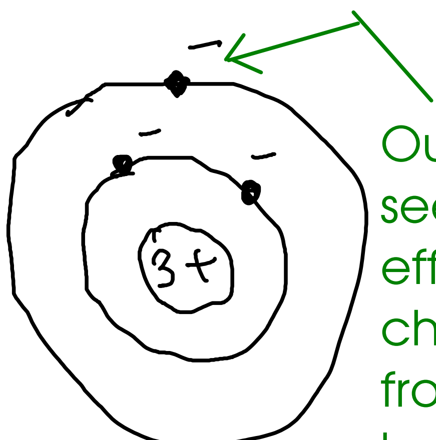
IONIZATION ENERGY (or FIRST IONIZATION ENERGY)

- 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 energy 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!

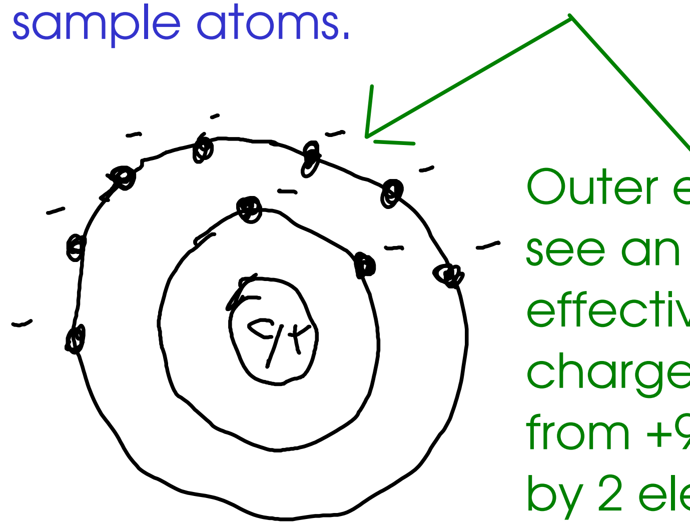
- As you go ACROSS A PERIOD ($\longrightarrow$), the ionization energy INCREASES.

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



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

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

PERIODIC TRENDS IN A NUTSHELL

LARGER IONIZATION ENERGY — SMALLER RADIUS

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

LARGER RADIUS — SMALLER IONIZATION ENERGY

 DESCRIBING CHEMICAL BONDING

"octet rule"

- a "rule of thumb" (NOT a scienfitic law) predicting how atoms will exchange or share electrons to form chemical compounds
- atoms will gain, lose, or share enough electrons so that they end up with full "s" and "p" subshells in their outermost shell.

- Why "octet"? An "s" subshell can hold two electrons, while a "p" subshell can hold six. 2+6 = 8

IONIC COMPOUNDS

- When atoms react to form IONS, they GAIN or LOSE enough electrons to end up with full "s" and "p" subshells.

example:

$$Al + 3\,Br \rightarrow AlBr_3$$

Al: $[Ne]3s^2 3p^1$

Br: $[Ar]3d^{10}4s^2 4p^5$

Br: $[Ar]3d^{10}4s^2 4p^5$

Br: $[Ar]3d^{10}4s^2 4p^5$

$1s^2 2s^2 2p^6$

Al^{3+}: $[Ne]$

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

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

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

Aluminum loses its outer three electrons, and each bromine gains one!

To save space, these electron configurations have been written with the "noble gas core" shortcut. Bromine's electron configuration is exactly like argon's - with the addition of some 3d, 4s, and 4p electrons!

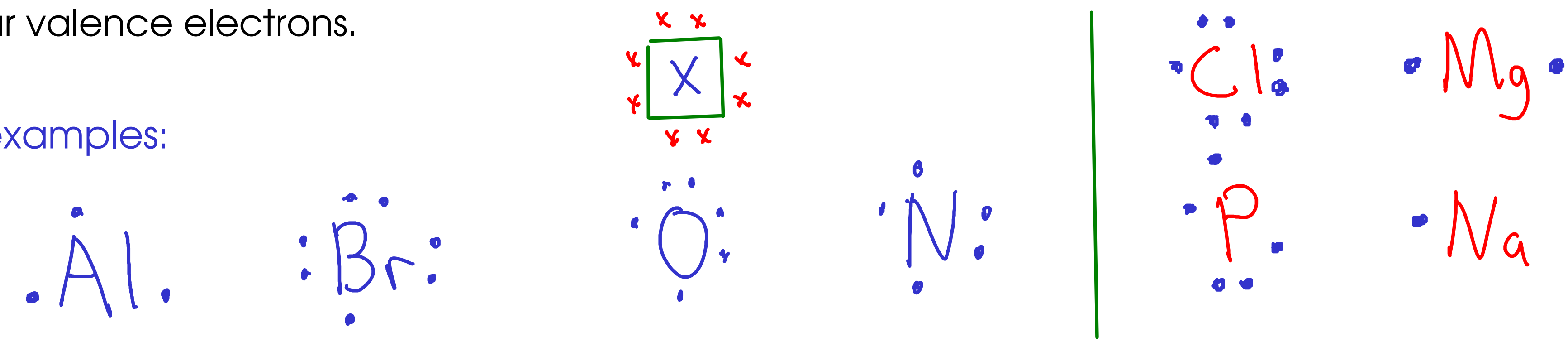
... but using electron configurations to describe how aluminum bromide forms is a bit cumbersome! Can we simplify the picture a bit?

LEWIS NOTATION / ELECTRON-DOT NOTATION

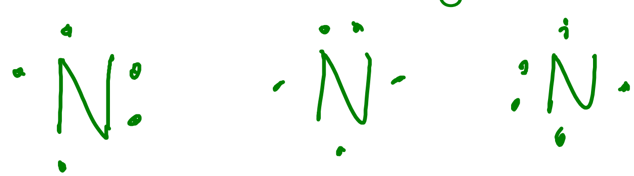
- Lewis notation represents each VALENCE electron with a DOT drawn around the atomic symbol. Since the valence shell of an atom contains only "s" and "p" electrons, the maximum number of dots drawn will be EIGHT.

- To use electron-dot notation, put a dot for each valence electron around the atomic symbol. Put one dot on each "side" of the symbol (4 sides), then pair the dots for atoms that have more than four valence electrons.

examples:

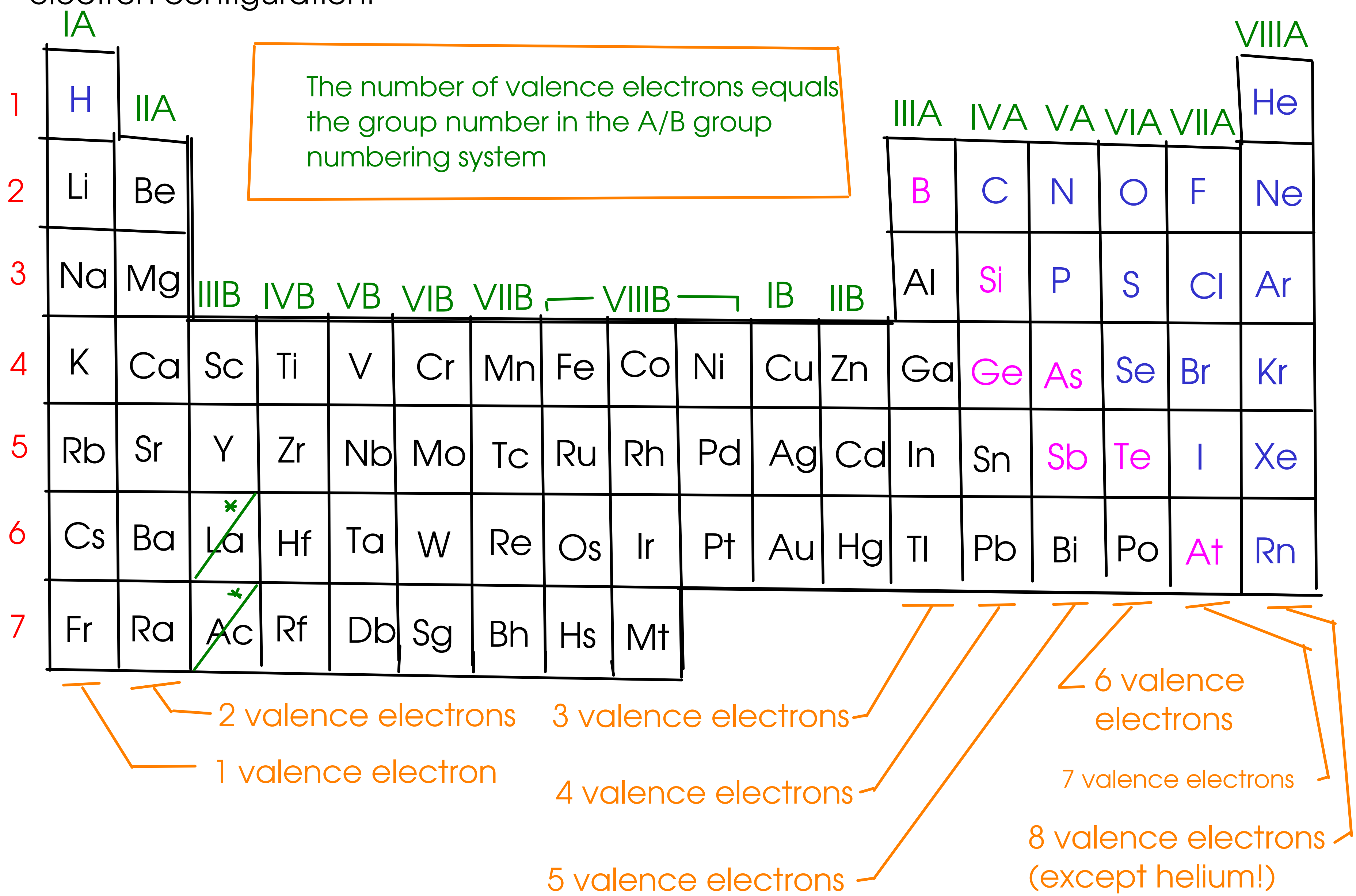


Which "side" you draw the dots on isn't important, as long as you have the right number of electrons and the right number of "pairs"

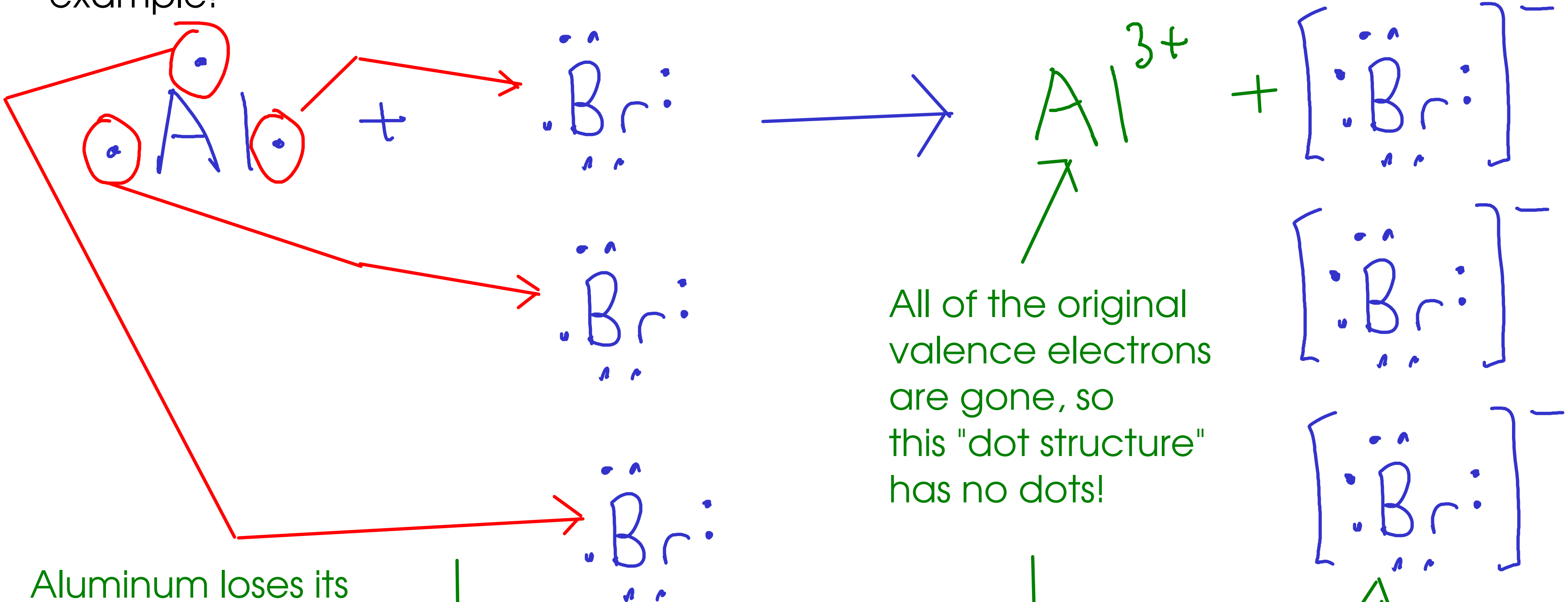


... . are all equivalent!

To draw a dot structure for an atom, you need to know HOW MANY valence electrons it has! You can determine this simply from the periodic table, WITHOUT writing the whole electron configuration!



... but how do we use this to describe a reaction that produces ions? Let's look at our previous example!



All of the original valence electrons are gone, so this "dot structure" has no dots!

Aluminum loses its outer three electrons (in the 3s and 3p subshells) (leaving it with full 2s and 2p subshells.

Each bromine atom requires one more electron to get a total of eight outer electrons (full "s" and "p" subshells)

Each bromide ion has eight outer electrons!

... this is a bit easier to follow than looking at all those letters and numbers in the electron configurations for these elements!

This is an OXIDATION-REDUCTION (electron transfer) reaction. Dot notation makes the transfer of electrons very obvious.