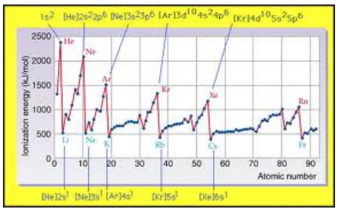
Trends

def: Patterns which repeat themselves.



- The periodic table has repetition of patterns from one period to the next called: **periodicity of properties.**
- **Examples are**: boiling point, melting point, reactivity.

Example of a pattern   Not a pattern

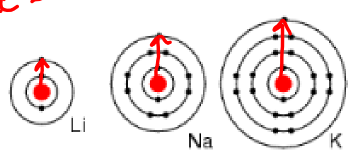


Important trend variables

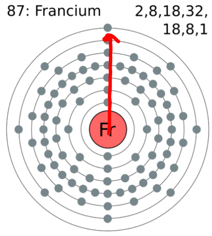
1. The number of orbits

The more orbits an element has, the greater distance there is between the protons in the nucleus and the valence electrons on the last orbit. This will cause a weaker attraction between protons (nucleus) and valence electrons.

more orbits
= Less attraction

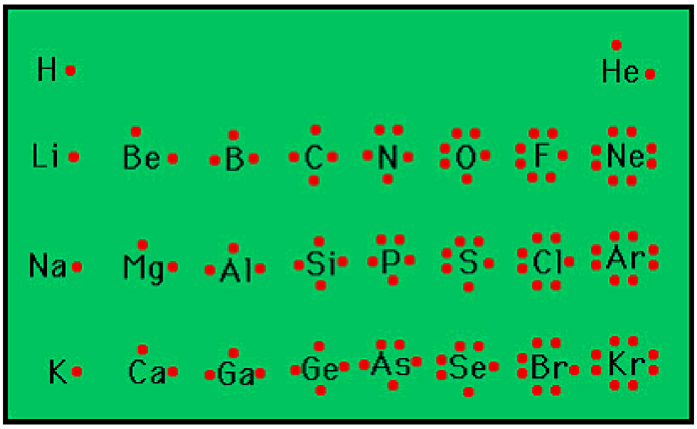


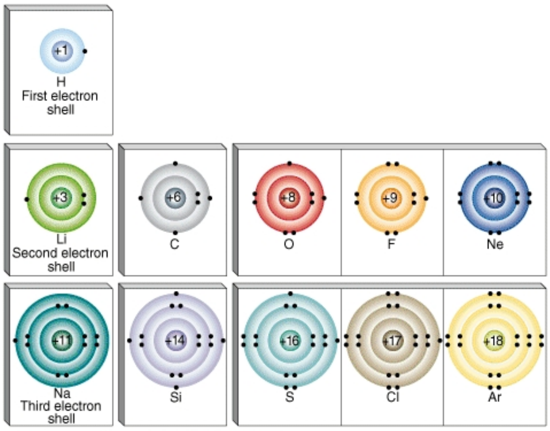
Periodic table group IA elements: Li, Na, and K



2.The number of valence electrons

The more valence electrons an element has on its last orbit, the more attraction there will be between the valence electrons and the protons in the nucleus.





- More orbits = Less attraction
- More ve = More attraction

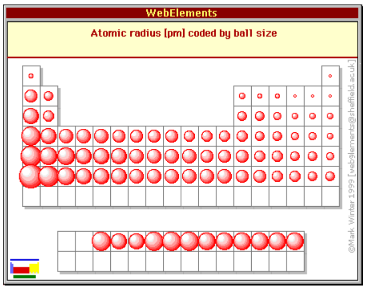
**These 2 variables will cause the trend to increase or decrease as it goes across the period.

Trend Examples

family 2 atomic radius trend.mp4

1- Atomic Radius (size of the atom)

Trend is ← ↓



Why do elements get bigger as you go down the group?

As you go down the group, the elements will have MORE ORBITS.

btw

MORE ORBITS = more distance between valence electrons and protons (weak attraction).

VE

Weak attraction = atom has more space to increase in size.

Why do elements get bigger as you move towards the alkali metals?

As you go from across the period from right to left, the elements contain less valence electrons.

Less Valence electrons = weaker attraction between valence electrons and protons.

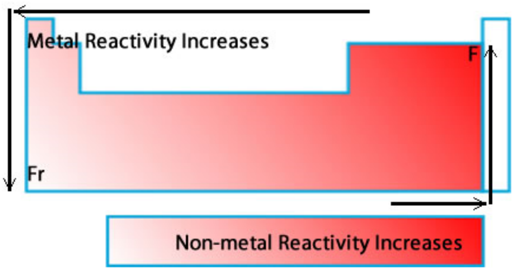
Weaker attraction means the atom has more space to increase in size.

More orbits = weaker attraction = bigger

***As the # of protons and electrons increases, the protons power of attraction for the electrons increases (atoms shrinks and tightens).

The atomic radius decreases from the 1st element to the last in the period (from left to right) because there are more valence electrons so they are more attracted to protons.

2- Reactivity



Metal: Why as you go down the group?	More orbits = less attraction between the valence electrons and protons = it gets easier to "get rid" of the valence electrons so there is a large chemical reaction when electrons are donated.
Metal: Why as you go to the alkali (right to left)	Less valence electrons means there is less attraction to protons so higher tendency to react
Non-metal: Why as you go up the group?	Less orbits = more attraction between valence electrons and protons (there are more electrons closer to nucleus) = strong chemical reaction when electrons are accepted
Non-metal : Why as you go to the halogens (to the right)	More valence electrons = more attraction between the valence electrons and protons = strong reaction when accepting the electron (less work to accept so react faster and bigger)

Which metal has the strongest chemical reaction?

Francium

Which non-metal has the strongest chemical reaction?

Fluorine

Perioidic Trends in Ionization Energy.mp4

3- ionization energy	4- electronegativity
now much energy is required to remove valence electron(s) from the atom. (stronger to the right of the table and up the group)	The degree to which an atom tends to gain electrons to form a chemical bond (increases from left to right of the table and up the group)

Why stronger as you go to the noble gases? (to the right)	Why greater as you go up the group?
The more valence electrons an atom has, the more attraction to the protons in the nucleus. More energy is required to REMOVE an electron. **Stronger ionization:** atoms need more energy to remove more electrons. **Higher Electronegativity:** **A**s you move to the halogens, the easier it becomes to gain electrons (Strong pull from nucleus to gain electron)	LESS ORBITS **Strong Ionization:** Less orbits = valence electrons are closer to the nucleus so there is more attraction. It takes more energy to remove the electrons. **Higher Electronegativity:** Less orbits = stronger pull from nucleus to gain the electrons to become stable More orbits means weaker pull from nucleus to gain electrons

- Group 8 has strongest ionization energy because HIGH amount of energy would be required to lose electrons
- Group 8 has no electronegativity because they do not attract electrons since their orbits are full.

a) Which element has the weakest ionization energy?
Francium (one valence electron with extremely low attraction to nucleus- not a lot of energy required to lose electron)

Strongest ionization energy?

He (one orbit with full shell - attraction to nucleus is extremely strong because electrons are very close to protons and holds on to all electrons)

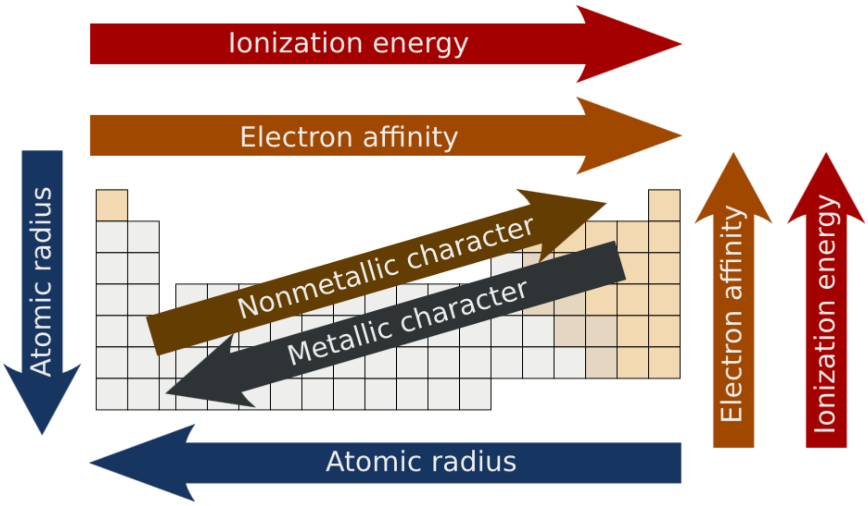
b) Which element has the weakest electronegativity energy?

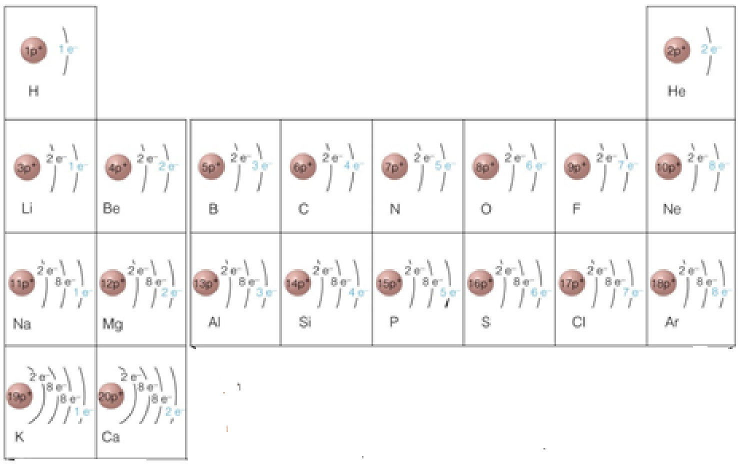
Francium - less likely to have enough energy to gain 7 electrons and too many orbits so weak attraction to nucleus

Strongest electronegativity energy?

Fluorine -strong ability to attract electrons because it only needs one to become stable and the attraction between protons and electrons is very strong because they are closer (pull from nucleus to become stable is stronger)

All trends put together





1	2	13	14	15	16	17	18
H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca	Ga	Ge	As	Se	Br	Kr
Rb	Sr	In	Sn	Sb	Te	I	Xe
Cs	Ba	Tl	Pb	Bi	Po	At	Rn
Fr	Ra						

Warm up questions:

1- Explain why F has a greater ionization than Cl.

2- Explain how Mg is smaller than Na.

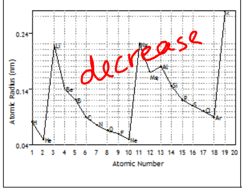
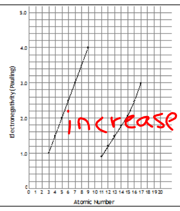
3- Explain why K is bigger than Na, but smaller than Rb.

4- Explain why S is more reactive than P, but less than Cl.

5- Explain why S is more reactive than Se, but less than O.

Past exam questions

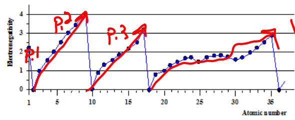
1. The graphs below show the measurement of atomic radius and the measurement of electronegativity of certain elements as a function of their atomic number.

Graph 1 Atomic Radius versus Atomic Number	***Graph 2*** Electronegativity versus Atomic Number
	

According to the graphs, which of the statements below is TRUE?
A) Both atomic radius and electronegativity increase from left to right across a period.
B) Both atomic radius and electronegativity ~~decrease~~ ↑ from left to right across a period.
C) The atomic radius increases ↓ and electronegativity decreases ↑ from left to right across a period.
D) The atomic radius decreases ✓ and electronegativity increases ✓ from left to right across a period.

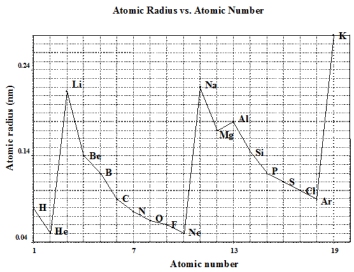
2. The graph below shows the electronegativity of some elements. Describe: the progression of this property for elements within the 3rd period on the periodic table.

As atomic # ↑



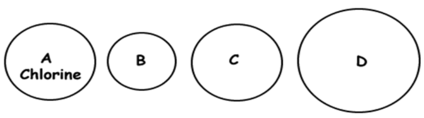
V.E ↑
Electronegativity is stronger

3. Consider the graph below



Based on this graph, which of the following statements is correct?
A) The atomic radius increases across the period and decreases down a group.
B) The atomic radius decreases across the period and increases down a group.
C) The atomic radius increases across the period and increases down a group.
D) The atomic radius decreases across the period and decreases down a group.

4. Given the representation of a chlorine atom, which circle might represent an atom of sulfur?



Attachments

family 2 atomic radius trend.mp4

Perioidic_Trends_in_Ionization_Energy.mp4